Chemical Bonds
Formation of Compounds from atoms

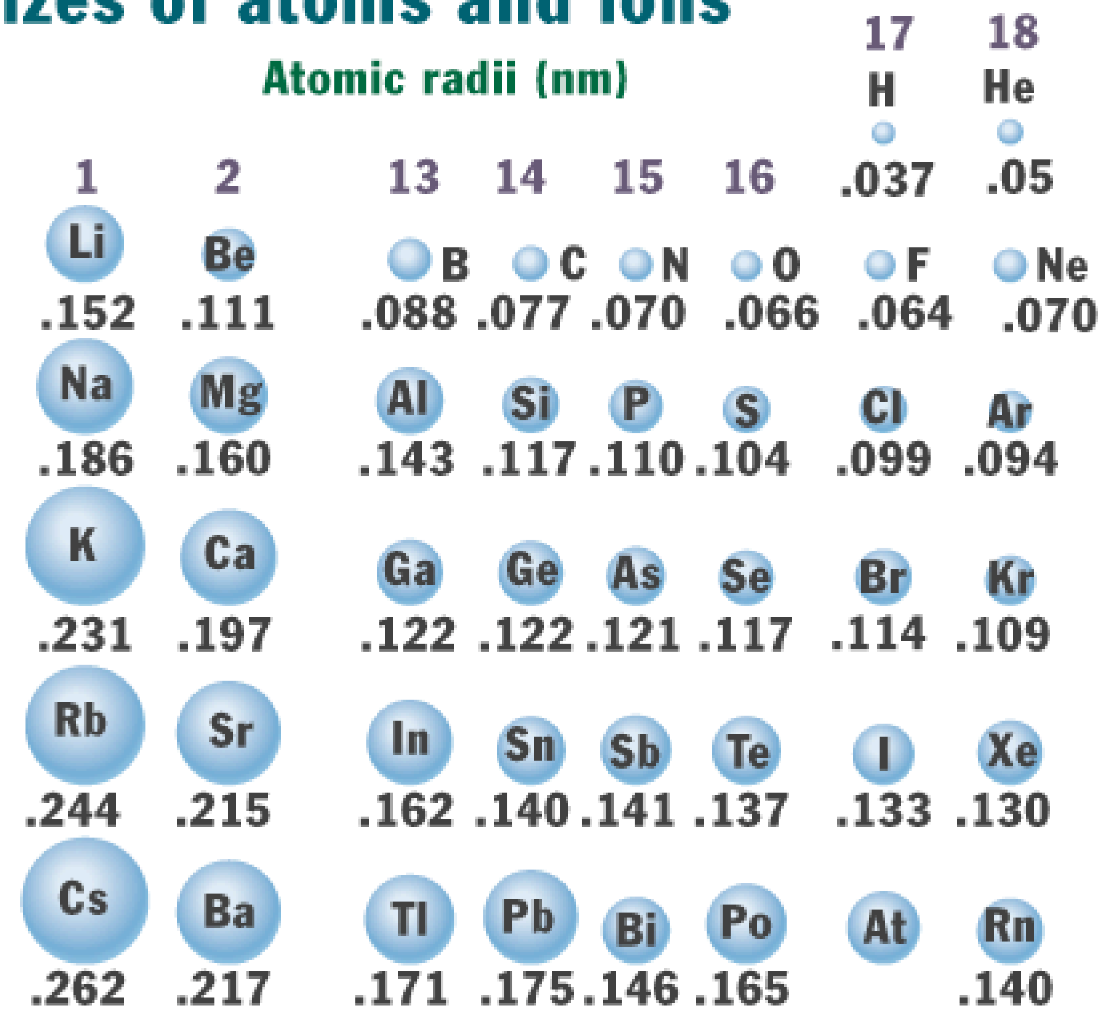
Preparation for College Chemistry
Columbia University
Department of Chemistry

