5. Chemical Bonding: The Covalent Bond Model
5.1 The Covalent Bond Model

Almost all chemical substances are found as aggregates of atoms in the form of molecules and ions produced through the reactions of various atoms of elements except the noble-gas elements which are stable mono-atomic gases.

Chemical bond is a term that describes the attractive force that is holding the atoms of the same or different kind of atoms in forming a molecule or ionic solid that has more stability than the individual atoms. Depending on the kinds of atoms participating in the interaction there seem to be three types of bonding:

Gaining or Losing Electrons:

Ionic bonding: Formed between many ions formed by **metal** and **nonmetallic** elements.

Sharing Electrons:

Covalent bonding: sharing of electrons between **two atoms** of **non-metals**.

Metallic Bonding: sharing of electrons between **many atoms** of **metals**.

	Ionic Compounds	Covalent Compounds	Metallic Compounds
1.	**Metal** and **non-meal** element combinations.	**Non-metal** and **non-meal** elements combinations.	**Metal** of one type or, combinations of two or metal elements combinations.
2.	**High melting brittle** crystalline solids.	**Gases, liquids, or waxy, low melting soft** solids.	**Conducting, high melting, malleable, ductile** crystalline solids.
3.	Do not conduct as a solid but conducts **electricity** when **molten**.	**Do not** conduct electricity at **any state**.	Conduct electricity at **solid and molten** states.
4.	Dissolved in water produce conducting solutions (**electrolytes**) and few are soluble in non-polar solvents.	Most are soluble in non-polar solvents and few in water. These solutions are non-conducting (**non-electrolytes**).	Insoluble in any type of solvents.

The differences in these three bonding types are mainly due to the number of valence election of the interacting atoms compared to noble gas elements. Noble gases show no inherent tendency to form any type of bonding apparently due to their closed valence shell electron configurations. Non-metals need only few electrons to achieve a closed shell. We focus mainly on covalent bonding in this chapter.

Covalent bond: bond in which one or more pairs of valence electrons are shared by two atoms in achieving a noble gas in contrast to complete transfer of electrons in forming ionic bonds.

Covalent bonding can be visualized with the aid of the **octet rule** and the **Lewis structure** we learned previously in chapter 4.

A molecular or covalent compound is made up of two or more **atoms** which are held together by a **covalent bond**. Even though covalent bonds are mainly formed between two atoms of non-metals, there could be many metal and nonmetal combination that are held by covalent bonds.

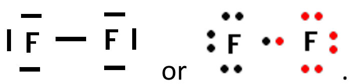
5.2 Lewis Structures for Molecular Compounds

Two atoms in molecule or ion in which each atom **shares one** valence electron with the other atom to from a single bond that keeps the two atoms together in forming a covalent bond.

Lewis Theory of Covalent Bonding:

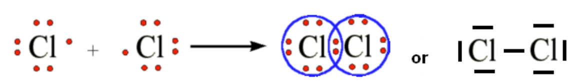
The idea that the stability of noble gas electron configurations and the realization of the connection of reactivity of elements to achieve **octet of valance electrons** through formation of electron-pair bonds with other atoms of main group elements except duet for hydrogen.

E.g. In fluorine molecule, F_2 each atom F atom has 7 valence electrons and shares one valence electrons to form a covalent bonding pair. Bonding pair is represented by a line.

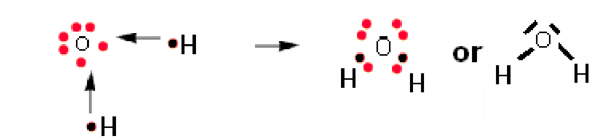


Lewis Electron-Dot (Line) Formulas: A simple way of writing out a formula that shows the disposition of the shared valance electron pairs between the different atoms in a molecule.

In chlorine molecule, Cl_2 each atom Cl atom has 7 valence electrons and shares one valence electrons to form a covalent bonding pair. Bonding pair is represented by a line.



In water, H_2O molecule, one oxygen atom with 6 valence electrons shares two valence electron with two hydrogen atoms each with one electron forming two covalent bonding pair. Bonding pair is normally represented by a line.



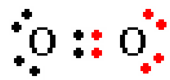
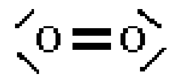
Bond pairs: An electron pair ( )shared by two atoms in a bond.
Lone pair: An electron pair ( ) found solely on a single atom.

