SCH 3U1

Unit 1 Matter, Chemical Trends, and Chemical Bonding (18 periods)

Day 1 - diagnostic test, atomic theory, atomic structure review (atomic #, mass # atomic mass) pg36
- comparison of subatomic particles table 2.1 pg 35
- calculating # of p, n and electrons(complete charts) pg 37 #1, bohr-rutherford diagrams
- define isotopes and radioisotopes pg 37 ; fraction of an isotope versus abundance
- assignment page 39 (1,2, and 3)

Day 2 - periodic law (pg 40), arrangement of electrons(pg 43), valence e's, review of periodic table layout (pg 41) and review group names, metals, nonmetals and metalloids
- lewis structures to represent valence e's(pg 46)
- assign page 42 # 2; page 46 (3,5,6); page 48 (3,6 and 11)

Day 3 - periodic table trends (size, ionization energy, electron affinity)pg 49
- complete assignment 2 A page 50
- do practice problems with students page 52, 55,
- assignment page 60 (2-5) page 61 (6, 11)

Day 4 - Video : reactivity of metals ;Teacher demos , review sheets, alien periodic table

Day 5 - Quiz chp 2 ; chemical bonding, ionic and covalent definitions (pg 70), electronegativity (pg71) and predicting bond type (pg72); properties of ionic and covalently bonded compounds(pg 67) , Complete pg 68 Ionic or Covalent
- Assign page 74 (4,5,6) pg 76 (2,3) Read careers in chemistry pg 77.

Day 6 - Using diagrams to represent ionic bond formation (pg 75)
- octet rule ; practice problem page 76 # 2; account for properties of ionic compounds(pg 78).

Day 7/8 - Use diagrams to represent covalent bonding; define pure covalent bond; state the diatomic elements; Polar covalent bonds (page 86 (11 -13 draw lewis structures for single, double and triple bonds. Identify and sketch the 5 basic shapes of molecules;
Watch video www.teachersdomain.org : molecular shapes
-do practice problems page 81 (6, 7) and page 82 (8-10)
- assign page 84 (1,2,3,4 and 6) page 94 (1, 4)

Day 9 - Do lab modeling molecules page 92. Assignment

Day 10 - Research Assignment : Computer lab.

Day 11 - Writing formulas and naming compounds; know valence chart page 97) do practice problems pg 97(14,15)page 99 (16, 17)pg 100(18,19) complete handouts given. Know common names for compounds and uses see page 102

Day 12 - Binary compounds multivalent cations pg 103 (20,21), pg 106 (2,4); binary acids

Day 13 - Know the following polyatomic ions (see charts pages 97,98) Naming using the stock system and alternative system, naming acids, prefix method for naming compounds containing 2 or more non-metals, hydrated salts pg 103 (20,21) pg 105(23) pg 106 (1,3)

Day 14 - Naming using the stock system and alternative system, naming acids, prefix method for naming compounds containing 2 or more non-metals
pg 105(24) do review page 107 (2,3,7,8,9,14,15,16,17,18,19,22)

Day 15 - Unit Review pg. 154 (8,9,11,28,29,30,31,32,40,43,44-47,49) pg 647 (8-21)

Day 1

Review Quiz :

True or False? Each of the statements below is either true or false. If the statement is true, write the letter T beside its number. If false, write the letter F and rewrite the statement so that it is true.

1. A particle that gains electrons is called a cation.

2. The symbol 27/13Al represents an atom that has 27 protons and 13 neutrons.

3. Elements that have the same number of energy levels have similar chemical properties.

4. The Lewis structure for an atom of neon, Ne, has 8 dots around the element symbol.

5. The atom has a sharply defined outer boundary.

6. Elements in the periodic table are arranged in order of increasing atomic mass.

7. In the periodic table, elements that have similar properties are grouped in horizontal rows.

8. Of the elements, helium, iron, sulfur, calcium, and potassium, the one that would react in a way that is most like sodium is calcium.

9. Neutrons and electrons have nearly identical masses.

10. In the chemical notation $\frac{A}{Z}$ X, *A* is the atomic symbol and *Z* is the mass number.

11. Neutral atoms have the same number of protons and electrons.

12. There is no limit to the number of electrons that can occupy the energy levels surrounding an atom.

13. Given the choice of aluminum, magnesium, oxygen, helium, and hydrogen, the element that is most stable is oxygen.

14. Hydrogen-1, hydrogen-2, and hydrogen-3 are elements that have the same number of protons but different numbers of electrons.

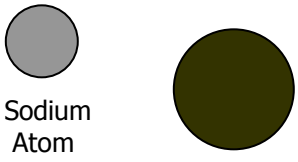
15. The Lewis structure for an element that has an atomic number of 12 would have 2 dots, with both dots placed together on the same side of the element symbol.

16. A fluorine ion has 7 valence electrons.

Models of Matter

- All matter is composed of tiny particles called _____
- John Dalton devised the first modern concept of the atom.
- Scientists have revised Dalton's Theory based on more recent discoveries:

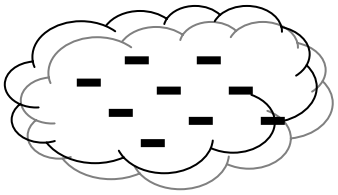
Dalton's Atomic Theory
" Billiard Ball Model" (1803)



Sodium Atom

Copper Atom

Thomson's Atomic Theory
"Plum Pudding Model" (1897)

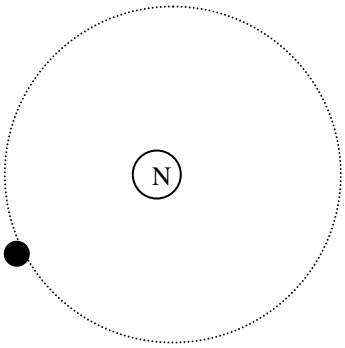


Electrons embedded in a

_____ charged cloud

Rutherford's Atomic Theory "Nuclear Atom Model" (1911)

_____charged nucleus
with a negatively charge
electron orbiting it.



Complete table below :

Model	Weakness

- **Mosely** - atoms of different elements were bombarded with electrons and the wavelengths of the X-rays emitted by the atoms were studied.
- He found that the wavelengths were related to an integer he called the _____number (Z)

Review of Atomic Number and Mass Number

A ← MASS NUMBER (#p + #n)

X ← Atomic Symbol

Z ← ATOMIC NUMBER (#p)

#neutrons (n) = A - Z

Element Name	Chemical Notation	Number of Protons	Number of Electrons	Number of Neutrons
Aluminum				
Beryllium				
		5		
			18	
Carbon				
			17	18
		19		
	$^{4}_{2}He$		2	
Lithium				
				0
		12		
Neon				
	N			
	$^{16}_{8}$			8
		15		
		28		
			14	
				16
	Na			

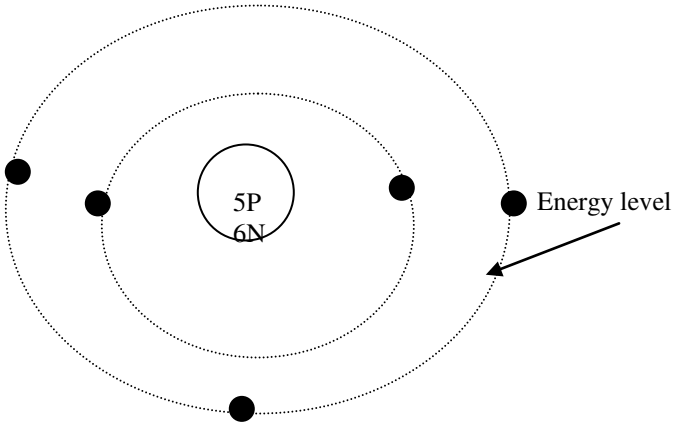
- **comparison of subatomic particles table 2.1 pg 35**

Properties of Subatomic Particles						
Subatomic Particle	Symbol	Actual Charge (in Coulomb)	Relative Charge	Actual Mass (in grams)	Relative Mass	Mass (in atomic mass units*)
Electron	e^-	1.60×10^{-19}	-1	9.109×10^{-28}	1/1837	0
Proton	p^+	1.60×10^{-19}	+1	1.673×10^{-24}	1	1
Neutron	n^0	0	0	1.675×10^{-24}	1	1

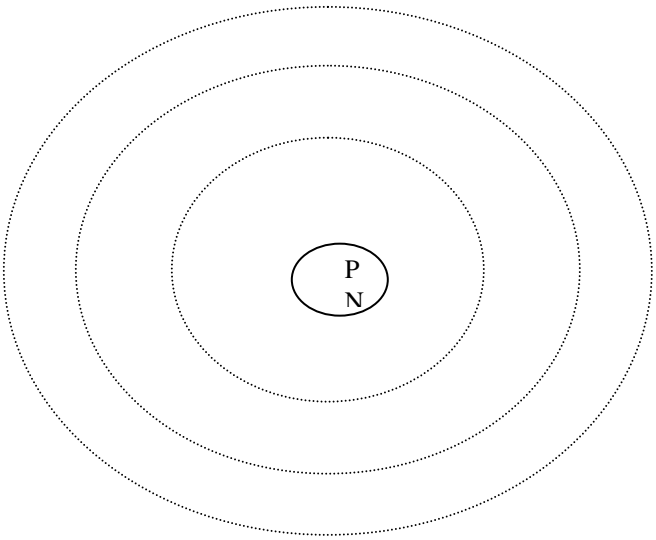
<u>Orbits – Niels Bohr</u> : Solar System Model (1913)

- The Bohr-Rutherford model of an atom is a two-dimensional model, which shows the electrons orbiting around the nucleus. These ORBITS are also referred to as _____ LEVELS. Each orbit can only hold a certain number of electrons.

	Energy Level/Orbit (n)	Maximum Number of Electrons ($2n^2$)
Lowest energy ←	1 (K) – closest to the nucleus	2
	2 (L)	8
	3 (M)	8 (used for convenience, really 18)
Highest energy ←	4 (N)	8 (used for convenience, really 32)



Example #1: Draw a Bohr diagram for Na (Z = 11)



- **define isotopes and radioisotopes pg 37**

ISOTOPES - _____

- Most elements have at least 2 naturally occurring isotopes.
- For this reason, the atomic mass of an element is not a whole number but rather a decimal.
- The _____ _____of an element is a number which represents the average weighted mass of all the isotopes of an element.
- Periodic table states *Average Atomic Masses*, which take into account *relative mass* of isotopes of each element, as well as *relative abundance*.

Symbol	Isotope	Protons	Neutron	Electrons	Abundance
35Cl	Chlorine-35	17	18	17	75.77%
37Cl	Chlorine-37	17	20	17	24.23%

Calculating Average Atomic Mass

When calculating average atomic mass, the fraction of each isotope must be determined.

$$\text{fraction of an isotope} = \frac{\text{abundance (\%)}}{100\%}$$

Once the fraction of each is known, the following formula can be used:

$$\text{average atomic mass} = (\text{relative atomic mass A} \times \text{fraction of A}) + (\text{relative atomic mass B} \times \text{fraction of B})$$

RADIOISOTOPES (pg 38)

- The nuclei of some atoms are _____
- Henri Becquerel, Marie Curie, and Pierre Curie were awarded the Nobel prize in 1903 when they discovered that the radioactive element _____ would decompose into the elements Polonium & Radium.
- The _____ atom (mass # 238) releases a great deal of energy (in the form of radiation) as it breaks down into 2 lighter atoms; Polonium (mass # 209) and Radium (mass # 222). Isotopes which are radioactive are known as **radioisotopes**
- An element may have one or more isotopes which are not radioactive in addition to a radioisotope

There are approximately 2000 known isotopes. Most of these isotopes are of radioactive elements that emit invisible rays. These elements have unstable nuclei and emit rays in an attempt to become stable.

Radioactive isotopes have an application in medical therapy. Cobalt-60 sources were made for the treatment of malignant tumors. These cobalt-60 sources produce intense beams of _____rays, which destroy tumors.

Many radioactive isotopes have been used as tracers. Thus, a small amount of a radioactive isotope can be mixed with the non-radioactive element. This radioactive "tag" enables doctors to trace the movement of an element through the human body by using a Geiger counter, which detects the rays from the radioactive atoms. (i.e. a solution of sodium iodide mixed with iodine-131 can be injected to measure the uptake of iodine by the thyroid)

Radioisotopes can be used to control pests. Male insects that have been exposed to a nonlethal dose of radioactivity become sterile and therefore cannot fertilize eggs laid by the females. Population drops.

The _____ radiation from cobalt-60 is used to sterilize fruits, vegetables and grain, as well as medical supplies and wool. For example, potatoes that are exposed will not sprout and therefore have a longer shelf life.