Trends in the Periodic Table

Lewis Structures

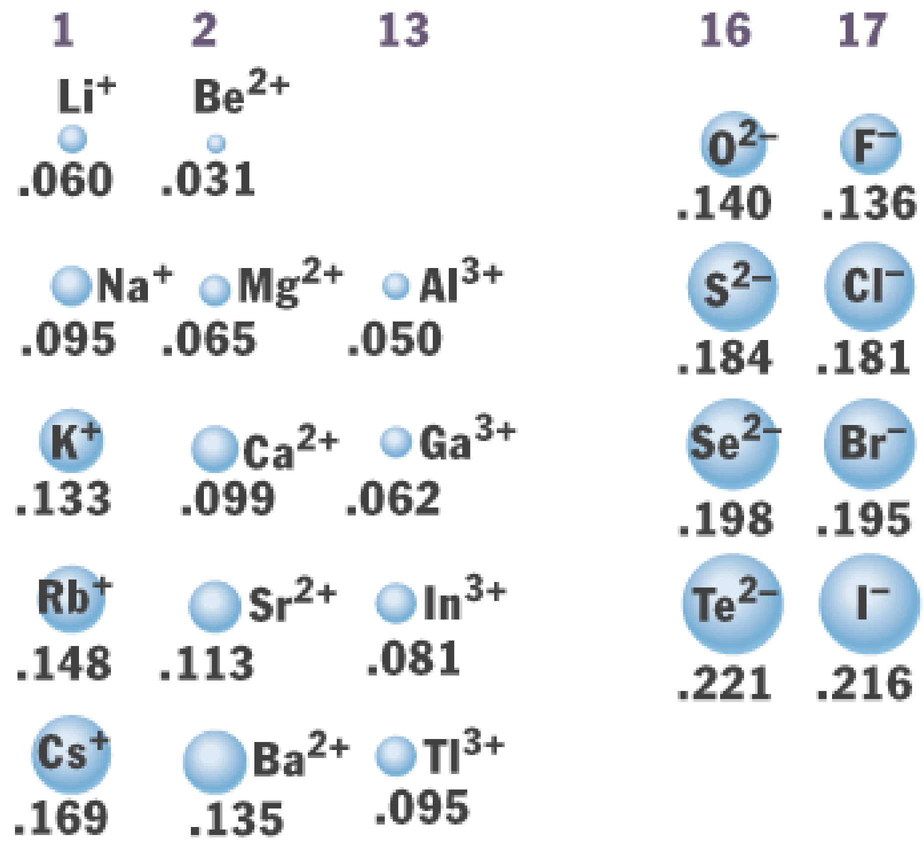
VSEPR Model

Sizes of atoms and ions

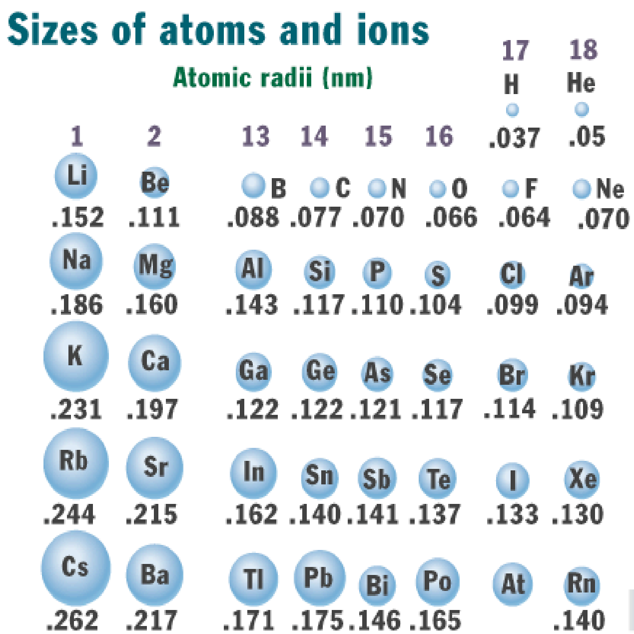


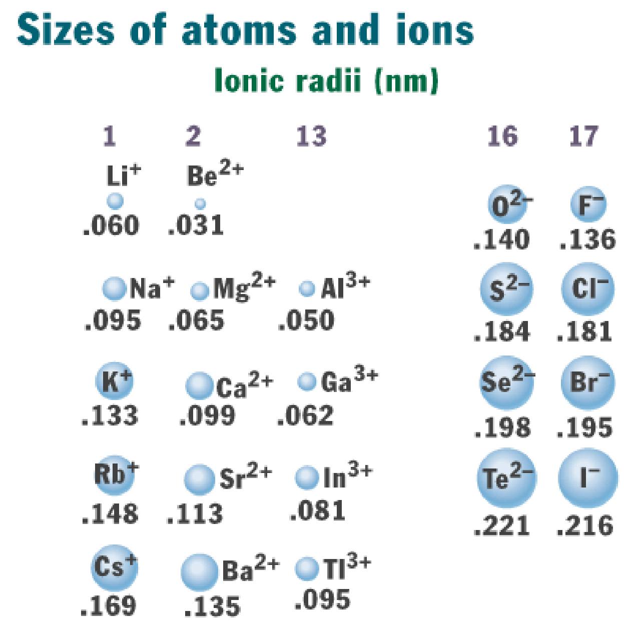
Sizes of atoms and ions

Ionic radii (nm)



Atomic and Ionic Radii





Decreases going across a period from left to right, increases going down group

Ionization Energy

Minimum energy necessary to remove an electron from a neutral gaseous atom in its ground state (IE > 0, ground state stable system)

$$X_{(g)} \longrightarrow X^{+}_{(g)} + e^{-} \qquad \Delta E = IE_1$$

$$X^{+}_{(g)} \longrightarrow X^{2+}_{(g)} + e^{-} \qquad \Delta E = IE_2$$

First Ionization Energies

Ionization energy increases

1	2	13	14	15	16	17	18
						H 1312	He 2372
Li 520	Be 900	B 801	C 1086	N 1402	O 1314	F 1681	Ne 2081
Na 496	Mg 738	Al 578	Si 786	P 1012	S 1000	Cl 1251	Ar 1520
K 419	Ca 590	Ga 579	Ge 762	As 944	Se 941	Br 1140	Kr 1351
Rb 403	Sr 550	In 558	Sn 709	Sb 832	Te 869	I 1009	Xe 1170
Cs 376	Ba 503	Tl 589	Pb 716	Bi 703	Po 812	At	Rn 1037

Electron Affinity EA

Electron attachment energy. Energy released when a gaseous atom in its ground state gains a single electron.

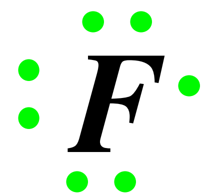
$$X_{(g)} + e^{-} \longrightarrow X^{-}_{(g)}$$

