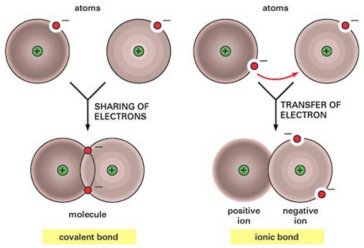
Introduction to Chemical Bonding

A chemical bond is the force that holds two atoms together in a molecule (or formula unit).

The three kinds of bonds are ionic, covalent, and metallic.



Types of Chemical Bonding

1. **Metal with nonmetal:**

   electron transfer and ionic bonding

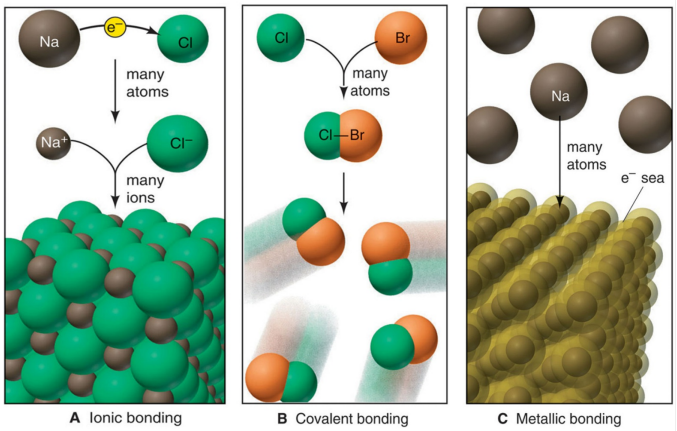
2. **Nonmetal with nonmetal:**

   electron sharing and covalent bonding

3. **Metal with metal:**

   electron pooling and metallic bonding

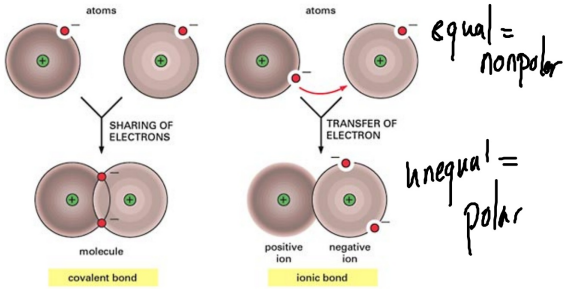
$^{+}NaCl^{-}$ Three Models of Chemical Bonding



A Ionic bonding **B** Covalent bonding **C** Metallic bonding

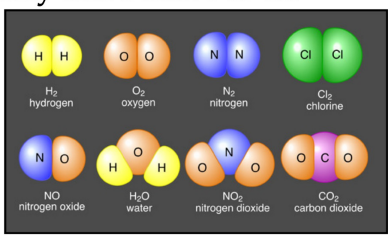
b Covalent Bonding

Covalent bonds occur when electrons are shared by 2 atoms. If the electrons are shared equally, the bond is called nonpolar covalent, and if the electrons are shared unequally, the bond is called polar covalent.



Molecules vs. Compounds

A molecule is defined as a structure of 2 or more atoms bound chemically and behave as an independent unit.



A compound is defined as a substance composed of two or more elements that are chemically bound in fixed proportions.

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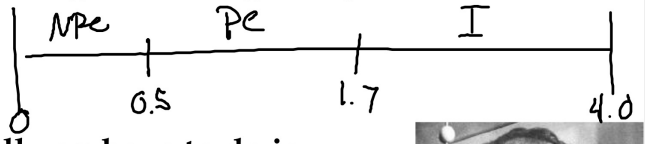
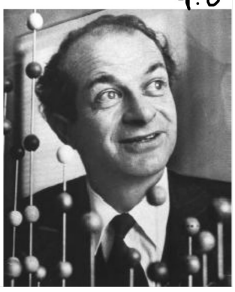
Classifying the Bond Type Mathematically

Pauling Electronegativity Scale:



All you have to do is subtract (and take absolute value).

Values found on good periodic table.



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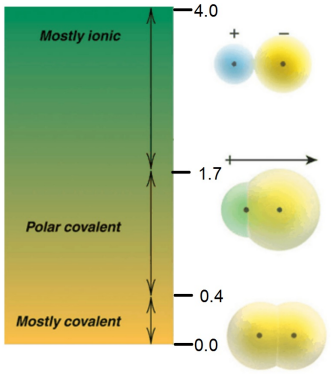
Electronegativity Difference Scale

Bond Polarity

ΔEN	IONIC CHARACTER
>1.7	Mostly ionic
0.4-1.7	Polar covalent
<0.4	Mostly covalent
0	Nonpolar covalent

ΔEN



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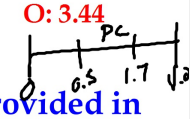
Bond polarity Practice