Carbon-14 is a radioactive nuclide constantly produced in the atmosphere. It has a half-life of 5730 years, and undergoes beta emission, decaying to nitrogen-14. During photosynthesis, green plants absorb carbon in the form of carbon dioxide. A percentage of this carbon dioxide is made from carbon-14. Once the plant dies, photosynthesis stops, and no more radioactive carbon dioxide is absorbed. However, decay of the carbon-14 continues. Careful measurements of the amount of radioactivity of carbon-14 in a once-living plant yields the approximate time in history when the plant died.

Deuterium is the name given to hydrogen with a mass number of 2. It boils at 101.41 degrees Celsius and melts at 3.79 degrees Celsius. The rates of chemical reactions involving deuterium (heavy water) are slower than similar reactions involving plain water. Heavy water is used as a coolant and as a moderator in CANDU nuclear reactors. A moderator slows neutrons enough to ensure the reactor fuel (uranium) can capture them more readily, causing further nuclear reactions.

Assignment pg 37 # 1 , page 39 (1,2, and 3)

Day 2 - periodic law (pg 40), arrangement of electrons(pg 43), valence e's, review of periodic table layout (pg 41) and review group names, metals, nonmetals and metalloids
- lewis structures to represent valence e's(pg 46); bohr-rutherford diagrams
- assign page 42 # 2; page 46 (3,5,6); page 48 (3,6 and 11)

Organizing Matter

- As scientist were discovering different elements, it became apparent that they needed a way to organize them. At first they were organized by their atomic mass (find 3 places on the table where the mass decreases left to right)
- Mendeleev – the "founder" of the periodic table, did this.

periodic law –

energy levels – represented as horizontal rows also called _____; maximum e's for each level is n = 1 ____n=2 ____n=3 ____n=4 _____ general formula ____

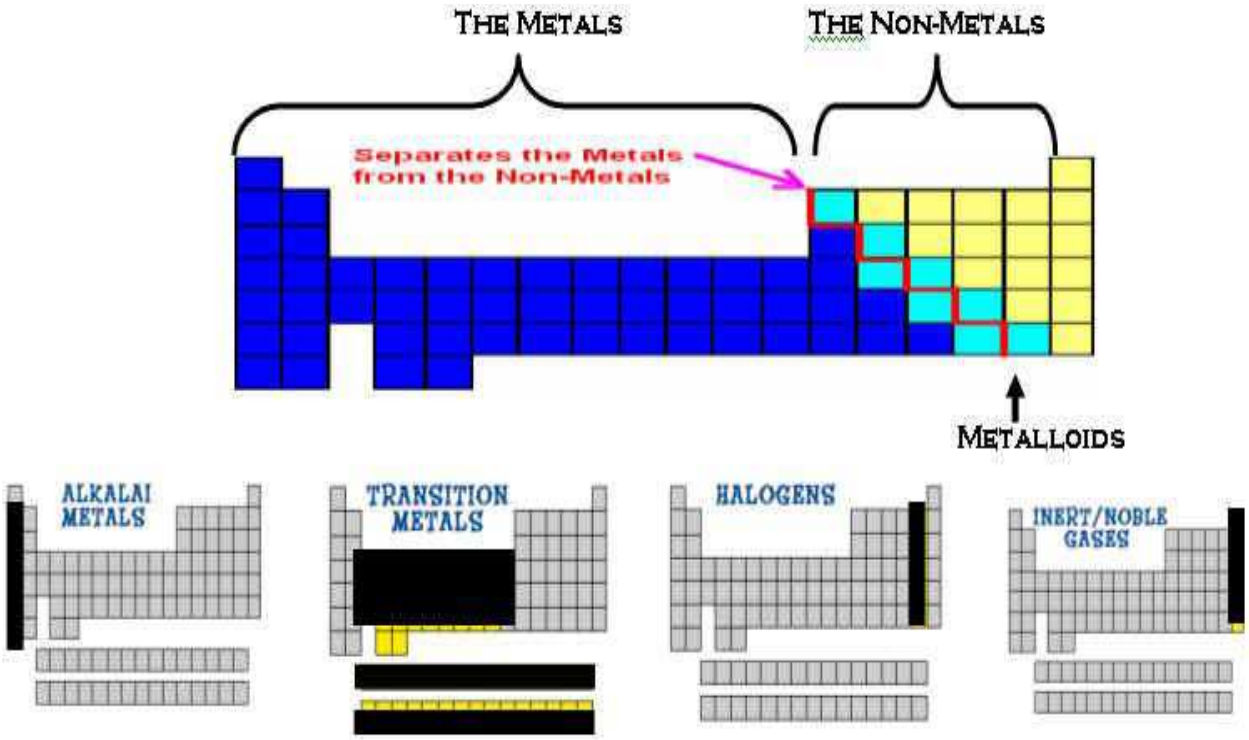
valence e's - _____ e's ; obtained by the vertical columns called _____; these are numbered from 1 - ____ ; the valence e's is the group # but if a double digit like 13 the # of valence e's is the 2nd number

Main group elements – groups _____(Representative elements)

Transition elements – groups _____

Alkali metals – group ___ ; alkaline earth metals – group ____; halogens – group _____; noble gases- group _____

- maximum e's in outer shell is _____; known as _____ rule ; very stable
- in the table below place above the roman numerals #'s from 1 to 18.



Lewis Dot Diagrams (Electron Dot Diagrams)(pg 46)

Because it is very time consuming to write out the entire Bohr diagram for an atom, we use a shortened version which only considers those electrons in the VALENCE SHELL (outermost shell). These electrons are called VALENCE ELECTRONS.

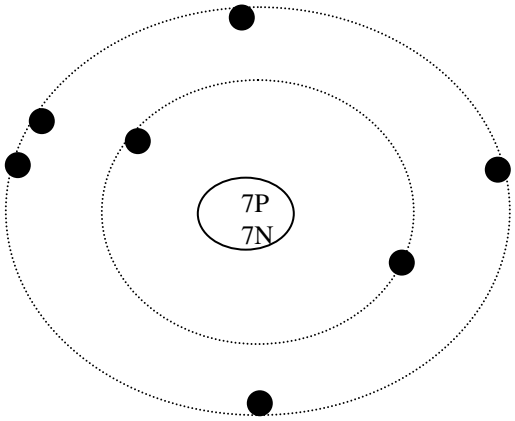
Na

Method:

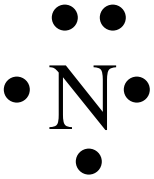
1. The valence electrons are represented by DOTS. The rest of the atom (core) is represented by the element symbol.
2. The symbol has 4 sides (max. 2 electrons per side).
3. Each side receives a single electron before they start to pair up.

Examples:

Nitrogen (Rutherford-Bohr)



Nitrogen (Lewis dot)



page 42 # 2; page 46 (3,5,6)

2.

3. **BOHR-RUTHERFORD DIAGRAMS**

$^{1}_{1}H$							$^{4}_{2}He$
$^{7}_{3}Li$	$^{9}_{4}Be$	$^{11}_{5}B$	$^{12}_{6}C$	$^{14}_{7}N$	$^{16}_{8}O$	$^{19}_{9}F$	$^{20}_{10}Ne$
$^{23}_{11}Na$	$^{24}_{12}Mg$	$^{27}_{13}Al$	$^{28}_{14}Si$	$^{31}_{15}P$	$^{32}_{16}S$	$^{35}_{17}Cl$	$^{40}_{18}Ar$
$^{39}_{19}K$	$^{40}_{20}Ca$						

5.

6.

page 48 (3,6 and 11)

Day 3 - periodic table trends (size, ionization energy, electron affinity)pg 49
- complete assignment (see below)
- do practice problems with students page 52, 55,
- assignment page 60 (2-5) page 61 (6, 11)

Trends in the Periodic Table **Name:** _____

Purpose
Determine the trends, if they exist, for atomic size and ionization energy in the Periodic Table.

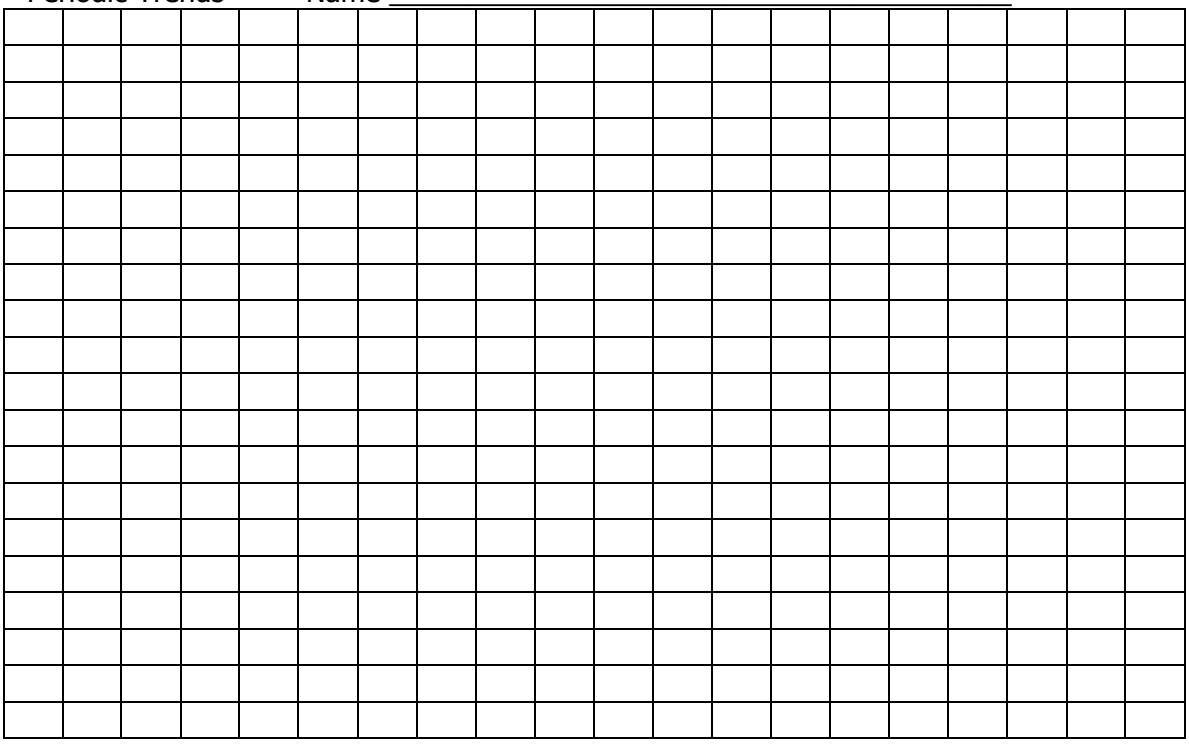
Materials
Graph paper

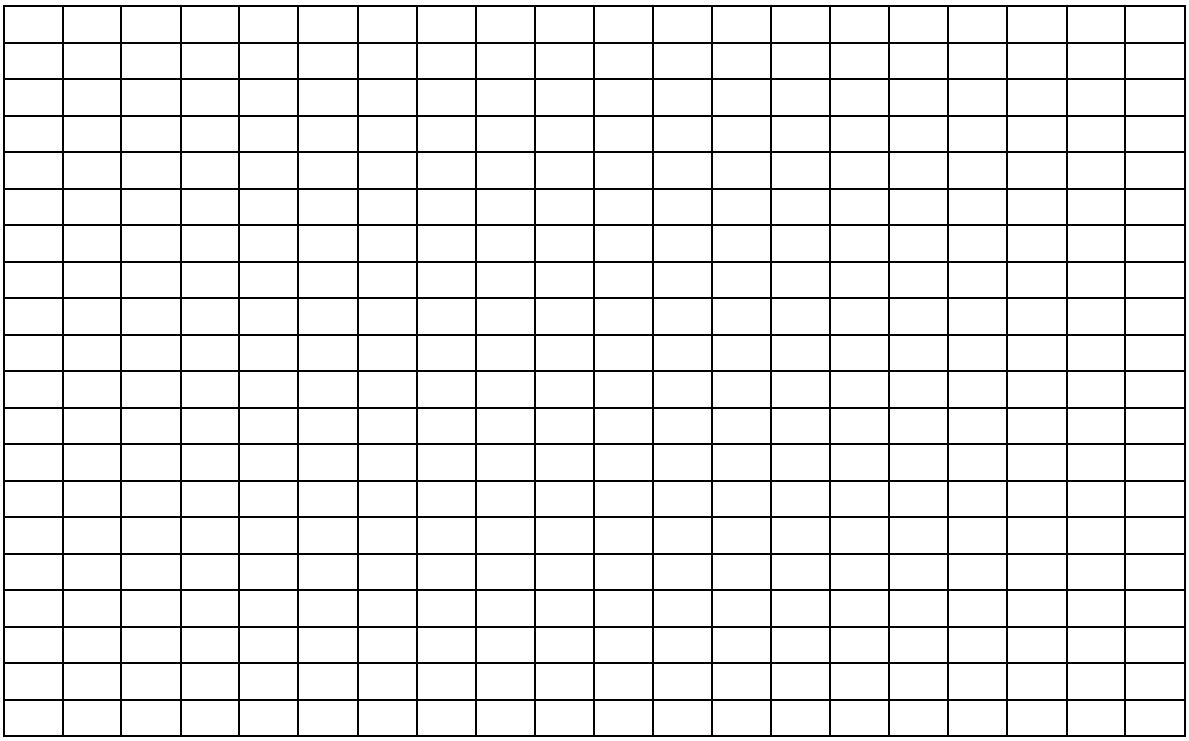
Procedure
1. Use the information from the section of the periodic table. Be sure to give each graph a title and to label each axis.
2. For elements 3-20, make a graph of atomic radius as a function of atomic number. Plot atomic number on the X axis and atomic radius on the Y-axis.
3. For elements in Family 1A (1) and Family 2A (2), graph period number vs. atomic radius. Use a different color or symbol for each line.
4. For elements 3-20, make a graph of ionization energy as a function of atomic number. Plot atomic number on the X-axis and ionization energy on the Y-axis.
5. For elements in Family 1(1A) and Family 2 (2A), graph period number vs. ionization energy. Use a different color or symbol for each line.

