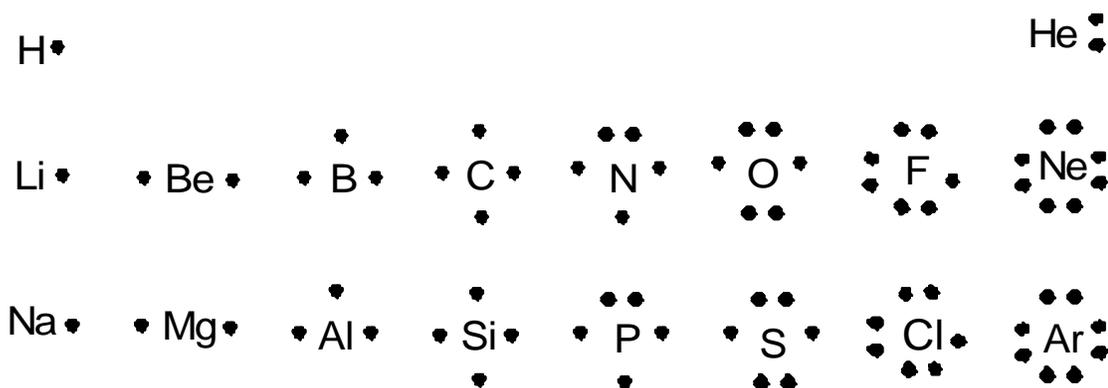


Chapter 9: Chemical Bonding I: General Concepts

9.1: Lewis Dot Symbols

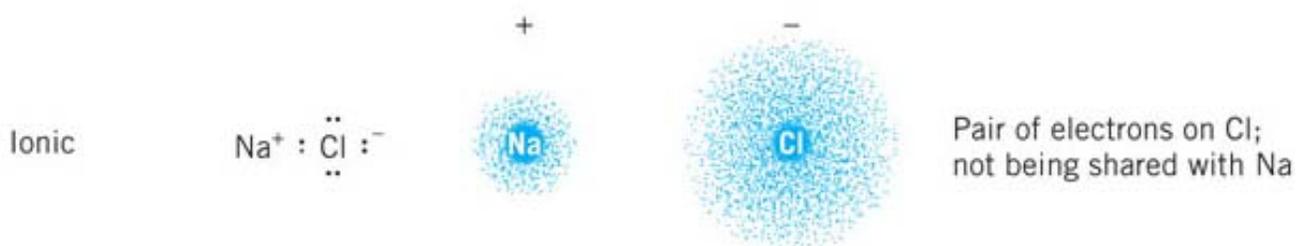
- Lewis Structure:** - sometimes refer to as Lewis Dot Diagram.
 - shows electrons of valence electron (*s* and *p* orbitals) as dots for all atoms of representative elements (Groups IA to VIIIA).
 - max of eight electrons around each atomic symbol, and hence Lewis structure cannot be used on transition metals.

Lewis Structure for Atoms in the First Three Rows of the Periodic Table of Elements



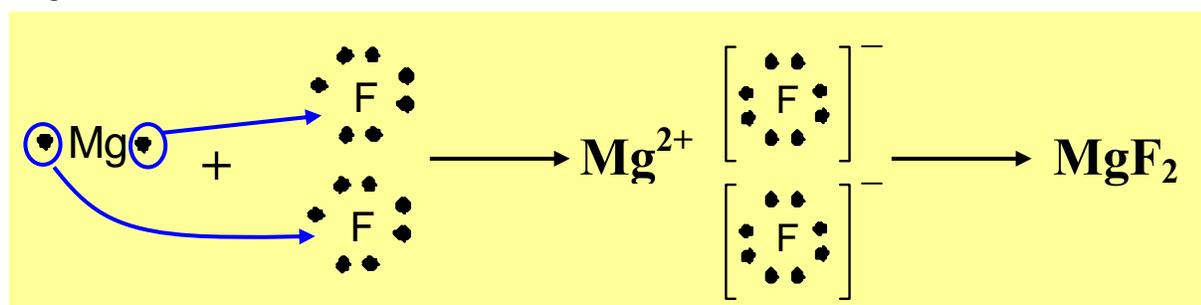
9.2: The Ionic Bond

- Ionic Bonds:** - a chemical bond (attraction) that is characterized by the **transfer of electrons** to form ions between metals and non-metals.
 - the resulting attraction between the cation(s) and anion(s) is the ionic bond within an ionic compound.