$$EA \begin{cases} > 0 \\ < 0 \end{cases}$$

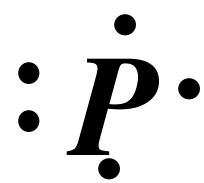
Lewis Structures of Atoms

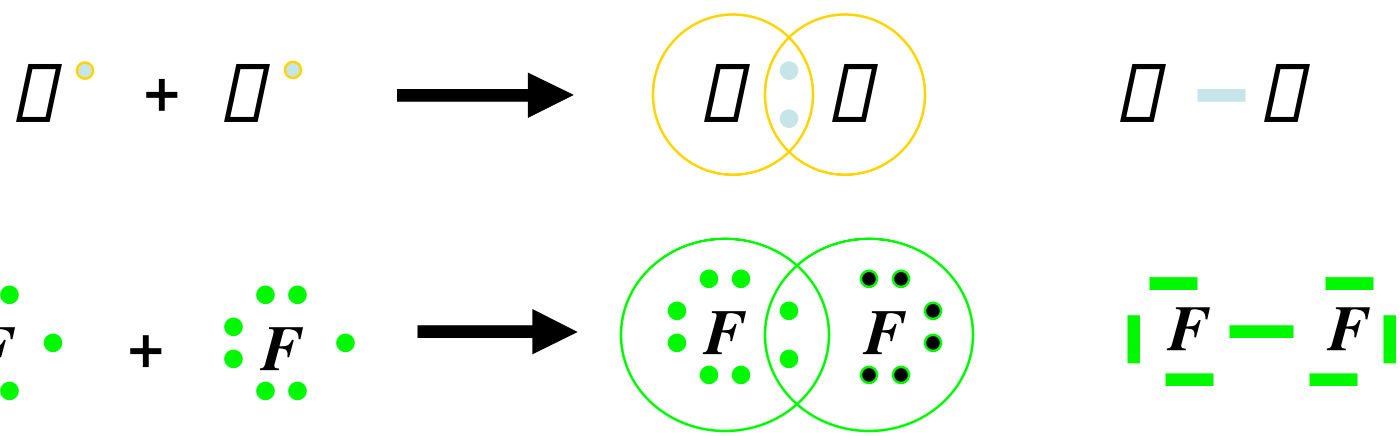
Gilbert Lewis

F $2s^2p^5$

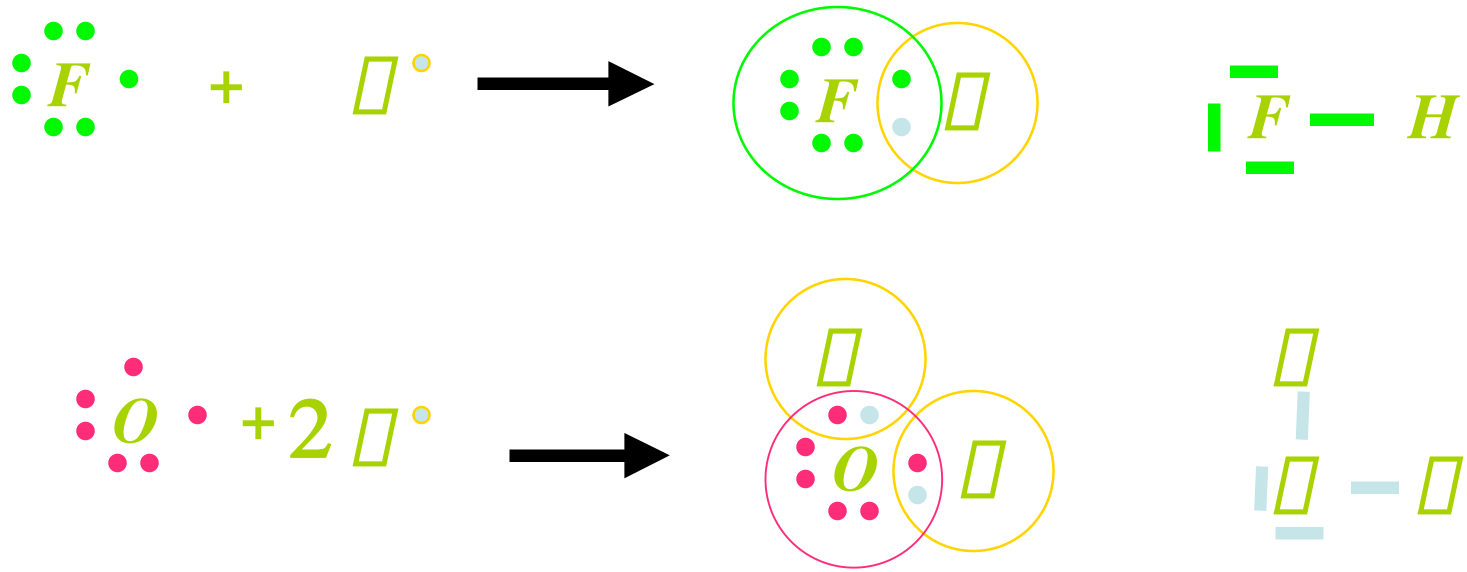


P $2s^2p^3$





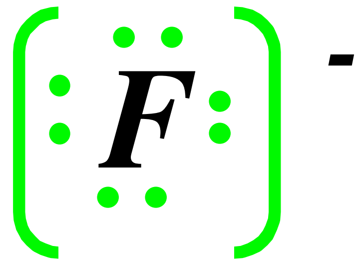
The Octect Rule



In H_2O and HF, as in most molecules and polyatomic ions, nonmetal atoms except H are surrounded by 8 electrons (an octet). Each atom has a noble gas electronic configuration (ns^2p^6).

