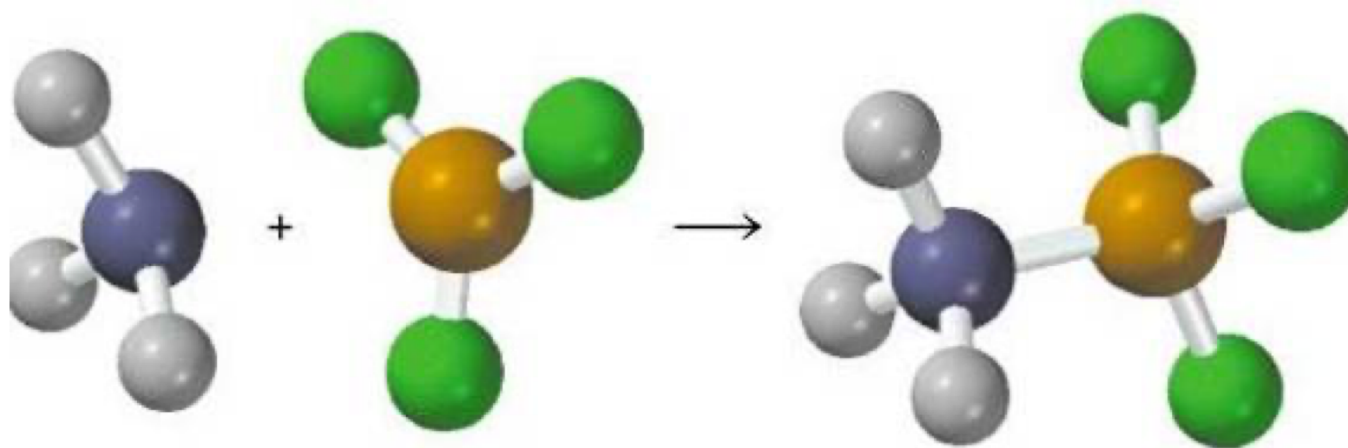


Chapter Six

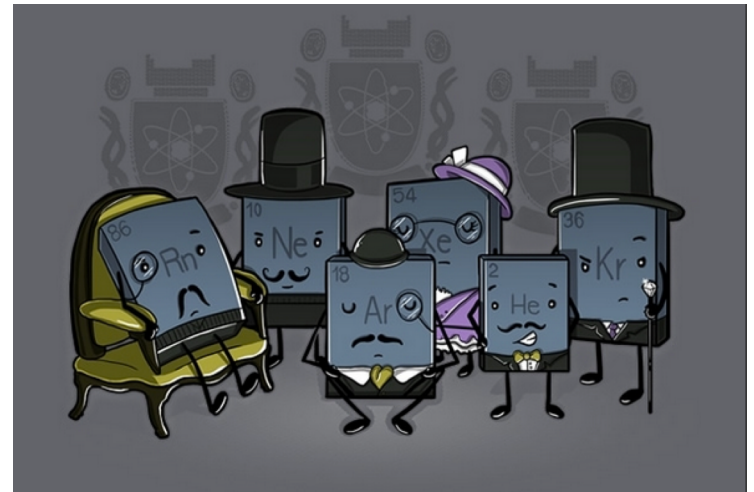
Representing Molecules



Effect of Valence Electrons on Elements

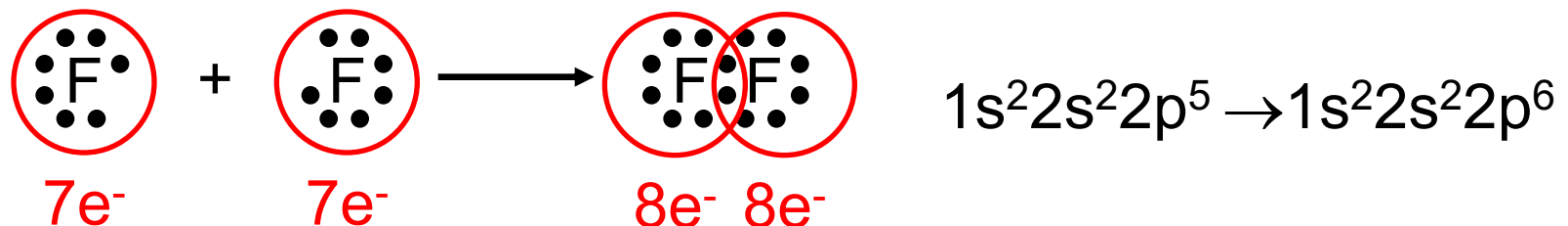
Octet Rule:

- Elements most stable with 8 valence electrons ($2s + 6p$)
- Noble gases have 8 valence electrons
 - No e^- want to be added or removed
 - Why they are so unreactive
- Main group elements form ions to become isoelectronic with the noble gases
 - Same electron configuration
- He & H follow duet rule
 - 2 e^- ; too small for $8e^-$



Lewis Structures

Lewis structures represent covalent bond formation

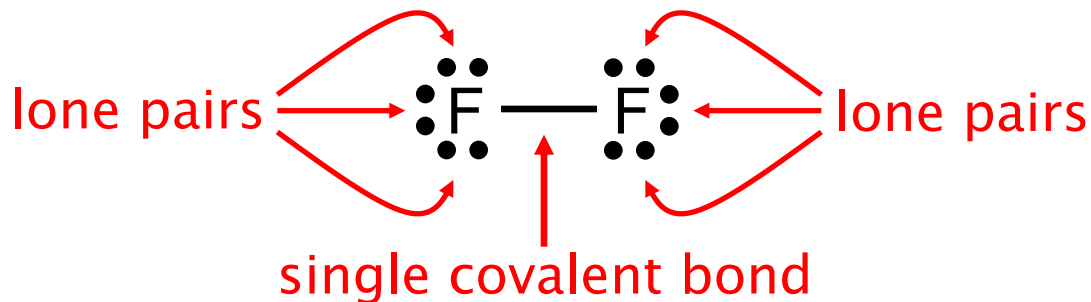


Bonding Pairs: Electrons shared by both atoms

- Represented by a dash (-) between bonded atoms

Lone Pairs: Non-shared electrons count for 1 atom

