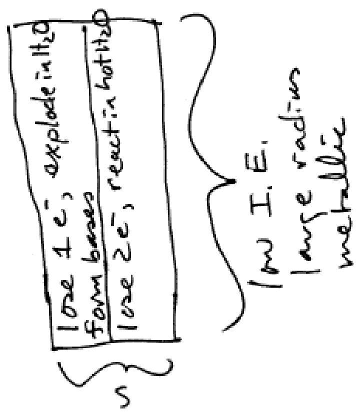
LECTURE 7. AN INTRODUCTION TO THE CHEMICAL BOND

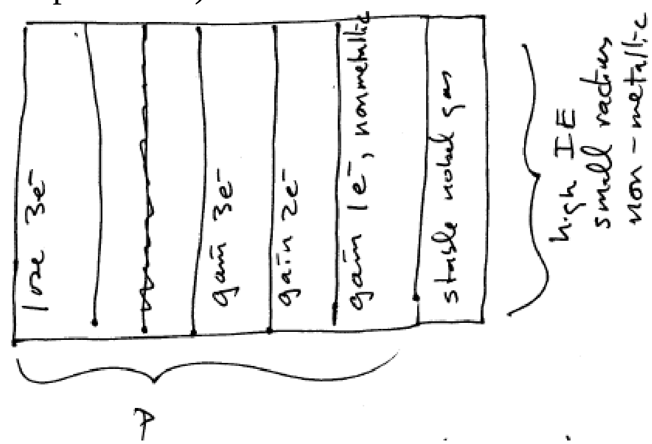
A summary:

- We use the concept of valence electrons surrounding atoms to create 2-dimensional structures in which species with chemical bonds are more stable than bare atoms.
- We base the stability argument for chemical bonds on the generation of species that have achieved stable noble gas configurations (s2p6)
- One kind of bond is ionic- electrons are donated or accepted to create anions and cations that have s2p6 configurations.
- A second kind of bond is covalent and occurs when non-metal atoms share electrons.
- Many compounds achieve s2p6 configurations by sharing electrons in overlapping p orbits to form double and triple bonds as well as resonance structures.
- Bonds can form even when s2p6 configurations are not possible. Atoms like boron have 3 valence electrons and usually can form only 3 covalent bonds with just six electrons. Compounds that contain an odd number of electrons form radicals that can't satisfy octet. Larger atoms (n=3 and higher) have available d orbits to hold five and six bonds.
- Many compounds can have multiple lewis structures that satisfy the octet rule. It is possible to evaluate the formal charge of atoms in the resonance structure to find the lowest energy, most stable structures.
- Electronegativity is a measure of the pull that an atom has for electrons in a chemical bond. Calculation of the difference in EN assists in distinguishing ionic and covalent character of bonds.
- Trends in bond energy and bond length can be predicted. Multiply charged bonds are higher energy and shorter.

Some last periodic table trends as we say goodbye to elements and take on bonding.

As we leave atoms and take on bonding, we need to have
Impact on materials – main group (s and p valence)

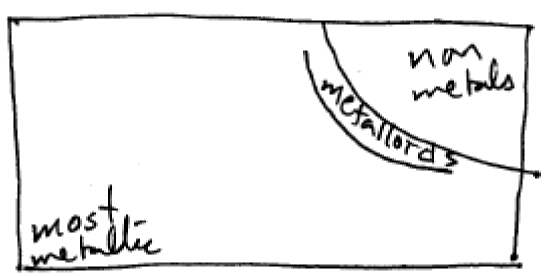




These main group materials exhibit very distinct changes by group. This makes for easy chemistry to teach. Trends are actually present.

Speaking of trends, a new one, metallic character is in the main group elements. Note that Cs is most metallic, that the entire left side of table is metallic. The right side is more non-metallic but in groups as you go down, metallic character increases.

Example: O_2 is nonmetal gas; S is nonmetal gas, Te and Po are metalloids.



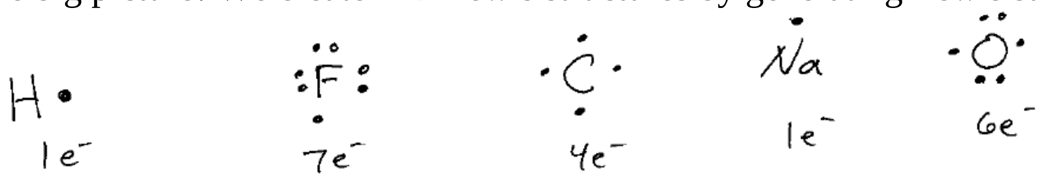
Transition metals:

d block elements are a mess. There are no trends, properties are similar, they all look like metals. This makes teaching them awful in this course.

- Why are transition metals so similar?
    - The d electrons are inner electrons and since they are what change, nothing much happens to valence e- and properties are the same.
- Can this be good?
    - Similar atomic radii means we can mix and match to form allows with neat properties – like superconductors.
    - Lots of oxidation numbers mean e- transfer in biological systems.

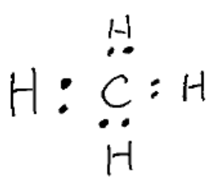
On to Bonding—A Comparison of Covalent and ionic Bonds

The big picture! We create 2-D Lewis structures by generating Lewis structures of atoms with valence electrons assigned.

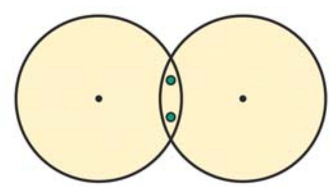


And mixing and matching to form chemical species in which the duplet or octet is satisfied.

A Bond: What happens when two electrons are placed in an orbit and the overall energy is lowered. Example:

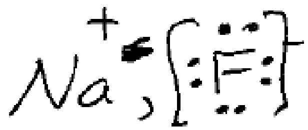


H= has n=1 shell filled, C has n=2 filled



6 Shared electron pair

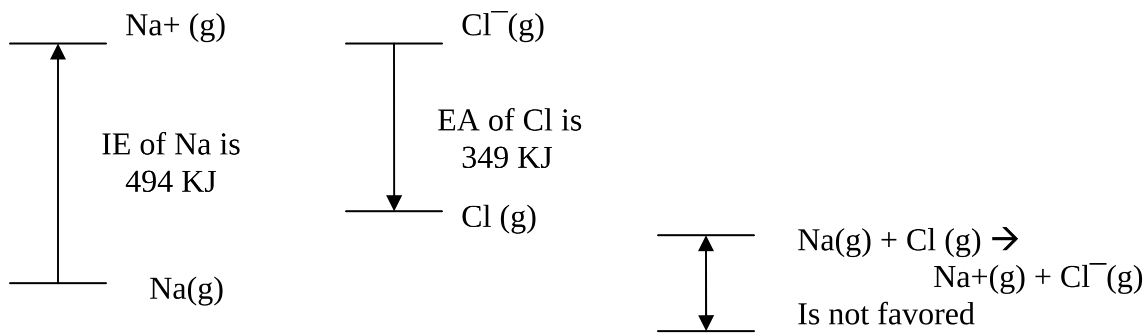
There are two kind of bonds. Ionic in which there is a donation of electron and covalent where there is a sharing.



Na^+ has n=2 shell filled , F^- has n=2 shell filled. **This is an ionic bond.**

C has n=2 filled O has n=2 filled. **This is a covalent bond.**

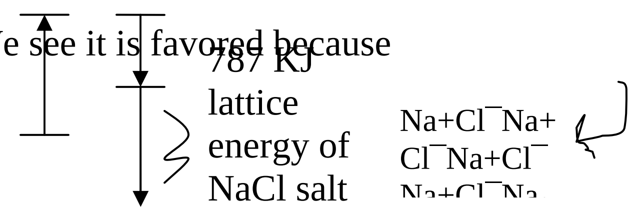
Theory of ionic bonds:



We want to know if $Na(g) + Cl(g) \rightarrow Na^+, Cl^-$ is energetically favored.