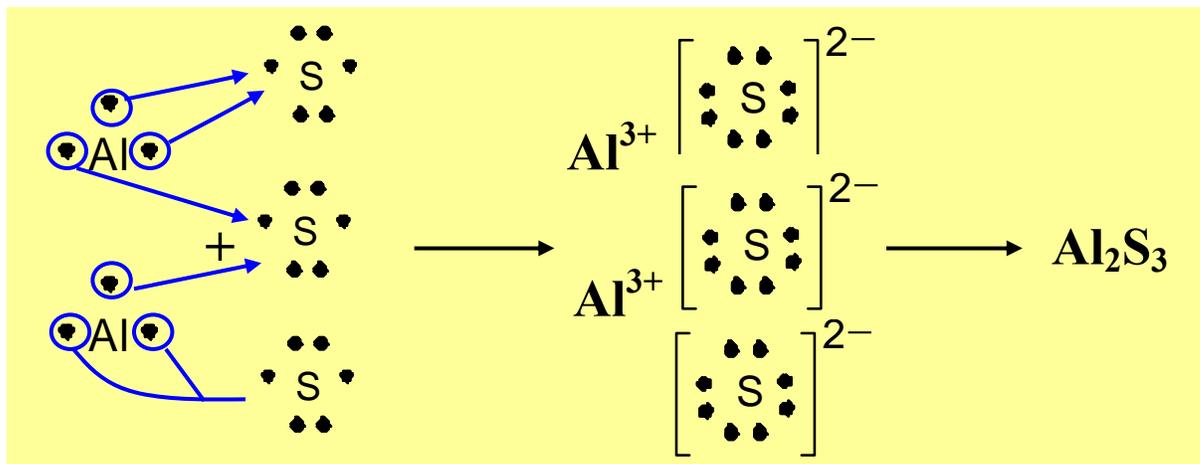


Example 1: For each ionic compound, draw the Lewis dot diagram for each atom and ion involved.

a. Mg and F

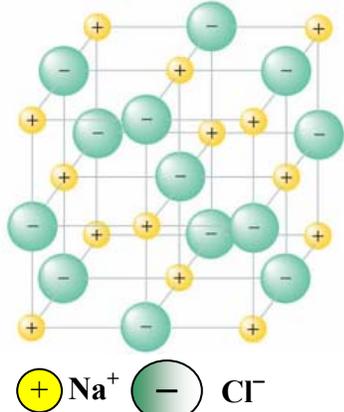


b. Al and S

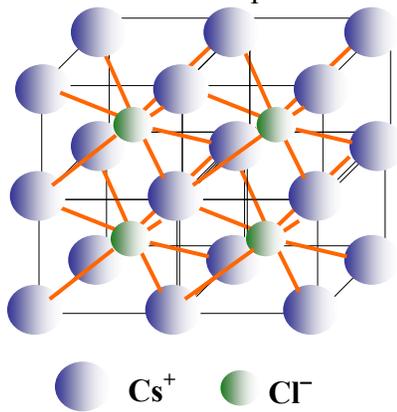


Properties of Ionic Crystalline Solids

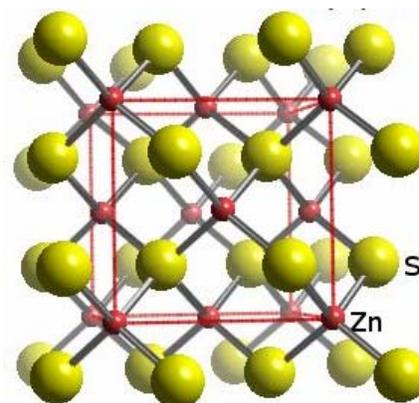
- 1. Ionic Compounds have a definite Crystalline Structure and are Poor Conductors of Electricity and Heat in their Solid Form.** Conduction of electricity and heat requires ions to move freely within the solid. The lattice structures of the solid ionic compounds do not allow ions to move freely.



