4.1 Types of Chemical Bonds

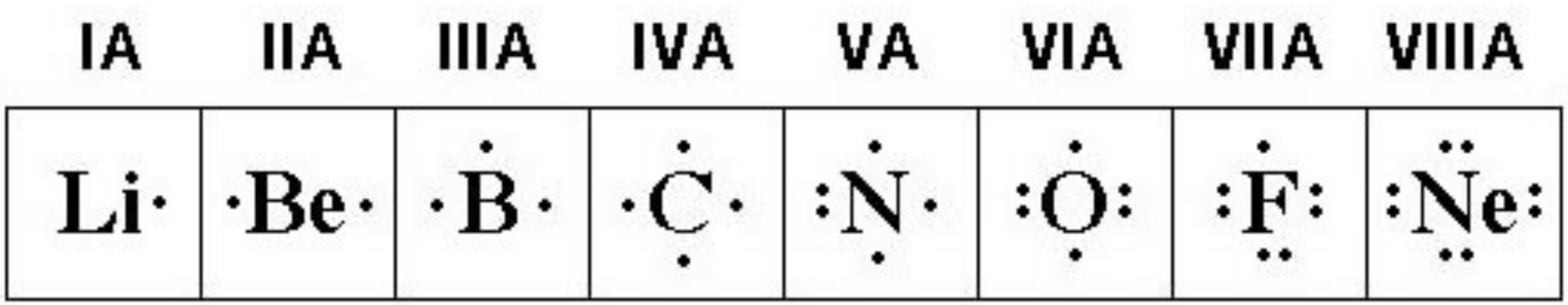
Lewis Theory of Bonding

- Atoms and ions are stable if they have a *full valence shell* of electrons
- Electrons are most stable when they are *paired*.
- Atoms *form chemical bonds* to achieve a full valence shell of electrons.
- A full valence shell of electrons may be achieved by an **exchange** of electrons *between metal and nonmetal* atoms.
- A full valence shell of electrons may be achieved by the **sharing** of electrons between *nonmetal atoms*.
- The sharing of electrons results in a *covalent bond*.

The Octet Rule

- Many atoms tend to form bonds until it is surrounded by eight valence electrons
  – Example: Hydrogen follows the "Duet" Rule
- Used to make predictions about molecular shape (discussed later)

Figure 1: Valence Electrons by Group



Lewis Structures for Ionic Bonds

- Exchange of electrons between metal (loses) and nonmetal (gains)

Examples:

1) LiF

2) CaF_2

Lewis Structures for Covalent Bonds

- Sharing of electrons between 2 nonmetals

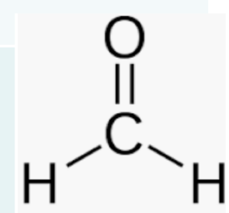
Examples:

1) H_2

2) F_2

← Lone electron pairs:
A pair of valence electrons that are not involved in bonding

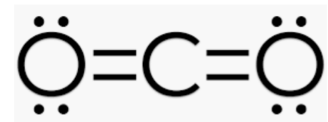
Steps for Drawing Lewis Structures

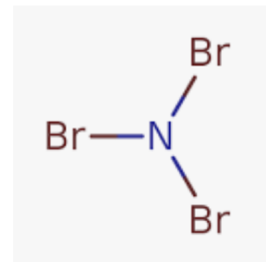
Steps for Drawing Lewis Structures	Example: CH_2O (Formaldehyde)
1. Identify the central atom, which is the element with the highest bonding capacity (or least electronegative) and arrange the remaining elements around it.	
2. Add up the valence electrons available in each atom. Adjust for ions by adding if a negative charge and subtracting if a positive charge.	
3. Connect each surrounding atom to the central atom with a line that represents a covalent bond (2 electrons)	
4. Place pairs of the remaining valence electrons as lone pairs on the surrounding atoms. Follow the octet rule (or duet rule for hydrogen).	
5. Determine how many electrons are still available by subtracting from total valence electrons.	
6. Place the remaining electrons on the central atom in pairs.	
7. If the central atom does not have a full octet, move lone pairs from surrounding atoms to create multiple bonds.	

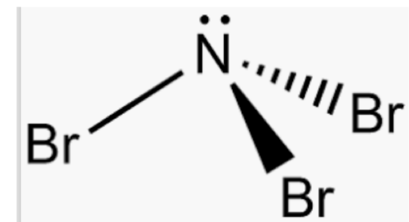
Practice

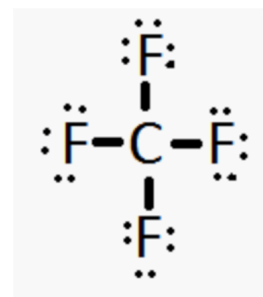
Draw the Lewis Structure for each of the following molecules.