5.3 Single, Double, and Triple Covalent Bonds

Single covalent bond: An electron pair ( ) shared by two atoms in a bond. Always there has to be a single covalent before forming multiple covalent bonds.

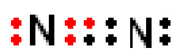
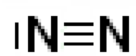
Multiple Covalent Bonds in Lewis Structures: For every pair of electrons shared between two atoms, a single covalent bond is formed. Some atoms can share multiple pairs of electrons, forming multiple covalent bonds.

Double covalent bond: Where two electron pairs are shared between two atoms. **For example**, oxygen, O_2 (which has six valence electrons) needs two electrons to complete its valence shell. In O_2, two pairs of electrons are shared, forming two covalent bonds called a **double covalent bond**.

Six valence electrons from each O Total of 12electrons or 6 electrons pairs. 	

Triple covalent bond: Found in unsaturated hydrocarbons with formula

For example, nitrogen, N_2 (which has 5 valence electrons) needs three electrons to complete its valence shell. In N_2, three pairs of electrons are shared, forming three covalent bonds called a **Triple double covalent bond**.

Five valence electrons from each N Total of 10 electrons or 5 electrons pairs. 	

Bond Length and Bond Energy

Bond Energy order: single=1 < double=2 < triple=3

Bond length: single (1pair) **>** double (2 pairs) **>** triple (3 pairs)

Bond lengths from periodic trends in atomic radii

Bond length is proportional to the sun of atomic radii forming the bond.

Atomic radii trend: Li> Be> B> C> N> O> F **decrease across a period**

Bond Length Order: Li-H> Be-H> B-H> C-H> N-H> O-H> F-H

5.4 Valence Electrons and Number of Covalent Bonds Formed

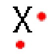
Predicting number of covalent bonds formed by representative elements:

Lewis dot symbols are useful in showing the ways in which non-noble gas electron configurations could be achieved by sharing electrons in forming covalent bonds. Number of covalent bonds formed is equal to the number of single electrons in the Lewis symbol of the representative element.

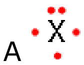
Group I A: alkali metals (Li, Na, K, Rb, Cs)

Common Lewis symbol of Group I A X• could only form a single covalent bond as found in NaH.

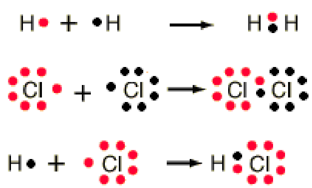
Group II A: alkali earth metals (Be, Mg, Ca, Ba)

Common Lewis symbol of Group II A  : alkali earth metals (Be, Mg, Ca, Ba) could form two covalent bond as found in $BeCl_2$.

Group III A: (B, Al, Ga, In)

Common Lewis symbol of Group III A : (B, Al, Ga, In) could form two covalent bond as found in NH_3.

Hydrogen, H_2 gas forms the simplest covalent bond in the **homo-nuclear diatomic molecule**, H_2. The halogens such as chlorine also exist as **homo-nuclear diatomic** gases by forming covalent bonds. The nitrogen and oxygen which makes up the bulk of the atmosphere also exhibits covalent bonding in forming diatomic molecules. Hydrogen and chlorine forms the simplest covalent bon d in the **hetero-nuclear diatomic molecule**, HCl.

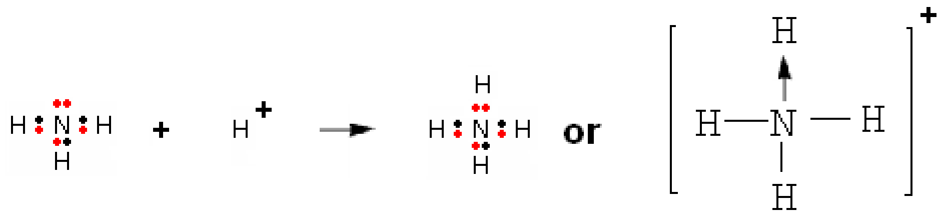


	Predicting Iionic charge of Representative Elements							
Periodic Table Group Number	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
Common Lewis symbols	X•	•X•	•X••	••X••	:X•••	:X:••	:X:•:	::X::
Lewis symbols 2nd period elements	Li•	•Be•	•B••	••C••	:N•••	:O:••	:F:•:	::Ne::
Number of covalent bonds formed	1	2	3	4	3	2	1	0

5.5 Coordinate Covalent Bonds

Coordinate Covalent bonds: Two atoms in molecule in which one atom **shares two** valence electrons with the other atom to from a bond that keeps the two atoms together.