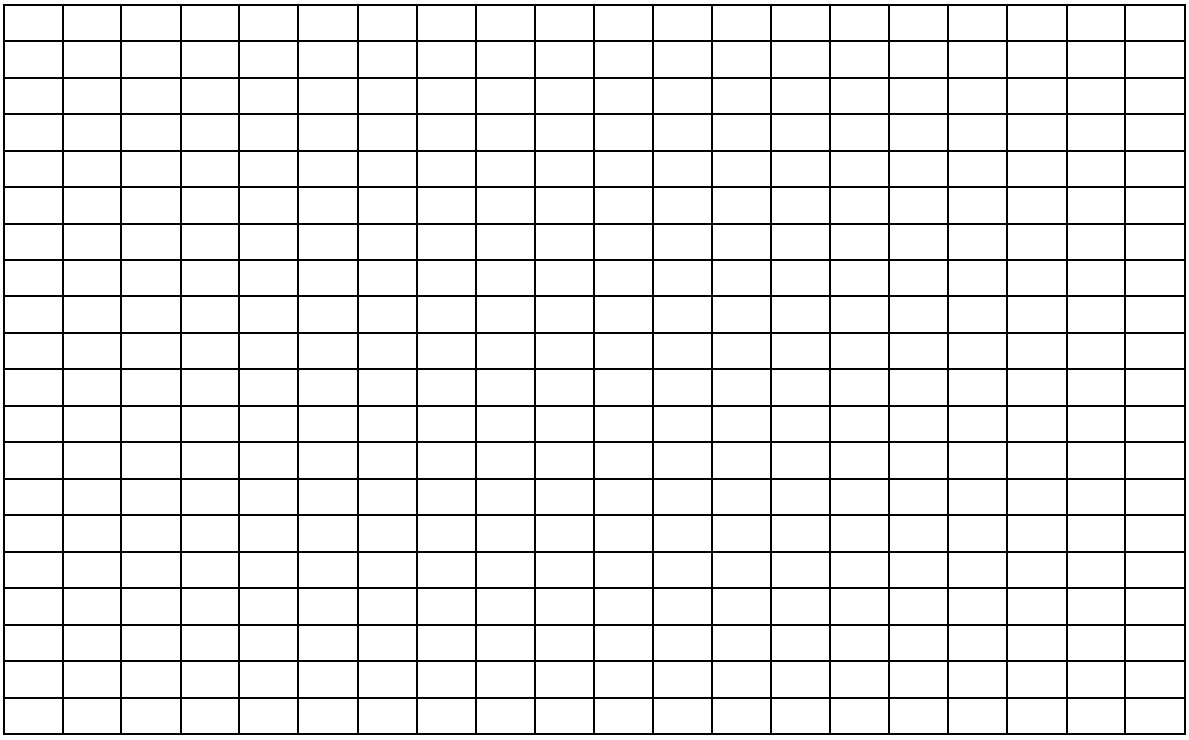
	IA (1)	IIA (2)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	VIIIA (18)
2	3 Li 1.23 124	4 Be 0.89 215	5 B 0.80 191	6 C 0.77 260	7 N 0.70 335	8 O 0.66 314	9 F 0.64 402	10 Ne 0.67 497
3	11 Na 1.57 119	12 Mg 1.36 176	13 Al 1.25 138	14 Si 1.17 188	15 P 1.10 242	16 S 1.04 239	17 Cl 0.99 299	18 Ar 0.98 363
4	19 K 2.03 100	20 Ca 1.74 141						
5	37 Rb 2.16 96	38 Sr 1.91 131						
6	55 Cs 2.35 90	56 Ba 1.98 120						

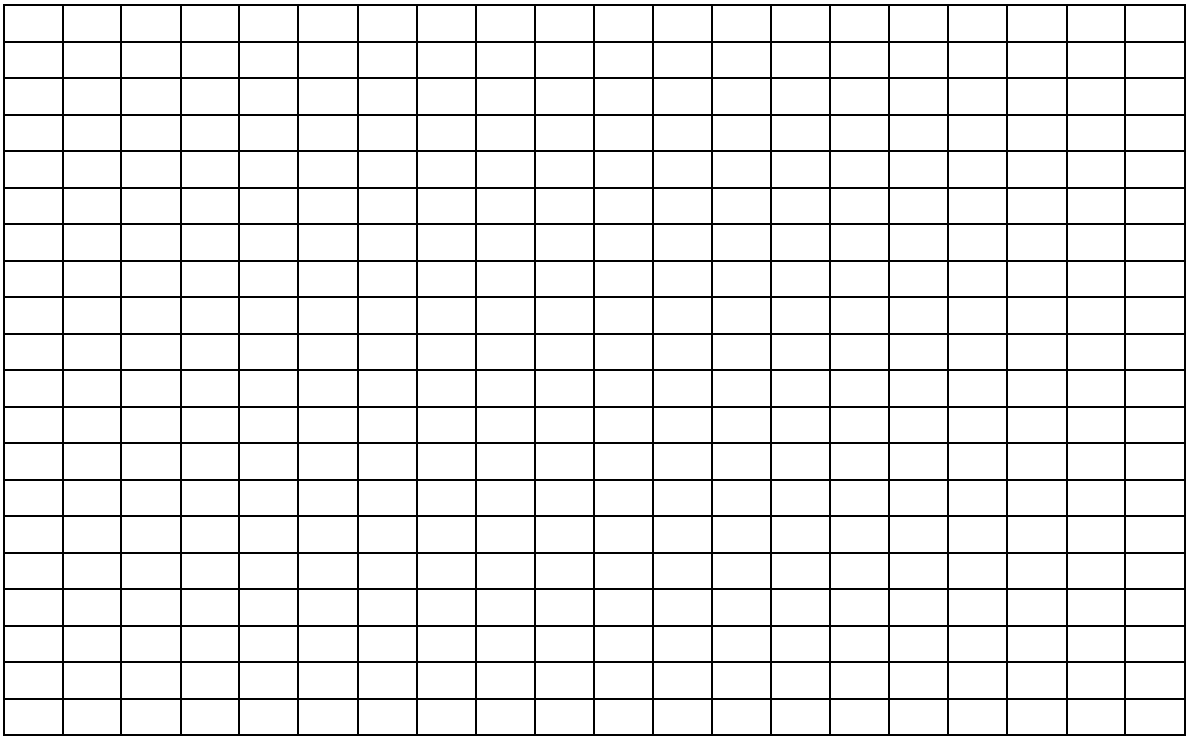
8	~~Atomic~~ number
O	~~Symbol~~
0.66	Atomic radius
314	Ionization energy

Periodic Trends Name _____









Analysis

1. What happens to the atomic radius as the atomic number increases across a period? Down a family?

2. What happens to the ionization energy as the atomic number increases across a period? Down a family?

Conclusion

1. Why does atomic radius change as it does?

2. Why does the ionization energy change as it does?

Lab Summary Paragraph

_____

_____

_____

_____

_____

_____

_____

_____

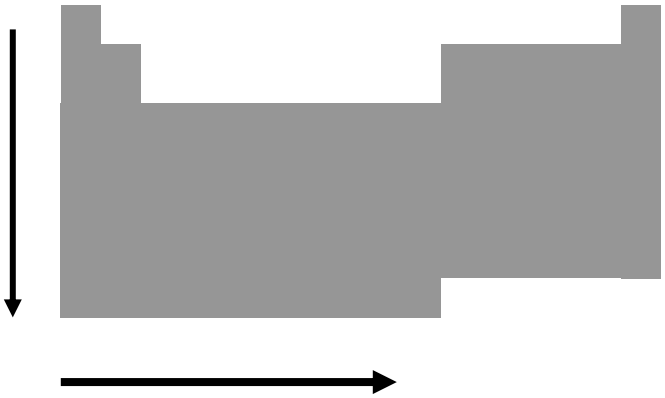
TRENDS IN THE PERIODIC TABLE (pg 49)

Periodic Trends: (Reactivity Trends)

Chemical Reactivity (non metals)

Down A Family: Reactivity _____

Across A Period: Reactivity _____



Chemical Reactivity: (Metals)

Down A Family: Reactivity _____

Across A Period: Reactivity _____

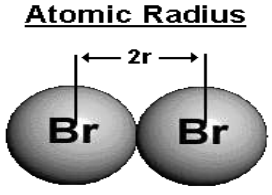
Which elements are more reactive?

a) Li or Be b) F or S c) C or Si d) He or F e) Rb or Ca f) Pb or Sn

g) As or Se h) Which is the most reactive metal? _____

i) Which is the most reactive nonmetal? _____

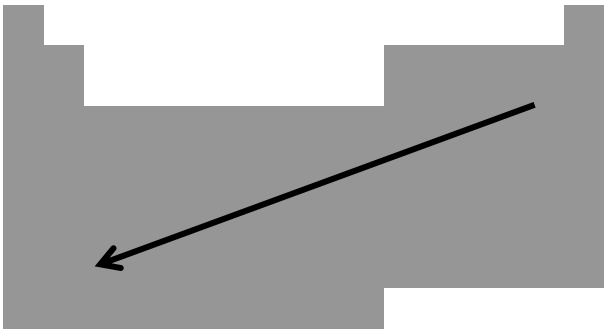
Atomic Radius(pg 52)



The atomic radius is one half the distance between the nuclei of two identical atoms in a diatomic molecule

a) - *going across a period* the A.R. _____
-electrons are added at the same energy level with some _____between electrons
-electrostatic attraction between positive nucleus and negative electrons is _____
-as positive charge of nucleus increases, attraction _____
-electrons are pulled closer to the nucleus

b) -*going down a family* the A.R. _____
-there are more _____ _____ so distance between positive nucleus and negative electrons _____; _____ forces of attraction
-inner electrons shield the outer electrons from the full positive charge of the nucleus
-large increase from Group 18 to group 1 (e.g. Ne to Na) because the single electron added to a new energy level is shielded by the _____ electrons in the lower level



