CHM 130
Lewis Dot Formulas
and Molecular Shapes

Introduction

A chemical bond is an intramolecular (within the molecule) force holding two or more atoms together. Covalent chemical bonds are formed by valence electrons being shared between two different atoms. Both atoms attain the noble gas configuration with eight electrons (octet rule) or two electrons in their outer shell.

Octet Rule -
Atoms bond in such a way that each atom acquires eight electrons in its outer shell.

Duet Rule -
Hydrogen only requires 2 electrons to fill its outer shell and have He's electron configuration. ($1s^2$)

Lewis Dot Formula (also called an electron dot formula) –
Shows the valence electrons, indicating the bonding between atoms. The following guidelines will help draw the electron dot formulas correctly.

Guidelines for drawing simple Lewis dot formulas (electron dot formulas):

(1) Draw a skeletal structure using single bonds to connect atoms to a central atom.

(2) Calculate the total number of valence electrons.

(3) Deduct one pair of valence electrons ($2e^-$ total) for each single bond drawn in Step #1, then use the remaining pairs to complete octets for the other atoms.

(4) If there are not enough electrons to complete an octet for each atom, because the molecule is short 2 e^-, then move a nonbonding electron pair (an unshared pair) between two atoms that already share an electron pair to create a double bond. If the molecule is short 4 e^-, then move nonbonding pairs to create either two double bonds or one triple bond.

(5) Check your work.

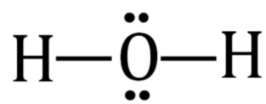
Example 1: Draw the electron dot formula and build the molecular model of water, H_2O.

(1) Draw a skeletal structure. Can hydrogen atoms be a central atom?

(2) H = 1 valence electron x 2 = 2 valence electrons; O = 6 valence electrons for a total of 8 valence electrons.

(3) 8 valence electrons - 4 electrons committed in single bonds = 4 electrons to distribute. Place two electron pairs on O to complete its octet.

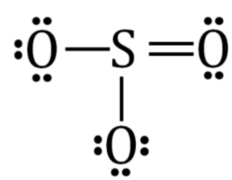
(4) Check. There are 2 electrons on each H and there are 8 electrons around the central O atom.

Molecule	# of valence e^-	Lewis Structure
H_2O	*2(1) + 6 = 8e-*	

The bonds depicted in a Lewis dot formula (an electron dot formula) can form three types of bonds depending on the number of pairs of electrons shared between the two atoms forming the bond. The three types of bonds are:

Single Bond – sharing one pair of electrons ($2e^-$ total)
Double Bond – sharing two pairs of electrons ($4e^-$ total)
Triple bond – sharing three pairs of electrons ($6e^-$ total)

Example 2: Draw the electron dot formula and build the molecular model of sulfur trioxide, SO_3.
(1) Draw a skeletal structure with a central S atom.
(2) Calculate the total number of valence electrons.
(3) Deduct a pair of electrons for each of the single bonds drawn in Step #1. Use the remaining pairs to complete octets for each of the other atoms.
(4) If there are not enough electrons to provide an octet for each atom, move a nonbonding (unshared) electron pair between two atoms that already share an electron pair.
(5) Check your work. [Hint: Did you need a double bond?]

Molecule	# of valence e^-	Lewis Structure
SO_3	*6 + 3(6) = 24e-*	

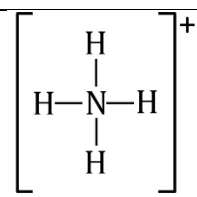
Example 3: Draw the electron dot formula and build the molecular model of hydrogen cyanide, HCN (hydrocyanic acid).
(1) Draw a skeletal structure that connects the atoms in the same order as shown in the formula.
(2) Calculate the total number of valence electrons.
(3) Deduct a pair of electrons for each of the single bonds drawn in Step #1. Use the remaining pairs to complete octets for each of the other atoms.
(4) If there are not enough electrons to provide an octet for each atom, consider moving nonbonding electron pairs between two atoms that already share an electron pair.
(5) Check your work. [Hint: Did you need a triple bond?]

Molecule	# of valence e^-	Lewis Structure
HCN	*1 + 4 + 5 = 10e-*	$H-C\equiv N:$

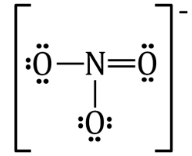
Lewis dot formulas (electron dot formulas) can also be drawn for ions. A monoatomic ion is a simple ion consisting of only one atom and having a positive or a negative charge. A polyatomic ion (or oxyanion) can contain two or more atoms held together by covalent bonds with an overall positive or negative charge. The following guidelines will help draw the electron dot formulas for ions correctly.