E.g. In ammonia ion fluorine molecule, NH_4^+, N atom has a lone pair of electrons and shares with the H^+ ion to from a single bond that keeps the two atoms together. Coordinate covalent bonding pair is represented by an arrow pointing to atom receiving the electron. Chemistry dealing with coordinative covalent bonds is called **Coordination Chemistry**.



Covalent Bonds in Lewis Structure of Molecules and Polyatomic Ions: Lewis electron dot structures are representations of the distribution of electrons in molecules and polyatomic ions. The procedure is described below.

5.6 Systematic Procedures for Drawing Lewis Structures

A Lewis structure can be drawn for any **molecule** or **ion** could be obtained by following six steps:

Steps for Writing Electron-Dot Lewis structures for Molecules and Polyatomic Ions

1. **Calculate the number of valence electron pairs** (including charges, if any). The total number of valence electrons for the molecule or ion is calculated by adding up the valence electrons for each of the atoms in the species (molecule or ion). The number of valence electrons contributed by an atom equals the group number of the element. If the species has a **charge**, you need to **add** (anions) or **subtract** (cations) electrons from the neutral atoms to give that charge. Add one electron for each negative charge and subtract one electron for each positive charge.

2. **Pick the central atom**: Lowest EN (**electronegativity**) atom, **largest atom**, and/or atom forming most bonds is usually the central atom.

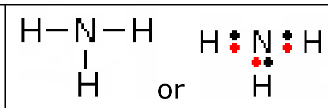
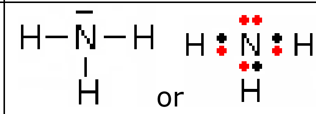
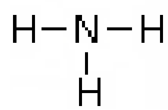
3. **Connect central atom to all terminal atoms.** Connect atoms by single bonds, using either a pair of **dots (:** or a **dash** (—). You may decide later that two or more of these bonds will add up to a double or triple bond.

4. **Fill octet to terminal atoms.** Distribute electron **dots (:)** or a dash (—) to the atoms surrounding the central atom to satisfy the octet rule in them. Calculate the number of remaining valence electrons pairs so far used.

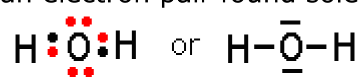
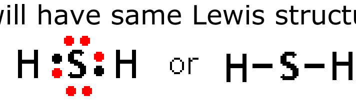
5. **Fill octet to central atom**: Distribute the remaining valence electrons to the central atom. If there arc fewer than eight electrons around the central atom, this suggests that you need to make double triple bonds taking **lone-electron pairs** (:)from terminal atoms. When you have a choice of atoms from which to obtain such a lone-pair, note that C. N, 0, and S atoms tend to form multiple bonds, so you take the lone pair from such an atom.

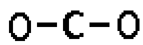
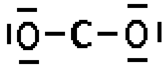
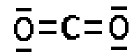
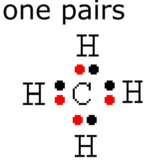
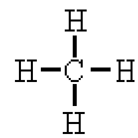
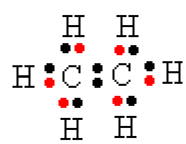
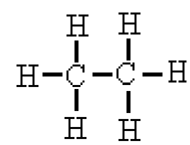
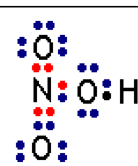
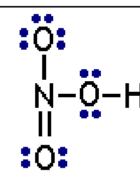
6. **Check that total number of valance electron pairs** in the structure match with the **dots (:** or a **dash** (—) or lines.

Draw Lewis Structure of NH_3

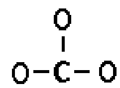
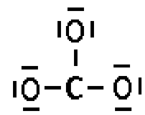
1) **Count Valence electrons**: 5 + 3 x 1 = 8 = 8/2 **4 electron pairs**	**4 electron pairs**	**4) Fill octet to terminal atoms.** Exception duet to H atoms.	
2) **N is the central atom**: Forms most covalent bonds	N	**5) Fill octet to terminal atoms.**	 **Lewis Structure**
3) Connect central atom to terminal Add terminal H atoms		**6) Check that total number in 5.**	3 bond pairs = 3 pairs 1 lone pairs = <u>1</u> pairs **4 electron pairs**

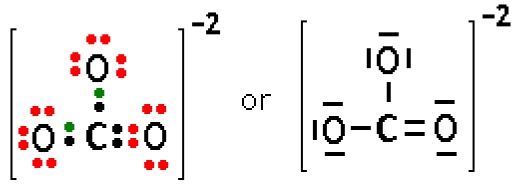
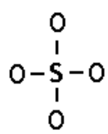
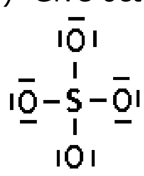
Problem: Draw the Lewis structures for the following molecules: H_2O

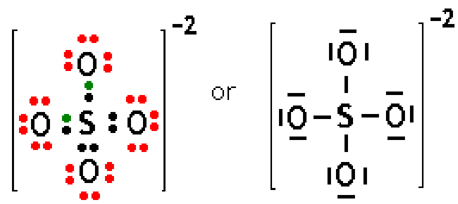
1) H_2O sum of valence electrons: 2 + 6 =8 ; 4 pairs **2)** Pick the central atom: usually the biggest atom and the atom forming most bonds. **O is the central atom** **3)** Connect central atom to other terminal atoms with a line representing a covalent bond (two electrons). **4)** Fill octet to central atom: A dot represents one electron and line represents two electrons. Check to see whether duet on H H–O–H	Count electrons and check valence electrons on C. 2 bond pairs = 2 x 2 = 4 2 lone pairs = 2 x 2 = <u>4</u> Total 8 = 4 electron pairs (an Octet) **Bond pairs**: an electron pair shared by two atom in a bond. E.g. two bond pairs between O-H in water. **Lone pair** : an electron pair found solely on a single atom.  E.g. two lone pairs found on the O atom at the top and the bottom.
H_2S has the same Lewis Structure as H_2O	
Since S is in the same **group VI A** as O. As a rule if an elements from the period 2 is replaced by another element from the period 3 and same group in the periodic table	They will have same Lewis structure. 
Draw Lewis structure of Nitrogen (N_2)	
sum of valence electrons: 5 + 5 =10 ; 5 electron pairs	**:N:::N:** triple bond between N and N
Draw Lewis structure of CO_2	

1) Valence electrons: 4 + 2 x 6 = 16 (8 pairs) 2) Central atom C;  3) Give octet to oxygen  Try to fill octet to C make double bonds	 or  4) Count electrons: 4 bond pairs = 4 pairs 4 lone pairs = <u>4</u> pairs 8 electron pairs
Draw Lewis structure of **CH_4**	
1) Sum of valence electrons: 4 (from one C) + 4 (from four H) = 8 electrons = **4 electron pairs** 2) Central atom is C 3) Octet on C atom and duet on H are already complete. 4) Count valence electrons on the Lewis structure:	bond pairs = 2 x 4 = 8 = **4 electron pairs** lone pairs = 0  or 
Draw Lewis structure of Ethane(CH_3CH_3)	
Draw two carbons and attach to them six hydrogens. Count valence electrons: 2C = 8 6H = <u>6</u> 14 (7 electron pairs)	Octet on each carbon and duet on each hydrogen  or 
Draw Lewis structure of nitric acid(HNO_3)	
H = 1 valence electron N = 5 valence electrons <u>O x 3 = 6 x 3= 18 valence electrons</u> Total of 24 valence electrons	 or 

5.7 Bonding in Compounds with Polyatomic Ions Present

<table>
<tr><td colspan="2">Draw Lewis structure of (CO_3^{2-}) 2 electrons are added to the total valence electrons</td></tr>
<tr><td>1) Valence electrons: 4 + (3 x 6) + 2= 24
(12 electron pairs)

2) Central atom C;

3) Give octet to 3 oxygens.

Try to fill octet to C make double bonds</td><td>
4) Count electrons:
4 bond pairs = 4 pairs
8 lone pairs = 8 pairs
12 electron pairs</td></tr>
<tr><td colspan="2">Draw Lewis structure of (SO_4^{2-}) 2 electrons are added to the total valence electrons</td></tr>
<tr><td>2) Valence electrons: 6 + (4 x 6) + 2= 32
(16 electron pairs)

2) Central atom S;
3) Give octet to 4 oxygens.


Try to fill octet to C make double bonds</td><td>
4) Count electrons:
4 bond pairs = 4 pairs
12 lone pairs = 12 pairs
16 electron pairs</td></tr>
</table>

5.8 Molecular Geometry

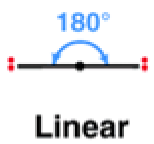
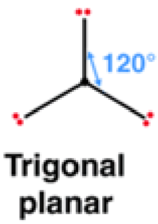
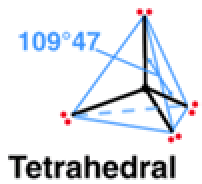
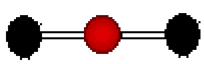
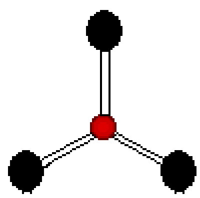
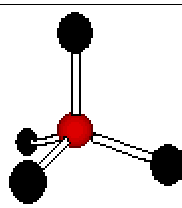
Lewis Structures and Molecular Geometry

Lewis structures are only helpful explaining the formation of covalent bonds and chemical stability of molecules. It does not provide information about the molecular shape or geometry as it is written and only account for octet rule. The valence shell electron pair repulsion (VSEPR) theory is an extension to Lewis theory of covalent bonding and helps to explain (or predict) molecular geometry which including **linear**, **trigonal planar**, and **tetrahedral** arrangements of terminal atoms to the central atom of a Lewis structure.

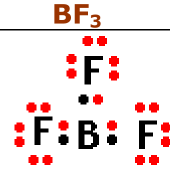
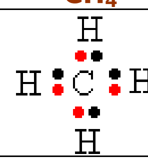
Valence-Shell Electron-Pair Repulsion Theory (VSEPR Theory)

In the theory, valence shell electron pairs are assumed to repel each other, assuming orientations to minimize repulsions and establish certain groups of molecular shapes.

Basic Geometries created by molecules with 2,3, and 4 electron around the central atom pairs.

<u>four things</u>	<u>three things</u>	<u>two things</u>
		
		

Molecules with no lone-pair electrons on central atom

Example	**BeF_2**	**BF_3**	**CH_4**
Lewis Structure			
Electron pairs around central atom	two	three	four
Type of Geometry only looking at bonding electron pairs	**Linear**	**Trigonal planar**	**Tetrahedral**

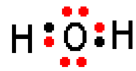
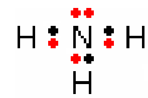
Molecules with lone-pair electrons on central atom

The molecular configuration is not exactly the same as the electron pair geometry of the central atom. The reason for this is that lone pairs of electrons not seen when you look at the bond connectivity which is the molecular geometry, although their effects are still seen (such as pushing bonding pair electrons closer together). This gives rise to more geometries (in addition to the 3 basic geometries: **linear**, **trigonal planar**).

The most common geometry with lone pairs is **angular structure** has four electron pairs surrounding the central atom. The water molecule (H:O:H) has two lone pairs and two bond pairs (4 total pairs on O). Another common geometry is **triangular pyramid** found in

NH_3 with only one lone pair of electrons.

