

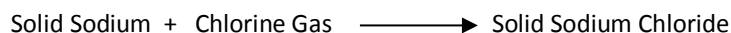
Unit 3: Chemical Bonding

Listed below are the learning objectives that you will be taught. At the conclusion of each lesson, check one of the four boxes to indicate your level of understanding for each individual objective.

Why Do Chemical Bonds Form

When sodium metal is heated and placed in a jar of chlorine gas, it burns with a bright flame. The result is a white solid, which has to be scraped from the sides of the jar.

The white solid is **sodium chloride**. It is formed by atoms of sodium and chlorine joining together, forming a **compound**. The reaction can be described like this:



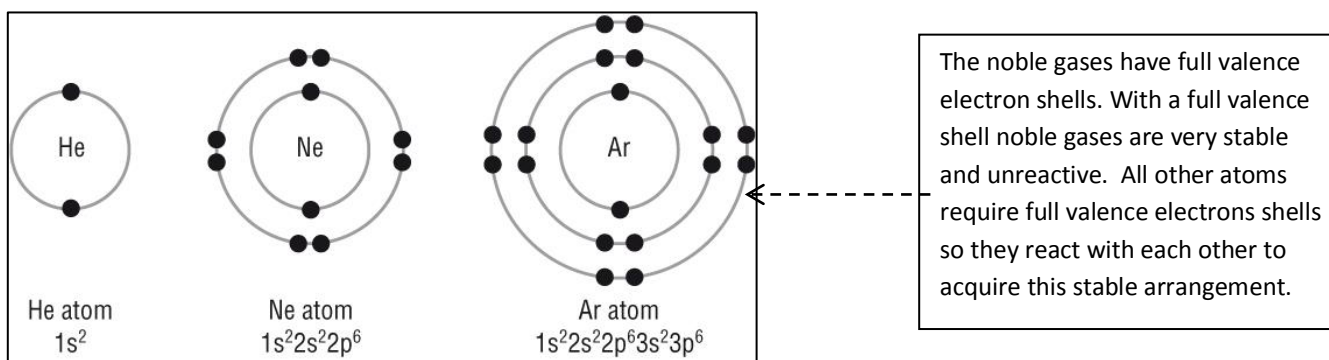
Like sodium and chlorine, most elements react to form compounds.

1. Define Chemical Bond

The losing, gaining or sharing valence electrons causes an electrostatic attraction between atoms. This attraction between atoms that holds them together in space is called a chemical bond. Without bonds the chair you are sitting in and the desk you are writing on would just fall apart into a pile of individual atoms.

2. Why do Chemical Bonds Form?

Bonds are formed during chemical reactions to help atoms fill their valence electron shells. Independent atoms have relatively high potential energy. Nature, however, favors arrangements in which potential energy is minimized. This means that most atoms are less stable existing by themselves with unfilled valence electron shells than when they are combined having full valence shells. By bonding with each other, atoms decrease in potential energy, thereby creating more stable arrangements of matter.



3. There are three types of chemical bonds

- Ionic Bond
- Covalent Bond
- Metallic Bond

Electronegativity

1. Define Electronegativity

Electronegativity is the ability of an atom to attract shared electrons to it. The higher the Electronegativity value the stronger the attraction for the shared electrons.

2. Electronegativity Trend

The Electronegativity increases as the atomic number increases in any given period. This is a result of the greater effective nuclear charge, Z_{eff} , which increases the attractive force of the nucleus and therefore it strengthens the attraction for the shared electrons.

Electronegative Values

Period	1A	2A	3B	4B	5B	6B	7B	8B	1B	2B	3A	4A	5A	6A	7A		
1	H 2.1																
2	Li 1.0	Be 1.5									B 2.0	C 2.5	N 3.0	O 3.5	F 4.0		
3	Na 0.9	Mg 1.2									Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0		
4	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 1.8	Ni 1.8	Cu 1.9	Zn 1.7	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
5	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
6	Cs 0.7	Ba 0.9	La* 1.1	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 1.8	Bi 1.9	Po 2.0	At 2.2
7	Fr 0.7	Ra 0.9	Ac† 1.1														

*Lanthanides: 1.1–1.3
 †Actinides: 1.3–1.5

The Electronegativity increases as the as the atomic number increases down a group. This is a result of an increase number of energy levels (n), which increases the distance over which the nucleus must pull on the shared electrons, reducing the attraction for electrons.

3. The electronegativity difference between two atoms can help determine the bond type

Type of bond		
Non-polar Covalent	Polar Covalent	Ionic

Lewis Dot Symbols

1. A Lewis Dot Symbol consists of the symbol of an element and one dot for each valence electron in an atom of the element. Elements in the same group have similar outer electron configurations and hence similar Lewis symbols. The transition metals, lanthanides, and actinides all have incomplete filled inner shells, and in general, we cannot write the Lewis dot symbols for them.