Simple Cubic Unit Crystalline Structure of NaCl



Body Centred Cubic Unit Crystalline Structure of CsCl



Face Centred Cubic Unit Crystalline Structure of ZnS

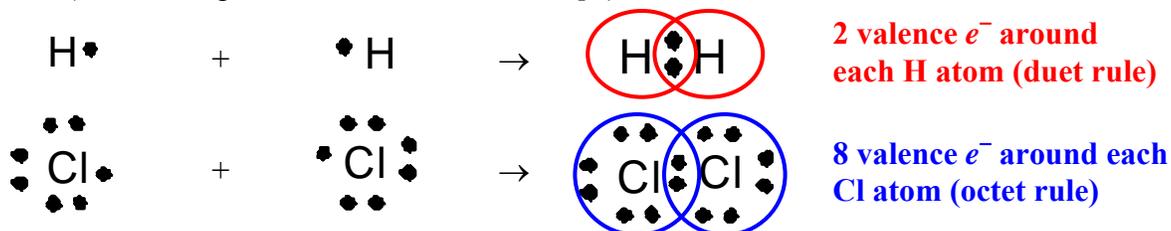
- 2. Ionic solids are generally High Melting Points (typically 300°C to 1000°C).** Since a strong force can only shatter the crystal but not bend it as in metals, the energy needed to completely break up the lattice structure (lattice energy) is very large and it is the same energy needed to melt the ionic compounds.
- 3. Ionic solids are Hard and Brittle.** The lattice structure of all ionic compounds holds the ions in definite positions. When the compound encountered a strong force, the close proximity of the ions stay close together. This causes the crystal to shatter, not bent like metal solid would.
- 4. Ionic solids can be Melted to form Liquids that are Electrical Conductors.** Ionic solids melt when the ions gain enough energy to break the lattice structure. They are move freely and can carry electrical charge through the liquid. This explains why a molten ionic substance conducts electricity, but a solid ionic material doesn't. The ions move through the liquid can carry charge from one place to another.
- 5. Soluble ionic solids dissolve to form solutions that are Electrical Conductors.** (Not all ionic substances are soluble in water.) Soluble ionic compounds form electrolytes (ions in aqueous form) that allow the conduction of electricity.

Assignment

9.1 pg. 306 #1 and 5; 9.2 pg. 306–307 #7, 9, 10, 13, 15, 16, 18, 20

9.4: The Covalent Bond

Covalent Bond: - an attraction force between two non-metal elements where the sharing of valence electrons is taken place so each atom can achieve the stable duet (for hydrogen) or octet (for other representative elements – ns^2np^6) conditions.

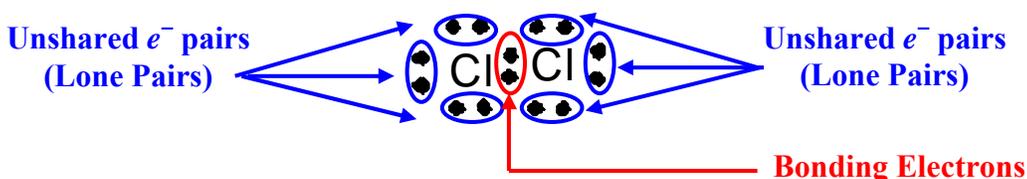


Covalent Compound: - a compound that consists of only covalent bonds.
- sometimes refers to as **molecular compound**.

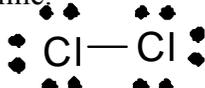
Types of Covalent Chemical Bonds:

1. **Single Covalent Bond:** - covalent bond where there are only **two valence electrons** being shared between two atoms (also refer to as **bonding electrons**).

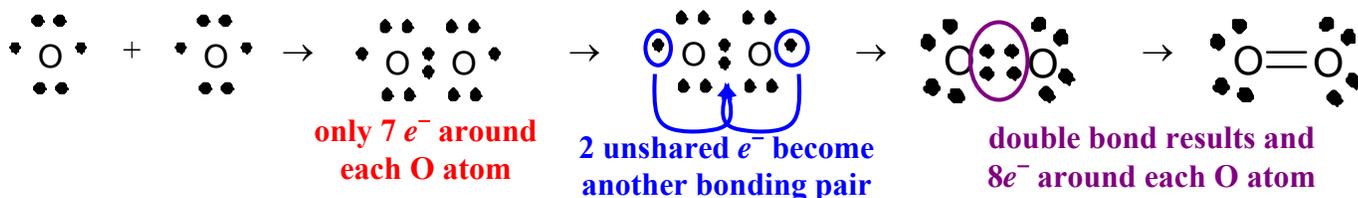
Unshared Electron Pairs (Lone Pairs): - pairs of electrons **not** involving in the **covalent bond**.