Examples:

1. H-H --> 2.20 - 2.20 = 0 NPC
2. H-Cl --> 3.16 - 2.20 = 0.96 PC
3. S-O ---> 3.44 - 2.58 = 0.86 PC

EN values:
H: 2.20
Cl: 3.16
S: 2.58
O: 3.44



Work on the practice problems provided in the handout. You may use the following key:

non-polar covalent = NPC
polar covalent = PC
ionic = I

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Lewis Electron-Dot Symbols

For main group elements -

1. **The A group number gives the number of valence electrons.**
2. **Place one dot per valence electron on each of the four sides of the element symbol.**
3. **Pair the dots (electrons) until all of the valenc electrons are used.**

Example:

Nitrogen (N) is in Group 5A (or 15) and therefore has 5 valence electrons.

Lewis electron-dot symbols for elements in Periods 2 and 3

Period	1A(1) ns^1	2A(2) ns^2	3A(13) ns^2np^1	4A(14) ns^2np^2	5A(15) ns^2np^3	6A(16) ns^2np^4	7A(17) ns^2np^5	8A(18) ns^2np^6
2	Li	Be	B	C	N	O	F	Ne
3	Na	Mg	Al	Si	P	S	Cl	Ar

The octet rule states that when atoms bond they ofte lose, gain, or share electrons to attain a filled outer l of 8 electrons.

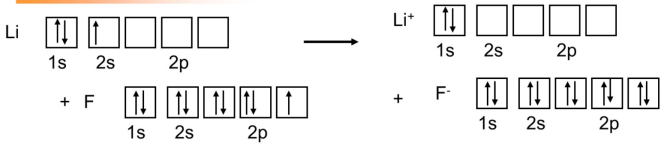
Exceptions: 2 for H and Li (Duet Rule)

Three Ways to Depict Bonding

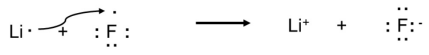
Electron configurations

Li $1s^2 2s^1$ + F $1s^2 2s^2 2p^5$ ⟶ Li^+ $1s^2$ + F^- $1s^2 2s^2 2p^6$

Orbital diagrams



Lewis electron-dot symbols



Our focus will be the Lewis Model

How to Draw Lewis Structures

Step 1: Determine the total number of valence electrons.

Step 2: Place atom with lowest electronegativity in center.

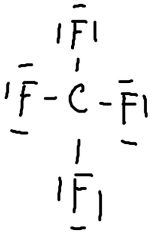
Step 3: Draw single bonds.

Step 4: Obey the octet rule (or the Duet Rule, Hydrogen only). Give each atom 8 electrons.

Examples and Practice

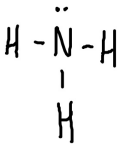
CF_4

$4+4(7) = 32$



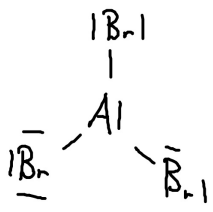
NH_3

$5+3(1)=8$



$AlBr_3$

$3+3(7)=24$

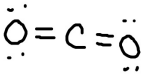


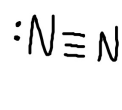
electron deficit

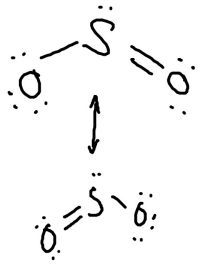
Structures with Multiple Bonds

Sometimes in order to get eight electrons around an atom I have to put multiple bonds (double and triple bonds) into the structure.

CO_2 **nitrogen gas** **sulfur dioxide** SO_2



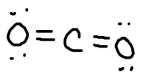


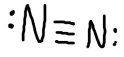


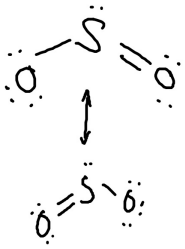
Generalized rules to Guide your Lewis Structures

Central atom	# of bonds	# of lone pairs
Group 13	3	0
Group 14	4	0
Group 15	3	1
Group 16	2	2
Group 17	1	3

CO_2 **nitrogen gas** **sulfur dioxide**







Rule Breaking Structures
Electron Deficient and Expanded Octets

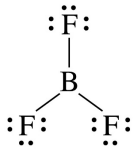
1A	2A											3A	4A	5A	6A	7A	8A
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg	3B	4B	5B	6B	7B		8B			1B	2B	Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt									

The octet rule occasionally fails for the shaded main-group elements. These elements, all of which are in the third row or lower, can use low-energy unfilled d orbitals to expand their valence shell beyond the normal octet. (only as a central atom)

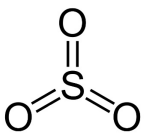
Structures that violate the Octet Rule

In some cases, structures will violate the octet rule

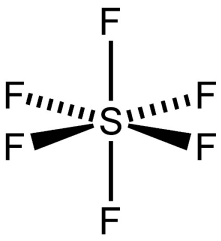
BF_3



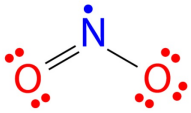
SO_3



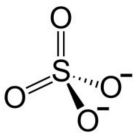
SF_6



NO_2



SO_4^{2-}



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Dative Bonds (coordinate-covalent bond)

Notice the change in molecular shape for the central atoms.