So that: $Na(g) + Cl(g) \rightarrow NaCl(s)$ is 642 KJ/mole

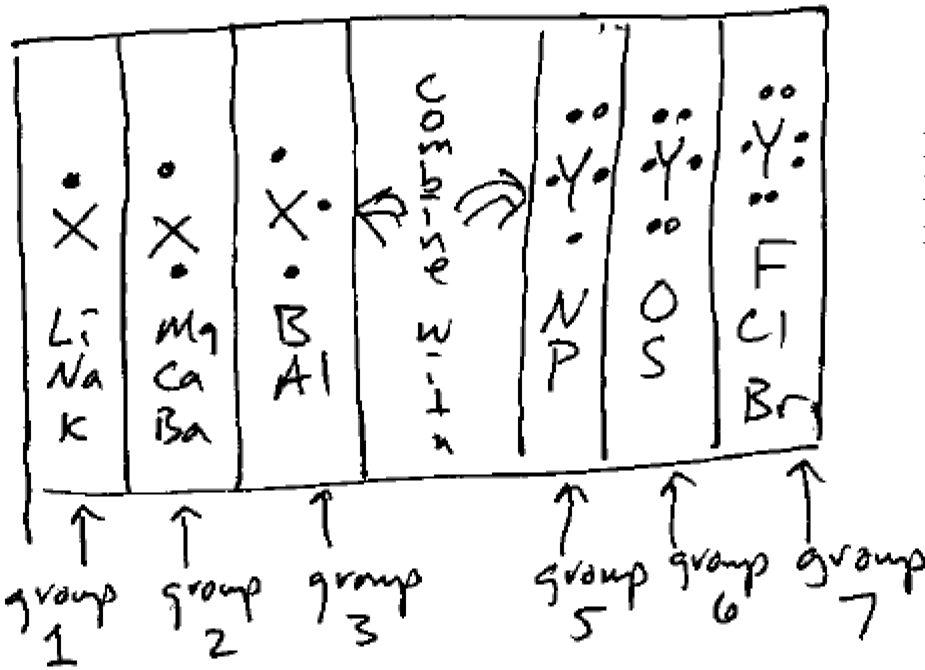
We see it is favored because



because of all that Coulomb attraction happens

So even though ion formation is not favored in the gas phase, it is favored as a solid because of all the Coulomb attraction when salt crystal is made.

What kind of ionic compounds form? Look at periodic table to see what matches up.

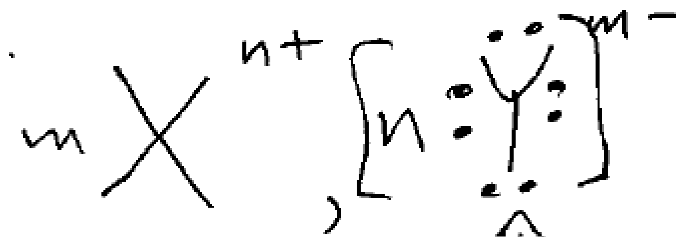


Find combinations of X and Y that yield ions with 8 electrons.

9 possible neutral forms of ionic salts

$NaCl \rightarrow$ X^{+}, Y^{-} $\leftarrow$ group 1and 7
$CaO \rightarrow$ $X^{++}, Y^{=}$ $\leftarrow$ group 2 and 6
$AlN \rightarrow$ X^{+3}, Y^{-3} $\leftarrow$ group 3 and 5
$Na_2O \rightarrow$ $2\ X^{+}, Y^{=}$ $\leftarrow$group 1and 6
$CaCl_2 \rightarrow$ $X^{++}, 2Y^{-}$ $\leftarrow$group 2and 7
$Li_3N \rightarrow$ $3X^{+}, 3Y^{-}$ $\leftarrow$ group 1and 5
$AlF_3 \rightarrow$ $X^{3+,}\ 3Y^{-}$ $\leftarrow$group 3and 7
$Ca_3N_2 \rightarrow$ $3X^{+2}, 2Y^{-3}$ $\leftarrow$ group 2 and 5
So as an example: $Al_2O_3 \rightarrow 2X^{+3}, 3\ Y^{=}$ $\leftarrow$ group 3 and 6

The general structure of Lewis ions is:

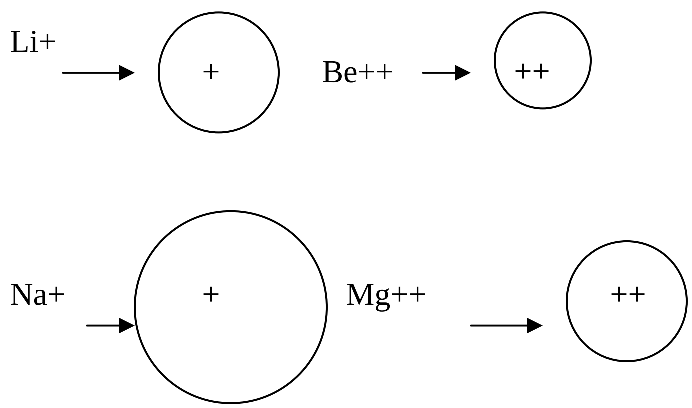


All e^{-} removed makes cation s^2p^6

e^{-} added makes anion s2p6

Time Out for an aside on Lattice Energy

We can develop a reasonable argument from charge density considerations that explains the magnitude of lattice energy. Consider four ions:



We can rank the charge density:

$Be^{+2} > Mg^{+2} > Li^{+} > Na^{+}$
Higher Charge — Lower Charge

When we take ions like this and combine them with anions, the higher the charge density, the closer the bonds and the higher the crystal lattice energy.

+ - - + Singly Charged	is weaker than	+2 -2 -2 +2 Multiply Charged

Example: rank the lattice energy of NaF, KCl, $MgCl_2$

high $MgCl_2 > NaF > KCl$ low
multiple charge — single charge low density