Chapter 7 – Chemical Bonding

7.1 Ionic Bonding

Octet rule: In forming compounds atoms lose, gain or share electrons to attain a noble gas configuration with 8 electrons in their outer shell (s^2p^6), except H and He want 2 outer electrons ($1s^2$). Basically atoms want to be like the noble gases which are stable and happy atoms.

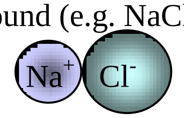
Chemical bond: what holds atoms or ions together in a compound. Chemical bonds form from the interaction of valence shell electrons.

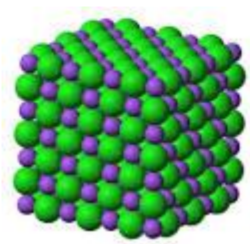
Ionic bond: electrostatic attraction between ***positively charged metal cations*** and ***negatively charged nonmetal anions*** (opposites attract).

Recall that metals lose electrons and become (+) cations while nonmetals gain electrons and become (-) anions. These ions have opposite charges and attract. This attraction holds them together and creates the ionic bond that holds the ionic compound together.

Formula unit: smallest unit of an ionic compound (e.g. NaCl, Al_2O_3, etc.)

Example: the formula unit for NaCl





Example: the compound NaCl

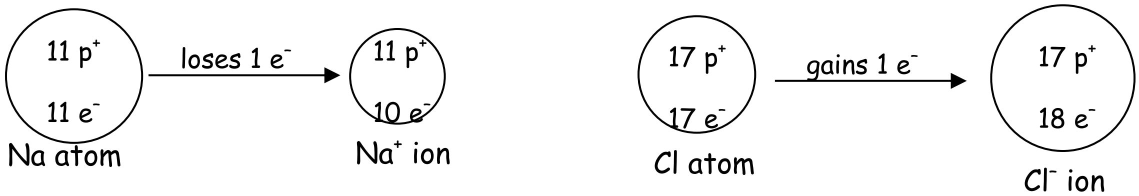
An ionic compound is actually a network of ions, with each cation surrounded by anions, and vice versa. In the 3D representation of NaCl at the right Na^+ ions are shown in purple and Cl^- ions are shown in green.

Ionic Radii: distance from the nucleus to the outermost electrons in an ion

⇒ ***cations*** have ***smaller radius*** than their corresponding atom since they have lost electrons, so there are more protons than electrons and the positive protons pull the negative electrons closer inwards

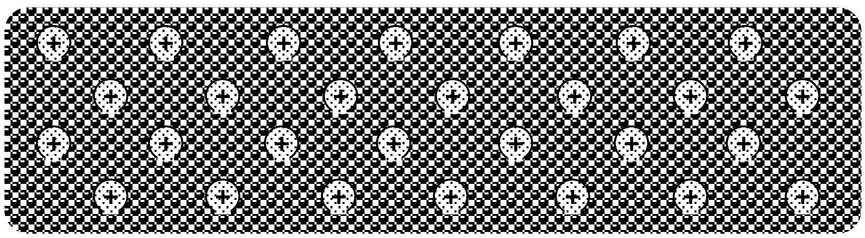
⇒ ***anions*** have ***larger radius*** than their corresponding atom since they have gained electrons, so there are more negative electrons than positive protons, and the additional electrons repel each other away

The representation below shows how the Na^+ ion is smaller than the Na atom because electrons were lost. And the Cl^- ion is larger than the Cl atom because electrons were gained.



7.2 *Metallic Bonding*

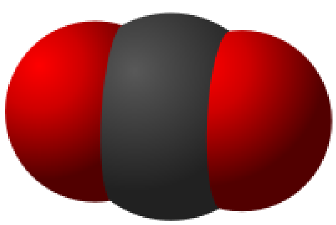
Metals exist as nuclei surrounded by a sea of electrons. The electrons in a metal are shared among all the nuclei, so the electrons are ***delocalized*** (i.e., they are not fixed to a specific atom) Thus metals conduct heat and electricity easily because electrons flow through the metal; metals are malleable and ductile because electrons act as a glue, holding the positively charged nuclei together.



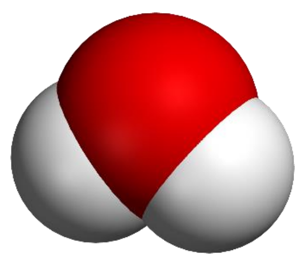
7.3 Covalent Bonding

Covalent bond: electrons are shared between two nonmetal atoms. Covalent, sometimes called molecular, compounds have covalent bonds. Diatomic elements have covalent bonds.

Molecule: smallest basic unit of a molecular compound (e.g. CO_2, H_2O)



CO_2

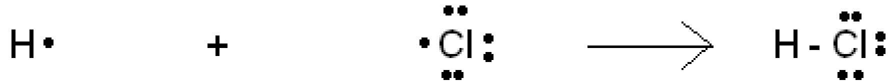


H_2O

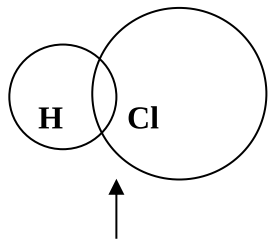
Nonmetal atoms form covalent bonds by *sharing electrons* to achieve a noble gas electron configuration. A single bond = 2 electrons shared, double is 4, and triple is 6.

A covalent bond is achieved by overlapping the valence orbitals of the two atoms.

hydrogen atom + chlorine atom → HCl molecule

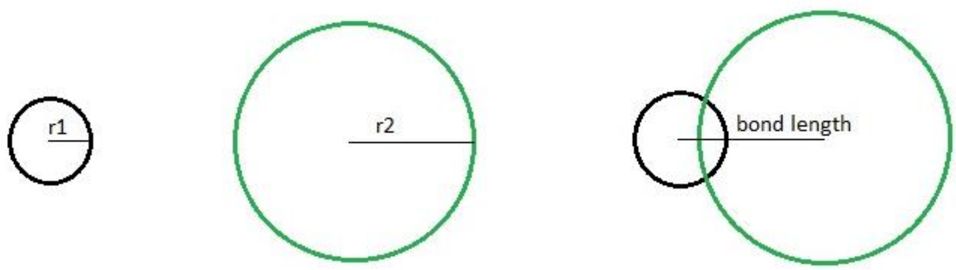


We can also represent the HCl molecule as follows:



This overlapping region is the covalent bond where electrons are shared.

Bond length: Actual distance from one nucleus to another when 2 atoms share electrons. The bond length for a covalent bond is less than the sum of the individual radii.



H atom | Cl atom | HCl molecule

r_1= 0.037 nm | r_2=0.099 nm | bond length=0.127 nm

yet $r_1 + r_2 = 0.136$ nm So we see that $r_1 + r_2 >$ bond length

Bond energy: amount of energy required to break a bond in a mole of gas

$$HCl(g) + \text{bond energy} \rightarrow H(g) + Cl(g)$$

Breaking bonds always takes energy, E is a reactant, it is absorbed. Forming bonds releases energy, E is a product, it is produced.

$$H(g) + Cl(g) \rightarrow HCl(g) + \text{bond energy}$$

Multiple Bonds

Single bond: *one pair of electrons* are shared by two atoms (H—H in H_2)

Double bond: *two pairs of electrons* are shared by two atoms (O=O in O_2)

Triple bond: *three pairs of electrons* are shared by two atoms (N≡N in N_2)

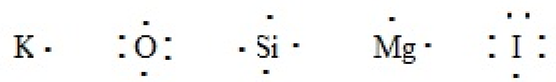
Note: Single bonds are the *longest and weakest*,

Double bonds are *shorter and stronge*r than single bonds, and

Triple bonds are the *shortest and strongest.*

7.4 Lewis Dot Structures of Atoms

Lewis dot structures show the symbol of an element surrounded by its valence electrons written as dots. First write down the element's symbol then determine the number of valence electrons. Assume the atom has four sides and distribute electrons with one electron per side before pairing electrons. There are a maximum of 2 e^- on each side!



7.5 Lewis Dot Structures of Covalent Molecules

Guidelines for drawing Lewis dot structures of molecules:

1. Calculate the total # of valence electrons for all atoms.
2. Surround the central atom with the other atoms and draw single bonds to them. The central atom will be underlined or in bold.
3. All atoms want an octet of 8 electrons, except H wants 2 electrons, around them.
4. Check and make sure you used the total # of electrons in your drawing.

   bonding electrons: electron pairs shared between two atoms

   nonbonding (lone pair) electrons: unshared electron pairs belonging to a single atom
5. If single bonds don't work, try double, then triple bonds until all atoms have an octet.

You will only be required to draw Lewis dot structures for molecules with one central atom.

YouTube video of $SeCl_2$ molecule AND now HCN molecule

7.6 Lewis Dot Structures of Polyatomic Ions

Polyatomic ion: a group of atoms bonded together that possess an overall charge. Ex: PO_4^{3-}

To draw Lewis dot structures of polyatomic ions you follow the same guidelines as above except you add or subtract electrons based on the charge.

- If the ion is positively charged, then electrons have been removed so you must subtract that # of electrons from the total electrons
- If the ion is negatively charged, then electrons have been added so you must add that # of electrons to the total electrons

Put brackets around the entire structure, and put the charge in the upper right-hand corner to indicate the charge belongs to the whole ion, not just to a single atom in the ion.

YouTube video of polyatomic ion PO_4^{3-}

7.7 VSEPR – Shapes and Bond Angles

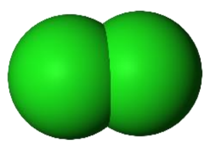
Valence-shell electron-pair repulsion (VSEPR) model

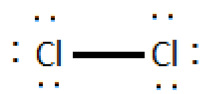
The shape of a molecule is largely determined by repulsions between electron pairs around the central atom in a molecule. The electrons repel each other because negative charges repel and move as far away from each other as possible.

Molecular Shape: three-dimensional arrangement of atoms in molecule. The shape is responsible for many physical and chemical properties. (mp, bp, density, etc.)

A= central atom B= outer atoms E = lone pair on central atom

If there are **only two atoms (AB)**, the molecule must be **linear**.

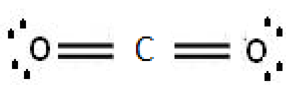


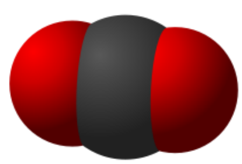


For molecules with **three or more atoms**, determine a generic formula—in the form **AB_x** (for x=2, 3, or 4)—to determine the molecular geometry (or shape) by looking at the table.

AB_2: linear – the two outer atoms are **180**˚ from each other

Ex. **$\underline{C}O_2$**

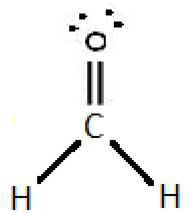


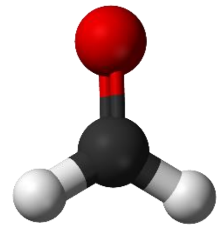


AB_3: trigonal planar

- three outer atoms at the corners of an equilateral triangle
- each outer atom is **120°** from the other two outer atoms

Ex. **$\underline{C}H_2O$**

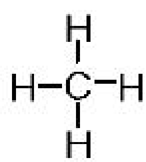


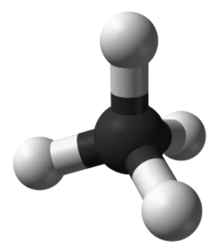


AB_4: tetrahedral

- (*tetra* = four) since four-sided, or four faces
- each outer atom is **109.5^o** from the other outer atoms

Ex. **$\underline{C}H_4$**.



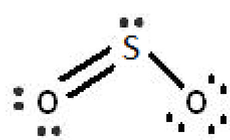


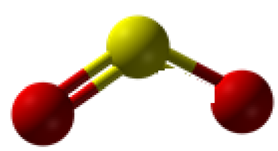
Lone pairs of electrons take up more space than bonded pairs of electrons because the bonded pair is shared by two atoms whereas the lone pair is held only by one atom.

AB_2E: bent (central atom and 2 outer atoms have a bent shape) Bond angle is < 120.

- central atom is bonded to two outer atoms (B) and has a **lone pair of electrons (E)**

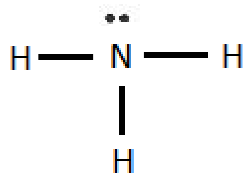
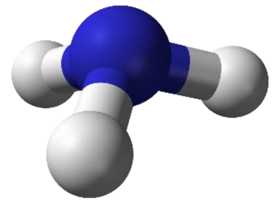
Ex. **$\underline{S}O_2$**.





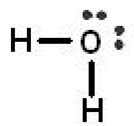
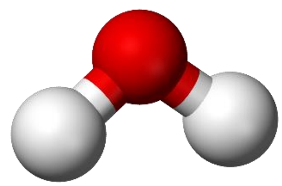
AB_3E: trigonal pyramid (central atom and 3 outer atoms make a pyramid) Bond angle is < 109.5^o.

- central atom is bonded to three outer atoms (B) and has a **lone pair of electrons (E)**

Ex. **<u>N</u>H_3**.  

AB_2E_2: bent (central atom and 2 outer atoms have a bent shape) Bond angle is < 109.5^o.

- central atom is bonded to two outer atoms (B) and has **2 lone pairs of electrons (E)**

Ex. **H_2<u>O</u>**.  

Given any covalent molecule or polyatomic ion, you should be able to determine the Lewis dot structure. Given the table of shapes and bond angles, you should be able to identify the shape and bond angles for any covalent molecule or polyatomic ion.

7.8 Bond Polarity

Nonpolar covalent bond: Bond between 2 nonmetal atoms with equal electronegativity values so the electrons are shared equally between the 2 atoms. No poles, no partial charges. Basically when nonmetals are bonded to themselves or C-H bonds. These are nonpolar bonds.

Polar covalent bond: Bond between 2 nonmetal atoms with different electronegativity values so the electrons are shared unequally between the 2 atoms. Has poles like a magnet with partial charges. Basically when two different nonmetals are bonded except C-H bonds which are nonpolar.

Electronegativity (EN): Ability of a **bonded** atom to attract the electrons in the bond towards itself. F is the most electronegative element.

- Elements are less electronegative the further away from F (Except for H, which has EN between B and C).
- Noble gas elements do not have EN values because they do not form bonds naturally.

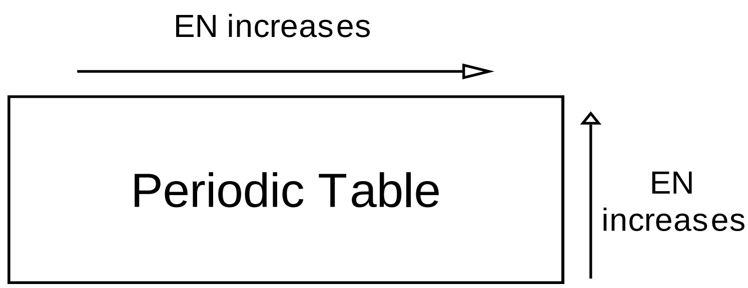
Molecules with two identical atoms always form nonpolar bonds: H_2, O_2, N_2, Cl_2, F_2, I_2, Br_2

Nonpolar covalent bonds can also occur between **different** atoms that have very similar EN values. e.g. C—H The C-H bond is considered Nonpolar.

In most covalent bonds between two different nonmetals, one of the two atoms holds electrons more tightly. So a **polar covalent bond** results between these two atoms such as these bonds: C - O, H - F, S - F, and C – N which are all polar covalent bonds.

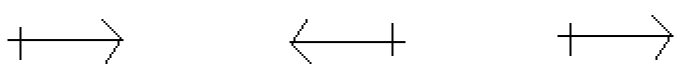
Electrons concentrate around the more EN atom in a molecule

- The more electronegative atom gains a partial negative charge, indicated with **δ -**
- Other atom gains a partial positive charge, indicated with **δ** +



Delta (δ) Notation for polar bonds and the polarity arrow: The delta notation and the polarity arrow are two way to indicate a polar bond. The more electronegative atom, due to the electons being held closer, is partially negative and thus gets the δ - sign next to it. The other atom by default is δ+. Meanwhile the arrow points towards the more EN atom and we cross the other end. Here are some examples:

δ+ δ -	δ - δ+	δ+ δ -
C – F	O – H	P – Cl



7.8 Molecular Polarity

- All nonpolar bonds = nonpolar molecule
- Polar bonds that don’t cancel out = polar molecule
- Polar bonds that do cancel out = nonpolar molecule

Molecular polarity depends on both the polarity of the individual bonds and on the shape around the central atom.

- Polar molecules will have an overall dipole (positive end and negative end)
- Nonpolar molecules have no net dipole (individual dipoles all cancel)

Nonpolar molecules are typically symmetrical, and the center of the + and - charge coincides.

Practice Problems

1. Draw Lewis dot structures for Na and Cl atoms. Now draw dot structures for NaCl, the ionic compound. How many protons and electrons are in Na^+ and Cl^- ions?

2. Draw the Lewis dot structures for each of the following molecules:

a. $\mathbf{H_2\underline{S}}$

c. $\mathbf{\underline{S}O_3}$

b. $\mathbf{\underline{C}H_2Br_2}$

d. $\mathbf{H\underline{C}N}$

3. Draw the Lewis dot structure for each of the following polyatomic ions:

a. $\mathbf{\underline{N}H_4^+}$

c. $\mathbf{\underline{P}O_4^{-3}}$

b. $\mathbf{\underline{N}O_3^-}$

d. $\mathbf{\underline{C}O_3^{2-}}$

4. For the following molecules or ions (where the central atom is underlined):

   i. Draw the Lewis dot structure.
   ii. Determine the shape of the molecule.
   iii. Determine the approximate bond angles.

a. $\underline{C}H_2F_2$

b. $\underline{O}F_2$

c. phosphite ion, $\underline{P}O_3^{-3}$

5. For each of the bonds below:

   i. Use delta notation (δ^+ and δ^-) to indicate which atom is more electronegative, and

   ii. Use a polarity arrow to point from the less electronegative atom to the more electronegative atom.

C—F   N—O   O—H   P —Cl   F—S

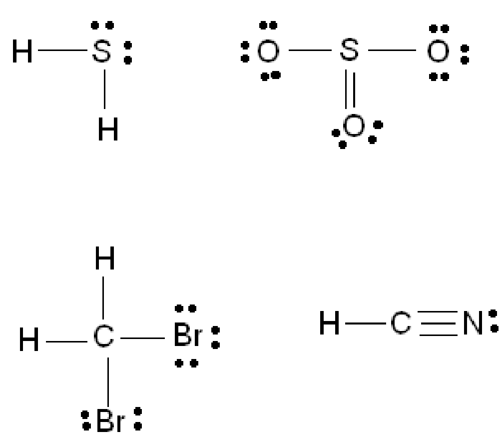
6. Identify the type of bond described for each of the following as ionic, polar covalent, nonpolar covalent, or metallic.

_____ a. The C–O bonds in CO_2.	_____ d. The C–C bonds in C_3H_8
_____ b. The bonds in F_2.	_____ e. The bonds in Ba metal.
_____ c. The bonds in K_2O.	_____ f. The bonds in H_2O.

7. Determine whether the following four molecules are polar or nonpolar: $\underline{C}O_2$, H_2O, $\underline{C}F_4$, and $\underline{C}HF_3$.

Answers

1. Na atom has one dot. Cl atom has 7 dots. Na^+ has no dots because it lost one electron. Cl^- has 8 dots because it gained one electron. Na^+ has 11 protons and 10 electrons. Cl^- has 17 protons and 18 electrons.



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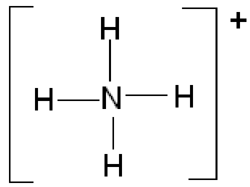
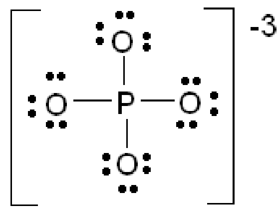
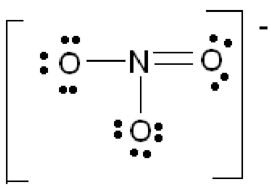
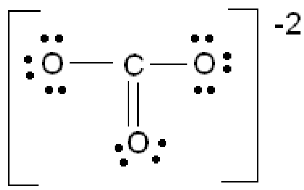
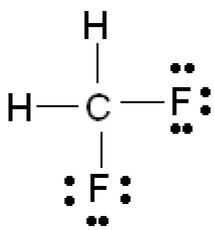
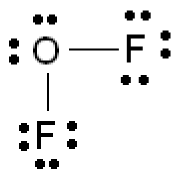


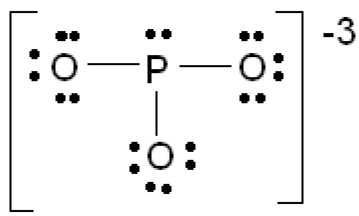


3. 



4. 





AB_4, tetrahedral, 109.5° AB_2E_2, bent, <109.5° AB_3E, trigonal pyramid, <109.5°

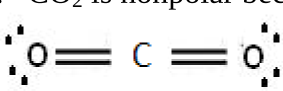
5.	δ^+ δ^-	δ^+ δ^-	δ^- δ^+	δ^+ δ^-	δ^- δ^+
	C—F	N—O	O—H	P —Cl	F—S

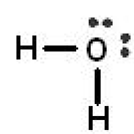
6 __PC___ a. The C–O bonds in CO_2. __NPC_ d. The C–C bonds in C_3H_8

__NPC___ b. The bonds in F_2. __Metallic e. The bonds in Ba metal.

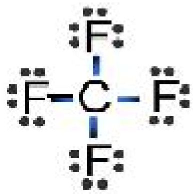
__Ionic___ c. The bonds in K_2O. __PC___ f. The bonds in H_2O.

7. CO_2 is nonpolar because the two polar bonds are equal and opposite so cancel out





H_2O is polar because the bonds are not opposite and don't cancel out

CF_4 is nonpolar because the bonds are all the same and cancel out, the outer atoms all the same. 

CHF$_3$ is polar because the bonds are not the same and don't cancel out, the outer atoms are

different.