Chemical Bonding

Lewis Structures, Polarity and Bond Classification

Elements of the Lewis Theory

1. Valence electrons play a fundamental role in chemical bonding
2. Sometimes bonding involves the TRANSFER of one or more electrons from one atom to another. This leads to ion formation and IONIC BONDS.
3. Sometimes bonding involves SHARING electrons between atoms, this leads to COVALENT BONDS.

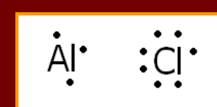
More Lewis Theory…

4. Electrons are transferred or shared such that each atom gains a more stable electron configuration
   - ☆ usually changes to Noble gas configuration

     Eg. Having 8 outer electrons
   - ☆ this arrangement (having 8 valence electrons) is called an OCTET

Electron Dot Diagrams

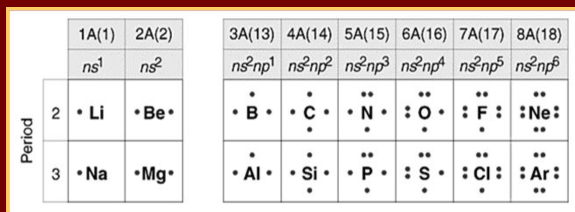
- Show the valence electrons of an atom / ion



- Chemical symbol represents the nucleus and the inner electrons
- Dots represent the valence electrons

Things to Note:

1. Since elements in the same family have the same number of valence electrons, their dot diagrams will look VERY similar (just different symbols)



2. Lewis dot diagrams only work well for representative elements
   - Transition metals, Lanthanides, and Actinides have incompletely filled inner shells (" d " or " f " orbitals), so we can't make simple Lewis diagrams for them

Back to Bonding…

- Remember, we have already talked about 2 types of bonding
  - –Ionic
  - –Covalent
- Now there is another type of bonding to know!

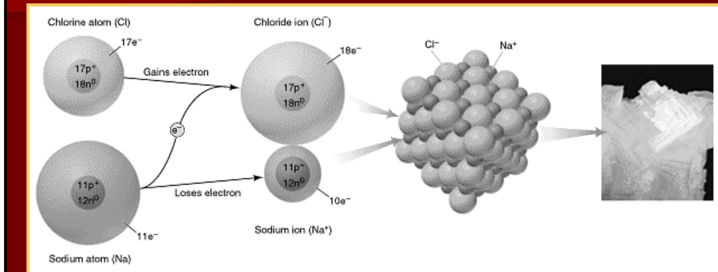
INTERMOLECULAR BONDING

1. Ionic Bonding

- Forces that hold ionic compounds together based on the electrostatic attraction of cations and anions
- 4 Steps to forming the bond…
  1. Form a positive ion – by loss of 1 or more electrons to become isoelectronic with noble gases
  2. Form a negative ion – by gaining 1 or more electrons to become isoelectronic (Lewis Diagram will show an OCTET of 8 electrons for the anion

3. Oppositely charged ions attract each other (electrostatic attraction) and form an ionic bond
4. An ionic crystal grows – cations are surrounded by anions and vice versa
   - formula unit is the smallest collection of ions that is electrically neutral
   - Lewis structures for ionic bonds represent one formula unit
   - ionic crystal is not a single molecule, but a collection of ions ("lattice" structure)

Example with NaCl:



Properties of Ionic Compounds

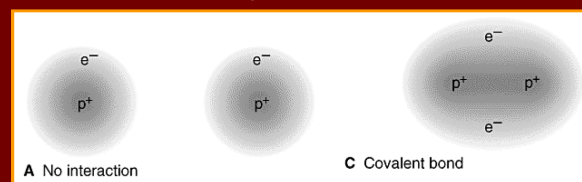
- Neutral overall (positives cancel out the negatives)
- No unique molecules (all bonded together)
- Decreased reactivity compared to atoms
- Conduct electricity when melted or dissolved

IONIC BONDS are **STRONG**, so ionic solids have **HIGH MELTING TEMPERATURES!**

2. Covalent Bonding

- Bonds formed by sharing of electrons between atoms
- Electrons are attracted by nuclei of both atoms involved

