Ionic Bonding and Lewis Structures

Formulas & Theorems Covered Today:

Homework:

Ionic Bonding

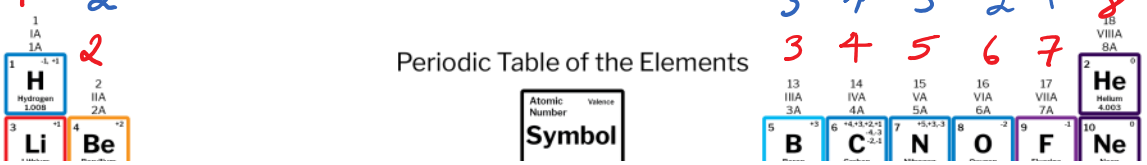
- Ionic bonding occurs when two ions of opposite charges are attracted to one another through electrostatic attractions.
- Ions are formed when atoms gain and lose electrons to become isoelectric to Nobel gases.
- In order to do that, atoms that tend do lose electrons will donate their electrons to other atoms that tend to gain electrons
- The atom that has lost an electron gains a net positive charge and becomes a cation while the atom that loses an electron gains a net negative charge and becomes an anion.

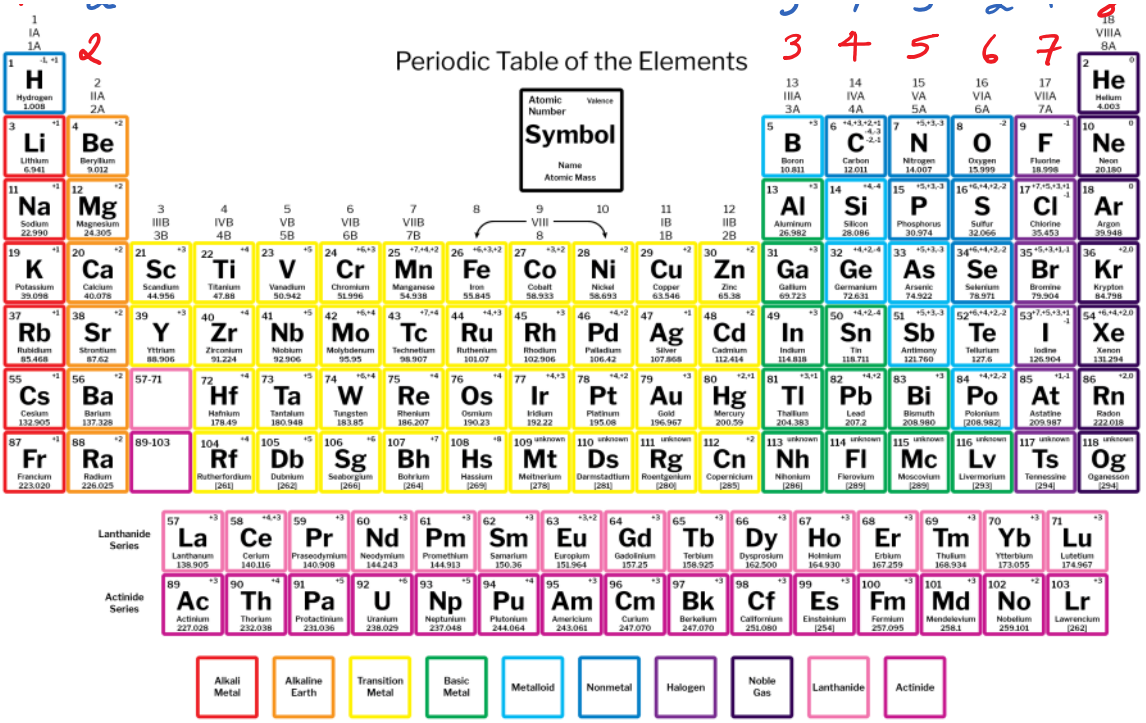
Valence Electrons

- As you learned last year and at the end of the last lesson, electrons that occupy the outer most orbital (shell) are called the valence electrons and subsequently, the outer orbital is called the valence shell
- These are of particular interest to chemists because they are the electrons that are involved, in one way or another, in chemical bonding.
- They also produce the chemical properties that you may have learned about last year.
- In order to be more efficient in our understanding of how atoms bond, and to be able to display how bonding occurs in a diagram, we have to simplify our Bohr-Rutherford diagrams into Lewis Structures
- Lewis Structures are named after Gilbert N. Lewis who, in a 1916 article called *"The Atom and the Molecule"* introduced these simplified structures to show how atoms bond together.
- When representing Lewis Structures of single elements, the symbol of the element is placed on the page and the equivalent number of valence electrons are added to the symbol.
- The electrons are added first at "12 o'clock" and then at 3, 6 and 9 o'clock in that order. (Note that if there are only 2 electrons that it would stop at 3 o'clock, and 3 electrons at 6 o'clock etc.)