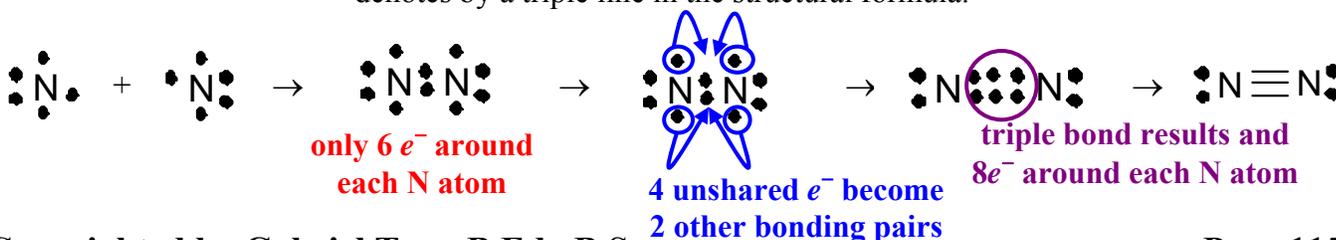
Structural Formula: - a Lewis dot diagram of a molecule where the bonding electrons are replaced by a short line.



2. **Double Covalent Bond:** - covalent bond where there are only **four valence electrons** being shared between two atoms.
- denotes by a double line in the structural formula.



3. **Triple Covalent Bond:** - covalent bond where there are only **six valence electrons** being shared between two atoms.
- denotes by a triple line in the structural formula.



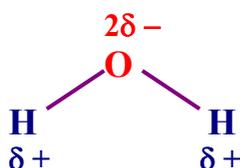
Coordinate Covalent Bond: - where an atom contributes an unshared electron pair (lone pair) to form double or triple bonds to achieve a stable octet around both atoms.
- sometimes refer to as **multiple bond**.

Note: In general, bond length decreases with multiple bonds.

Bond and Bond Type	Bond Length (pm)
C – C (single bond)	154
C = C (double bond)	134
C ≡ C (triple Bond)	120

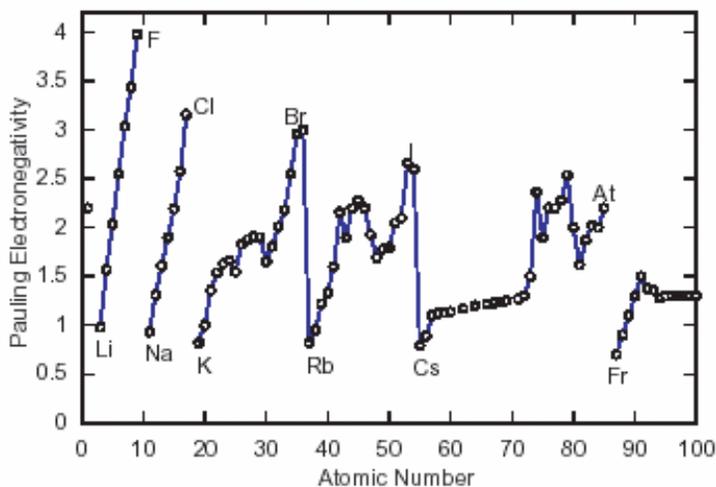
9.5: Electronegativity

Polar Covalent Bond: - a covalent bond where electrons are not evenly distributed, resulting in localized charges at either end.



Electronegativity: - first determined by **Linus Pauling**, it is a measure of the **capability of an atom within a molecule to attract shared electrons around itself**.

- the better the atom is able to attract electrons, the higher the electronegativity value.
- electronegativity of noble gases is 0 as their outer orbitals are filled and do not attract electrons.