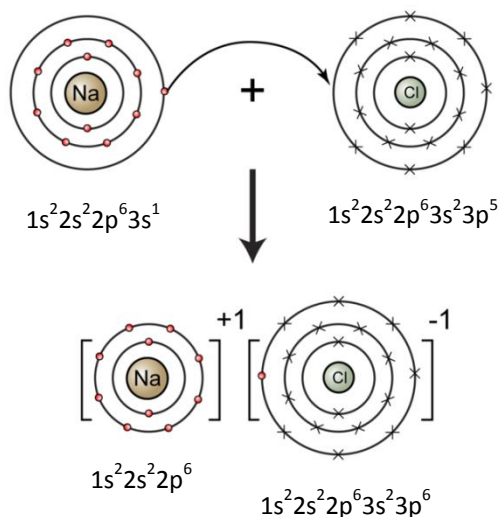
Group	1	2	13	14	15	16	17	18
Electron Configuration	[He]2s ¹	[He]2s ²	[He]2s ² p ¹	[He]2s ² p ²	[He]2s ² p ³	[He]2s ² p ⁴	[He]2s ² p ⁵	[He]2s ² p ⁶
Lewis Dot Symbol								

Ionic Bonding

1. Define Ionic Bond

Ionic bond is an electrostatic attraction between oppositely charged ions.

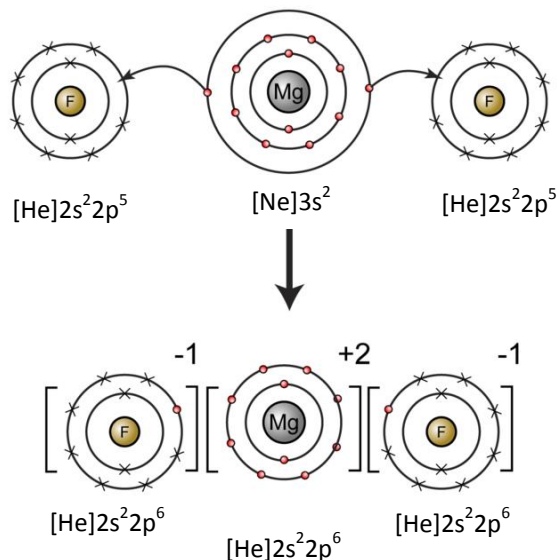
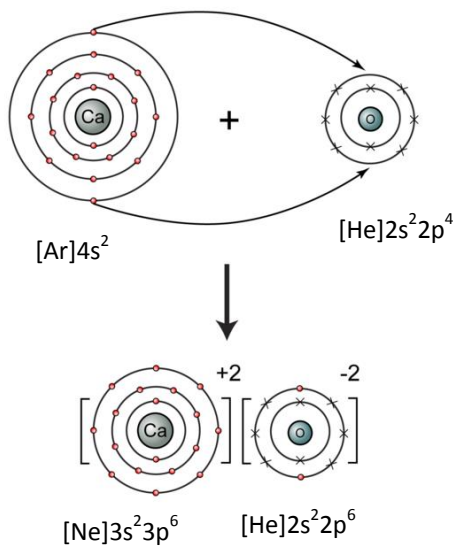
2. How are Ionic Bonds formed?



An ionic bond is created by:

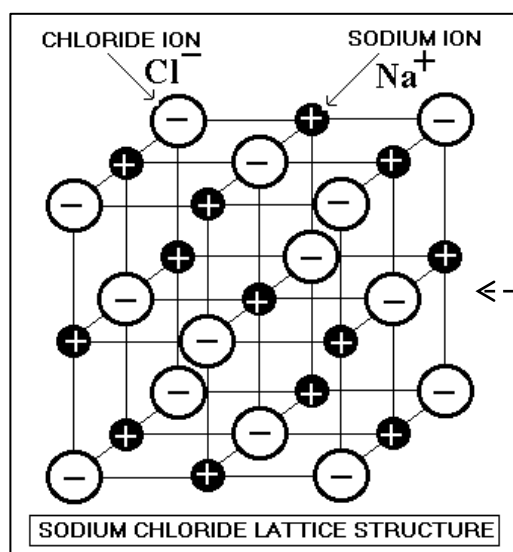
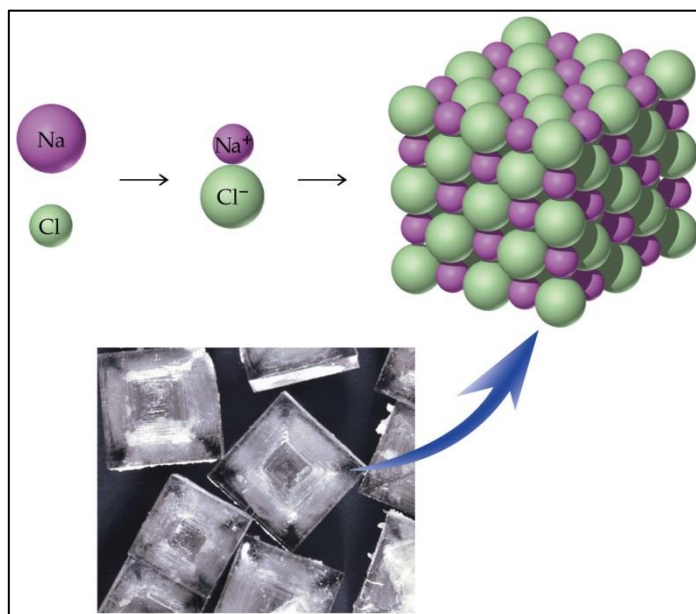
1. The metal atom transfers its valence electrons to the nonmetal atom. The metal becomes a positively charged cation.
2. The nonmetal atom gains valence electrons to fill its outer shell. The nonmetal becomes a negatively charged anion.
3. The cation and anion with opposite charges make, an attractive electrostatic (Coulombic) attraction that holds the two ions together creating an ionic bond.