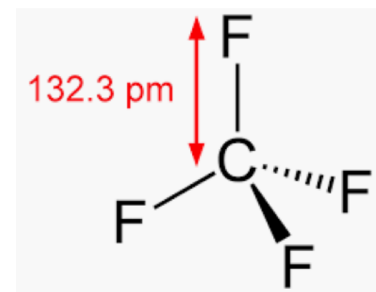
1. CO_2
2. NBr_3
3. CF_4
4. C_2H_4

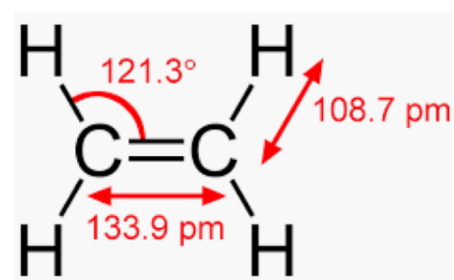








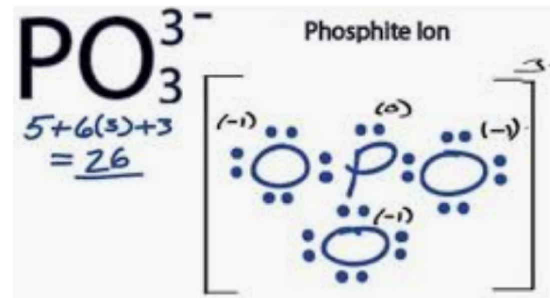


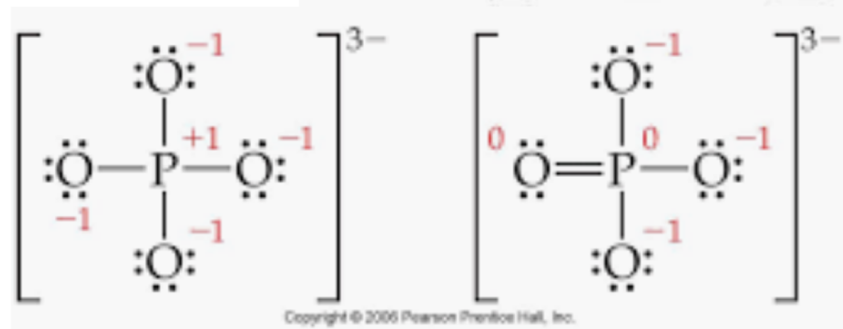


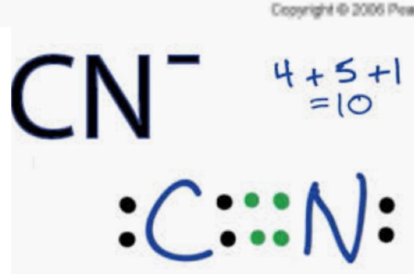
Practice: Polyatomic Ions

Draw the Lewis Structure for each of the following ions.

1. PO_3^{3-}
2. CN^-
3. NO_2^-
4. ClO^-







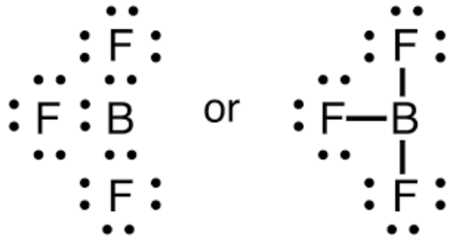
Exceptions to The Octet Rule

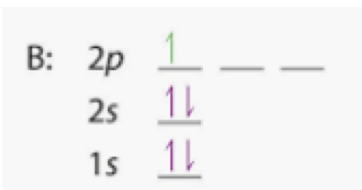
- There are some exceptions to this rule:
  1. Molecules with central atoms that are surrounded by fewer than 8 electrons (underfilled octets)
  2. Molecules with central atoms that are surrounded by more than 8 electrons (overfilled octets)