- Represented by a pair of dots ($\bullet\bullet$) around atom

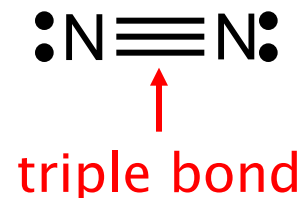
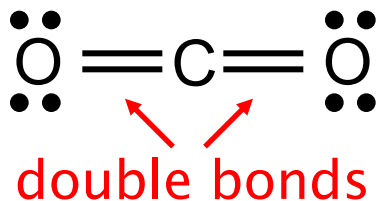


Multiple Bonds

More than one pair of electrons is shared between atoms so each atom can form an octet.

Single Bond:	1 shared pair:	1 dash (-)
Double bonds:	2 shared pairs:	2 dashes (=)
Triple bonds:	3 shared pairs:	3 dashes (\equiv)

Allows atoms in a molecule to share extra e⁻ if there are not enough for the central atom



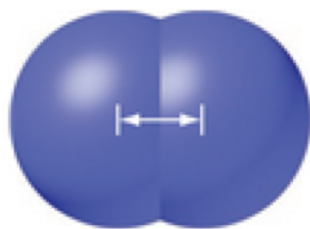
Multiple Bonds

The number of electrons shared impacts the length and strength of a covalent bond

In general:

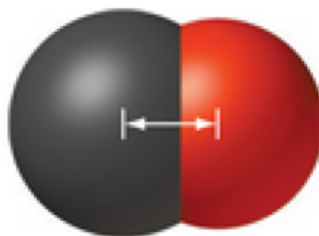
Single Bond:	longest	weakest (lowest bond energy)
Double bonds:	shorter	medium strength
Triple bonds:	shortest	strongest (highest bond energy)