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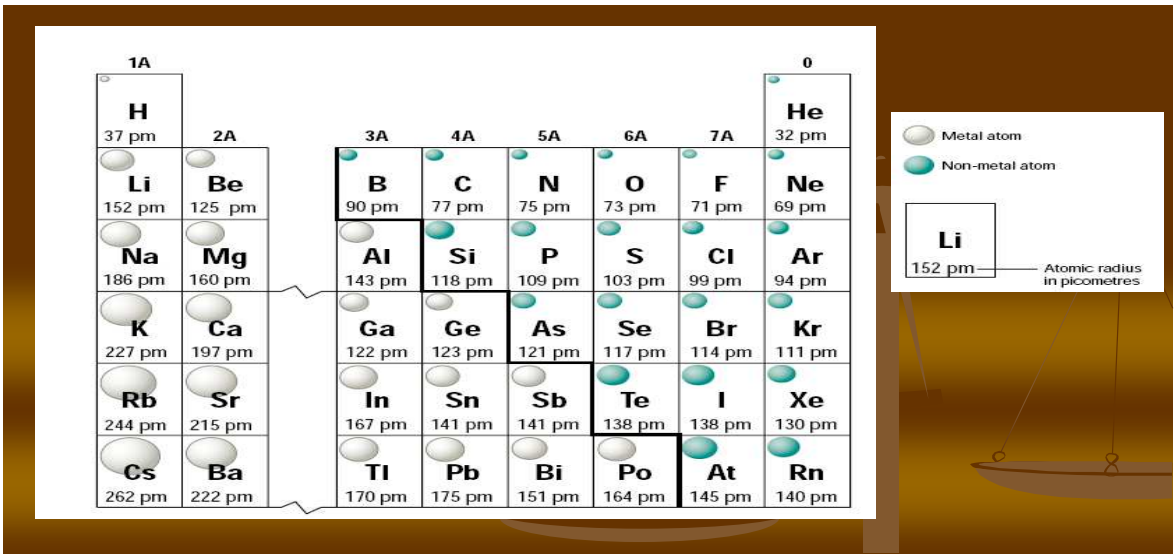
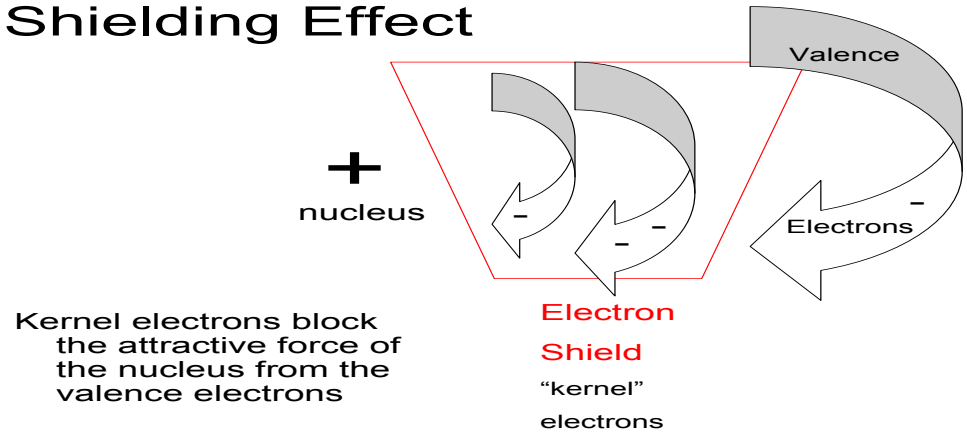
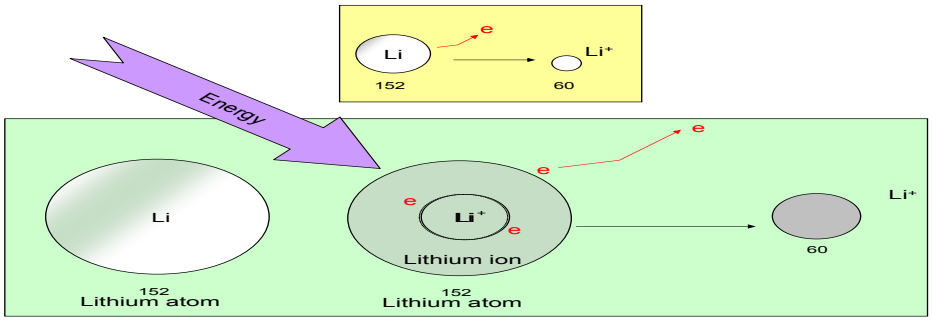




Ionization Energy (I.E.) (pg 53)

-the energy required to remove the _____ electron of an atom of an element in its _____state ; positive ion is _____ than the atom it came from

X (g) + energy (ionization energy) -------> X^+ (+ve charged ion cation) + 1 e^-



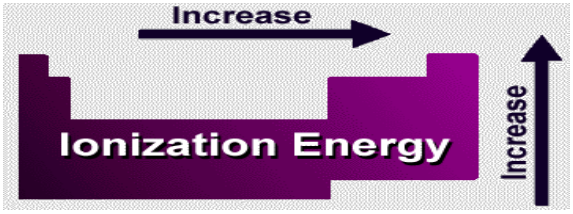
going across the period

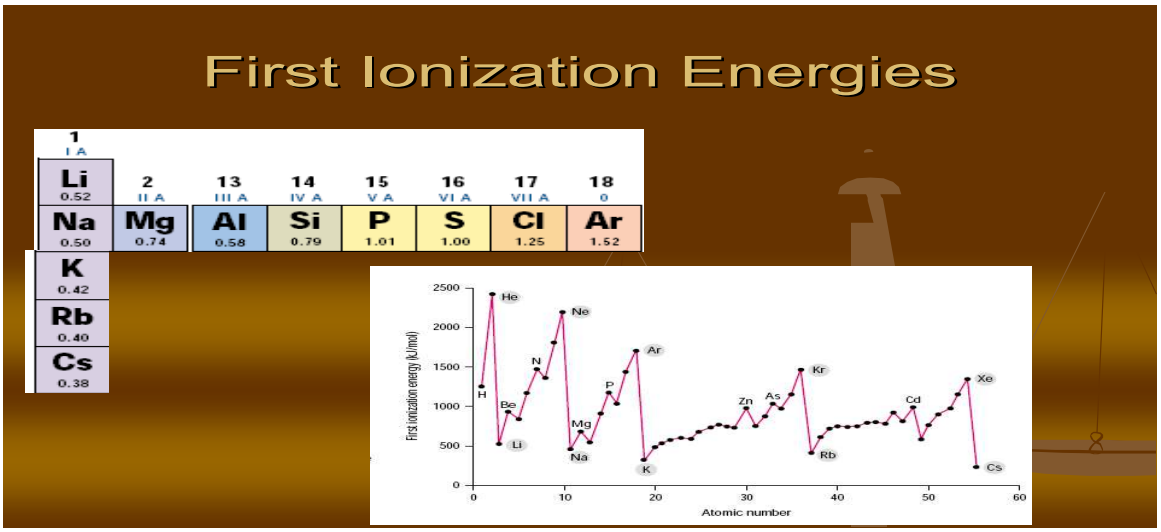
- I.E. _____
-energy level stays the same but number of protons and electrons _____
-electrostatic force of attraction _____;_____ energy is needed to remove an electron

going down a family

- I.E. _____
-more energy levels going down family
1. electrons are _____from the nucleus
2. _____by inner electrons

-so _____ in attractive forces;_____ energy to remove an electron





Electron Affinity

-the amount of energy _____when an atom of an element in the gaseous state _____ an e^- X (g) + 1 e^- --------> X^- + energy

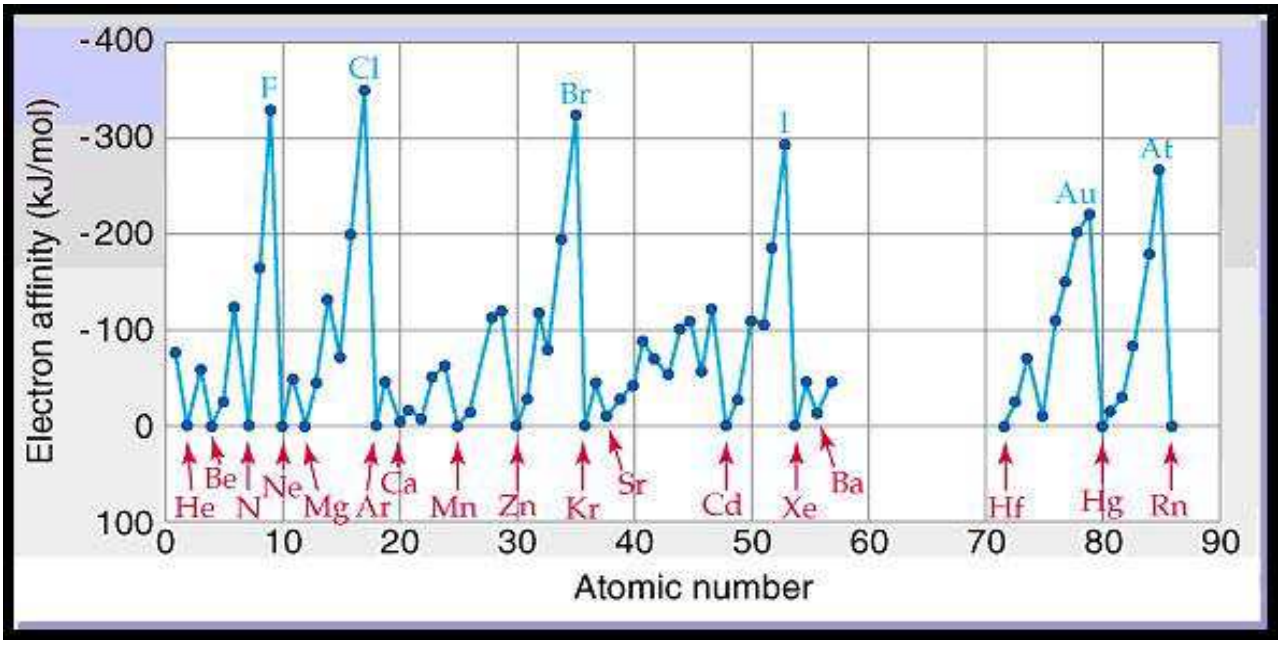
- negative ions are _____ than the atom it came from.

going across a period (L to R)

-the E.A._____, coming closer to noble gas configuration;_____ attraction by nucleus

going down a family

-E.A. _____; the e^- to be gained would be placed further away from the nucleus; _____ in attractive forces



Reactivity

metals -reactivity _____going down the family and _____ going across the period

non-metals -reactivity _____ going down the family

1. Metals react by _____electrons. The _____the ionization energy, the _____reactive the metal is.

2. Non-metals react by _____electrons. The higher the electron affinity, the _____these non-metals attract extra electrons.

3. Noble gases are very _____because they have stable electron arrangements and do not easily lose or gain electrons. Helium has _____ electrons in its full outer energy level and other have _____ electrons.

Do practice problems with students page 52 (7), pg55 (8,9), assignment page 60 (2-5)

TRENDS IN THE PERIODIC TABLE - SUMMARY SHEET

NAME	DEFINITION	TREND	EXPLANATION
ATOMIC RADIUS	Distance measured from the centre of the nucleus to the outermost e^- in pm or A^o	1. _____down a group 2. _____across a period from left to right	1. Increase in no. of energy _____ and electrons - more _____ 2. e^- held more tightly, _____in ENC (effective nuclear charge), _____ shielding _____radius
FIRST IONIZATION ENERGY	Energy required to remove the _____electron from a _____ (state) atom	1. _____down a group 2. _____across a period from left to right	1. _____in radius due to _____energy levels ,electrons _____tightly held 2. e^- held_____, _____in ENC _____ shielding , _____to remove e^-
ELECTRONEGATIVITY (relates to electron affinity)	The tendency to _____electrons	1. _____up a group 2. _____across a period from left to right not including Group _____	1. _____energy levels, _____atomic radius .˙. _____ attraction of electrons 2. _____ENC, _____attraction for electrons,_____ shielding
REACTIVITY METALS	The degree to which metals have a tendency to react with other substances by _____electrons	1. _____down a group 2. _____across a period from left to right	1. _____energy levels, _____atomic radius .˙. _____attraction of e^- -electrons _____easily removed 2. _____ENC, _____attraction for e^-, _____shielding, _____easily removed
REACTIVITY NONMETALS	The degree to which nonmetals have a tendency to react with other substances by _____electrons	1. _____up a group 2. _____across a period from left to right not including Group 18	1. _____energy levels, _____ atomic radius .˙. _____attraction of electrons 2. _____ENC, _____attraction for electrons, _____shielding _____ atomic radius

Day 4 - Video : reactivity of metals ;Teacher demos , review sheets, alien periodic table

<u>Teachersdomain.org video metals in hydrochloric acid</u>

An acid is a chemical compound that, when dissolved in water, has a pH of _____ than 7.0. Acidic compounds contain positively charged _____ ions—hydrogen atoms containing protons but no_____which readily react with other compounds. Examples of common acids are acetic acid, or vinegar, and sulfuric acid, which is used in car batteries. Monoprotic acids, which include hydrochloric acid, only give up one proton per acid molecule, while polyprotic acids can give up more than one proton per acid molecule.

When bare-metal surfaces are immersed in an acidic aqueous solution (a solution in which the solvent is water), the hydrogen ions in the acid attract and bond with negatively charged electrons from the metal to produce what is called an oxidation-reduction reaction. Hydrogen gas, which forms as two released electrons attach to two hydrogen ions, bubbles off, and metal atoms (now positively charged ions) are released into the solution. Through this oxidation-reduction reaction, the metal is said to corrode, or gradually wear away.

Metals consist of stacked layers of tightly packed, interacting atoms arranged in geometric patterns. Whereas non-metals tend to _____ electrons in chemical reactions, metals tend to _____ them. The gauge of just how readily a metal loses electrons is known as its redox potential, which is determined by its atomic structure.

Electrons inhabit various energy levels, or_____. The electron configuration shown in the periodic table indicates how many electrons are found in each shell, from innermost to outermost. Atoms of elements in the left-hand column have one electron in their outer shell, while atoms of elements in the right-hand column have eight electrons in their outer shell. Generally speaking, single electrons in an outer shell can easily be taken away from the atom with the application of very little energy. This makes atoms of elements in the left-hand column very reactive. On the other hand, it is very difficult to add or remove electrons from an atom that has eight electrons in its outer shell. The atoms of these elements, found in the column to the far right, are non-reactive. Knowing this helps explain why some metals, such as tin and copper, corrode less readily than others, such as magnesium and iron.