Metal atoms having low ionization energies are able to transfer their electrons to nonmetal atoms with very negative electron affinity values. When this transfer of electrons occur each atom now has a stable octet. It is important to remember that the lower the ionization energy for a metal the more reactive it is. Nonmetals have high ionization energies so they don't lose electrons very well but they do have very negative electron affinity values so when they gain electrons they release energy which is highly favorable.



3. Ionic Compound

An **ionic compound** must contain both metal and nonmetal ions. These ions form ionic bonds between themselves making a giant 3-D structure called a crystal lattice that maximizes the attractions and minimizes the repulsions between ions resulting in a minimal potential energy state. The crystal lattice contains the same ratio of ions as that shown in the chemical formula. The Coulombic forces are very strong in these solid crystal structures. Coulomb's law predicts that larger forces and stronger interactions result when ions have greater charges, and when ions can get closer together (i.e., when ions are smaller).



The crystal lattice of sodium chloride Left. Each ion is surrounded by six ions with opposite charge. There are equal numbers of sodium and chlorine ions.

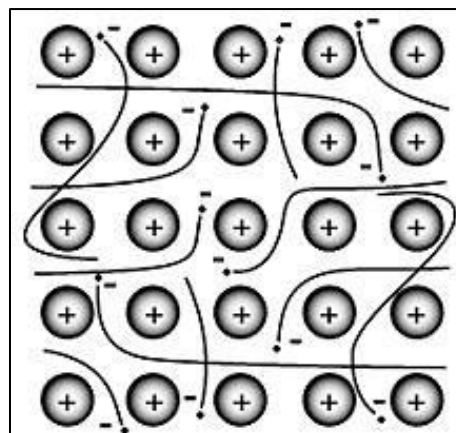
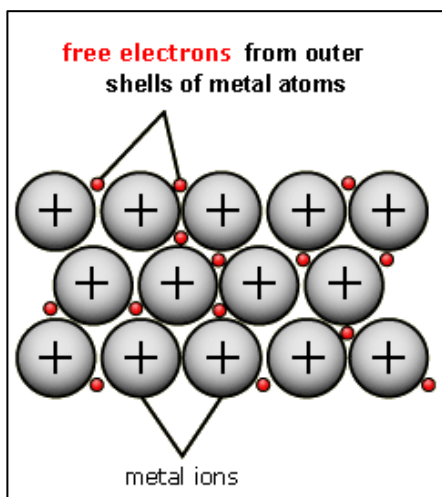
Metallic Bonding

1. What are delocalized electrons?

Metals have low ionization energies so they typically react with other elements by losing their valence electrons forming positive ions. The nuclei of these metal elements loosely hold onto their valence electrons. As a result if there is no nonmetals around to form ionic bonds then the valence electrons tend to “wander off”, or become delocalized.

2. How are metallic bonds formed?

Metals with delocalized electrons are no longer associated with any one metal nucleus but instead spread themselves throughout the metal structure. The metal atoms without these electrons become positively charged ions and form a regular crystal lattice (cation lattice) structure through which these electrons can move freely like a sea of electrons. The metal ions are held together by their attraction to the electrons between them. The strong forces of attraction are called metallic bonds.

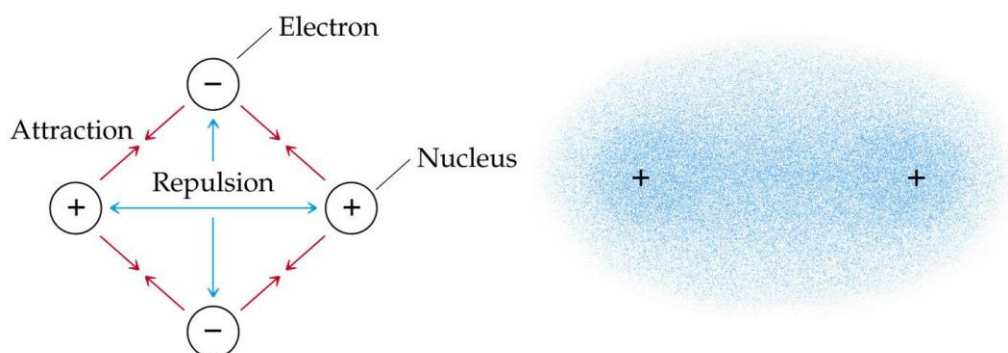


3. Metals and alloys can be represented as an array of positive ions surrounded by a “sea” of free-moving, delocalized electrons. This structure leads to many characteristic properties such as electrical and thermal conductivity, malleability, and ductility. **In general, the greater number of delocalized electrons the stronger the metallic bond.**