Bond length is measured as the distance between the nuclei of two bonded atoms



N_2

Bond length 1.10 Å



CO

Bond length 1.13 Å

Writing Lewis Structures: Hints & Tips

Electronegativity

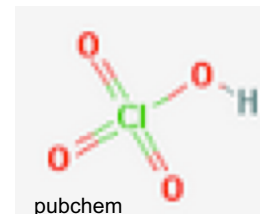
- Central atom usually has the **lowest** electronegativity (atom lower or to the left in periodic table)
- Terminal atoms (except H) have **higher** electronegativities

Terminal Atoms

- Bonded to only one other atom
- Hydrogen atoms are terminal atoms
- Halogens (col 17) are often terminal atoms

Oxoacids

- Hydrogen atoms are bonded to oxygen atoms in oxoacids



Hints & Tips for Drawing Lewis Structures Con't

- Final structure must include same number of valance e^- as sum of valance e^- from all atoms in the molecule
- Final structure must satisfy octet rule (unless it is an exception)
- Start with single bonds, try double then triple if necessary
- **Hydrogen** only wants one more e^-
 - forms **ONE SINGLE BOND**
 - will not be between two atoms
- **Carbon** usually does not have lone pairs
 - all 8 e^- must come from single, double, or triple bonds
- **Halogens** only want one more e^-
 - generally only form one single bond
- Molecules are often symmetrical
 - try single atom in middle with other atoms around it

Example: Write the Lewis Structure of HNO

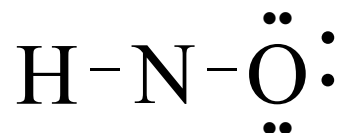
1. Add up the valence electrons in the structure

$$1(\text{H}) + 5(\text{N}) + 6(\text{O}) = 12 \text{ valence electrons}$$

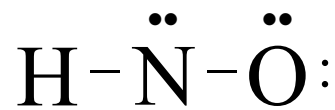
2. Arrange the atoms & place bonding electrons

H-N-O nitrogen less electronegative, put in center

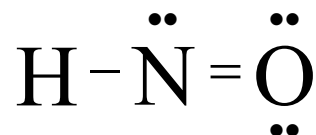
3. Place e- pairs around terminal atoms to get 8



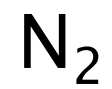
4. Place remaining electron pairs on central atom



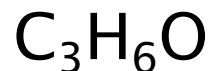
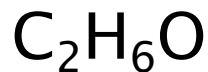
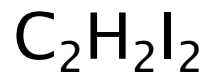
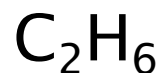
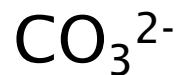
5. Add double bond to finish nitrogen octet (8)



Drawing Lewis Structures:



Drawing Lewis Structures: Polyatomic ions & simple organic molecules



Orgo – degrees of unsaturation:

$\text{C}_n\text{H}_{2n+2}$ – all single bonds

C_nH_{2n} – one double bond (or ring)

$\text{C}_n\text{H}_{2n-2}$ – 1 triple bond, 2 double bonds, 2 rings, double bond + ring

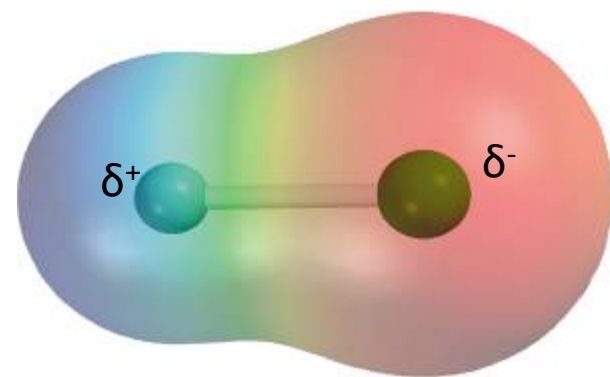
– 2 for each double bond/ring

– 4 for each triple bond

Electronegativity & Polar Covalent Bonds ¹¹

Electronegativity

- The ability of an atom to attract e^-
- F is the most electronegative atom
- Nonmetals - high electronegativities