- Once you have added the 4th electron, any subsequent electrons are added in the same order starting at 12 o'clock, pairing up with the electron that is already in that position

Ex.

Na	Mg	Al	Si	P	S	Cl	Ar
Na	Mg	Al	Si	P	S	Cl	Ar
valence — 1							
valence e — 1	2						
	2				3	4	3	2	1	0
					3	4	5	6	7	8





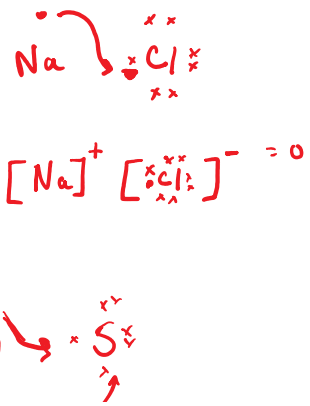
- After completing the Lewis Diagram for the elements, you will notice that every atom on the periodic table with the exception of the Nobel gases will have electrons in the structure that are unpaired.
- These unpaired electrons represent the valence number that is assigned to that element
- A valence numbers is another way of saying that this is the number of ionic bonds that that particular element can make
- For example, if we look at magnesium atoms, they have 2 valence electrons, both of which are unpaired resulting in a valence number of 2
- Similarly sulfur atoms have 6 valence electrons but only have 2 electrons that are unpaired electrons resulting in a valence number of 2

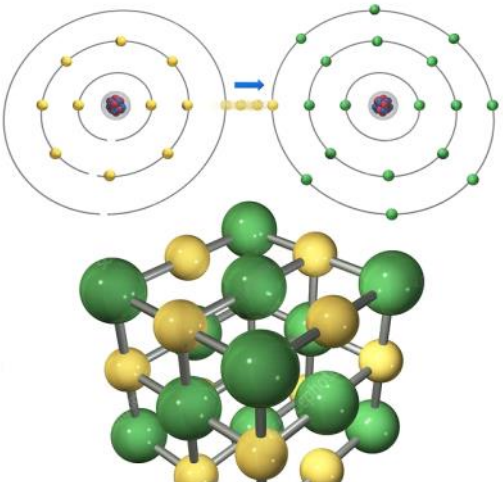
Lewis Structures of Ionic Compounds

- When creating Ionic Lewis structures, you have to complete them in 2 steps
- In the first step, you must show the transfer of electrons (where they are coming from and where they are going to)
- In the second step, you must depict the ionic compound as it appears in ion form
- Remember that ionic compounds are made of charged particles that form some type of chemical array as is depicted to the right

- Ex.

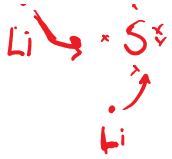
Na + Cl

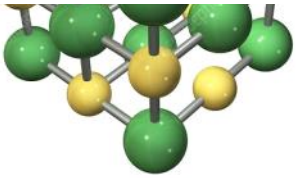




Li + S

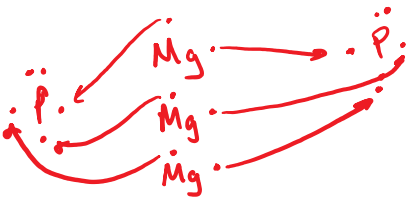
Li + S





$[Li]^{+}[S]^{2-}[Li]^{+} = 0$

Mg + P



$[P]^{3-}[Mg]^{2+}[Mg]^{2+}[P]^{3-}[Mg]^{2+} = 0$