Underfilled Octets

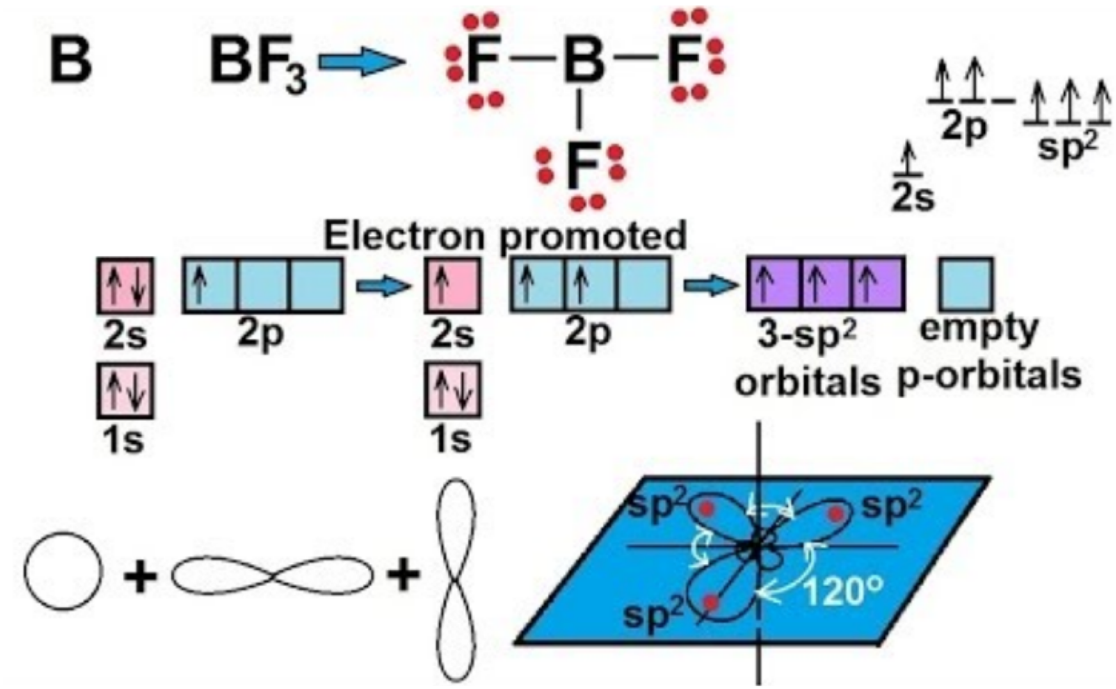
- Boron forms compounds in which the boron atom has fewer than 8 electrons around it
- it is considered electron deficient and will react with electron-rich molecules like H_2O and NH_3

Figure 2: Lewis structure for Boron trifluoride




B:
2p
2s
1s


1s2
2s2
2p5


B
BF3
F
B
F
F
2p
sp2
2s
Electron promoted
2s
2p
1s
2s
2p
1s
3-sp2
orbitals
empty
p-orbitals
sp2
sp2
sp2
120°

Exceptions to The Octet Rule

<u>Overfilled Octets</u>

- Atoms that exceed the octet rule use nearby vacant d orbitals to fill with extra electrons
- Therefore, this does not happen for elements in periods 1 or 2
- The sulfur atom can have up to 12 electrons surrounding it
- Sulfur hexafluoride (SF_6) is a very stable molecule. The sulfur atom uses the 3s, 3p and 3d orbitals to accommodate 12 electrons

