NS G7 (pg 1 of 2) Compare and Contrast

Name__________________Per____

Ionic Bonding − Molecular Bonding – Metallic Bonding

With two different types of elements: metals and nonmetals, there are <u>three possible combinations</u> of those two types of elements:

A. **metals + nonmetals** combine to form **ionic compounds** and will the focus of this unit

B. **nonmetals + nonmetals** combine to form **molecular compounds** which are held together by covalent bonds. Electrons are shared rather than transferred. We will name *binary molecular compounds* in this unit, and then come back to understand the bonding mechanism further in a later unit.

C. **metals + metals** metals do not *chemically* combine to form compounds, they simply *mix* together to form **alloys** just as we formed the alloy in LAD G4. The type of bonding that holds metal atoms frozen together is discussed below.

Use the chart below to compare and contrast ionic compounds with molecular compounds.

Compare and Contrast:	**Ionic Compounds**	**Molecular Compounds**
Made of?	metal + nonmetal	nonmetal + nonmetal
What are the electrons doing?	transfer electrons metals lose and become + ion (cation) nonmetals gain and become − ion (anion)	electrons are shared NO ions are formed
Formulas?	always written in lowest whole # ratio	may not always be lowest ratio
order of formula?	metal (positive ion) always first nonmetal (negative ion) always second	central atom usually written first, though there are many exceptions, and it really doesn't matter
Using roman #'s?	metals require roman numeral to indicate the charge (group 1, 2, & 3 and Ag, Zn, Cd do NOT require Roman #)	no roman numerals used since the atoms do NOT have charges.
Using prefixes?	prefixes NOT used (except within a few polyatomic ions)	prefixes are used to indicate the number of atoms present in the compound
Naming?	metal keeps its name nonmetal ends in –ide polyatomics end in –ate and –ite (a few in –ide)	for binary compounds, the second element listed ends in –ide
Particles are called?	individual particles are called formula units ("ionicules")	individual particles are called molecules
Polyatomic ions	end in –ite and –ate (a few end in –ide)	NA (not applicable)
Diatomic molecules	NA (not applicable)	diatomic gases: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2

What about bonding metals with metals?

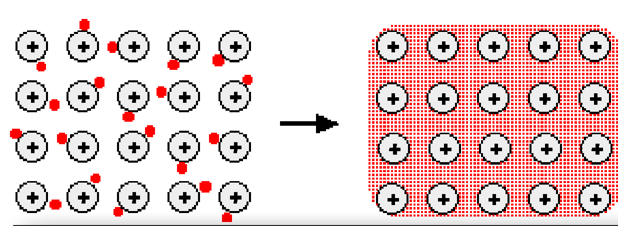
Let's use the bonding in solid sodium metal as an example. Sodium has the electronic structure $1s^22s^22p^63s^1$. When sodium atoms stick together as a solid lump of metal, the electron in the 3s atomic orbital of one sodium atom overlaps space with the corresponding electron on neighboring atoms to form a "sea" of loose valence electrons.

The 3s valence orbitals on all of the atoms overlap with nearby atoms to give a vast number of valence orbitals that are loose and mobile. The electrons can move freely within these valence orbitals, and so each s^1 electron is somewhat detached from its parent atom. The electrons are said to be **delocalized** (not local to its own atom). The metal is held together by the strong forces of attraction between the positive nuclei and the delocalized electrons. This is sometimes described as "an array of positive ions in a sea of valence electrons."

This "sea of loose valence electrons" model is used explain to *metallic properties* such as

- good conductors of electricity
- good conductors of heat
- low ionization energies
- easily oxidized
- malleable and ductile
- shiny luster

When viewing the diagram to the right, beware! Is a metal made up of atoms or ions? Metals are made of **atoms.** Each positive center in the diagram represents all the inner core of electrons. The valence electron hasn't been lost, but the electron is NOT tightly attached to any particular atom and thus is loose and free moving. Sodium metal is therefore written as Na, not Na^+.



The left view represents the electrons as localized – each e- close with its own inner core of e^- and nucleus.
In the right view the solid red area is representing the loose valence electrons – delocalized electrons – the "sea of electrons."

more on diamonds and graphite on the back →

 Bonding in Diamonds & Graphite

Carbon Conducts Electricity?

Carbon is not a metal. Why does graphite, an allotrope of carbon, conduct electricity? Yet diamonds are an allotrope of carbon and diamonds do not conduct electricity. What is going on?

Notice that the carbon atoms (not on the edge) in diamonds each have four bonds to other carbon atoms. Carbon's four valence electrons are locked in place in each of these four bonds, thus there are no loose electrons that can move to carry electricity. The four covalent bonds holding the entire diamond network together is what makes diamonds so very hard.

Notice in the graphite structure each carbon (not on the edge) has only three bonds. So where is the fourth valence electron? In the layers between the graphene sheets. When you rub your pencil across the page, you are depositing graphene sheets onto the paper. Pure graphite is very soft, but the graphene sheets are very, very strong, perhaps even forty times stronger than diamonds. You pencil is not made of pure graphite, which is very soft, thus the hardness or softness of your pencil is caused by different amounts of graphite, clay, and glue in the mixture. The second diagram draws a comparison between the loose electrons in metals, with the loose electrons in graphite.