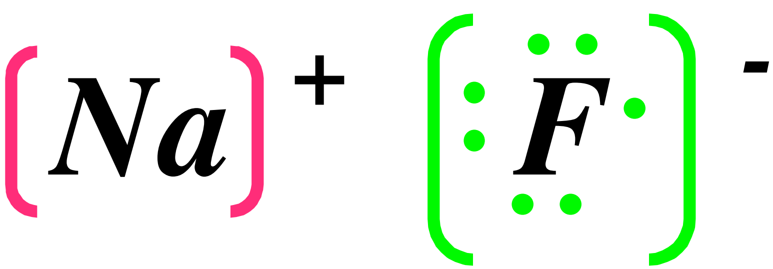
Ionic Bond. Electron Transfer

F^- $2s^2p^6$

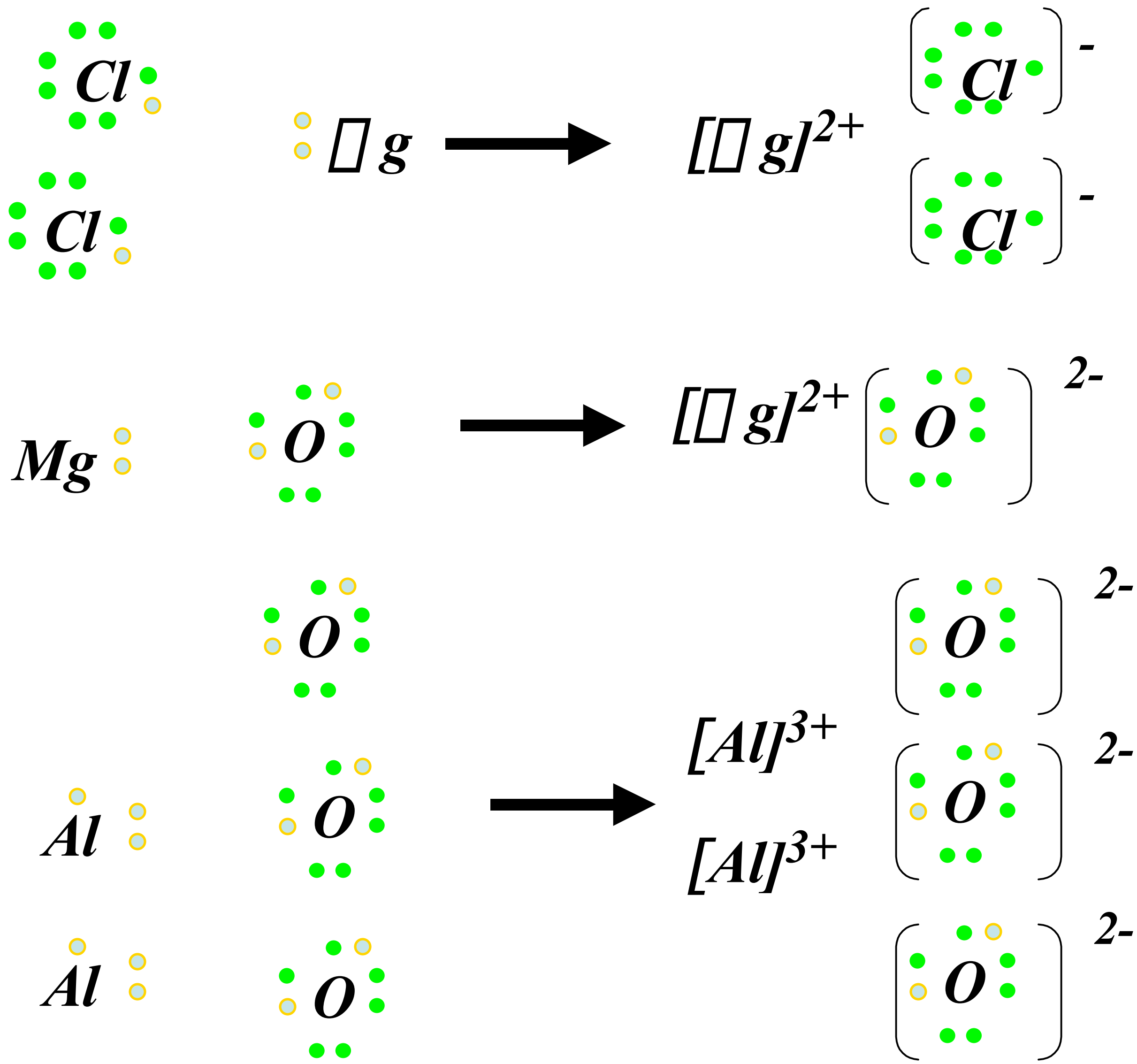


Na^+ $2s^2p^6$

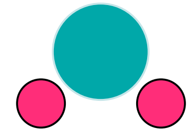
$[Na]^+$

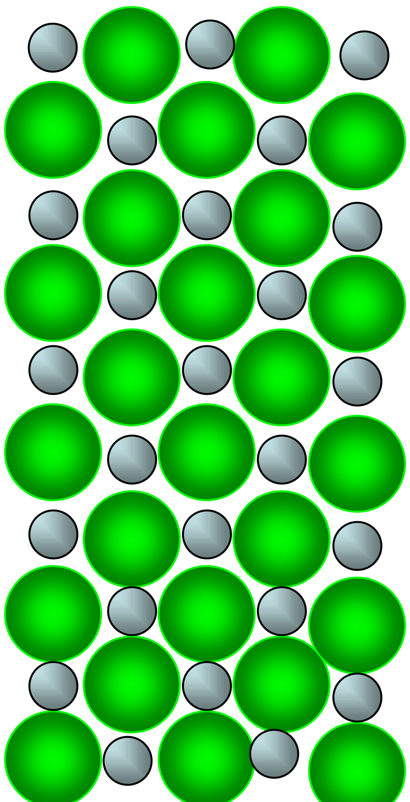


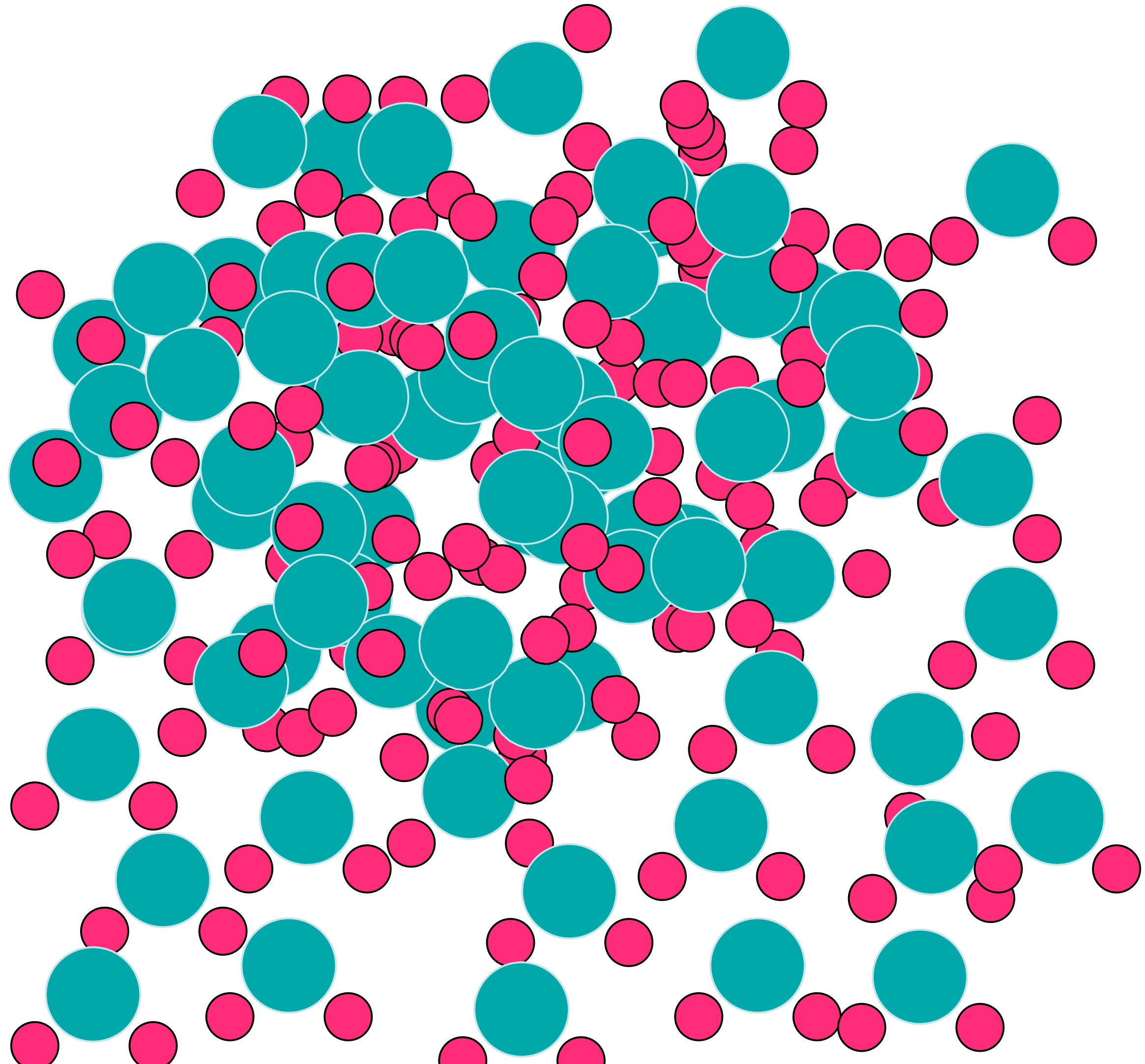
Ionic Bond. Electron Transfer



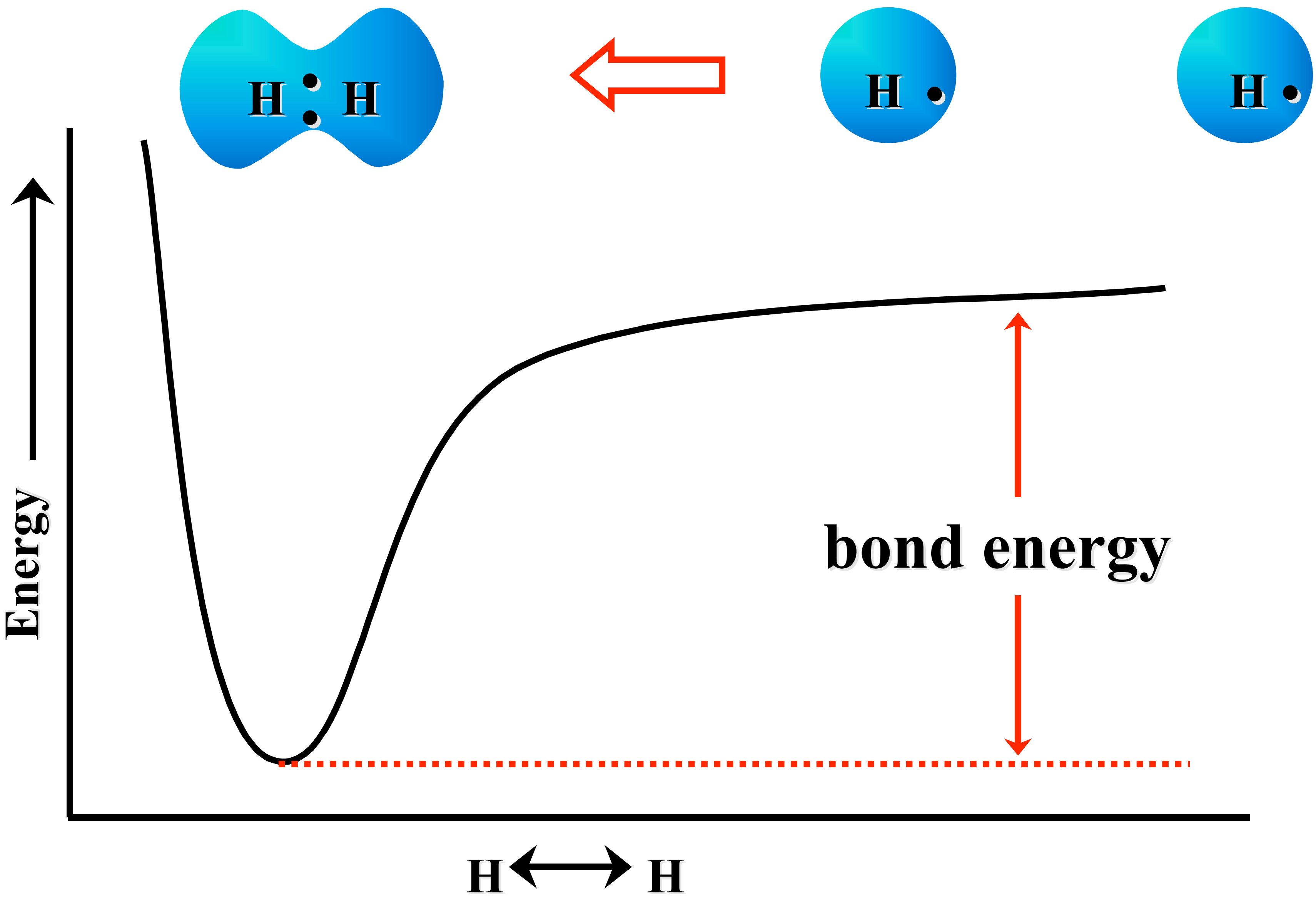