Figure 3: Lewis structure for sulfur hexafluoride

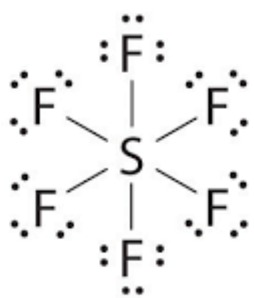
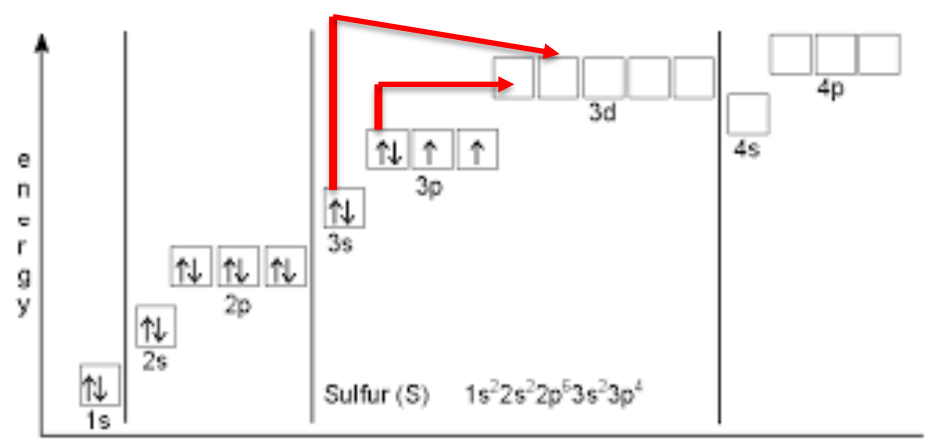
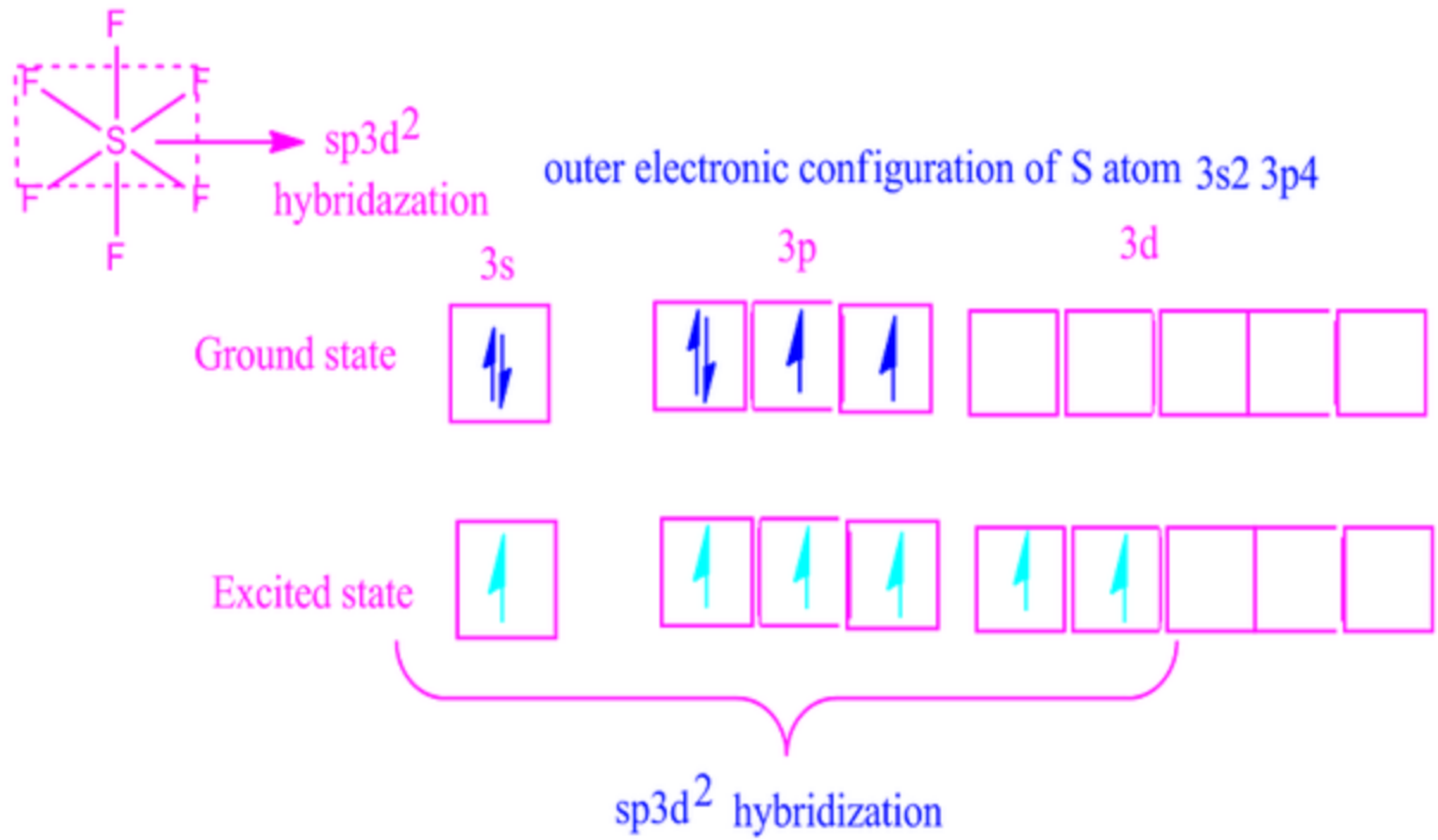


Figure 4: Energy Diagram for Sulfur atom



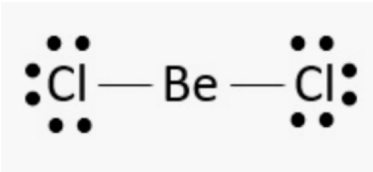


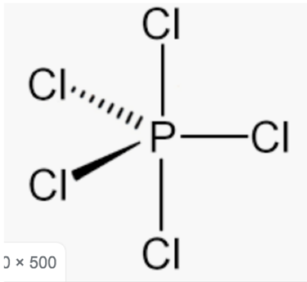
F
F F
S
F F
F
sp3d2
hybridazation
outer electronic configuration of S atom 3s2 3p4
3s
3p
3d
Ground state
Excited state
sp3d2 hybridization

Practice: Exceptions to the Octet Rule

Draw the Lewis Structure for the following molecules.

1. $BeCl_2$
2. PCl_5





Homework

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