Polar Covalent Bonds

- Differences in electronegativity result in unequal sharing of electrons between atoms
- More electronegative atom has a partial neg. charge (δ^-)
- More electropositive atoms has a partial pos. charge (δ^+)

Percent Ionic Character

- Measure of polarity of bond
 - 100% ionic is full transfer of electron, no sharing
 - 100% covalent is equal sharing, H_2 , Cl_2 , etc.

Electronegativities of Common Elements

Increasing electronegativity

Increasing electronegativity																					
1A																	8A				
H 2.1	2A															3A	4A	5A	6A	7A	
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	3B	4B	5B	6B	7B	8B			1B	2B	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0					
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 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 3.0				
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	Xe 2.6				
Cs 0.7	Ba 0.9	La-Lu 1.0-1.2	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 1.9	Bi 1.9	Po 2.0	At 2.2					
Fr 0.7	Ra 0.9																				

Electronegativity Difference determines bond type



F & Na: $4.0 - 0.9 = 3.1 = \text{ionic}$

C & H: $2.5 - 2.1 = 0.4 = \text{nonpolar covalent}$

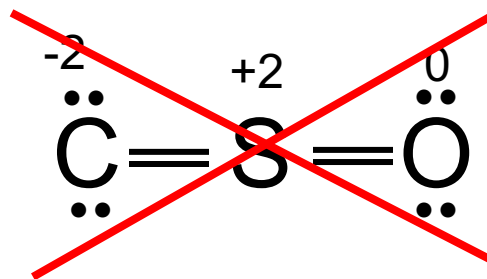
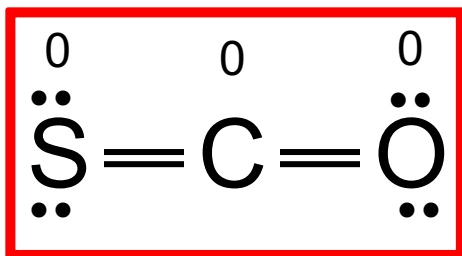
Formal Charge

Difference between the # of valence e⁻ in a free atom & the # of e⁻ assigned to that atom in a Lewis structure.

$$\text{F.C.} = \# \text{ ve}^- - (\# \text{ of lone e}^- + \# \text{ bonds})$$

Get #ve⁻ from group #

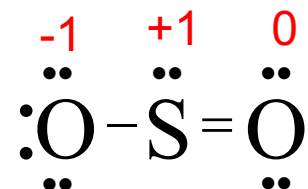
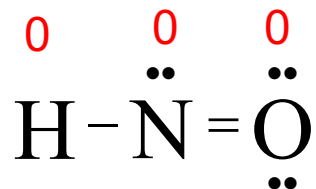
Molecule is most stable if formal charge is 0 for each atom.



- Most likely Lewis structure has lowest formal charges
- Negative F.C. must be on more electroneg. atom
- Sum of formal charges:
Molecules = 0
Polyatomic ions = charge

Calculating formal charge

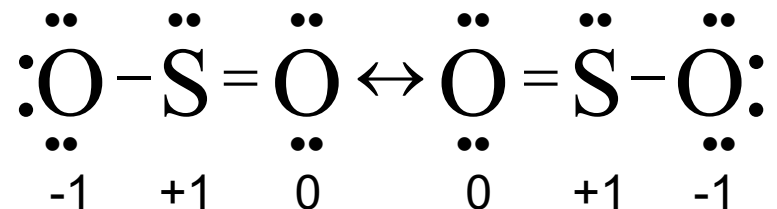
$$\text{F.C.} = \# \text{ ve}^- - (\# \text{ of lone e}^- + \# \text{ bonds})$$



Resonance

If a molecule or ion can be represented by 2 or more Lewis structures that differ only in electron location, the true structure is a blend of those structures.