Ionicity vs. Covalency







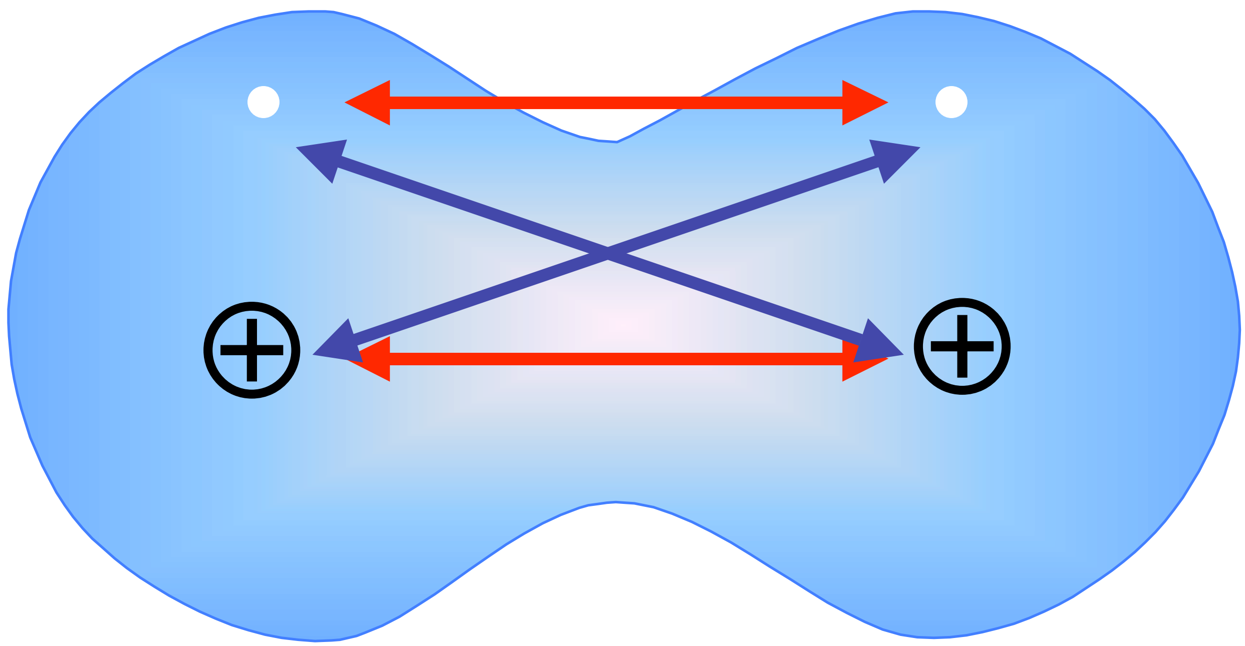
Potential Energy Diagram



Electrostatic forces in the H_2 molecule

electron repulsion (destabilization)

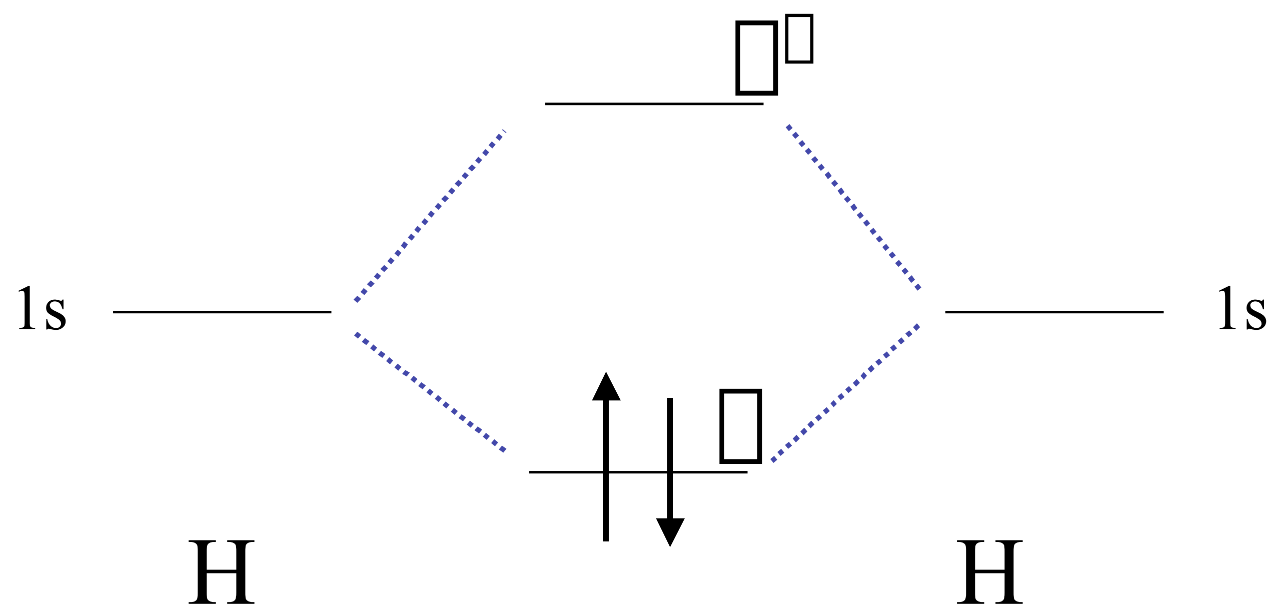
nuclear repulsion (destabilization)



electron-nuclear attraction (stabilization)

H_2 Electron Configuration:

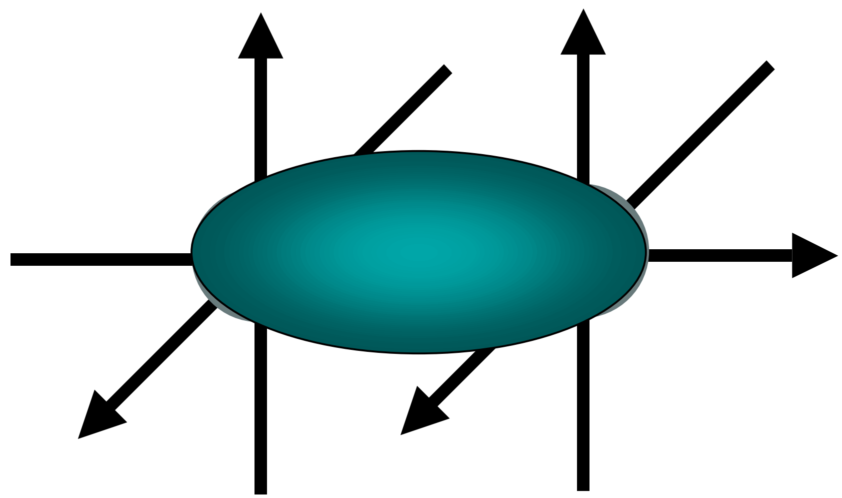
Bonding and Non-Bonding Orbitals



Two s atomic orbital = Two σ molecular orbital (MO)

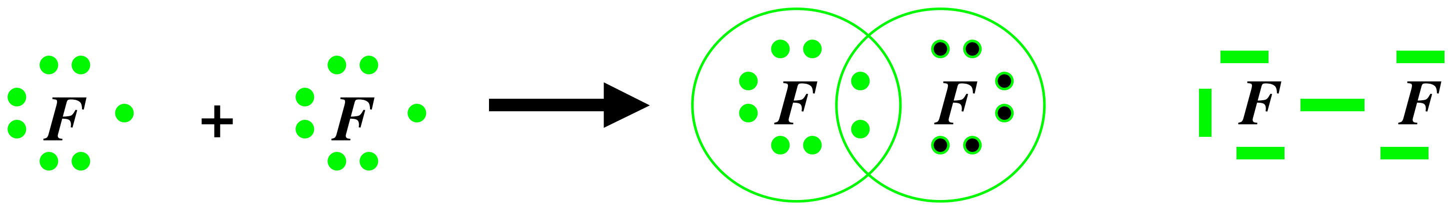
One bonding, one antibonding.