Because some metals are more prone to corrosion, engineers have learned to use certain techniques during an item's manufacture to protect it from structural decay. One method involves plating, or painting over a metal with a coating. This technique prevents electrochemical reactions from occurring. When exposed to oxygen in air or water, for example, two elements, aluminum and chromium, form a self-renewing, microscopically thin layer, known as an oxide film, which protects the underlying metal from corrosion. The alloy stainless steel, known for its sleek, shiny surface and tremendous strength, contains a high concentration of chromium.

Metal	Observation	Ranking (lowest = 1)
Cu		
Mg		
Zn		
Fe		
Ag		
Ni		
Pb		
Sn		

Questions :

Explain using electron structure why silver does not react in HCl?

What determines the reactivity of zinc?

How would you classify the level of reactivity of the various metals in HCl?

Can you explain why some metals reacted in HCl while others did not? Could the Periodic Table help?

How would you explain what happens on the molecular level in terms of electron movement?

Activity Series :

Most reactive to least reactive : _____

Teacher Demo : Reactivity of lithium and sodium in water

Observations :

page 61 (6, 11)

Student Worksheet: Atomic Structure

1.Complete the following table

Element Name	Chemical Symbol	Z number	A number	# of protons	# of neutrons	# of electrons
		36	84			36
				35	45	35
		53	127			54
		27			32	27
Zinc						
	Cd^{2+}		112			
				38	50	36
	X^{2-}				75	54
Calcium ion						
	X^{3+}			26	56	

Student Worksheet: Alien Periodic Trends

The following alien elements belong together in families as grouped below. The elements as listed ARE NOT necessarily in order. The Letters ARE NOT the normal Symbols for the element, therefore the alien periodic table. ZRD, SIFP, JXBE, LHT, QKA, WOV, YMC, GUN

The assignment is to arrange these elements in the proper periodic form, according to the information given below.

Fill in the answers in the periodic table provided. Place the symbol, the Lewis dot diagram, oxidation state in the appropriate box of the periodic table. Label each period and group on the periodic table as well

1. U has a total of six electrons
2. A is the second most common element in the atmosphere
3. E is a noble gas
4. S is an alkali metal with a 1+ charge
5. O is a halogen with a 1- charge
6. O has an atomic number larger than V but smaller than W
7. The outer most electrons of L are found in the S^2 orbital
8. C has five electrons in its outer energy level
9. The boiling point of T is more than H but less than that of L
10. M has one less proton than that of A
11. N is found in the third energy level
12. R has the largest atomic radii of the group
13. F is a gas at room temperature
14. Atom B contains 10 protons
15. Q has an atomic radius slightly larger than that of K
16. Y contains more metallic properties than either M or C
17. X has an atomic number one higher than F
18. D has highest ionization energy of its group
19. P is the most reactive element in its family
20. J has the greatest density in its family
21. Atoms of I are larger than those of S

Day 5 - Quiz chp 2 ; chemical bonding, ionic and covalent definitions (pg 70), electronegativity (pg71) and predicting bond type (pg72); properties of ionic and covalently bonded compounds(pg 67) , Complete pg 68 Ionic or Covalent

- Assign page 74 (4,5,6) pg 76 (2,3) Read careers in chemistry pg 77.

CHEMICAL BONDS (pg 70)

1. Metals react chemically by _____electrons thus forming _____ ions.

2. Non-metals react chemically by _____ electrons thus forming _____ ions.

3. Chemical bonds in compounds containing a metal & non-metal element consist of the electrostatic attraction between these oppositely charged _____. These are called _____ **bonds**

4. Metal atoms lose all of their electrons in the outer energy level when they react, resulting in a noble gas electron arrangement. Non-metal atoms gain electrons when they chemically react so as to complete their outer energy level. As a result, the non-metal atom obtains a noble gas electron arrangement.

5. Atoms may also _____ electrons to achieve a noble gas configuration. This is called a _____**bond.**

6. Covalent bonds will occur between two non-metals, although they can sometimes occur between a non-metal and a metal

ELECTRONEGATIVITY (pg 71)

Electronegativity is a measurement of an atom's relative attraction for a pair of electrons in a chemical bond. Each element has been assigned an electronegativity value (EN). This value follows the same trend as electron affinity.

Trend Across a Period

Electronegativity _____ across a period of the periodic table because there is an _____ in positive nuclear charge which will increase the pull on the electrons.

Trend Down a Group

Electronegativity _____ down a group of the periodic table because the number of shielding electrons increases, _____ the attraction of the bonding electrons by the nucleus.

PREDICTING BOND TYPES

1. An **ionic bond** will occur when one of the bonding atoms has _____than half the maximum number of electrons (1, 2, or 3). These bonds generally occur between a metal and a non-metal atom.

2. A **covalent bond** will occur when both bonding atoms have half or more of their maximum number of valence electrons (4, 5, 6, or 7) Covalent bonds generally form between the atoms of 2 non-metals.

3. The _____ in electronegativity between 2 bonding atoms may be used as an aid in determining whether the bond is ionic or covalent.

   a) If the difference in electronegativity (ΔEN) between 2 bonding atoms is greater than 1.7, the bond is likely to be _____. The greater the difference, the more ionic the bond becomes.

   b) If ΔEN is less than 1.7, the bond is likely to be _____.

   c) If the ΔEN is 1.7, the bond may possess both ionic and covalent characteristics; _____ % ionic

d) If the ΔEN is 0, the bond is non-polar covalent. As the ΔEN value increases from 0, the bond gains a _____ nature.

Table of Electronegativities

			H 2.1						
D									
E									
C									
R									
E									
A									
S									
E									
	Li 1.0	**Be 1.5**		**B 2.0**	**C 2.5**	**N 3.0**	**O 3.5**	**F 4.0**
	Na 0.9	**Mg 1.2**		**Al 1.5**	**Si 1.8**	**P 2.1**	**S 2.5**	**Cl 3.5**
	K 0.8	**Ca 1.0**		**Ga 1.6**	**Ge 1.8**	**As 2.0**	**Se 2.4**	**Br 2.8**
	Rb 0.8	**Sr 1.0**		**In 1.7**	**Sn 1.8**	**Sb 1.9**	**Te 2.1**	**I 2.5**
	Cs 0.7	**Ba 0.9**		**Tl 1.8**	**Pb 1.8**	**Bi 1.9**	**Po 2.0**	**At 2.2**

INCREASE

Intermolecular & Intramolecular Forces of Attraction

Intermolecular: The force of attraction between the molecules of a compound. These forces are relatively weak.
e.g. force of attraction between water molecules in water

Intramolecular: The force of attraction between the atoms in a molecule.
e.g. the covalent bonds holding the two H and one O atoms together within a water molecule

Properties of Covalent Compounds (pg. 67)

1. Forces of attraction between molecules are much _____ compared to those between ions. Thus, much _____ energy is required to separate the molecules in a covalent compound. This results in _____ melting and boiling points.

2. Covalent compounds exist as solids, liquids, or gases, depending on the relative strength of the intermolecular attraction. In the cases of liquids and gases, the intermolecular bonds are quite _____compared to solids. If the covalent bond is polar the intermolecular forces of attraction will be higher, increasing melting and boiling points.

3. Covalent compounds _____conduct electricity in the liquid state because the liquid _____ possess charged particles.

4. Covalent compounds are generally _____ in water. The polar water molecules have very little or no effect on the uncharged molecules of a covalent compound.

5. Since covalent compounds are not normally soluble in water, they do not form solutions that conduct electricity. This is due to the absence of charged particles in the water.

Properties of Ionic Compounds

1. Ionic compounds have relatively _____melting and boiling points. Bonds between oppositely charged ions are extremely strong making it difficult to separate ions.

2. Ionic compounds usually exist as _____ at room temperature due to the relatively strong forces of attraction between the ions. The 3-dimensional arrangement of alternating cations and anions usually results in the formation of a crystalline structure called a lattice. Thus ionic compounds are usually crystalline, rather than amorphous.

3. Ionic crystals _____ conduct electricity in the solid state because of the strong bonds preventing movement of charged particles. However, in the liquid state attractive forces are weaker, thus allowing for the movement of charged particles. Ionic compounds _____ conduct electricity in the liquid (molten) form.

4. Ionic crystals are _____ in water. Since water is a polar molecule, it attacks a lattice and pulls it apart. Once removed from the lattice, an ion is quickly surrounded by water molecules.

5. When ionic crystals dissolve in water, the quantity of charged particles in the solution increases dramatically. So, solutions which contain dissolved ionic crystals are quite good conductors of electricity and are referred to as _____.

Complete pg 68 Ionic or Covalent

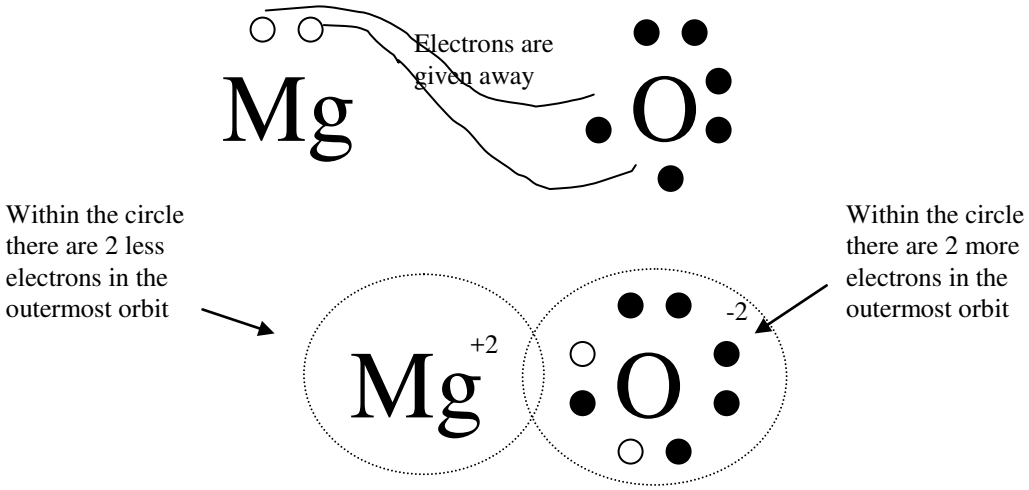
#1 _____ _____bonding **#2 _____ _____ bonding**

#3 _____ _____bonding **#4 _____ _____ bonding**

#5 _____ _____bonding

Assign page 74 (4,5,6) pg 76 (2,3) Read careers in chemistry pg 77.

Day 6 - Using diagrams to represent ionic bond formation resulting in lewis structures(pg 75)
- octet rule ; practice problem page 76 # 2; account for properties of ionic compounds(pg 78). pg 78 (4,5)Watch animation teachersdomain : ionic bonding ; ionic bonding worksheet



Formation of ionic compound : NaCl

Steps :

1. Metal loses e's forming positive ion
2. Non-metal gains e's forming a negative ion
3. Multiply equations by appropiate factors so that e's lost = e's gained
4. Oppositely charged ions attract

1. $Na° + energy ----> [Na]^+ + 1\ e$

2. $Cl + 1\ e \text{ -----> } [Cl]^- + energy$

3. $Na^+ + Cl^- \text{ -----> } NaCl + energy$

Try Magnesium chloride :

1. $Mg : + energy \text{ ----> } [Mg]^{2+} + 2\ e$

2. $Cl + 1\ e \text{ -----> } [Cl]^- + energy$

3. Multiply equ. 2 by then add

4. $Mg^{2+} + 2\ Cl^- \text{ ----> } MgCl_2 + energy$

Try aluminum oxide :

To determine the formulas of simple ionic compounds ; place the oxidation #'s of each element and then cross them over
+3 -2
e.g. Al O ------>

Try magnesium nitride

Electron dot diagrams : Ionic Compounds e.g. [Na]$^{+}$ [Cl]$^{-}$

Name	Formula	Lewis Structure
Sodium oxide		
Magnesium oxide		
Aluminum Chloride		
Sodium nitride		
Aluminum phosphide		

octet rule – atoms try achieve _____ electrons in the outer ; in the case of H _____

page 76 # 2,3; pg 78 (4,5) Reading assignment : account for properties of ionic compounds(pg 78).