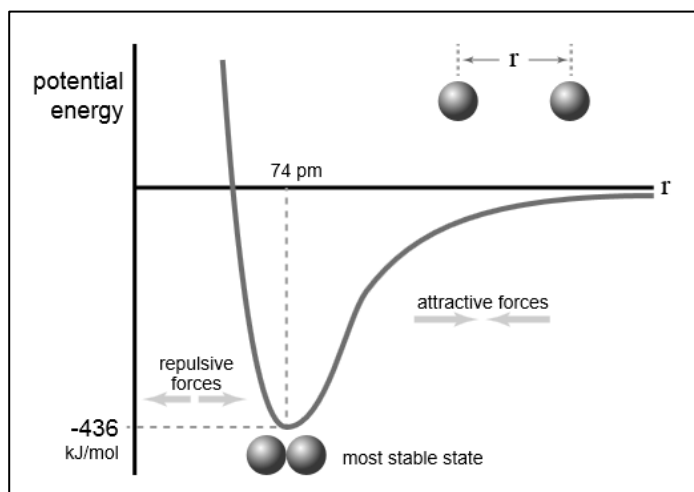
Covalent Bond

1. Define Covalent Bond

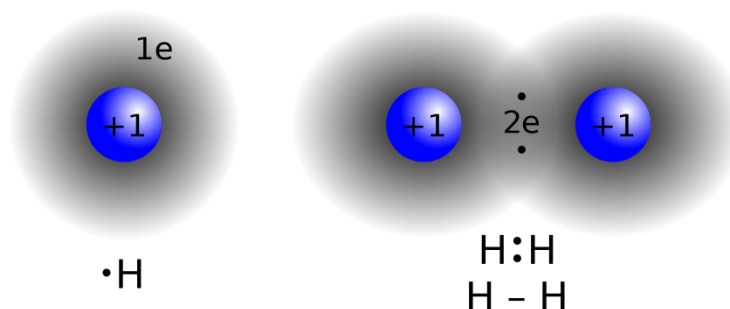
A covalent bond is formed by the electrostatic (Coulombic) attraction between a shared pair of electrons and the positively charged nuclei. Covalent bonds usually occur between nonmetal atoms that share their valence electrons to complete their outer shells (stable octet) forming molecules and polyatomic ions.



2. Covalent bonds are formed by minimizing the potential energy between atoms.



The relative strength of attraction and repulsive forces is related to Coulomb's Law. A Covalent bond is formed by balancing the attractive and repulsive forces between the two atoms. This balancing act will minimize the potential energy between the two atoms making a very stable state. The distance between the two nuclei forming a covalent bond is known as the **bond length**. The bond length occurs at the minimum potential energy where the attractive and repulsive forces are balanced.

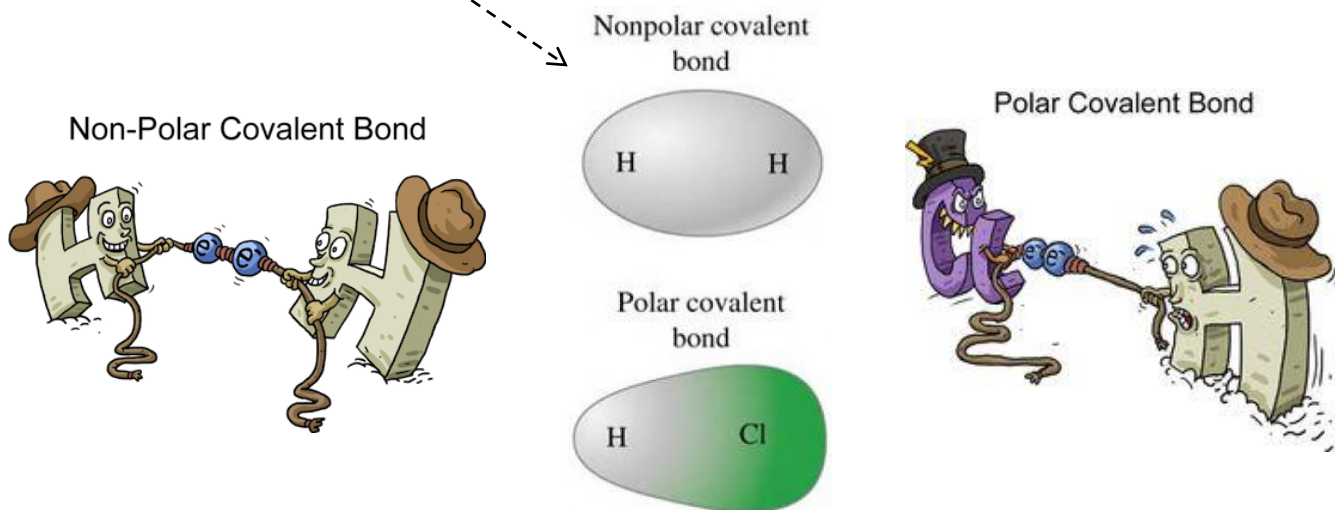


3. Polar vs. Nonpolar Covalent Bonds

A nonpolar covalent bond occurs when the two atoms sharing electrons have identical or nearly identical electronegativity values. This means that electrons will be shared equally between the two atoms.