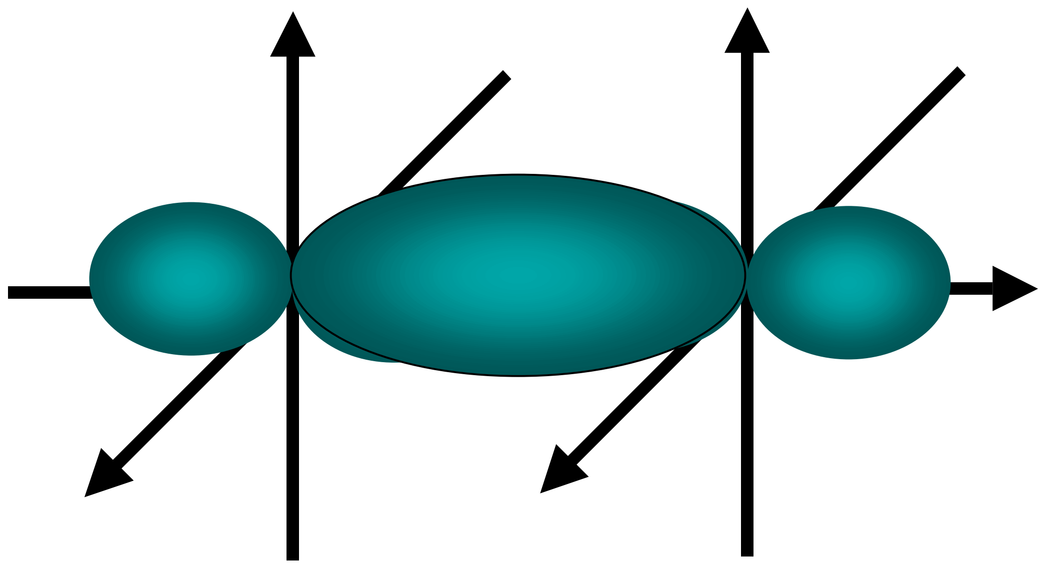
Covalent Bond. Sharing e$^-$



Only bonding MO shown

Collinear orbitals form σ bond

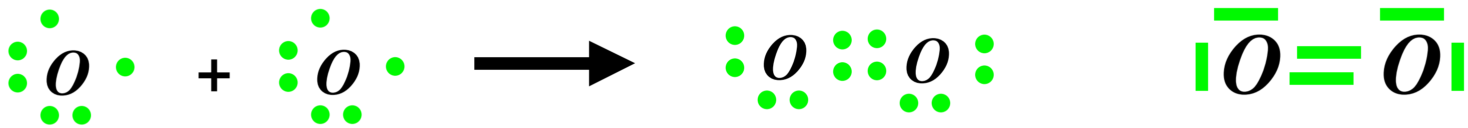


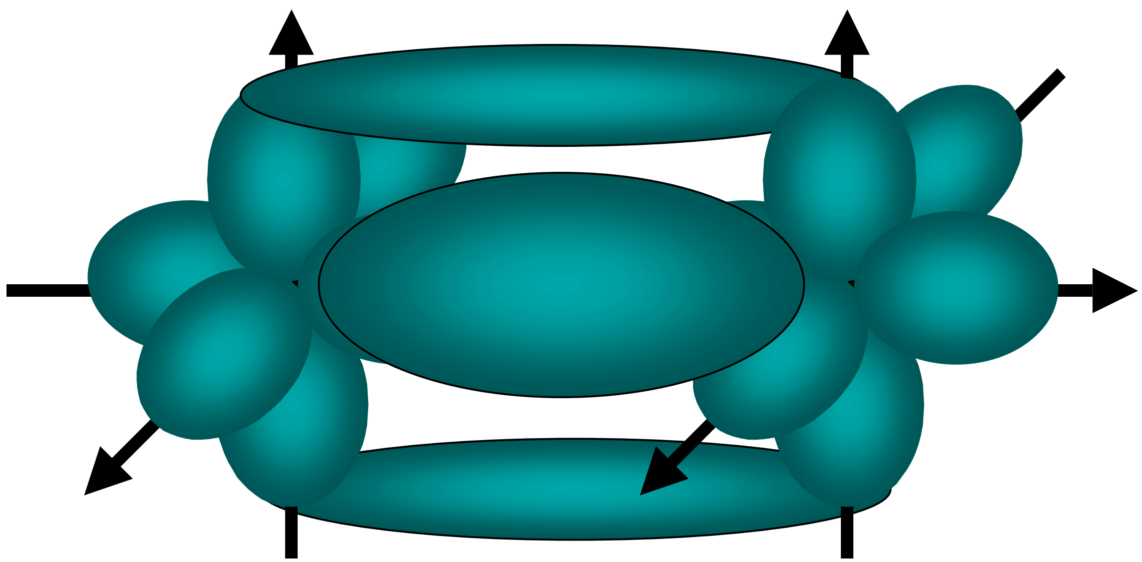


Two p AO = Two σ MO

Only bonding MO shown

Coplanar orbitals form □ bond

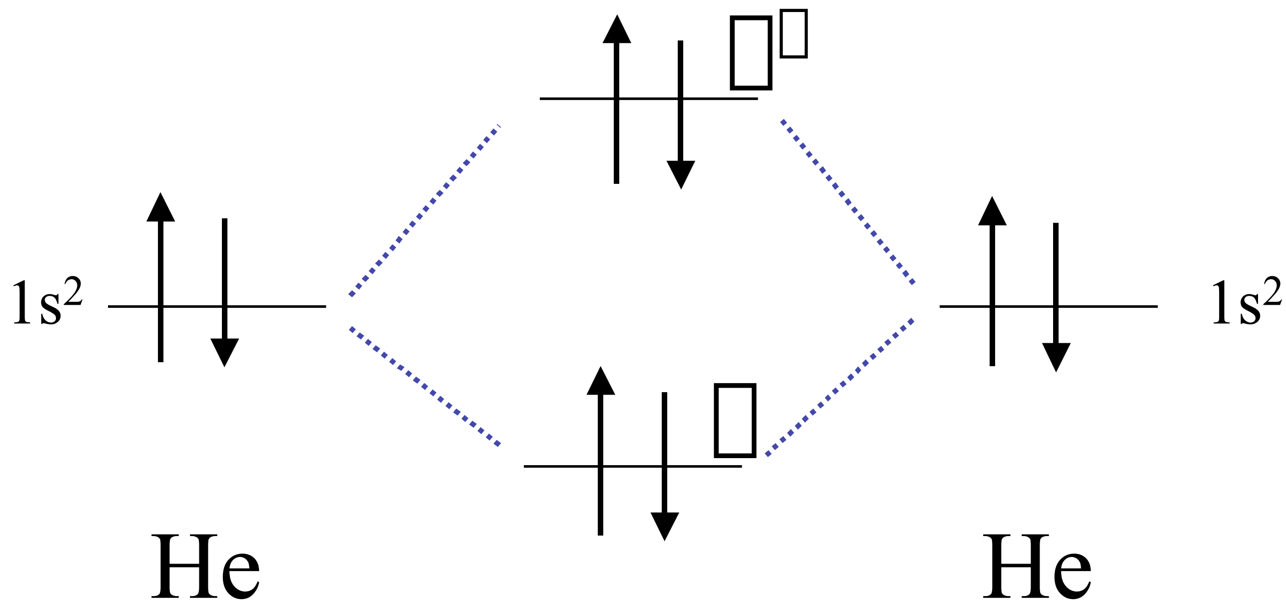




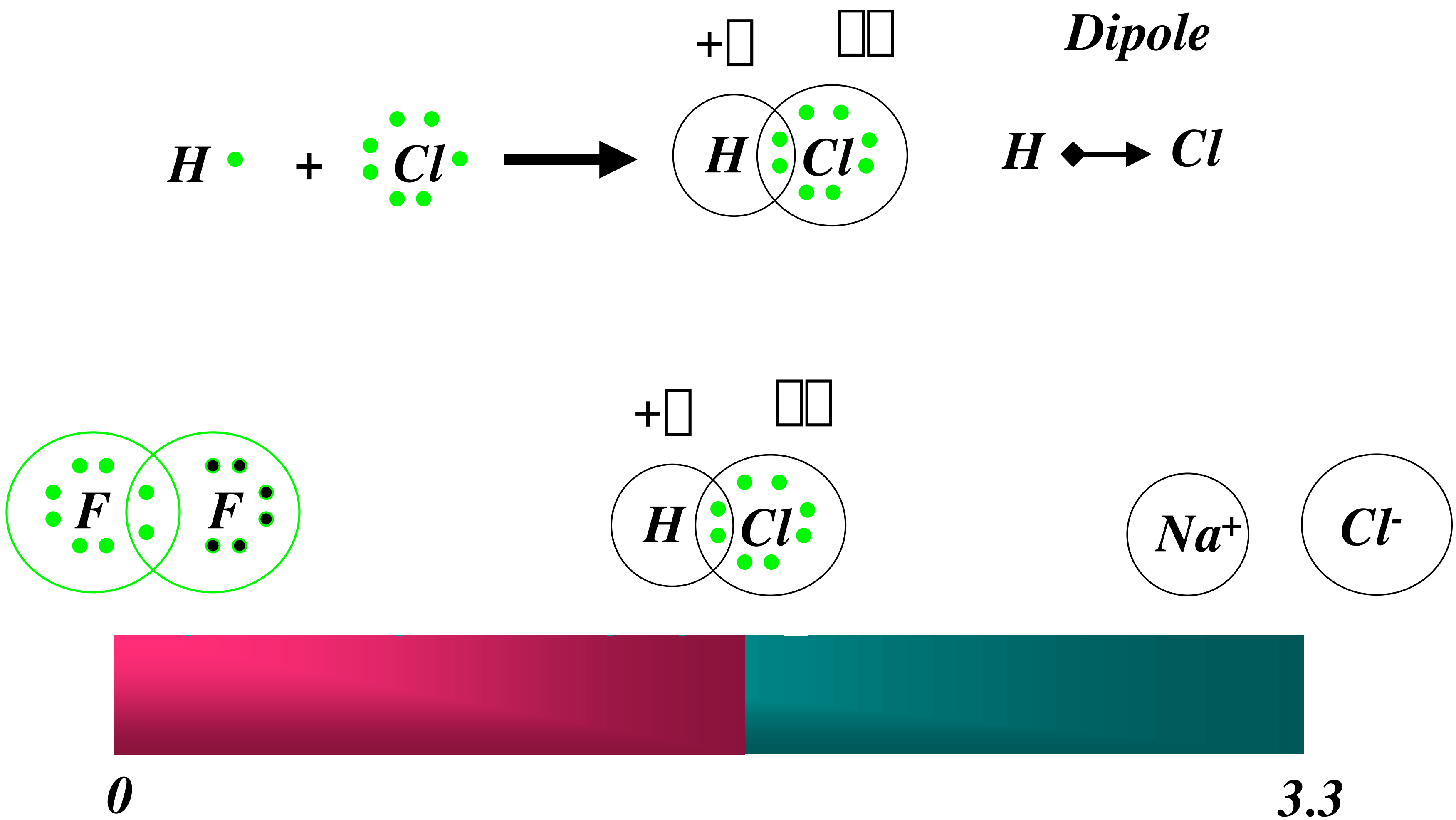
Four p AO = Two □ MO, and two □□□

Only bonding MO shown

He_2 Electron Configuration



Linus Pauling Electronegativity



Lewis Structures of Compounds

Count valence electrons available.

number of valence electrons contributed by nonmetal atom is equal to the last digit of its group number in the periodic table.
(H = 1)

Add electrons to take into account negative charge.