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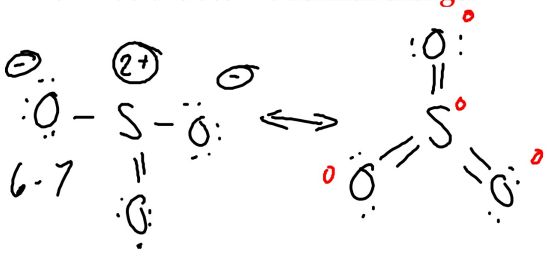
Formal Charge

SO_3

The charge an atom is assigned in a molecule in known as the atom's formal charge.



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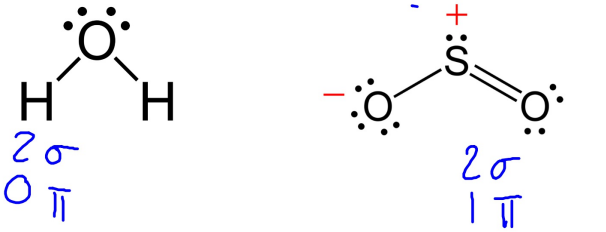
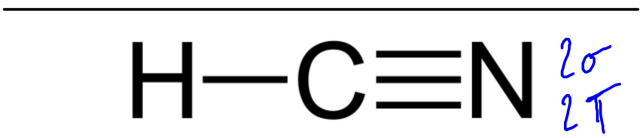
Sigma and pi bonds





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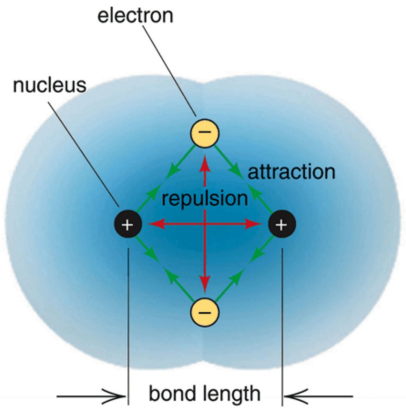
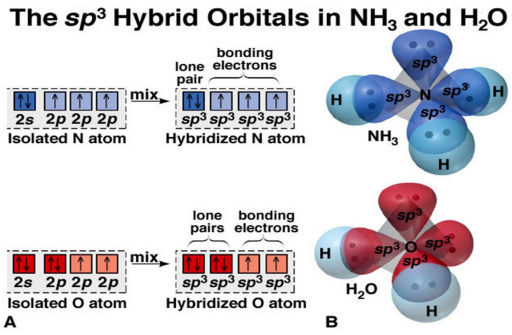
Bond Length and Bond Strength

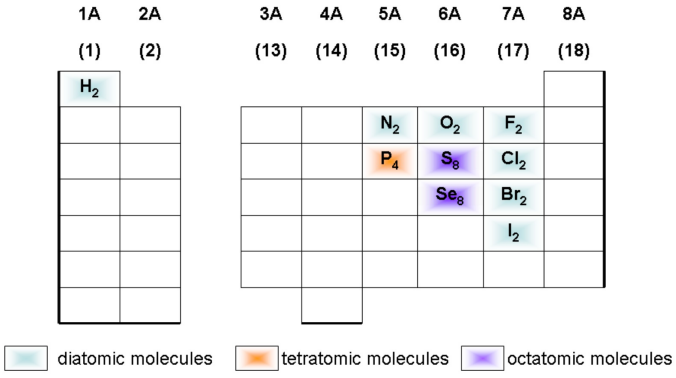


Hybridization

an atomic orbital postulated to form during bonding by the mathematical mixing of specific combinations of nonequivalent orbitals in a given atom.



Diatomics and More Review



Theories of Bonding

Valence Bond Theory

Valence Shell Electron-Pair Respulsion Theory

Molecular Orbital Theory

Crystal Field Theory

Ligand Field Theory

Valence-Shell Electron-Pair Repulsion Theory (VSEPR)

VSEPR theory states that each group of valence electrons around a central atom is located as far as possible from the others, to minimize repulsions.

A "group" of electrons is any number of electrons that occupies a localized region around an atom. The following are all considered to be one (1) electron group:

(1) single bond
(2) double bond
(3) triple bond
(4) lone pair
(5) free radical