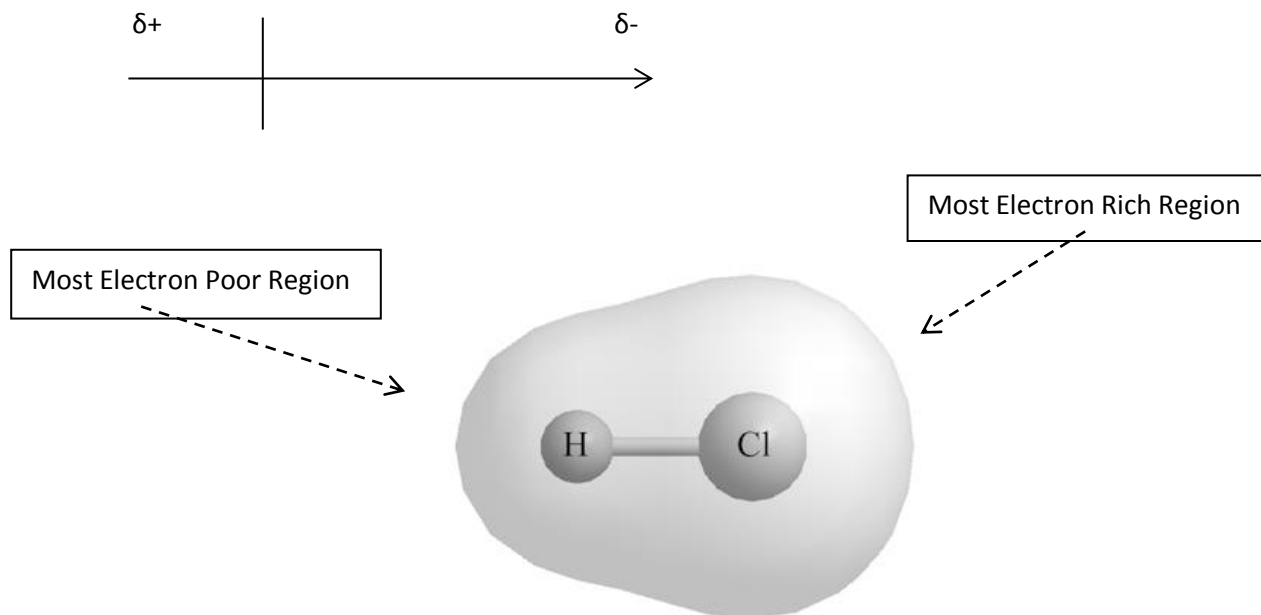
A polar covalent bond occurs when the two atoms sharing electrons have different electronegativity values. This means that electrons will not be shared equally between the two atoms. The atom that has the highest electronegativity value will spend more time with the shared electrons than the other atom. Atoms that have polar covalent bonds build up a partial negative and positive charge creating two oppositely charged poles.

Two Hydrogen atoms have identical electronegativity value that results in an equal sharing of electrons. The electron density is equally distributed so there will be no oppositely charged poles.



Hydrogen—chlorine covalent bond will have an unequal sharing of electrons. Chlorine has a higher electronegativity value than hydrogen resulting in an unequal distribution of electron density. This unequal distribution of electron density will create two oppositely charged poles. The higher electronegative atom will be assigned the partial negative charge because of the larger electron density. The other atom is assigned a partial positive charge.

The uneven distribution of charge is called a dipole, and dipole size increases with increasing difference in electronegativities. An arrow pointing from the partial positive charge (δ^+) toward the partial negative (δ^-) charge is the usual convention used to show the existence of a dipole.



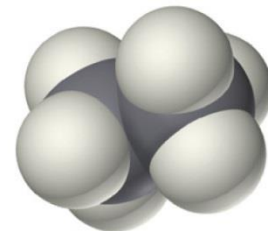
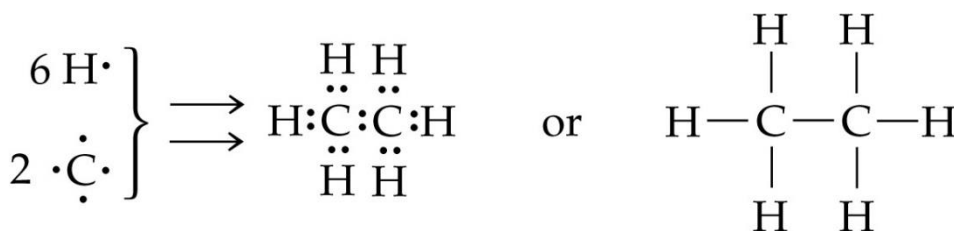
In diatomic molecules, the magnitudes of the two partial charges will be the same.

4. The electronegativity difference between two atoms can help determine the bond type

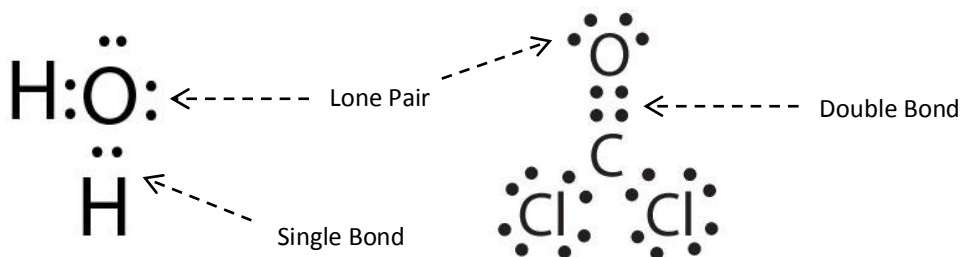
Type of bond		
Non-polar Covalent	Polar Covalent	Ionic

All bonds have some ionic character, and the difference between ionic and covalent bonding is not distinct but rather a continuum. Electronegativity difference alone does not determine ionic vs covalent. Compounds with a metal and nonmetal will be ionic and compounds with only nonmetal atoms will be covalent.

Lewis Dot Structures

Ethane, C₂H₆How to Draw a Lewis Dot Structure

- Count the total number of valence electrons for each atom in the molecule.
 - If the total charge is -1, -2, or -3, ADD 1, 2, 3 valence electrons, respectively.
 - If the total charge is +1, +2, +3, SUBTRACT 1, 2, 3 valence electrons, respectively.
- Determine the central atom of the molecule. The central atom is usually the one that needs the most electrons to fill its outer shell.
- Make a covalent bond between the central atom and the other atoms in the molecule by creating a bonding pair of electrons represented by two dots. The two dots in the bond represents the two valence electrons that are being shared between the two atoms.
- Subtract two electrons for each single covalent bond drawn in step 3 from the total number of valence electrons calculated in step one.
- Distribute the remaining electrons as lone pairs around the atoms.
 - Start from the outside atoms and work inward.
 - Attempted to achieve an octet/duet on each atom.
- If an atom is deficient of its octet, convert lone pairs from neighboring atoms into bonding pairs of electrons, thereby creating double and/or triple bonds.



1. Lewis Dot Structure Practice

H ₂	O ₂	HCl
N ₂	H ₂ S	CH ₄
NH ₃	CS ₂	HCN

VSEPR Theory (Molecular Shapes)

1. What does VSEPR Represent?

V

S

E

P

R

VSEPR is a model used to predict the shapes of covalent molecules and polyatomic ions from Lewis diagrams. VSEPR is based upon valence shell electrons repelling each other (via repulsive Coulombic interactions) so that they are as far apart as possible and, therefore, minimize that repulsion.

2. What are the steps to apply the VSEPR model?

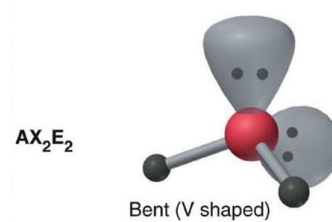
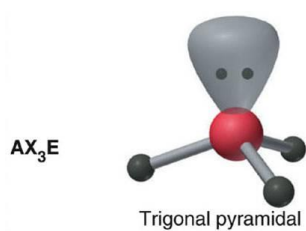
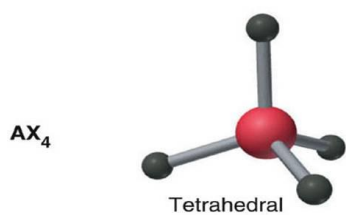
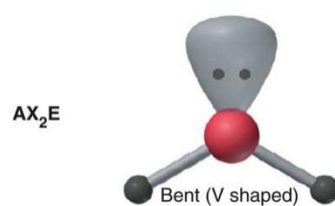
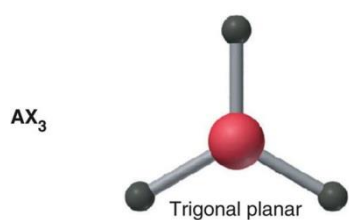
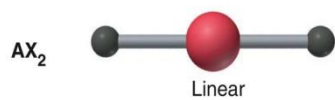
When applying the VSEPR model it is important to remember these three key ideas

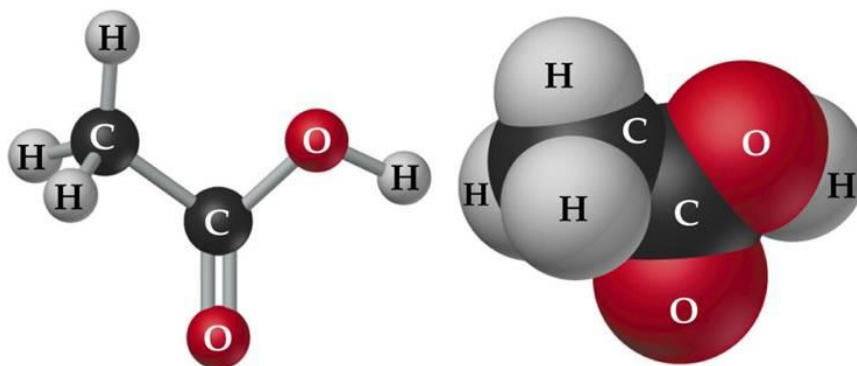
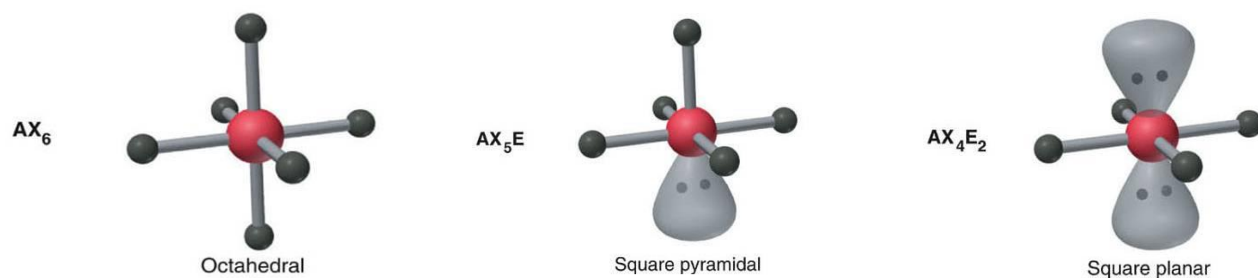
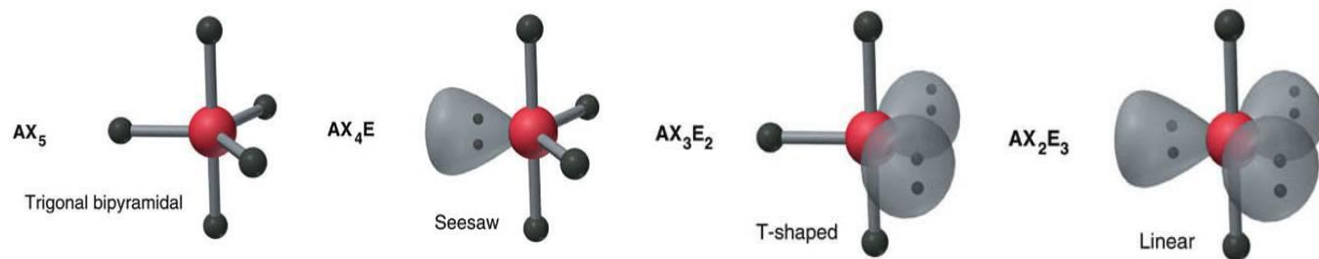
- 1) Count the number of electron pairs around the central atom (bonding pairs and lone pairs). Treat double and triple bonds as though they were single bonds.
- 2) Electron pairs repel each other (Coulombic interactions) and thus position themselves as far apart as possible.
- 3) Lone pair repels another lone pair or a bonding pair more strongly than a bonding pair repels another bonding pair.

Using these three key ideas you should be able to predict the approximate angle between bonding pairs which will help to identify the geometry.

3. Learning the VSEPR geometry.

- 1) For this class you will need to learn the geometric names and bond angles of the VSEPR system. This will require some memorization on your part. It is vital that you learn these geometric shapes.

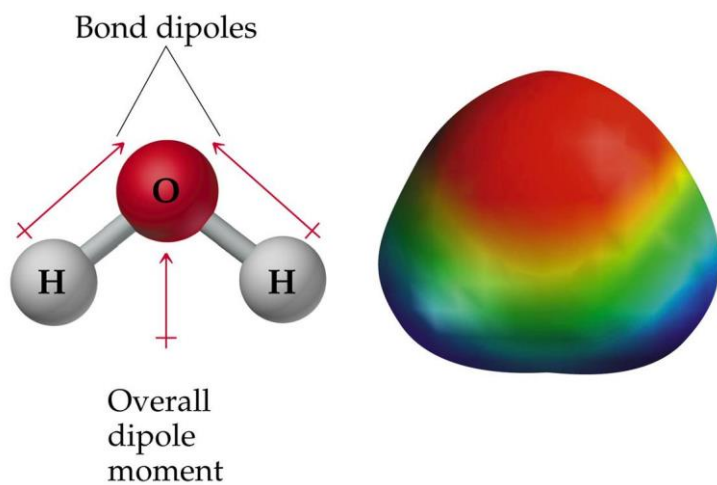
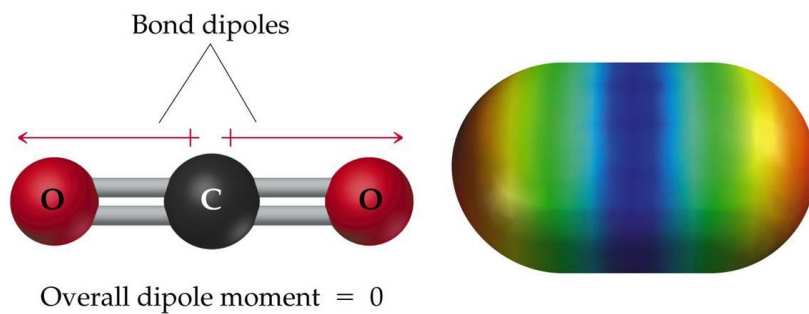




Polar/Nonpolar Molecules

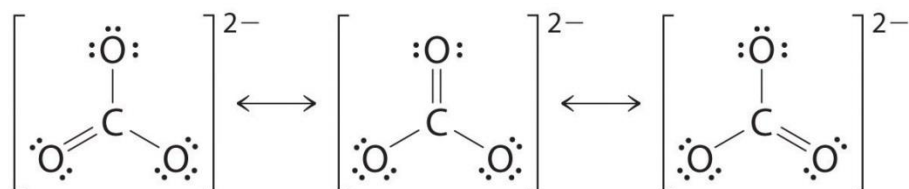
1. How to determine if a molecule is polar or nonpolar?

Some molecules have polar bonds but no dipole moment. This is because the individual dipoles cancel out due to symmetry and geometry. Molecules that **are symmetrical** are usually **nonpolar**.



Resonance Structure

1. Resonance is used when more than one feasible Lewis structure exists.
2. Resonance is often occurs when a double bond can be positioned in more than one place.
3. Bond lengths and strengths found in these species are “averages” of the single and double bonds present.
4. An example is the carbonate ion, as shown in the below diagram. Bond lengths and strengths are found to be intermediate between single and double bonds.

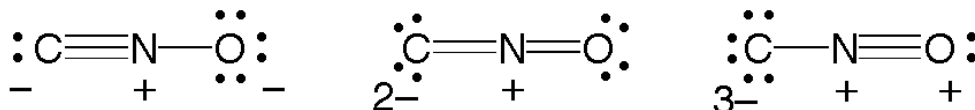
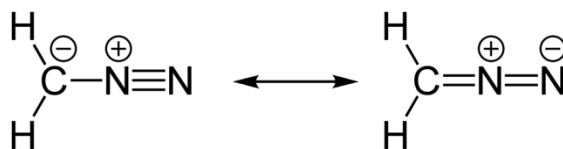


5. The formal charge of each atom within a structure can be calculated using the formula below.

Formal Charge on an atom within a Lewis structure	=	The number of valence electrons around that atom in the free atom.	-	The number of non-bonding electrons around that atom in a Lewis structure	- $\frac{1}{2}$	The number of bonding electrons around that atom in the Lewis structure
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6. The formal charge of an atom within a Lewis structure is generally used in one of two ways:
 - 1) To suggest where charge may most reasonably lay.
 - 2) To help select the most plausible structure from a set of resonance structures.

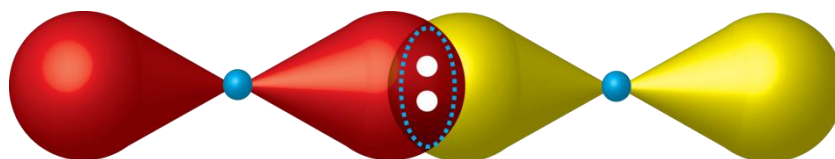
In order to determine which structure is most likely, choose the structure with zero formal charges, and/or formal charges with absolute values as low as possible, and/or keep any negative formal charges on the most electronegative atom.



CCl_4	NF_3	SO_3
H_2S	SO_4^{2-}	NH_4^+
NO_3^{1-}	ClF_3	SCN^{1-}
SeF_4	Cl_4	XeF_4

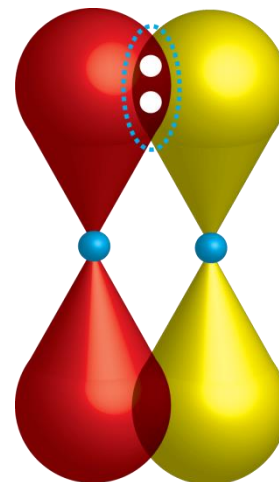
Valence Bond Theory (VBT)

1. Valence bond theory is based on the concept that bonds are formed by the overlap of atomic orbitals.
 - 1) Orbitals overlap to form bonds between atoms.
 - 2) One electron from each of the bonded atoms accommodates the overlapping orbital.
 - 3) Both electrons in the overlap are attracted to the nucleus of each atom. This is why electron pairs are located between two atoms in Lewis structures.
2. Sigma bonds (single bonds) are formed when an overlap occurs end-to-end of the bond
 - 1) The overlap of two s orbitals such as in H-H.
 - 2) The overlap of an s and p orbital such as in H-Cl.
 - 3) The overlap of two p orbitals such as Cl-Cl.



3. Pi bonds occur in double and triple bonds. Sideways overlap occurs with electron density concentrated above and below the plane of the nuclei of the bonding atoms.
 - 1) The sideways overlap of two p orbitals such as in C-C.
 - 2) Double bonds contain 1 sigma and 1 pi bond.
 - 3) Triple bonds contain 1 sigma and 2 pi bonds.

In the box below draw a double bond using valence bond theory



4. Sigma bonds are generally stronger than pi bonds, so they have larger bond energies. However, multiple (double and triple) bonds contain a combination of sigma and pi bonds, meaning that they tend to be shorter and stronger than sigma (single) bonds alone.

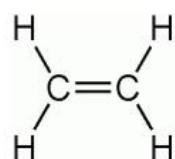
Hybridization

1. Hybrid orbitals are formed by mixing of s and p orbitals

2. How to form sp^3 hybrid orbitals?

3. How to form sp^2 hybrid orbitals?

4. How to form sp hybrid orbitals?



Ethylene (Ethene)