∴ Electrons spend most of the time between the two atoms - forming the bond



- Covalent bonds are VERY strong
- Bonds between MOLECULES vary in strength though, so melting points vary
- Covalent compounds can form **MULTIPLE BONDS**
  - use #'s (bond order) to describe how many bonds are being made

  1 = single ; 2 = double ; 3 = triple

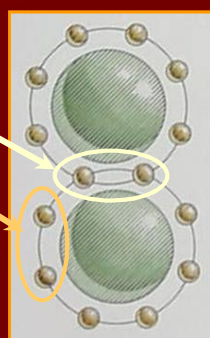
  (bond order shows how many pairs of electrons are shared)

Octet Rule:

- Atoms tend to form bonds until they are surrounded by 8 valence electrons
- Exception #1 – Hydrogen will form bonds to have 2 valence electrons
- There are other exceptions to the octet rule (more to come later)

Bonding Terminology:

- Bonding Electrons – the electrons that are shared in the bond
- Lone Pairs / Non-bonding Electrons – electrons not involved in bonding



Bond Classification using Electronegativity

- Type of bond is based on the electronegativity difference of the atoms involved in the bond
- Between 1.6 and 2.0 need to look at atoms in bond to classify (if a metal is present, then bond is ionic)

Bond Type	Electronegativity Difference
Non-polar Covalent	≤ 0.4
Polar-Covalent	between 0.4 and 1.6
Ionic	≥ 2.0

Electronegativities of Atoms

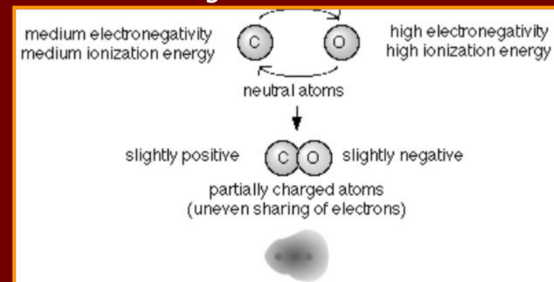
Increasing electronegativity →

Decreasing electronegativity ↓

				H 2.1												
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 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La-Lu 1.0-1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac 1.1	Th 1.3	Pa 1.4	U 1.4	Np-No 1.4-1.3										

What is a Polar-Covalent Bond?

- Electrons in the bond are shared unequally because of electronegativity differences between bonding atoms
- the electrons spend more time on the atom that is more electronegative



More on Polar Bonds…

- Polar molecules have partial charges on them
  - The atom where the electron spends more time is partially negative
  - The atom where the electron spends less time is partially positive
- The **greater** the electronegativity difference, the **more polar** the bond (extreme case is an ionic bond!)

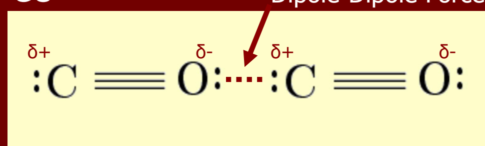
3. Intermolecular Bonding

- Bonding between molecules!
- Two Types
  - Dipole-Dipole Forces
  - London Forces

3a. Dipole-Dipole Forces

- Formed with polar covalent molecules
- The partial positive (δ+) and partial negative charges (δ-) attract each other forming electrostatic bonds (like WEAK ionic bonds… but are NOT true bonds)
- Ex. CO

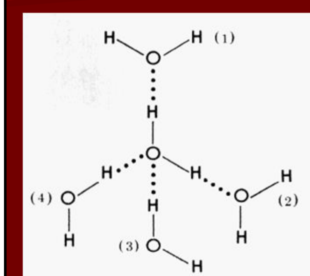
Dipole-Dipole Force



Special Case: Hydrogen bonding

- Occur when H bonds with F, O, or N (large electronegativity difference, so stronger than regular dipole-dipole forces)
- Common example: attraction between water molecules