- Electrons are moving around the molecule
- Neither bond is completely single or double (1.5)



Resonance Structures

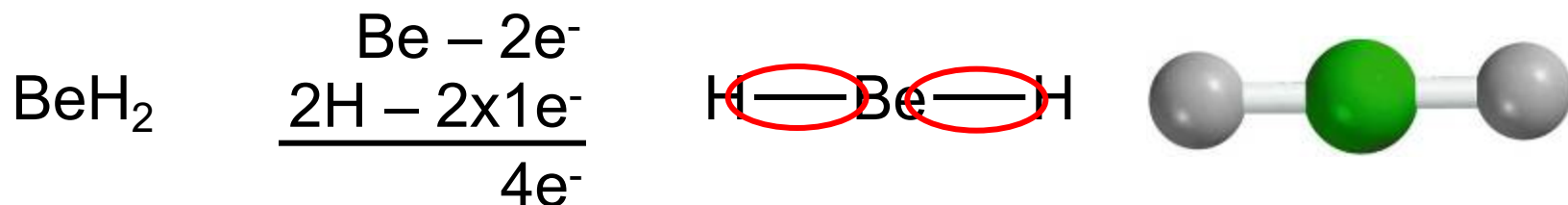
- Equivalent Lewis structures for a single molecule
- Formal charges will usually be present

Delocalization

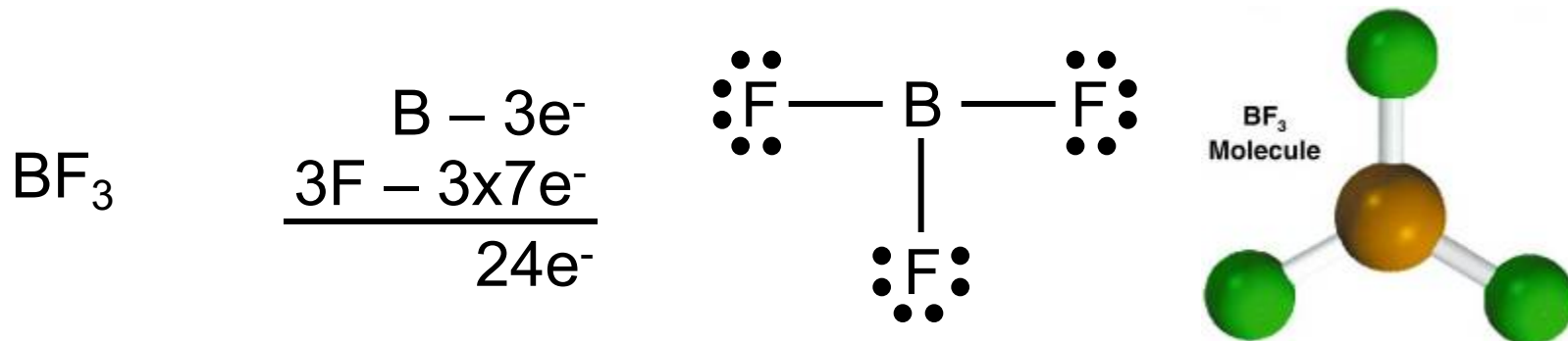
- Electrons are shared by more than two atoms
- Stabilizes the molecule

Exceptions to the Octet Rule: Incomplete Octet

Not enough electrons for central atom to have 8



Terminal atoms unwilling to donate more electrons
- Would destabilize terminal atoms & create formal charge