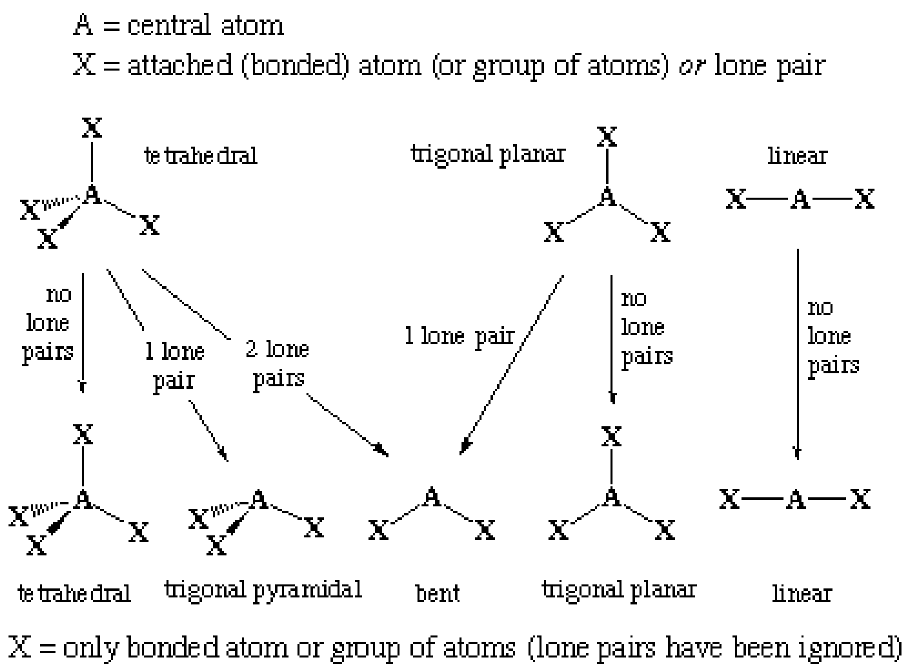
Geometries with lone pair of electrons

Example	H_2O	NH_3
Lewis Structure		
	AX_2E_2	AX_3E
Electron pairs around central atom	four	four
Lone pair electron pairs around central atom	two	one
Type of Geometry only looking at bonding electron pairs	**Angular** or **Bent**	**Triangular pyramid**

A = central atom; X= terminal atoms E= lone pair of electrons

Chemistry at a Glance: The Geometry of Molecules

Summary of geometries without/with lone pairs of electrons



5.9 Electronegativity

Electronegativity is useful in describing the electron distribution in a covalent bond. It is a measure of the tendency of an atom to attract a bonding pair of electrons. The electronegativity scale created by **Linus Pauling** is commonly used in predicting bond polarity. Fluorine (the most electronegative element) is assigned a value of 4.0, and values range down to cesium and francium which are the least electronegative at 0.7.

Periodic trend in electrongativity

As you go down a group, electronegativity decreases and crossing a period it increases.

increasing → F

increasing ↑

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	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9					

5.10 Bond Polarity

The electron pairs shared between two atoms in a covalent bond are not necessarily shared equally **except in homonuclear diatomic molecules**. The degree of electron pair sharing in a covalent bond leads to three types of covalent bonds.

a) **Non-polar covalent bond**
b) **Polar covalent bond**
c) **Ionic bond** (considered as an extreme case of covalent bonding)

Pure covalent bond

Equally shared electron pair in a covalent bond is called a pure covalent bond as found in homonuclear diatomic molecule, Cl_2 the shared electron pairs is shared equally.

Polar covalent bond is one in which **one atom has a greater attraction for the electrons than the other atom** as found in heteronuclear diatomic molecule, HCl the shared electron pairs is shared is attracted more towards the Cl because of its greater electronegetivity.

Bond Properties: Bond Polarity and Electronegativity

Bond polarity is a useful concept for describing the sharing of electrons between atoms

a) **Nonpolar covalent bond** is one in which the electrons are shared equally between two atoms

b) **Polar covalent bond** is one in which **one atom has a greater attraction for the electrons than the other atom**.

c) **Ionic bond:** If this relative attraction is great enough, then the bond is an **ionic bond**

Electronegativity Difference		
0-0.5	**0.6-1.6**	**1.7, or greater**
Non-polar covalent	**Polar covalent bond**	**Ionic**
C-H	O-H	Na-Cl

Polar covalent bonds

In the O-H bond **O** has an electronegativity of **3.5**, H has an electronegativity of **2.1**.

The difference in electronegativity is 3.5 - 2.1 = **1.4**

1.4 is less than 1.7, so the resulting bond is **polar covalent**.

Oxygen is the more electronegative so it will have a greater share of the bonding electrons and therefore a partial negative charge, $O^{\delta-}$

Hydrogen is less electronegative so it will have a lesser share of the bonding electrons and therefore a partial positive charge, $H^{\delta+}$

Since the bond has two 'poles' or 'ends' it is sometimes referred to as a dipole.

The polar covalent bond can be represented as:

$^{\delta-}O - H^{\delta+}$ or $\underset{<----+}{\mathbf{O - H}}$ Or $^{\delta-}\mathbf{O} \text{—:—} H^{\delta+}$

Where the arrow head points towards the most electronegative atom.