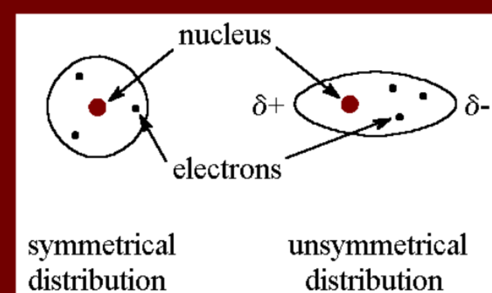
Hydrogen Bonds Between Water Molecules



- Oxygen is more electronegative so is partially negative (δ-)
- Hydrogen is less electronegative, so is partially positive (δ+)
- The attraction between the two partial charges is shown with dotted lines **between** the water molecules
- Because water is polar it can dissolve ionic compounds
  - The partial charges of water attract the ions of the ionic compund

3b. London Forces

- Formed by temporary (instantaneous) charges on an atom when electrons move to unsymmetric positions around the nucleus



- London forces are present between ANY molecules (including polar molecules) when they are close together
- Are the WEAKEST type of bond (unsymmetric electron location isn't always occurring to cause the attraction)

Writing Lewis Structures for Ionic Compounds

- Draw the Lewis Dot Diagram for each of the ions involved - including the ion charge
- Place the ions beside each other
- Remember: the metal will have no valence electrons and the non-metal will have a full valence shell

Example: NaCl

$$Na\cdot + \cdot\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}: \longrightarrow [Na^{+}][:\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}:^{-}]$$

Your turn to try…

- Draw Lewis Structures for the following ionic salts
  - KBr
  - $MgCl_2$
  - Li_2S
  - K_3P
- NOTE: If you count the valence electrons of the atoms involved and then count how many electrons are in your diagram, they should be the same #!

Drawing Lewis Structures for Covalent Compounds

- Steps to follow:

  1. Count the total number of valence electrons for the molecule and adjust for positive or negative charges on the molecule

  2. Determine which atoms are bonded together and put 2 electrons between them to represent the bond

3. – place remaining valence electrons to complete the octets of the atoms around the central atom(s). If any remain, place them in pairs on the central atom(s).

4. If the central atom has less than 8 electrons, have a neighboring atom share electrons with the "defficient" atom by putting an extra pair of electrons into a shared bond (repeat if needed)

5. If desired, you can replace the bonding pair(s) of electrons with dashes to represent the bonds

Lewis Structures for Covalent Compounds that Obey the Octet Rule

- Examples to do together:
  - O_2
  - HOPO
  - C_2H_4
  - NH_4^+
  - CO_3^{2-}

Lewis Structures for Covalent Compounds that DON'T Obey the Octet Rule

- Electron Deficient Molecules:
  - Be, B, and Al are common exceptions to the octet rule
    - Be can only share 4 electrons
    - B and Al can share up to 6 electrons only
      - Ex. BF_3

- Expanded Octet Examples:
  - Elements in the 3rd + 4th periods frequently have more than 8 valence electrons when covalently bonding (extra electrons are in d-orbitals)
  - Write the Lewis structure the same way, except extra electrons will be placed on the central atom
    - P and S are common examples
    - Ex. PCl_5, SF_4

What is on the TEST?

- **History of the Atom**
  - The scientists and their discoveries (Dalton, Bohr, Chadwick, etc.)
- **The Atom**
  - Atomic number and atomic mass
  - # of protons, neutrons, and electrons
  - Isotopes
- **The Periodic Table**
  - Development (including scientists)
  - Divisions within the Table
  - Metals, non-metals, and semiconductors
- **The Electronic Structure of the Atom**
  - Theory
  - Electron configurations for atoms and ions.
  - Core notation
  - Exceptions
  - Valence electrons
- **Periodic Trends**
  - Atomic and Ionic Radii
  - Ionization Energy
  - Electron Affinity and Electronegativity
- **Lewis Theory**
  - Elements of the Theory
  - Ionic Bonds
  - Covalent Bonds
    - Including Polarity
  - Hydrogen Bonding
  - London Forces
  - Writing Lewis Structures
    - Simple Ionic Compounds
    - Structures that Obey the Octet Rule
    - Structures that DON'T obey the Octet Rule