Chemical Bonding

In nature, most elements are not found as individual atoms.   How are most elements found?
Chemically combined with other elements in compounds.

These atoms are held together by:
Chemical Bonds

Which family of elements is the exception to this?
Noble Gases

Chemical Bonds:
Attractive forces between nuclei and electrons that hold atoms together in compounds.

Types of Chemical Bonds we will study:

1. **Ionic Bonds**
2. **Covalent Bonds**
    a. **Polar Covalent Bonds**
    b. **Nonpolar Covalent Bonds**
3. **Metallic Bonds**

The type of bonding that occurs depends on the **Difference in Electronegativity (*Δ EN)*** of the atoms in the bond.

To understand why chemical bonds occur, recall the **Rule of Octet**:
Atoms tend to gain, lose, or share e- to acquire a full set of valence electrons. (2 or 8)

Lewis Dot Diagrams

These are helpful in visualizing the electrons in a chemical bond.

To draw a lewis dot diagram for an element:

1. Use the chemical element symbol to represent the element's nucleus, and core electrons.

2. Use dots around the four sides of the symbol to represent the valence electrons of the element.

3. If the element has more than 4 valence electrons, place one dot on each of the four sides before you add a second dot to any side.

4. The dots without pairs represent electrons that will form bonds.

5. Examples:

1. Ionic Bonds

A. Definition:

Chemical bonds resulting from the transfer of 1 or more electrons from 1 atom to another, forming a cation and an anion. The bond is the attraction between the anion and cation.

B. Occur when:

Δ EN ≥ 2.0. Metals with low EN and nonmetals with high EN.

C. Examples:

NaCl ***KF*** ***CaO***

D. Form Ionic Compounds:

i. Structure of Ionic Compounds:

a. Contain ***Cations*** and ***Anions*** held together by ***Ionic*** bonds.

b. **Monatomic ions**:

Ions that contain 1 atom only. Examples:

c. **Polyatomic ions**:

Atoms that are covalently bonded together, having a net charge.

Examples: ClO_4^- PO_4^{3-} SO_4^{2-} CO_3^{2-}

$Na^+ \; Cl^-$

d. Use Lewis dot diagrams to represent the ionic bond between Na and Cl. $Na\cdot \;\; \cdot\ddot{Cl}: \rightarrow Na^+ - :\ddot{Cl}:^-$ $NaCl$

$Ca^{2+} \; F^-$ 1 2

e. Use a Lewis dot Diagram to represent the ionic bond between Ca, and F. What must the ratio of Ca to F be in order to satisfy the rule of octet? $:\ddot{F}\cdot \; \cdot Ca \cdot \; \cdot\ddot{F}:$ $:\ddot{F}:^- - Ca^{2+} - :\ddot{F}:^-$ CaF_2

$Al^{3+} \; O^{2-}$ 2 3

f. Try Al and O. What is the ratio of Al to O?

$:\ddot{O}\cdot \; \cdot Al \cdot \; \cdot\ddot{O}\cdot \; \cdot Al \cdot \; \cdot\ddot{O}:$ $:\ddot{O}:^{2-} - Al^{3+} - :\ddot{O}:^{2-} - Al^{3+} - :\ddot{O}:^{2-}$

Al_2O_3

g. What is the sum of all of the charges on the ions in d,e,and f? Why?

0; The sum of charges in a compound must = 0.

ii. Ionic compounds don't exist as distinct molecules – rather, they exist as:

Crystalline arrays of ions

http://images.google.com/imgres?imgurl=http://z.about.com/d/chemistry/1/7/q/p/Sodium-chloride-3D-ionic.jpg&imgrefurl=http://chemistry.about.com/od/sciencefairprojects/ig/Science-Fair-Project-Pictures/Sodium-Chloride-Ionic-Crystal.-Mmo.htm&usg=__V73BKCmiW_VuqA9-85qe2TIMWc8=&h=476&w=500&sz=917&hl=en&start=2&um=1&tbnid=_i4UBNiskqqhLM:&tbnh=124&tbnw=130&prev=/images%3Fq%3Dsodium%2Bchloride%26hl%3Den%26safe%3Dactive%26um%3D1

iii. **Formula Unit**:
A formula unit shows the kinds and numbers of ions in the smallest representative unit of a substance. In the case of an ionic compound, it is the lowest whole number ratio of cation to anion in a unit cell of an ionic crystal lattice. (the thing, not the formula)

iv. **Empirical Formula**:
The lowest whole number ratio of atoms or ions in a compound.

v. Bond Energy in an ionic bond:
The sum of all of the attractions between all of the cations and anions in a crystal lattice. It is very high because of the sheer number of ions in the crystal.

vi. Structure of Ionic compounds causes the properties of ionic compounds:

a. ***Crystalline brittle solid with a high melting point.***

b. ***Soluble in H_2O: H_2O is Polar, its charges pull ions apart***

c. ***Conductive when dissolved in H2O (Ions are moving charges).***
http://highered.mcgraw-hill.com/sites/0035715985/student_view0/chapter4/animations_center.html

d. ***Conductive when molten, but not when solid.***

Write the Empirical Formula for Ionic compounds that form from

1. Ca and Cl ___ $CaCl_2$ ___
2. Na and O ___ Na_2O ___
3. Al and N ___ AlN ___
4. Li and F ___ LiF ___
5. Mg and S ___ MgS ___

2. Covalent Bonds: ***Occur between NONMETALS***

A. Definition: ***Bonds formed when electrons are shared.***

B. Ionic Character in a covalent bond: Bonds are not usually purely ionic or purely covalent. There is some degree of ionic character to most covalent bonds. The degree is determined by the ***Electronegativity*** of the elements in the bond. In a covalent bond, the electrons will be found closer to the atom that attracts it more strongly. This element is more ***Electronegative***.