Amended guidelines for drawing Lewis dot formulas (electron dot formulas) for polyatomic ions:
(1) Draw a skeletal structure of the molecule.
(2) Calculate the total number of valence electrons. If the ion is negatively charged, then add electrons for each charge; if the ion is positively charged, then subtract electrons for each charge.
(3-5) The other rules are the same.

Example 4: Draw the electron dot formula and build the molecular model of ammonium ion, NH_4^+.

Molecule	# of valence e^-	Lewis Structure
NH_4^+	*5 + 4(1) - 1 = 8e-*	

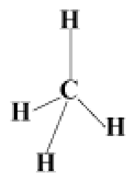
Example 5: Draw the electron dot formula and build the molecular model of nitrate ion, NO_3^-.

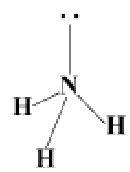
Molecule	# of valence e^-	Lewis Structure
NO_3^-	$5 + 3(6) + 1 = 24e^-$	

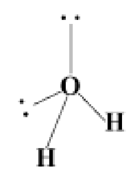
The **Valence Shell Electron Pair Repulsion Theory (VSEPR)** states the electron pairs surrounding an atom tend to repel each other and the shape of the molecule is the result of this electron pair repulsion. This model portrays bonding and nonbonding electron pairs as occupying specific positions around the central atom in the molecule.

Electron Group Geometry indicates the arrangement of bonding and nonbonding electron groups around the central atom without differentiation. All groups are considered equivalent. The three molecules below all contain four electron groups and no distinction is made between bonding and nonbonding electron groups in describing their electron group geometry. Make a molecular model of each and observe the orientation of the four electron groups in each molecule.

CH_4 – tetrahedral electron pair geometry
NH_3 – tetrahedral electron pair geometry
H_2O – tetrahedral electron pair geometry


methane


ammonia


water

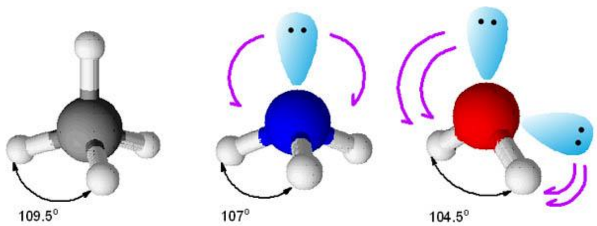
Molecular Geometry, or molecular shape as it is sometimes called, indicates the arrangement of atoms around the central atom as a result of electron group repulsion. There are two considerations in describing the molecular shape of a molecule:

(1) **The number of regions of electron density around the central atom**. This number dictates the electron group geometry around the central atom and the approximate bond angles between atoms because of repulsion between electron groups. Three possible electron geometries are considered in this course:

Two regions of electron density = **linear** electron geometry (180˚ bond angle)
Three regions of electron density = **trigonal planar** electron geometry (120˚ bond angle)
Four regions of electron density = **tetrahedral** electron geometry (109.5˚ bond angle)

(2) **The number of bonding and nonbonding groups of electrons around the central atom.** The atoms in a molecule are positioned in space because of the repulsion between electron groups around the central atom (i.e.; the electron geometry). The shape of the molecule depends on how many of the electron groups around the central atom are bonded to atoms and how many are unshared. Bond angle is defined as the angle formed by any two atoms bonded to a central atom. The nonbonded electron pairs create a larger cloud of electrons and create a greater repulsion than do the bonded electron groups. The effect of nonbonded pairs is to compress the bond angle. The effect is seen in the three possible bond angles considered in this activity:

109.5° – formed by a tetrahedral molecular shape, for example CH_4, in which all electrons are bonding electrons.
107° – formed by a trigonal pyramidal molecular shape, for example NH_3, in which three groups are bonding and one pair is unshared.
104.5° – formed by a (tetrahedral) bent molecular shape, for example H_2O, in which two groups are bonding and two pairs are unshared.



Another consideration is molecular polarity. **Polar bonds** are a type of covalent bond in which the electrons that are shared are drawn more closely to one of the atoms than the other. **Electronegativity (E_n)** is a numerical value that describes the ability of an atom to attract electrons in a chemical bond. The general trends in electronegativity are as follows:

Left to right - increase in E_n (nonmetals have high E_n, metals have low E_n)
Down a group - decrease in E_n (small atoms have high E_n, large atoms have low E_n)
F = 4.0 and Cs = 0.7

The degree of polarity can be calculated by taking the difference in E_n between the two atoms involved in the bond. For CHM130 the following guidelines will be used:

If the ΔE_n is < 0.4, the bond is considered to be nonpolar (the electrons are shared equally within the bond)
If the ΔE_n is between 0.5 and 2.0, the bond is considered to be polar (the electrons are shared but not equally within the bond)
If the ΔE_n is > 2.0, the bond is considered to be ionic (the electrons are completely transferred from one atom to the other)

For example: H-Cl H = 2.1 & Cl = 3.0. The difference in E_n = 3.0 - 2.1 = 0.9. Therefore, the bond between H and Cl is polar and the electron pair is attracted more towards the Cl than the H, because Cl has a higher electronegativity. This makes the Cl partially negative and the H partially positive.