- **Watch animation teachersdomain : ionic bonding**

Lewis dot – Ionic Bonding worksheet

Use Lewis diagrams to show the bonding that is taking place between the following atoms:

1. K and F

2. Li and P

3. Al and O

4. Sr and S

5. C and F

6. Na and I

7. Ca and Br

8. Ga and Cl

9. Rb and As

10. Mg and O

11. Ba and N

12. Cs and Se

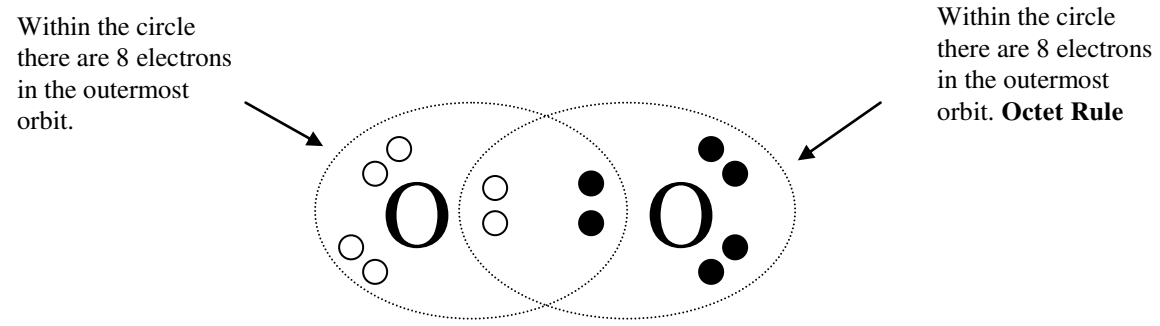
13. Ca and P

14. K and Si

15. Be and I

Day 7/8 - **Use diagrams to represent covalent bonding; define pure covalent bond; state the diatomic elements; draw lewis structures for single, double and triple bonds. Indicate 5 basic shapes of molecules**
-do practice problems page 81 (6, 7) and page 82 (8-10)
- **assign page 84 (1,2,3,4 and 6) Complete pg 86 (11,13) ; page 94 (1, 4) : video covalent bonding;** Watch video www.teachersdomain.org : molecular shapes

Covalent Bonding



i. Pure Covalent or non-polar covalent (pg 81)

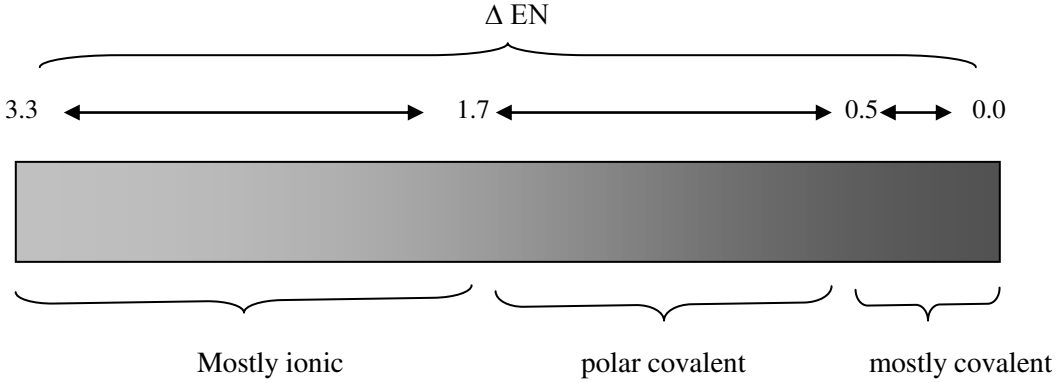
- _____ sharing of e's between atoms e.g. H_2 H : H
- bonds occur due to simultaneous attraction of both nuclei for the pair of e's
- by sharing each H has _____ e's making H_2
- the following exist in diatomic form : N_2, O_2, F_2, Cl_2, Br_2, I_2.

ii. Polar Covalent Bonding (pg 85)

- _____sharing of e's between atoms; e's spend more time closer to the more electronegative atom e.g. HCl
H - 2.2 and Cl -3.16 ; shared pair are closer to _____ thus it has a slight _____ charge and the other atom slight _____.

H Cl

- the degree of polarity can be determined by calculating the electronegativity difference



Arrange the following from least polar to most polar :
H - O , S - O , C- F , C - I, N - Cl,

_____ _____ _____ _____ _____

Complete pg 86 (11,13) Students are watch program www.teachersdomain.org : covalent bonding

Assignment

1) For each of the following, classify the bonding as being covalent, polar covalent or ionic:

a) BrCl	d) SiF_4	g) CaO	j) NaI	m) $MgCl_2$
b) CH_4	e) Cs_2S	h) OCl_2	k) H_2S	n) CCl_4
c) NH_3	f) Cl_2	i) Na_20	l) KBr	o) N_2

2) Determine which one of the following compounds has the bonds with the greatest polarity in it:

HF, CO_2, FBr, H_2O, F_2O, HI _____

3) Arrange the following compounds in order of decreasing polarity of their bonds:

HBr, H_2O, HF, CO_2, HI, HCl, SI_2 _____

4) For each of the following, show the three steps involved in the formation of the ionic compound and draw its Lewis structure:

a) lithium nitride

d) strontium chloride

g) magnesium sulphide

b) barium chloride

e) rubidium sulphide

h) potassium bromide

c) potassium oxide

f) calcium nitride

i) magnesium iodide

Shapes of Molecules Electron dot diagrams and Lewis structures (pg 87)

1. Tetrahedral (AB_4) A = central atom B = bonded atoms

-

e.g. CH_4 C - ___ valence e's (needs ____) (__ bonds)
4H - 4 x __ val. e (needs ___) (____ bond each)

Total val. e = _____ (_____ dots)

Electron dot :

Lewis Structure :

Drawing :

Bond angles

* atom with the greater need is the central atom

VSEPR Theory

- **states that the arrangement of atoms around a central atom is determined by the repulsion between the electron pairs in the valence shell of the central atom**
- **a lone pair of electrons will repel more (take up more space) than a shared pair of electrons**

2. **Pyramidal shape (AB_3E) (3 bonds and 1 lone pair)--> pg 89**

e.g. NH_3 N - _____ val e (needs _____) (__ bonds)
3H - 3 x ___ val. e (needs _____) (__ bond)

Total val. e = _____

Electron dot :

Lewis Structure :

Drawing :

Bond angles

3. **Trigonal Planar (AB_3) (3 bonds or 3 bond sites**)

note : 3 bonds will not obey octet rule ; one of the bonds is a double consisting of 4 e's

e.g. $COCl_2$ C - 4 val e(needs___) (___ bonds)
O - 6 val e (needs ___) (___ bonds)
2 Cl - 2 x ___ val e (needs ___) __ bonds

_____ val e

Electron dot :

Lewis Structure :

Drawing :

Bond angles

4. **Bent or Angular Shape (AB_2E_2) (2 bonds and 2 lone pair)(pg. 88)**

e.g. H_2O O - _____ val e (needs _____) (___ bonds)
2H - 2 x ___ val e (needs ___) (___ bonds)

Total val. e =

Electron dot :

Lewis Structure :

Drawing :

Bond angles

5. **Linear (AB or AB_2)**

e.g. Cl_2 - Cl - _____ val e (needs ___)(__ bond)
Cl - _____ val e (needs ___) (__ bond)

Total =

Electron dot :

Lewis :

Drawing :

Bond angles :

Multiple Bonds :

1. O_2 O - 6 val e (needs___)
O - 6 val e (needs___)

Total val e =

Electron dot :

Lewis :

Drawing :

Bond Angles :

- bond is a _____ bond consisting of _____ e's _____ from each atom.

2. HCN H - ___ val e (needs ___)
C - ___ val e (needs ___)
N - ___ val e (needs ___)

Total = _____

Electron Dot :

Lewis :

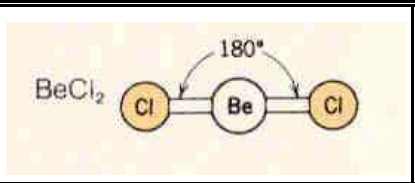
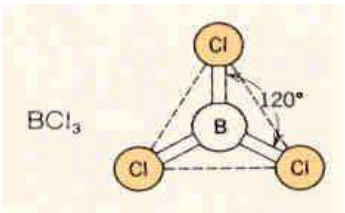
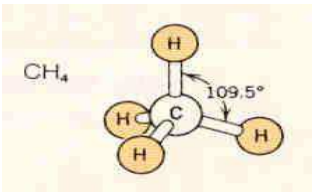
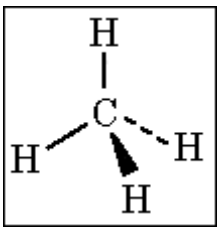
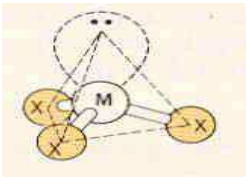
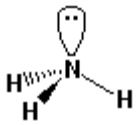
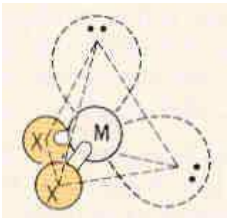
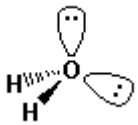
Drawing :

Shape :

Bond angles :

page 81 (6, 7) and page 82 (8-10)

SHAPES OF MOLECULES

shape	no. of electron pairs		example	bond angle
	shared	unshared		
linear (AB or ____)	**1/2**	**0**		**180°**
_____ (AB_)	**____**	**_____**		**120°**
tetrahedral ( )	**____**	**____**		**109.5°** 
_____ ( )	**3**	**_____**	**NH_3** 	**<109.5° due to lone pair repulsion** **107.3°** 
angular (bent) ( )	**____**	**_____**	**H_2O** 	**<107.3° due to 2 lone pair repulsion** **104.5°** 

Draw the following Lewis diagrams of the following covalent molecules, and then draw the structural diagram.

Name	Formula	Use	Electron Dot Diagram	Lewis Diagram
Hydrogen	H_2	Space shuttle fuel		
Silicon dioxide	SiO_2	Anti-caking agent		
Fluorine	F_2	Reacts to form fluorides		
Hydrogen chloride	HCl	Dissolves in water to form hydrochloric acid		
Water	H_2O	Universal solvent		
Hydrogen sulphide	H_2S	Poison rotten egg gas		
Ammonia	NH_3	Refrigerant at ice rinks		
Methane	CH_4	Home heating fuel		

Hydrogen peroxide	H_2O_2	Disinfectant		
oxygen	O_2	Gas required for aerobic metabolism		
Carbon dioxide	CO_2	Fire extinguisher fuel.		
Carbon disulphide	CS_2	Toxic solvent		
Phosphine	PH_3	Poisonous gas that smells like rotting fish		
Chloroform	$CHCl_3$	First anesthetics. Carcinogenic solvent.		
Hydrogen cyanide	HCN	Insecticide		
Ethane	C_2H_4	Hydrocarbon		

Complete the following Chart :

Molecule	Lewis Structure	Shape Name	Drawing
NCl_3			
HClO			
N_2Cl_4			
COS			
C_2H_2			
Cl_2O			
NH_2Cl			
H_2CO			
H_2O_2			
CF_4			
N_2F_2			
OF_2			

assign page 84 (1,2,3,4 and 6)

Watch video www.teachersdomain.org : molecular shapes

Day 9 Lab: Structure of Molecules Names

Molecule	# of valence electrons	Electron dot diagram	Lewis Diagram	# of bond pairs	# of lone pairs	Name of shape	Molecular Sketch
HCl							
NH_3							
H_2O							
CH_4							
HCN							
O_2							
CO_2							
N_2							
CH_2O							
HOCl							

CCl_4									
H_2O_2 (2 central atoms)									
CH_3OH (2 central atoms)									
N_2H_4 (2 central atoms)									
N_2H_2 (2 central atoms)									
C_2H_6 (2 central atoms)									
C_2H_4									
C_2H_2									

Day 10 - Research Assignment

1. Analyze on the basis of research , the properties of a commonly used but potentially harmful chemical substance (e.g. fertilizer, pesticide, a household cleaning product, materials used in chemical batteries) and how that substance affects the environment, and propose ways to lessen the harmfulness of the substance (e.g by reducing the amount used, by modifying one of its chemical components) or identify alternative substances that could be used for the same purpose. Consider the following issue ; many household cleaning products contain corrosive substances that can accumulate in the environment. There are now many green cleaners that do not contain these substances, although some of these products may not be as environmentally friendly as claimed.

Questions to research:

Why is it more environmentally friendly to use latex rather than oil-based paint.

Why should paint never be poured down the drain.

What properties of some common pharmaceuticals allow them to stay in the water systems and influence the growth and development of organisms. What are some ways in which this impact can be reduced.