Ionic bond (considered as an extreme case of covalent bonding) In **heteonuclear compounds** such as NaCl the 3*s* electron is stripped from the Na atom and is incorporated into the electronic structure of the Cl atom - and the compound is most accurately described as consisting *of individual* Na^{+} *and* Cl^{-} *ions*

For most covalent substances, their bond character falls between these two extremes

5.11 Molecular Polarity

Lewis Structures and Polarity

A molecule containing only non-polar bonds is always non-polar molecule. Having a polar bonds do not always lead to molecular polarity. In a molecule with polar covalent bonds there is a possibility that they might add up to give a overall stronger molecular polarity or they might cancel each other to produce a non-polar molecule. Concepts of molecular polarity are better understood when viewed from the perspective of a Lewis structure. A polar molecule has at least one polar covalent bond. The electronegativity difference helps to assess the degree of polarity of a bond. A molecule containing polar bonds may be polar or non-polar, depending on the relative position of the bonds.

Molecular Polarity

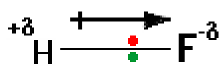
Molecules composed of covalently bonded atoms may also be polar or non-polar. For a molecule to have polarity, it must, of course, have polar bonds. But the key factor for determining the polarity of a molecule is its shape. If the polar bonds (**dipoles**) are **symmetrical** around the central atom, they offset each other and the resulting molecule is non-polar. However, if the dipoles are **not symmetrical** (**asymmetric**) around the central atom, the electrons will be pulled to one end of the molecule. The resulting molecule is polar. Ball and stick models are often used to demonstrate molecular shape. In this exercise you will build several covalent molecules and predict each molecule's polarity on the basis of its molecular shape.

The geometry of the molecule affects its **dipole moment** or **polarity** - the asymmetric distribution of positive and negative charges.

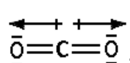
Polar Molecules will orient in a magnetic field When two elements with different electronegativities are bonded together, one obtains a polar bond.

Example: HF

Depending on the molecular shape, polar bonds can give rise to a polar molecule, one with an overall uneven distribution of electron charge. We can see this charge imbalance when the molecules are placed in an electric field. Polarity is a key feature of a molecule because it can influence physical, chemical, and even biological properties.

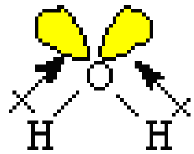
For a **hetero-diatomic molecule**, a polar bond *must* lead to a polar molecule. Consider **hydrogen fluoride**, 

Note the polar arrows with the negative end pointing to F atom. You can see that the F end of the molecule is much more negative than the H end, and thus HF is highly polar.

If a molecule has more than two atoms, its shape can affect the polarity in a crucial way. For example, in carbon dioxide, 

since oxygen is more electronegative than carbon, each bond is highly polar. But the linear molecular shape (**symmetrical**) makes the bond polarities cancel each other, so the CO_2 molecule is **non-polar**. Notice that the orientation of the molecules is random, whether the field is off or on.

In water, H_2O, the V-shape of the molecule allows the bond polarities to reinforce each other, so water is highly polar, as the large polar arrow shows.



The case of boron trifluoride is non polar similar to that of CO_2. Because the three highly polar bonds point to the corners of an **equilateral triangle**, the bond polarities cancel each other, and BF_3 is nonpolar.

Like BF_3, ammonia has four atoms and three polar bonds, but the **trigonal pyramidal** shape means that the bond polarities reinforce each other. Thus, ammonia is **highly polar** for the same reason that water is. Substituting a hydrogen for one of the chlorines gives chloroform, **$CHCl_3$** another tetrahedral molecule, but now the polar bonds reinforce each other. As a result, chloroform is highly polar.

Using VSEPR molecular structures predict the polarity (net dipole moment or zero dipole moment) of the following molecules: H_2O, NH_3, CO_2, and SO_3

Molecule	**Shape**	**Electron pair distribution**	**Polarity**
H_2O	bent	asymmetric	polar
NH_3	Trigonal pyramid	asymmetric	polar
CO_2	liner	symmetric	Non-polar
SO_3	Trigonal planer	symmetric	Non-polar

Chemistry at a Glance: Covalent Bonds and Molecular Compounds

5.12 Naming Binary Molecular Compounds

Most covalent compounds are formed by the reactions of non-metals with another non-metal. Covalent compounds exist as molecules and are named using prefixes that denote the number of each element present in the compound.

A covalent compound is given systematic name similar to an ionic compound following certain set of certain rules. Before the rules are made common names was given without following systematic rules: E.g. water for H_2O. The "shorthand" symbol for a covalent compound is its formula. Formula gives types atoms and numbers each atom in the chemical compound. Many familiar covalent compounds have common names. It is useful to correlate both systematic and common names with the corresponding molecular formula.