Ex.
OCl^- ion: 6 (O) + 7 (Cl) + 1 (charge) = 14 valence e^-
CH_3OH molecule: 4 (C) + 4(H) + 6 (O) = 14 valence e^-

Lewis Structures of Compounds

Draw skeleton structure using single bonds

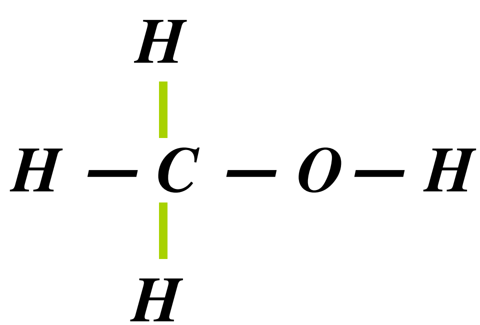
Note that carbon almost always forms four bonds.

Central atom is written first in formula.

Terminal atoms are most often H, O, or a halogen.

Ex.

$O - Cl^-$



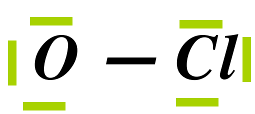
Lewis Structures of Compounds

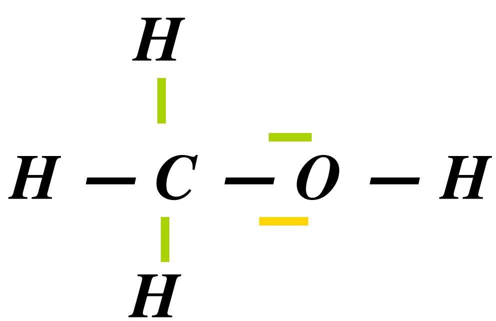
Subtract two electrons for each single bond

O-Cl⁻ ion: $14 - 2 = 12$ ***valence*** e^{-} ***left***

CH_3OH molecule: $14 - 10 = 4$ ***valence*** e^{-} ***left***

Distribute remaining electrons to give each atom a noble gas structure (if possible).





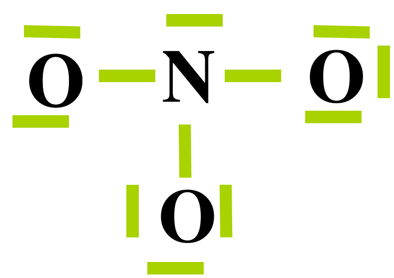
Lewis Structures of Compounds

Too Few Electrons? ***Form multiple bonds***

Ex. ***What is the structure of the NO_3^- ion?***

$$\textit{valence } e^- = 5(N) + 18\ (3O) + 1(\textit{charge}) = 24\ e^-$$

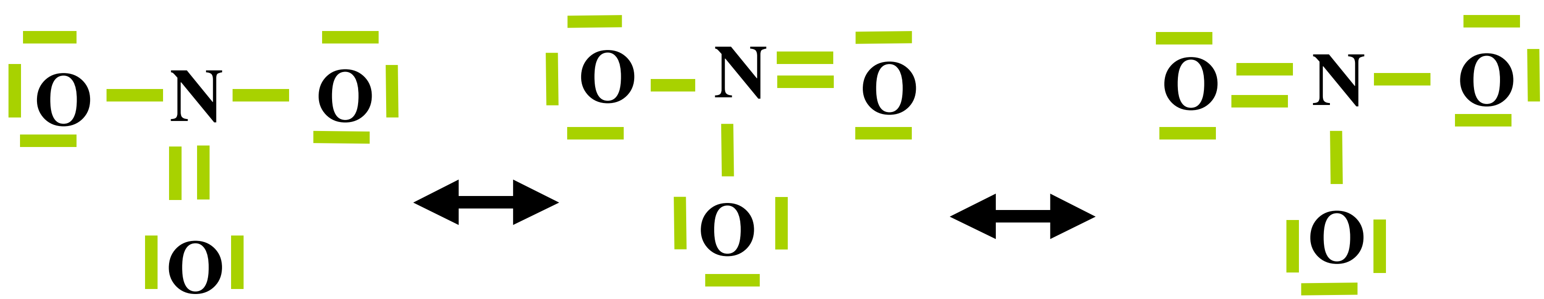
Skeleton:



Nitrate Ion (cont.)

valence e^- left = 24 - 6 (3 single bonds) = 18 e^-

Adding a double bond and rearranging:



I *II* *III*

Resonance Structures

Molecular Geometry

Molecular Geometry. VSEPR

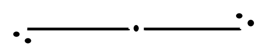
1. ***Electron pairs (lone and bonding pairs) around a central atom tend to be oriented so as to be as far apart as possible to minimize their repulsions***
2. ***The molecular geometry is hence determined by the relative locations of the electron pairs***
3. ***The SN (Steric Number) of the central atom is used to find the geometry that applies***

$$SN = \left(\begin{array}{c}\#\ \text{atoms bonded to} \\ \text{central atom}\end{array}\right) + \left(\begin{array}{c}\#\ \text{lone pairs on} \\ \text{central atom}\end{array}\right)$$

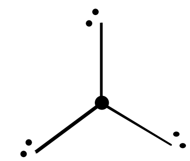
Molecular Geometry

In XY_n molecules in which there are no lone pairs, the SN is used to predict geometry

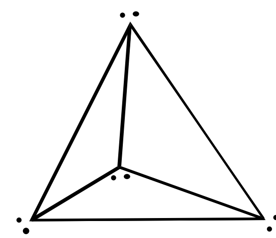
BeF_2 linear (SN = 2)



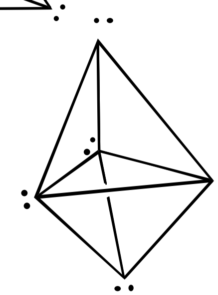
BF_3 trigonal planar (SN = 3)



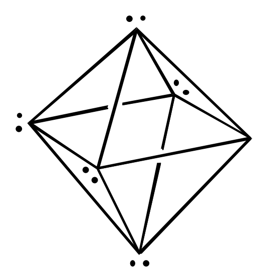
CF_4 tetrahedral (SN = 4)



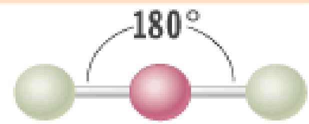
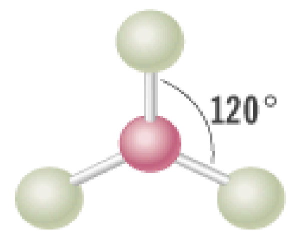
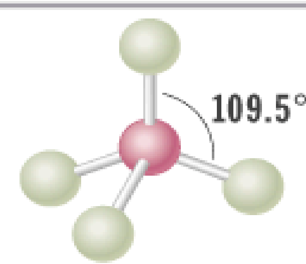
PF_5 triangular bipyramid (SN = 5)



SF_6 octahedral (SN = 6)



Molecular Geometry

Species type	Orientation of electron pairs	Predicted bond angles	Example	Ball and stick model
AX_2	Linear	180°	BeF_2	
AX_3	Triangular planar	120°	BF_3	
AX_4	Tetrahedron	109.5°	CH_4	

VSEPR model

Molecule	Lewis Str.	Pairs of e-	electron arrangem.	Molecular Shape
H_2S	$H - \underline{\overline{S}} - H$	4	tetrahedral	bent
CCl_4	Cl Cl — C — Cl Cl	4	tetrahedral	tetrahedral