2. Evaluate the risks and benefits to human health of some commonly used chemicals (e.g. chemical additives in foods, pharmaceuticals, cosmetics and perfumes; household cleaning products. Consider the issue : Artificial sweetners, such as aspartame, are used as sugar substitutes to reduce calories in processed foods and beverages. Although such sweetners may benefit people who are watching their weight, or those with diabetes, some experts say that their harmful effects on human health may outweigh their benefits

Questions to Research :

1. How can the use of non-stick cookware help reduce the amount of fat in our diet. What risks are associated with the use of such cookware

2. What are the risks and benefits of using sunscreens containing PABA.

3. What are the risks and benfits of using insect repellant containing DEET

4. Outline some of the concerns related to the overuse of aspartame

Day 11 - **Writing formulas and naming compounds; valence chart page 97 do practice problems pg 97(14,15)page 99 (16, 17)pg 100(18,19) complete handouts given. Know common names for compounds and uses see page 102**

Formula writing and Naming

Writing formulas and naming binary compounds; refer to valence chart page 97. Do practice problems Complete page 97 (14,15)pg 99 (16,17), pg 103 (20,21), pg 106 (2,4)
complete handouts given.

Classical System

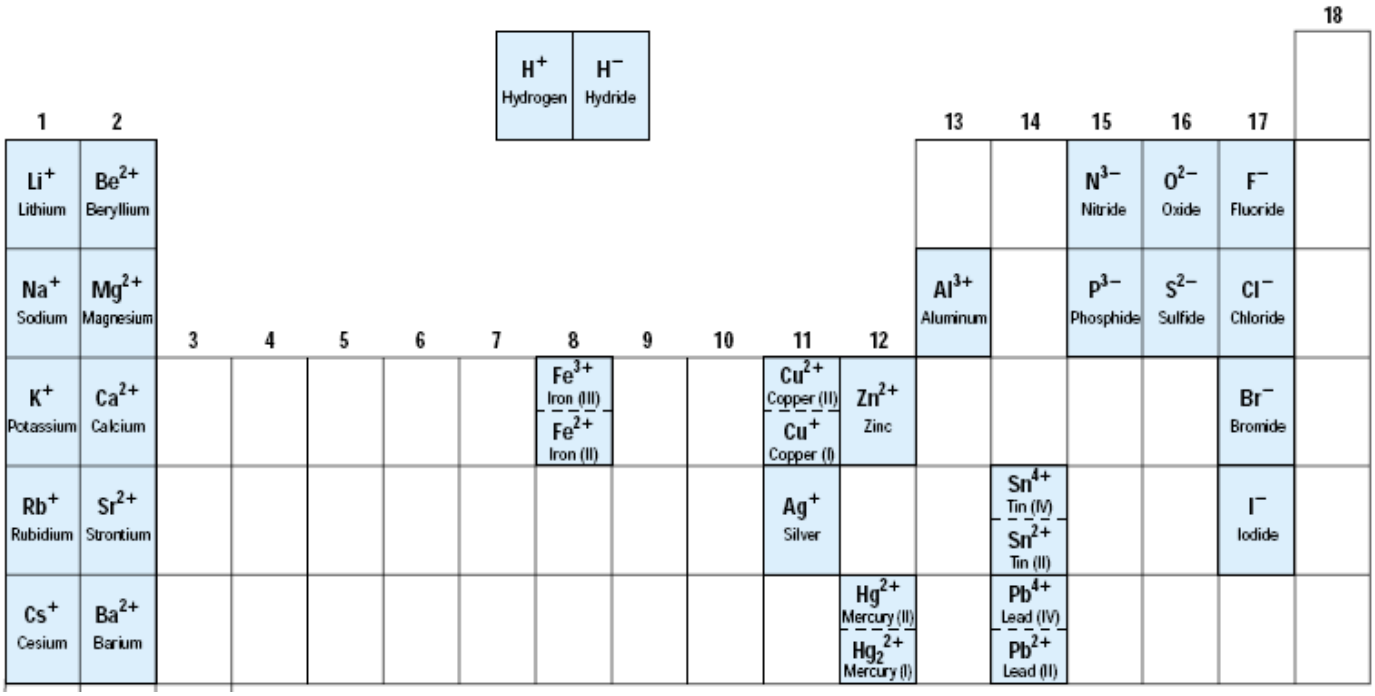
- Derived from _____ and _____ words to represent cations.
- Not commonly used today, but should recognize it.
- Word ending (suffix) and beginning (prefix) contained much information about compound.

Metal	Ion of Lower Charge	Ion of Higher Charge
iron	ferrous, Fe^{2+}	ferric, Fe^{3+}
copper	cuprous, Cu^{+}	cupric, Cu^{2+}
tin	stannous, Sn^{2+}	stannic, Sn^{4+}
lead	plumbous, Pb^{2+}	plumbic, Pb^{4+}
mercury	mercurous, Hg_2^{2+}	mercuric, Hg^{2+}

Stock System (IUPAC)

- Instead of suffixes and prefixes, Roman Numerals were used to represent charges.
- Metals form postive cations.
- Non metals form negative anions.

I	1
II	2
III	3
IV	4
V	5
VI	6



Writing Formula for Binary Ionic Compounds

- Formed by _____ oppositely charges ions.
- Overall charge of compound is _____

- When naming, cation name remains the same. Ending of anion changes to _____
- If multivalent ion, Roman Numeral in parenthesis indicates charge of cation.
- If you can reduce subscripts, do it!

Ex. **Binary Ionic Compounds**

Iron (III) Sulphide
Fe and S
Fe^{3+} and S^{2-} (use cross over rule)
_____

Tin (IV) Oxide
Sn and O
Sn^{4+} and O^{2-}
_____ _____

Naming binary compounds

- use the reverse cross over rule with a check. If the anion has the correct charge then continue. If there is no match use the correct charge of the anion and work out the charge of the cation (multivalent ions)

3+ 2-
Fe_2O_3 **-** charge on O matches thus name is iron(III) oxide

1+1- **2+2-**
SnO **-** charge on O is 2- **SnO** name is tin(II) oxide

<u>Cross-over-rule for writing Formulas</u>

1. write down the symbols of the elements in the order given in the name

2. Write valences above elements symbol

3. Divide valences by the highest common multiple

4. Cross-over valences

5. Drop all 1's and unnecessary brackets

Examples:
sodium oxide Na_2O calcium sulfide CaS magnesium bromide $MgBr_2$

Exercise:
For all exercises you will give the name if formula provided and the formula if the name is provided.

calcium nitride _____ silicon oxide _____

aluminum carbide _____ aluminum bromide _____

silver sulphide _____ zinc silicide _____

sodium fluoride _____ potassium bromide _____

barium iodide _____ magnesium chloride _____

LiCl _____ BaO _____

K_2S_____ Al_2O_3 _____

BINARY COMPOUNDS

PART A: Name the following compounds.

NaCl _____ $CaCl_2$ _____

Ca0 _____ $MgBr_2$ _____

CaS _____ Ag_2S _____

H_20 _____ AlI_3 _____

Na_20 _____ Al_4C_3 _____

Mg_3N_2 _____ H_2S _____

AlN _____ SiC _____

Al_20_3 _____ KBr _____

PART B: Write the chemical formula for each of the following.

aluminum carbide _____ hydrogen oxide _____

aluminum oxide _____ silver sulphide _____

silicon carbide _____ calcium chloride _____

sodium chloride _____ sodium oxide _____

magnesium bromide _____ calcium sulphide _____

hydrogen sulphide _____ magnesium nitride _____

aluminum iodide _____ calcium oxide _____

aluminum nitride _____ sodium hydride _____

page 97 (14,15)pg 99 (16,17)

Day 12 Binary Compounds with Elements having more than one valence value

Rules for Binary Compounds

- The name of the binary compound always ends in **"ide"**.
- Whenever the first mentioned element has more than one valence value, this must be indicated in the name. It will always be the first element. The second element has a valence value equal to the value for its group on the periodic table.
- **Valence values of transition metals must be memorized.**
  On the Periodic table, above each symbol are a number of values which the legend calls oxidation states. In many cases these also correspond to valence values.

There are three, 3, ways of doing this, and the different methods must not be mixed.

Method 1 - Roman numeral method (IUPAC)

- The valence value to be used is indicated by using uncrossed Roman numerals
- It is placed in brackets immediately following the name of the first element.

Examples:

iron(III) chloride	$FeCl_3$	tin(IV) iodide	SnI_4
phosphorus(III) oxide	P_2O_3	mercury(II) oxide	HgO

Exercise:
For all exercises you will give the name if formula provided and the formula if the name is provided.

iron(III) chloride _____

SbF_3 _____

MnO_2 _____

sulphur(VI) oxide _____

bismuth(v) phosphide _____

mercury(II) chloride _____

BiF_5 _____

ZnO _____

tin(IV) oxide _____

phosphorus(V) chloride _____

copper(I) bromide _____

antimony(V) sulphide _____

arsenic(III) oxide _____

mercury(I) sulphide _____

Pb_3N_2 _____

NiI_2 _____

Co_2Se_3 _____

SnO_2 _____

copper(II) sulphide _____

BINARY COMPOUNDS

Write the chemical formula or the chemical name for each of the following using the indicated method.

PART A: STOCK/IUPAC METHOD

Pb0 _____ As_2S_5 _____

Fe_20_3 _____ CuI_2 _____

Sn0 _____ $SbCl_3$ _____

P_20_3 _____ $Mn0_2$ _____

mercury(I) chloride _____ Iron(II) oxide _____

antimony(III) iodide _____ phosphorus(V) oxide _____

tin(II) oxide _____ copper(II) bromide _____

PART B: -OUS/-IC METHOD

$SnCl_4$ _____ Sb_20_3 _____

$CuBr_2$ _____ $FeBr_2$ _____

As_2S_5 _____ HgI _____

mercurous chloride _____ antimonous chloride _____

phosphoric sulphide _____ ferric oxide _____

phosphorous oxide _____ mercuric chloride _____

stannous bromide _____ cuprous iodide _____

phosphoric oxide _____ arsenic sulphide _____

stannic fluoride _____ ferrous oxide _____

Binary Acids - consist of 2 elements only H is bonded to a non- metallic atom ; when dissolved in water form an acidic solution

- to name these use hydro _____ (name of non-metal ending changed to ic) followed by the word acid; complete table below :

Binary compound (gas)	Formula	Binary Acid (aq)	Formula
hydrogen chloride		hydrochloric acid	
	HBr (g)		
		hydroiodic acid	
			HF (aq)
	H_2S (g)		
Hydrogen phosphide			

BINARY COMPOUNDS - REGULAR			
Write Formulas		**Write Names**	
1.	sodium chloride	17.	CaO
2.	calcium fluoride	18.	AgCl
3.	barium bromide	19.	Ca_3N_2
4.	lithium carbide	20.	H_2O
5.	silver iodide	21.	$SiBr_4$
6.	potassium oxide	22.	Al_2S_3
7.	aluminum bromide	23.	Na_2O
8.	calcium nitride	24.	AlF_3
9.	radium oxide	25.	NaCl
10.	boron fluoride	26.	KBr
11.	hydrogen sulphide	27.	BaS
12.	rubidium hydride	28.	AlN
13.	cesium oxide	29.	BAs
14.	magnesium sulphide	30.	HBr (aq)
15.	calcium carbide	31.	$ZnCl_2$
16.	zinc oxide	32.	MgI_2

Part A : NOMENCLATURE **NAME _____**

Beside each of the following compound names, give the correct symbol formula (binary)

potassium oxide		potassium fluoride	
sodium oxide		sodium chloride	
silver oxide		silver bromide	
copper(I) oxide		cupric iodide	
calcium oxide		calcium nitride	
magnesium oxide		magnesium sulfide	
mercury(I) oxide		ferrous sulfide	
iron(II) phosphide		copper(I) fluoride	
lead (II) oxide		aluminum sulfide	
copper(II) bromide		iron (III) nitride	
aluminum oxide		calcium fluoride	
iron (III) oxide		aluminum phosphide	

Complete pg 99 (16,17), pg 103 (20,21), pg 106 (2,4)