A **binary covalent compound** (a molecule) is consisting of only **atoms** of two elements in which one or more electron pairs are shared to form covalent bonds. When naming these compounds, its composition must be considered. In naming covalent compounds, some of the features used in naming ionic compounds are flowed but no the part that deal with ionic charge. The most electronegative element is normally named similar to the anion part of the ionic compound by adding a prefix –ide to end of the element name. E.g CO_2, is named carbon dioxide.

Binary molecular compounds are composed of only two elements. Examples are H_2O, NO, SF_6 etc. Sometimes these compounds have generic or common names (e.g., H_2O is "water") and they also have systematic names (e.g., H_2O, dihydrogen monoxide). The common name must be memorized. The systematic name is more complicated but it has the advantage that the formula of the compound can be deduced from the name. Since there are no charges prefixes has to be used in front of the element name to indicate number of atoms of each element in the compound. If there are four chlorine it would be tetra-chloride.

Prefixes used:

1	**mono**	**6**	**hexa**
2	**di**	**7**	**hepta**
3	**tri**	**8**	**octa**
4	**tetra**	**9**	**nona**
5	**penta**	**10**	**deca**

Compound	**Systematic name**	**Common name (if it has one)**
NF_3	nitrogen trifluoride	
NO	nitrogen monoxide **note: for first element we don't use mono- prefix**	nitr**ic** oxide
NO_2	nitrogen dioxide	
N_2O	dinitrogen monoxide	laughing gas nitr**ous** oxide
N_2O_4	dinitrogen tetraoxide	
PCl_5	phosphorous pentachloride	
SF_6	sulfur hexafluoride	

S_2F_{10}	disulfur decafluoride	
H_2O	dihydrogen monoxide	water
H_2S	dihydrogen monosulfide	hydrogen sulfide
NH_3	nitrogen trihydride	ammonia
N_2H_4	dinitrogen tetrahydride	hydrazine
PH_3	phosphorous trihydride	phosphine

Problem: Give the names of following formulas from the names of following covalent compounds names:

a. H_2S
b. CS_2
c. PCl_5
d. P_2O_5

Answer:
a. hydrogen sulfide
b. carbon disulfide
c. phosphorus pentachloride
d. diphosphorus pentoxide

Names of Acids and Bases
Binary acids: made up of only two elements - hydrogen and one other element.
Naming binary acids:
Begin with the prefix hydro.
Determine the "stem" - part of the name of the element that combines with hydrogen.
Add the suffix ic.
Examples:
HF -hydro fluor ic - **hydrofluoric acid**
HCl - hydro chlor ic - **hydrochloric acid**
HBr - hydro brom ic - **hydrobromic acid**
HI -hydro iod ic - **hydroiodic acid**
Ternary acids: made up of **three elements** - hydrogen, oxygen, and another element.
Naming ternary acids:
Acids made up of three elements including hydrogen
Determine the "stem" - part of the name of the third element.
The most common acid is given the suffix ic.
Add the prefix per for the acid with one more oxygen.
The suffix ous is given to the acid with one less oxygen.
Add the prefix hypo for the acid with two less oxygen atoms.
Examples:
$HClO_4$ - per chlor ic - perchloric acid - one more oxygen atom.
$HClO_3$ - chlor ic - chloric acid - the most common form of the acid.
$HClO_2$ - chlor ous - chlorous acid - one less oxygen atom.
HClO - hypo chlor ous - hypochlorous acid - two less oxygen atoms.
HNO_3 -nitric acid
HNO_2 -nitrous acid

H_2SO_4 - sulfuric acid
H_2SO_3 -sulfurous acid
H_3PO_4 -phosphoric acid
H_3PO_3 -phosphorous acid
H_3BO_3 -boric acid

Examples:
P_2O_5 - this is named diphosphorus pentoxide, because there are two phosphorus atoms and five oxygen
atoms.
CO - this is carbon monoxide (you need the "mono-" because there's only one oxygen atom).
CF_4 - this is carbon tetrafluoride, because there's one carbon atom and four fluorine atoms.

Chemical Connections: Nitric Oxide: A Molecule Whose Bonding Does Not Follow "The Rules"; Molecular Geometry and Odor