Free Radicals and Expanded Octets

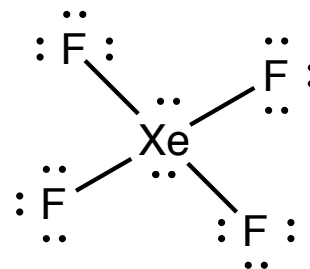
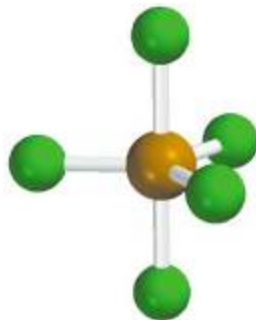
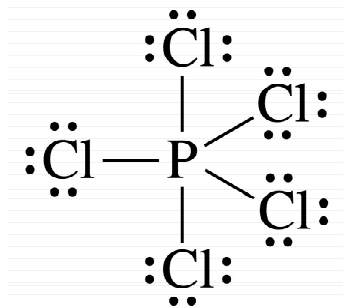
Free Radicals:

- Molecules with an odd number of valence electrons (N)
- **Extremely reactive**, odd electron wants to be part of a pair



Compounds with expanded valence shells

- Central atom has more than eight electrons
- May have lone pair electrons as well as bonding pairs
- Must be in row 3 or above to have an expanded octet



Drawing Lewis Structures: Exceptions to the Octet Rule



Coordinate Covalent Bonds

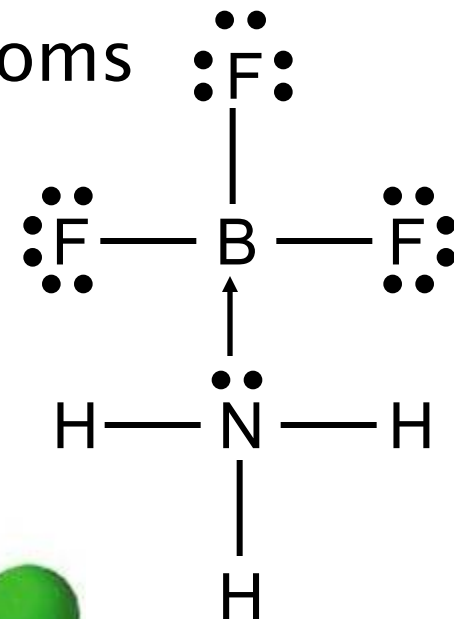
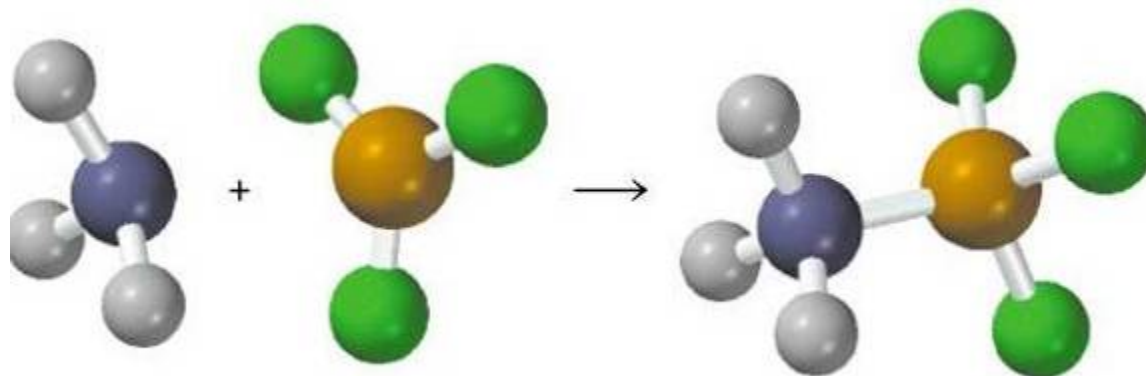
1 atom provides both electrons

Electrons are then shared between 2 atoms

Ex: BF_3 and NH_3

B needs 2 electrons to fill octet

N has a lone pair to share



Seen often with transition metals – can accept electrons into empty d-orbitals