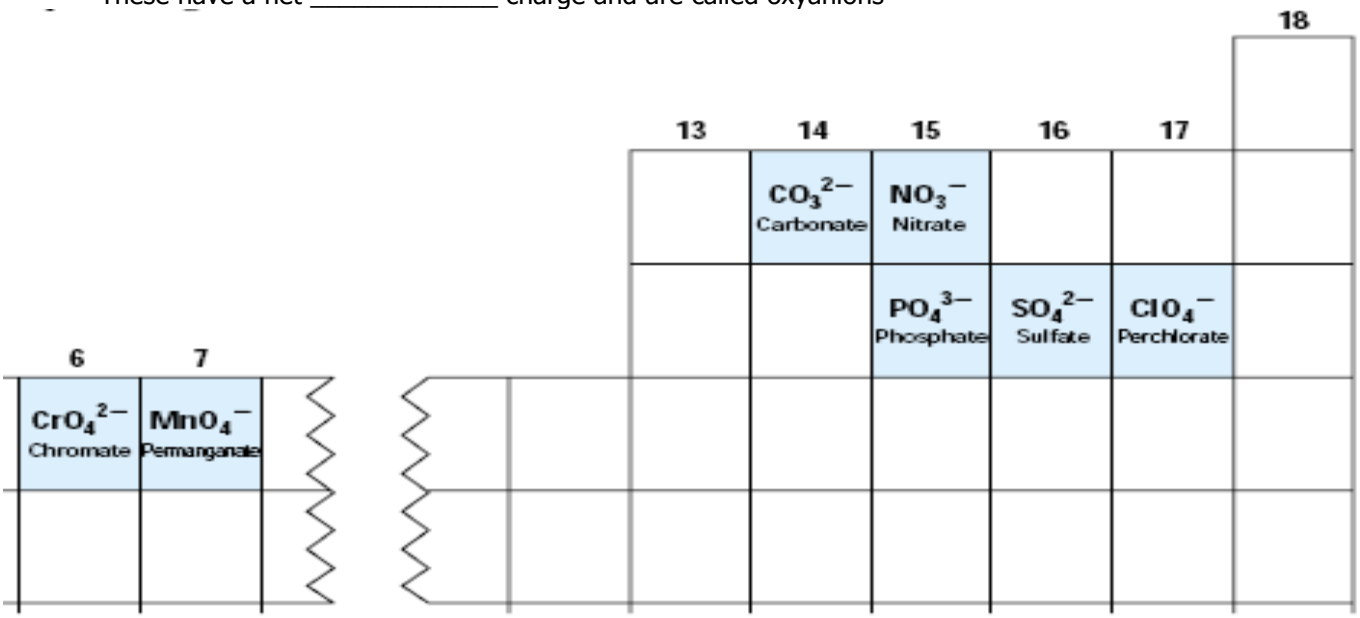
Name: _____ Date: _____

Nomenclature: Binary Ionic Compounds

#	Formula	Description (for interest only)	Name
1	$CaCl_2$	used on roads to keep down dust	
2		dietary supplement for iodine	potassium iodide
3	MgO	in fire bricks	
4		used to de-hair hides	aluminum chloride
5	Na_2S	aluminum ore (bauxite)	
6	Al_2O_3	used in antiperspirants	
7		black powder	lithium nitride
8	CaO	quicklime	
9		heart stimulant	barium chloride
10		table salt	sodium chloride
11	ZnO	UV blocker	
12	BaF_2	used in embalming	
13		ignites on contact with air	magnesium hydride
14	$MgCl_2$	11% of salt in sea water	
15		in soldering paste	zinc chloride
16	Ag_2S	tarnish on silverware	
17		potash fertilizer	potassium chloride
18	CaF_2	used to flourinate drinking water	
19		zinc ore	zinc sulphide
20		white solid	gallium fluoride

Day 13 Polyatomic Ions

- Almost all polyatomic ions contain oxygen and only one other element.
- These have a net _____ charge and are called oxyanions



Different forms of Oxyanions

- Sometimes elements form more than one oxyanion
- To distinguish between just two – a suffixes are added
- To distinguish between more than two – a prefixes are added

	Prefix	Suffix
One more oxygen atom:	*per-*	*-ate*
"Normal" atom:		*-ate*
One less oxygen atom:		*-ite*
Two less oxygen atoms:	*hypo-*	*-ite*

nitrate	NO_3^-
nitrite	NO_2^-

perchlorate	ClO_4^-
chlorate	ClO_3^-
chlorite	ClO_2^-
hypochlorite	ClO^-

sulfate	SO_4^{2-}
sulfite	SO_3^{2-}
thiosulfate	$S_2O_3^{2-}$

hydroxide	OH^-
cyanide	CN^-
acetate	$CH_3CO_2^-$

Naming Binary Compounds with Oxyanions

$Mg(ClO_2)_2$

Mg^{2+} and ClO_2^-

_____

$Ca(ClO_4)_2$

Ca^{2+} and ClO_4^-

_____

$Fe(OH)_3$

_____

Use the reverse cross over rule with a check to see ions match their charges ; if the do not put in the charge of the anion

Fe^{3+} and OH^-

TERTIARY COMPOUNDS

Write the chemical formula for each of the following:		Write the Stock/IUPAC name for each of the following:	
silver phosphate	_____	Na_2SO_3	_____
sodium sulphite	_____	$Ba(NO_3)_2$	_____
potassium perchlorate	_____	$MnSO_4$	_____
magnesium phosphite	_____	Hg_2SO_4	_____
calcium sulphate	_____	$NaIO_4$	_____
ferric sulphate	_____	$CaCO_3$	_____
potassium nitrate	_____	$Fe_2(SO_4)_3$	_____
magnesium hypobromite	_____	KNO_3	_____
zinc chlorite	_____	*$NaClO_3$*	_____
calcium phosphate	_____	$Fe(NO_3)_2$	_____
sodium periodate	_____	$(NH_4)_2CO_3$	_____
tin(IV) bromate	_____	$Mg_3(PO_4)_2$	_____
manganese(IV) nitrate	_____	$PbCO_3$	_____
mercurous sulphite	_____	$Ca_3(PO_4)_2$	_____
cupric perbromate	_____	$Fe(NO_3)_3$	_____
calcium carbonate	_____	$KI0_3$	_____
tin(IV) iodite	_____	$(NH_4)_3PO_4$	_____
sodium phosphate	_____	$Al_2(SO_4)_3$	_____
calcium hypochlorite	_____	$MgSO_3$	_____
sodium chlorite	_____	$Sn(NO_2)_2$	_____
magnesium perchlorate	_____	$FeSO_3$	_____
aluminum bromite	_____	$AlPO_4$	_____
mercuric hypoiodite	_____	NaBrO	_____
cupric sulphate	_____	Na_3PO_3	_____

Hydrated Salts

- Some salts have water molecules associated with each formula unit.
- Ex. Copper (II) Sulfate pentahydrate
  $CuSO_4 \cdot 5H_2O$
- Use the following prefixes to identify # of waters.

Greek Prefix	Number	Greek Prefix	Number
mono-	1	*hexa-*	6
di-	2	*hepta-*	7
tri-	3	*octa-*	8
tetra-	4	*ennea-* (or *nona-*)	9
penta-	5	*deca-*	10

Acid Salts

- Salts whose anions contain one or more covalently bonded hydrogen atom.

Ex. NaHSO4

Formula	Name
HCO_3^-	hydrogen carbonate
HSO_4^-	hydrogen sulfate
HSO_3^-	hydrogen sulfite
HPO_4^{2-}	hydrogen phosphate
$H_2PO_4^-$	dihydrogen phosphate

Nomenclature and Formula Writing Write formulas and names for each of the following combinations of elements and radicals.

	Cl	**O**	**P**	**ClO_3**	**NO_3**	**$C_2H_3O_2$**	**PO_4**	**ClO_4**	**SO_4**
K									
Ca									
Mg									
Al									
Cu (I)									
Sn (II)									
Au (III)									
Ag									
Pb (IV)									

Beside each of the following compound names, give the correct symbol formula.

Potassium Nitrate		Copper(I) Nitrate	
Sodium Chlorate		Copper (II) Nitrate	
Calcium Carbonate		Iron (II) Nitrate	
Sodium Sulfate		Iron (III) Chlorate	
Aluminum Phosphate		Iron (II) Carbonate	
Calcium Nitrate		Ferric Sulfate	
Calcium Chlorate		Iron (II) Phosphate	
Calcium Sulfate		Aluminum Nitrate	
Calcium Phosphate		Aluminum Chlorate	
Magnesium Nitrate		Aluminum Carbonate	
Magnesium Chlorate		Aluminum Sulfate	
Magnesium Carbonate		Copper (I) Chlorate	
Magnesium Sulfate		Copper (II) Chlorate	

Using pg 667-669 Name the chemicals found in the following : baking soda _____, bleach _____ Epsom salts _____, laughing gas _____, milk of magnesia _____, muriatic acid _____

Day 14 - Naming using the stock system and alternative system, naming acids, prefix method for naming compounds containing 2 or more non-metals
pg 103 (20,21) , pg 105(23,24) , pg 106 (1-4) do review page 107 (2,3,7,8,9,14,15,16,17,18,19,22)

Oxy - acids

- name derived from corresponding ternary salt or radical e.g. H_2SO_4 is sulfuric acid derived from the radical SO_4^{2-}
- **sulfate thus ic acids -------> ate salts**
- **the # of H's relate to the charge on the anion e.g. SO_4^{2-} has a -2 charge thus requires 2 H's**
- **ous acids --------> ite salts e.g. NO_2^- ----> nitrite ------> HNO_2 (aq) ------> nitrous acid**

Complete the following table :

Formula of radical	Name of radical	Formula of corresponding acid	Name of Acid
NO_3^-			**nitric acid**
	phosphite		
		H_2SO_4	
			sulfurous acid
CO_3^{2-}			
	acetate		
			phosphoric acid
	nitrite		
ClO_2^-			
			hypochlorous acid
		$HClO_4$	

Naming Covalently bonded molecular substances : Prefix Method

Binary Covalent Compounds

- Predicting formula is generally very difficult as they are variable.Ex. SCl, SCl2, SF4, SF6
- Prefixes are within name to show the number of specific atoms.

mono-	1
di-	2
tri-	3
tetra-	4
penta-	5

hexa-	6
hepta-	7
octa-	8
ennea- (or *nona-*)	9
deca-	10

Ex. Binary Covalent Compounds

- Less electronegative atom is written first
- If first element in compound has only one atom, 'mono' is not written.
- If second element in compound has only one atom, 'mono' is written.

S_2Cl_2 *di*sulfur *di*chlor*ide*
SCl_2 sulfur *di*chlor*ide*
SF_4 sulfur *tetra*fluor*ide*
SF_6 sulfur *hexa*fluor*ide*

CO → Carbon Monoxide
CO_2 → Carbon Dioxide
N_2O_2 → Dinitrogen dioxide

e.g. **CO_2 -----> carbon dioxide** **CO ------> _____** **N_2O ----> _____**

CCl_4 ---> _____ **N_2O_3 -----> _____** **SO_2 ----> _____**

P_2O_5 ----> _____

Complete the table below :

Formula	Name
N_2O_4	
	iodine monochloride
OF_2	
	dihydrogen oxide
PCl_5	
	sulfur hexafluoride
NO	
	phosphorous pentachloride
CS_2	
	nitrogen trifluoride
P_2O_3	

BINARY COMPOUNDS - prefix method			
Write Formulas		Write Names	
1.	carbon dioxide	26.	H_2O
2.	carbon monoxide	27.	SiO_2
3.	sulphur dioxide	28.	SO_2
4.	sulphur trioxide	29.	NO_2
5.	carbon tetrachloride	30.	CO
6.	nitrogen dioxide	31.	CCl_4
7.	diphosphorus pentoxide	32.	P_2O_3
8.	nitrogen monoxide	33.	As_2O_3
9.	silicon dioxide	34.	Cl_2O_7
10.	Dinitrogen tetroxide	35.	P_2O_5
11.	sulfur trioxide	36.	CBr_4
12.	phosphorus pentabromide	37.	SF_6
13.	dinitrogen trioxide	38.	SeO_2

14.	carbon tetrachloride	39.	$SiBr_2$
15.	tetraphosphorus hexoxide	40.	As_4O_6
16.	selenium disuphide	41.	As_2Br_5
17.	iodine heptafluoride	42.	N_2H_5
18.	dinitrogen pentoxide	43.	P_4S_{10}
19.	diboron nonoxide	44.	S_2O_7
20.	selenium dicarbide	45.	Br_3O_8
21.	phosphorus trifluoride	46.	As_3P_6
22.	dichlorine octoxide	47.	B_2S_5
23.	phosphorus pentafluoride	48.	CS_2
24.	selenium tetrafluoride	49.	B_4H_9
25.	Dinitrogen monoxide	50.	PBr_5

Mix and Match : Try these

1. H_2SO_3 (aq) _____
2. NO_2 _____
3. $Pb(NO_2)_2$ _____
4. H_2S(aq) _____
5. H_3PO_3 (aq)_____
6. HNO_2 (aq)_____
7. CuClO _____
8. $FeSO_4$ _____

Give formulas for :

5. nitric acid _____
6. acetic acid _____
7. carbon (II) oxide _____
8. phosphorous acid _____
9. carbon tetraiodide _____
10. Copper (II) phosphide _____
11. calcium nitride _____
8. nitrogen (IV) oxide _____
9. Phosphorous (III) oxide _____
10. sodium perchlorate _____
11. dinitrogen oxide _____
12. xenon tetrafluoride _____
13. silver phosphite _____
14. gold (I) nitrite _____

pg 103 (20,21) , pg 105(23,24) , pg 106 (1-4) do review page 107 (2,3,7,8,9,14,15,16,17,18,19,22)

Day 15 - Unit Review pg. 154 (8,9,11,28,29,30,31,32,40,43,44-47,49) pg 647 (8-21)