C. **Nonpolar Covalent Bonds**

i. Definition: ***Covalent bonds that form when electrons are shared equally.***

ii. Occur When: ***Difference in electronegativity of atoms bonding is very small:*** ΔEN ***<= 0.3***

iii. Examples: ***Diatomic Molecules:*** H_2 O_2 N_2 Cl_2 Br_2 I_2 F_2

D. **Polar Covalent Bonds**

i. Definition: ***Covalent bonds in which electrons are shared unequally.***

ii. Occur When: ***Difference in electronegativity is large (0.4- 1.9)***

iii. Examples: ***H – Cl,*** ***C - Cl,*** ***H - F*** ***Arrow points to more EN element***

E. Both types of Covalent bonds form **Molecular Compounds**:

i. Structure of Molecular Compounds:

a. Contain ***Nonmetal atoms*** held together by ***polarcovalent*** or ***nonpolar covalent*** bonds.

ii. **Molecules**: ***A group of atoms covalently bonded together.***

iii. **Molecular Formula**: ***The number and type of each atom in a molecule.***

vi. Structure of Molecular compounds causes the properties of molecular compounds:

a. Nonpolar molecules: ***molecules that have nonpolar bonds. These molecules are usually gases at room temperature, or liquids that become gases at low temperatures. Examples: H_2, O_2, Cl_2, N_2***

b. Polar Molecules: ***Molecules that have polar bonds. Since bonds are polar, the molecules have + and – charges separated by a distance (the bond length). Charges cause molecules to be mostly liquids and solids at room temperature. Solids have fairly low melting points, lower than ionic compounds. Example: Sucrose (table sugar) molecular formula $C_{12}H_{22}O_{11}$.***

v. **Structural Formulas**: ***Show type and # of atoms and the arrangement of covalent bonds in molecules.***

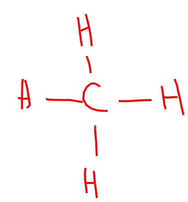
vi. **Lewis Structures**: **A type of structural formula. Atoms' nucleus and core electrons are represented by Chemical Symbol. Covalent bonds are lines between atoms. Nonbonding electron pairs (Lone Pairs) are lines outside atoms.**

Example:

molecular formula

CH_4

Structural/Lewis



Drawing Lewis Structures (Structural Formulas)

1. Write the number and type of each atom. (Molecular Formula)

2. Place the LEAST ELECTRONEGATIVE element in the center, and make a skeleton of the molecule. (DON"T put H in the middle – it can only form 1 bond) You want the skeleton to be compact, and symmetrical.

Ex: 1 O_2

O O

No "center" element

Ex: 2 H_2O

H O H

O is the center element

3. Calculate "**N**" N = the sum of number of electrons **N**eeded by each atom in the molecule to satisfy the octet rule. This number is 8 for most non hydrogen atoms, and 2 for H.

O: $N = 8 + 8 = 16$

H O H: $N = 2 + 8 + 2 = 12$

N=16 N=12

4. Calculate "**A**" A = the sum of the electrons **A**vailable for bonding. Add up all of the valence electrons for each atom in the molecule.

O O: $A = 6 + 6 = 12$ O is group 6A and has 6 valence e^-

H has 1 e^- H O H: $A = 1 + 6 + 1 = 8$

A=12 A=8

5. Determine the number of covalent bonds that link the atoms together using:

S=4 $S = 16 - 12 = 4$ **S= N- A** $S = 12 - 8 = 4$ S=4

Where **S** = the number of **S**hared electrons.
The number of chemical bonds = **S** / 2. (Since there are 2 electrons in a covalent bond.)

B=2 $B = \frac{4}{2} = 2$ $B = \frac{4}{2} = 2$ B=2

6. Place the correct number of covalent bonds between the atoms to link them together. O=O H – O – H

7. Now make sure every atom has an Octet of electrons around it: give "lone pairs" to each atom to fulfill the Octet rule. Remember Hydrogen only needs 2 electrons, so it gets a duet instead of an octet.

|Ō=Ō| H – Ō – H

8. Finally, count all the electrons on the molecule. It must add up to **A!**
Additional Lewis Structure Tips:

12 8

for #'s 12-25

Additional Lewis Structure Tips

1. If the compound is a polyatomic ion, the Available electrons (A) must be adjusted to account for the charge.

2. If the substance is an Oxyacid (Formula starts with H, then a nonmetal, then Oxygen) The skeleton always has the least EN element in the middle, surrounded by Oxygen, and the H is attached to an O. Don't put H on the central atom!

Examples of Oxyacids: H_2SO_4, HNO_3, H_3PO_4

3. When placing a double bond on an oxygen in an oxyacid, sometimes you will have a choice of which O gets the double bond. Do not place it on an O that has a H attached to it!

Lewis Structure Worksheet

1. H_2	10. C_2H_2
2. Cl_2	11. CH_2O
3. CO_2	12. H_2SO_4
4. HCN	13. H_2CO_3
5. CS_2	14. HNO_3
6. SO_2	15. $HClO_4$
7. CO	16. NF_3
8. C_2H_6	17. SO_4^{2-}
9. C_2H_4	18. NH_4^+

19. ClO_2^-	23. SiO_2
20. NO_3^-	24. C_3H_8
21. CO_3^{2-}	25. C_4H_8
22. ClO_3^-	26. CH_3COOH

See textbook and metalic bond WS

3. Metallic Bonds:

A. Definition:

B. Occur when:

C. Examples: