AP Chemistry Ultimate Review Packet
Unit 2: Molecular and Ionic Compounds

Types of Chemical Bonds

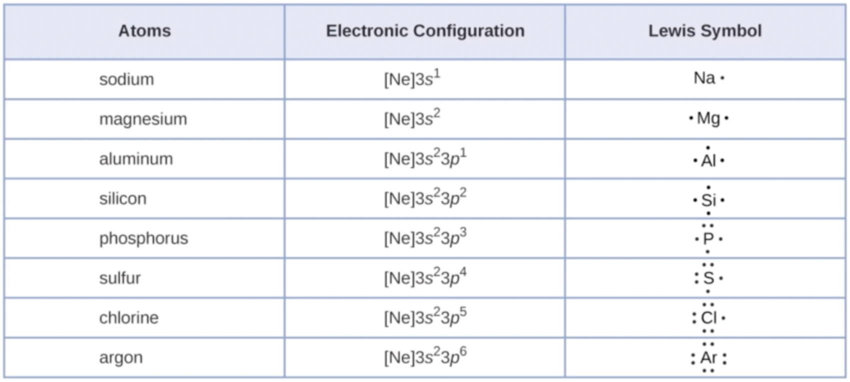
When atoms of two elements meet there are several possibilities depending on their electronegativities:

1) electronegativity difference > 1.8 → **ionic bond**
    - low ionization energy + high electron affinity = electron transfer
    - e.g. Na transfers an electron to Cl to produce Na^+ and Cl^-
    - favorable because each then has a full valence shell
    - formally charged ions form networks of ionic bonds
2) electronegativity difference < 0.4 → **nonpolar covalent bond**
    - two electrons are shared between two atoms
    - electron density is distributed roughly equally between the atoms
    - e.g. between atoms of the same element, or elements like C and H
3) electronegativity difference: 0.4 – 1.8 → **polar covalent bond**
    - two electrons are shared between two atoms
    - electron density is distributed unevenly
    - more electronegative element attracts most of the electron density
    - atom with electron excess has a partial negative charge (δ-)
    - atom with electron deficiency has a partial positive charge (δ+)
4) two metals: **metallic bond**
    - valence electrons are delocalized over a sea of positive metal cations
    - electron delocalization explains electrical conductivity, malleability, etc.

Lewis Diagrams

We can represent atoms and molecules using **Lewis symbols**

The dots are electrons and we place them in four coordination sites one at a time and then begin to pair them up (similar to Hund's rule)

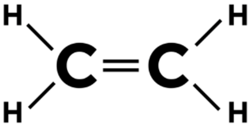


Atoms	Electronic Configuration	Lewis Symbol
sodium	$[Ne]3s^1$	$\text{Na}\cdot$
magnesium	$[Ne]3s^2$	$\cdot\text{Mg}\cdot$
aluminum	$[Ne]3s^2 3p^1$	$\cdot\dot{\text{Al}}\cdot$
silicon	$[Ne]3s^2 3p^2$	$\cdot\underset{\cdot}{\dot{\text{Si}}}\cdot$
phosphorus	$[Ne]3s^2 3p^3$	$\cdot\underset{\cdot}{\ddot{\text{P}}}\cdot$
sulfur	$[Ne]3s^2 3p^4$	$:\underset{\cdot}{\ddot{\text{S}}}\cdot$
chlorine	$[Ne]3s^2 3p^5$	$:\underset{\cdot\cdot}{\ddot{\text{Cl}}}\cdot$
argon	$[Ne]3s^2 3p^6$	$:\underset{\cdot\cdot}{\ddot{\text{Ar}}}:$

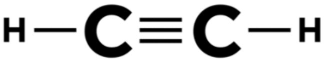
Lewis symbols can come together to form **Lewis diagrams** which represent compounds

1. Unpaired electrons on two atoms can come together to form a covalent bond (represented by a line, e.g. a chlorine-chlorine bond).

2. Adjacent atoms with two unpaired electrons each can make double bonds.



3. Adjacent atoms with three unpaired electrons each can make triple bonds.



4. Atoms like C, N, O, F, Ne will tend to be surrounded by 8 electrons (octet rule).

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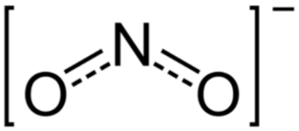
Atoms in a covalent compound can bear **formal charge** (+/-)

- Electrons contributed by an atom = typical valence → neutral atom
- More electrons than typical valence = formal negative charge (e.g. oxygen contributes 7 electrons vs. the usual 6)
- Fewer electrons than typical valence = formal positive charge (e.g. nitrogen contributes 4 electrons vs. the usual 5)

If multiple Lewis diagrams are possible for a compound these are **resonance structures**

- with these the pi electrons in double bonds and lone pairs are delocalized
- no resonance structure is an accurate representation of a molecule that is exhibiting resonance, we must look at the **composite resonance structure:**



- shows partial pi electron density distributed around the molecule

Resonance structures can have differing stabilities/contributing abilities:

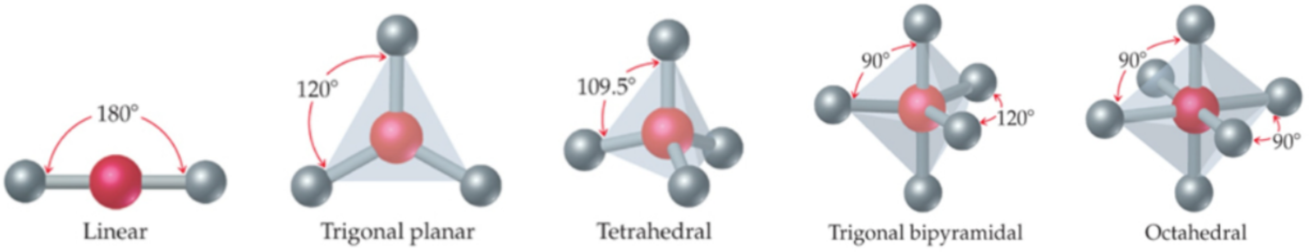
1. fewer formal charges is preferable
2. formal negative charges are preferred on more electronegative atoms
3. formal positive charges are preferred on less electronegative atoms
4. structures that maximize the number of octets are preferred

VSEPR Theory and Molecular Geometry

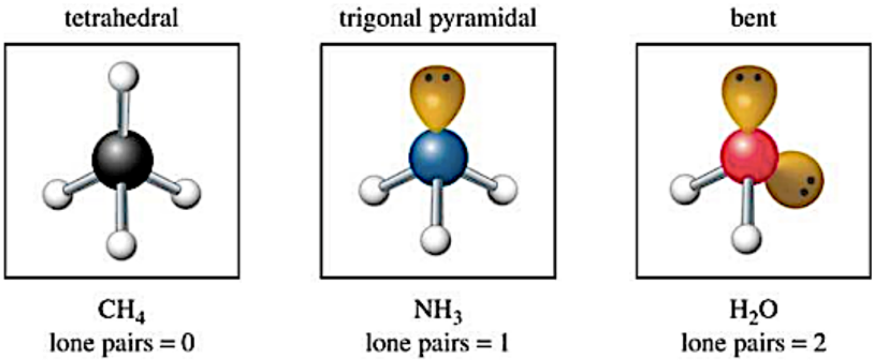
Molecules can be represented in three dimensions by **VSEPR theory**

- stands for valence shell electron pair repulsion
- operates on the principle that electron domains surrounding groups that are connected to a central atom will repel and spread out in space as much as possible
- both *atoms* and *lone pairs* qualify as electron domains
- we can refer to a table that correlates the number of electron domains around a central atom with its hybridization and the overall electron domain geometry

number of electron domains	hybridization of central atom	electron domain geometry
2	sp	linear
3	sp^2	trigonal planar
4	sp^3	tetrahedral
5	sp^3d	trigonal bipyramidal
6	sp^3d^2	octahedral

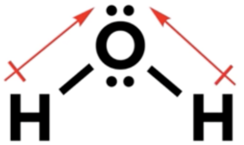


If atoms are replaced with lone pairs the electron domain geometry stays the same but the molecular shape changes (e.g. molecules with tetrahedral electron domain geometry below, and molecular shapes listed above)



By combining knowledge of polar covalent bonds and molecular geometry we can discern whether a covalent compound is *polar* or *nonpolar* overall

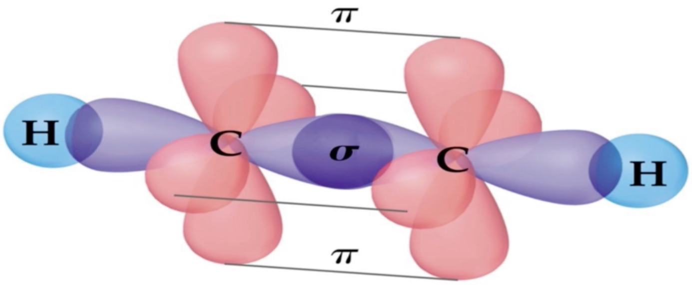
1. if there are no polar bonds the molecule is *nonpolar*
2. if there are only two atoms bound by polar bond(s) the molecule is *polar*
3. if there are multiple polar bonds between different sets of atoms, their bond dipoles must be combined to see if there is an overall **net dipole** (e.g. CO_2 the bond dipoles cancel, *nonpolar*, vs. H_2O, *polar*)



Sigma Bonds and Pi Bonds

Covalent bonds can be either sigma (σ) bonds or pi (π) bonds

- sigma bonds involve direct overlap of orbitals along the internuclear axis
- pi bonds involve lateral overlap of unhybridized p orbitals in the direction perpendicular to the internuclear axis
- double bonds involve one sigma bond and one pi bond
- triple bonds involve one sigma bond and two pi bonds (e.g. C_2H_2 below)



This molecule contains a triple bond between the carbon atoms

- the first bond is the **sigma bond** with direct overlap
- the second bond is a **pi bond** with lateral overlap in one plane
- the third bond is a **pi bond** with lateral overlap in another plane