The delta notation (δ) for polar bonds shows this partially charged situation by using δ^- to indicate the negative region and δ^+ to indicate the positive region.

δ^+ H - Cl δ^-

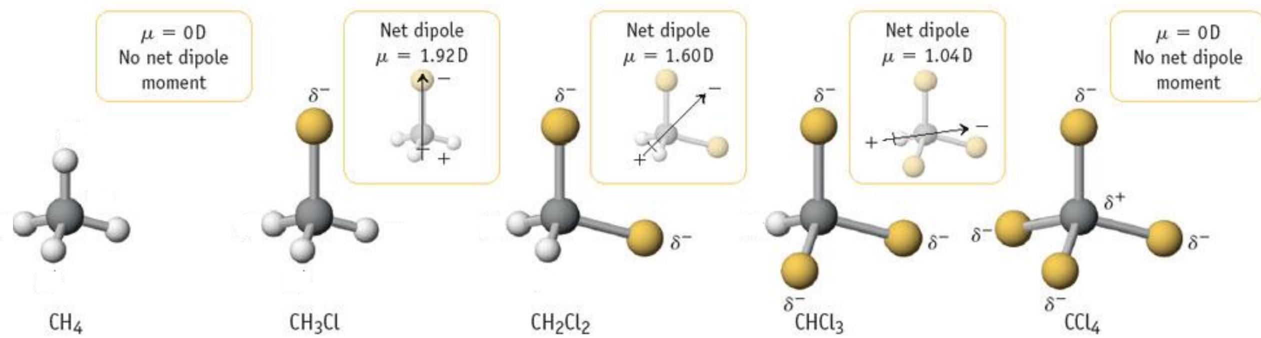
The following guidelines are for CHM 130 level ONLY:

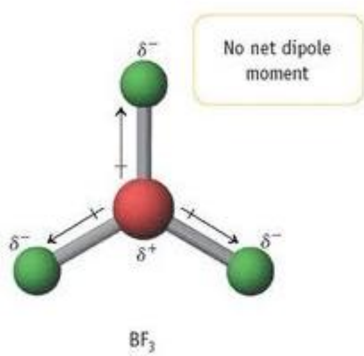
Polar Molecules are created when the polar bonds within the molecule add together to form a net dipole moment.

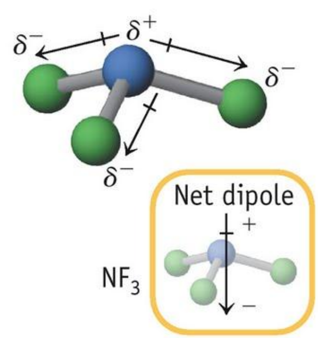
A molecule that contains only **nonpolar bonds** tends to be a **nonpolar molecule** (example: CH_4)
A binary molecule (2 atoms) will be nonpolar if the bond is nonpolar (example: Br_2) or polar if the bond is polar (example: HCl)
For molecules that contain 3 or more atoms and contain polar bonds, the molecular geometry of the molecule must be considered.

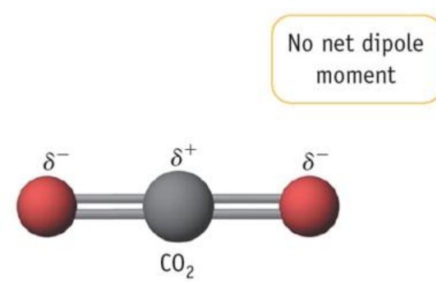
Molecular shapes that are **highly symmetric** (meaning no lone pairs on the central atom) tend to be **nonpolar**
All atoms bonded to the central atom must be the same (example: CCl_4, CO_2, BF_3)
Molecular shapes that are **asymmetric** (meaning they contain 1 or more lone pairs on the central atom) tend to be **polar**
Also applies to molecules that have different atoms bonded to the central atom (example: CH_3Cl, CH_2Cl_2, $CHCl_3$, NF_3, H_2O)

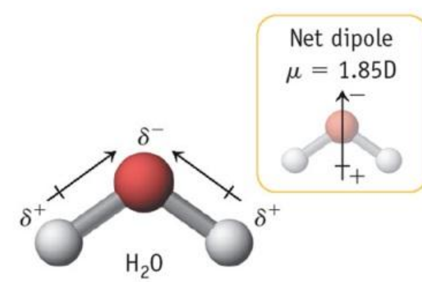
	Symmetric Shape (no lone pairs on the central atom)	**Asymmetric Shape** (the central atom has at least one lone pair)
Contains only **nonpolar bonds**	Nonpolar Molecule	Weakly Polar
Contains at least one **polar bond**	Nonpolar – if all bonds are the same Polar – if bonds are different	Polar







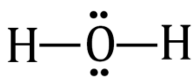
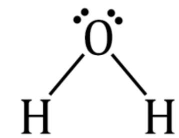




Name:____________________________________ Section & Date:__________________ Lab Partner:____

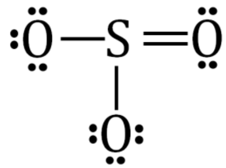
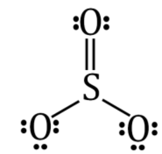
CHM 130LL Experiment #7 – Lewis Dot Structures Report Sheet

Part A – Molecules with single bonds

Molecule	# of valence e^-	Lewis Structure	Regions of e^- density	Electronic group geometry (name)	# of Bonding Regions	# of Lone Pairs	Shape of the molecule (drawing and name)	Polar or Nonpolar (P or NP)
H_2O	*2(1) + 6 = 8e-*		*4*	*Tetrahedral*	*2*	*2*	 *Bent*	*P*
Br_2								
HCl								
ICl								
CH_4								

Molecule	# of valence e^-	Lewis Structure	Regions of e^- density	Electronic group geometry (name)	# of Bonding Regions	# of Lone Pairs	Shape of the molecule (drawing and name)	Polar or Nonpolar (P or NP)
CH_2Cl_2								
HOOH								
NH_3								
N_2H_4								

Part B – Molecules with double bonds

Molecule	# of valence e^-	Lewis Structure	Regions of e^- density	Electronic group geometry (name)	# of Bonding Regions	# of Lone Pairs	Shape of the molecule (drawing and name)	Polar or Nonpolar (P or NP)
SO_3	*6 + 3(6) = 24e-*		*3*	*Trigonal planar*	*3*	*0*	 *trigonal planar*	*NP*
O_2								
C_2H_4								
HO<u>N</u>O			N	N	N	N		
SO_2								

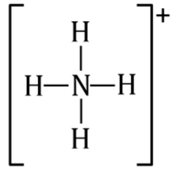
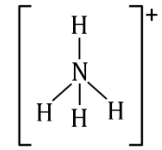
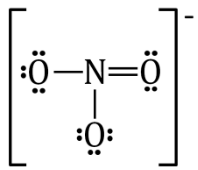
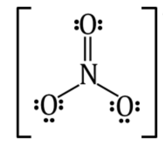
Part C – Molecules with triple bonds

Molecule	# of valence e^-	Lewis Structure	Regions of e^- density	Electronic group geometry (name)	# of Bonding Regions	# of Lone Pairs	Shape of the molecule (drawing and name)	Polar or Nonpolar (P or NP)
HCN	*1 + 4 + 5 = 10e-*	H—C≡N:	*2*	*Linear*	*2*	*O*	H—C≡N: *Linear*	*P*
N_2								
$HO\underline{C}N$			C	C	C	C		

Part D – Molecules with two double bonds

Molecule	# of valence e^-	Lewis Structure	Regions of e^- density	Electronic group geometry (name)	# of Bonding Regions	# of Lone Pairs	Shape of the molecule (drawing and name)	Polar or Nonpolar (P or NP)
CO_2								
$CH_2\underline{C}CH_2$								

Part E – Polyatomic Ions

Molecule	# of valence e^-	Lewis Structure	Regions of e^- density	Electronic group geometry (name)	# of Bonding Regions	# of Lone Pairs	Shape of the molecule (drawing and name)	Polar or Nonpolar (P or NP)
NH_4^+	*5 + 4(1) - 1 = 8e-*		*4*	*Tetrahedral*	*4*	*0*	 *Tetrahedral*	*NP*
NO_3^-	5 + 3(6) + 1 = 24e-		*3*	*Trigonal planar*	*3*	*0*	 *Trigonal planar*	*NP*
H_3O^+								
BrO_3^-								
CO_3^{2-}								
SO_4^{2-}								

Part F – Mixed

Molecule	# of valence e^-	Lewis Structure	Regions of e^- density	Electronic group geometry (name)	# of Bonding Regions	# of Lone Pairs	Shape of the molecule (drawing and name)	Polar or Nonpolar (P or NP)
SiH_4								
PH_3								
H_2S								
H_2Se								
CH_2O								
BF_3								