Several Notes on Trends in Electronegativity

1. In general, **Electronegativities INCREASE as one move to the right of a period (up to and including halogens)**. This is because of the increase in electron affinity of the non-metals. These non-metals like to form anions to fill the valence orbitals. Metals tend to the high ionization energy because they like to give away electrons to form cations.
2. **Electronegativities DECREASE Down a Group**. This is due to the fact there are more orbitals as the number of row increases. The outer electrons are, of course, further away from the nucleus. Hence, it is more difficult for the protons of the nucleus to attract electrons into the valence orbitals.

Relative Bond Polarity: - the difference in electronegativity in a molecule.

- 1. Ionic Bonds have High Bond Polarity.**
- 2. Non-Polar Covalent Bonds have No Bond Polarity.**
- 3. Polar Covalent Bonds have Some Bond Polarity.**

Example 1: Using Figure 9.5 of the textbook, determine the relative bond polarity of Cl₂, HCl, and LiCl. Order them from least to greatest and classify them.

$\begin{array}{r} \text{Cl} - \text{Cl} \\ 3.2 \quad 3.2 \end{array}$	Difference in Electronegativity = 0 (Non-Polar Covalent Bonds – No Bond Polarity)
$\begin{array}{r} \text{H} - \text{Cl} \\ 2.2 \quad 3.2 \end{array}$	Difference in Electronegativity = 1 (Polar Covalent Bonds – Some Bond Polarity)
$\begin{array}{r} \text{Li} - \text{Cl} \\ 1.0 \quad 3.2 \end{array}$	Difference in Electronegativity = 2.2 (Ionic Bonds – High Bond Polarity)

Assignment

9.4 pg. 307 #29 and 30

9.5 pg. 307 #32, 34, 35, 37, 38

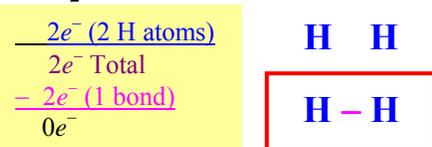
9.6: Writing Lewis Structures

Procedure to draw Lewis Dot Diagram

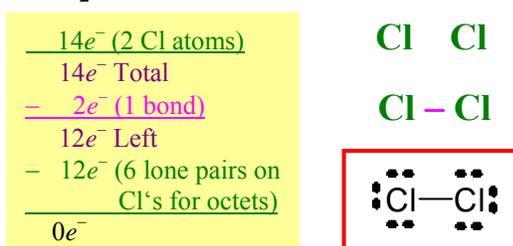
- 1. Add up all the valence electrons available from all atoms and extra charges from complex ions (ex: charge of 1- means one more electron; charge of 1+ means one less electron).**
- 2. Decide on the central atom. Usually, it is the atom with the most available bonding sites (carbon and silicon has the most bonding site – 4 valence e⁻) or it is element with the least number of atom. Place a pair of electrons to form a single bond between each atom.**
- 3. Arrange any remaining electrons around all atoms such that the octet rule is followed for the second and third period atoms and duet rule is satisfied for the first period atoms. This may sometimes mean setting up multiple bonds (moving lone pairs to the central atom so that it satisfies the octet rule).**
- 4. Final Lewis structure of Polyatomic Ion must include a big square bracket and the indication of net charge.**

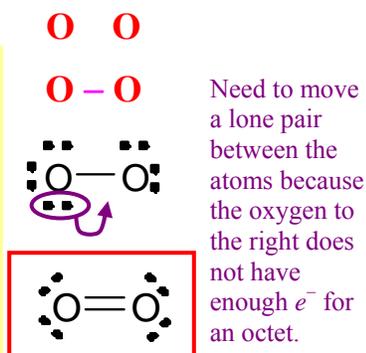
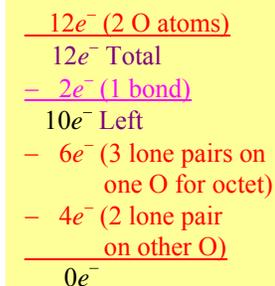
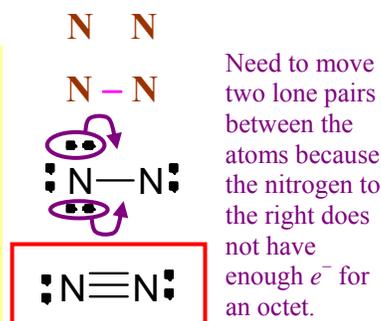
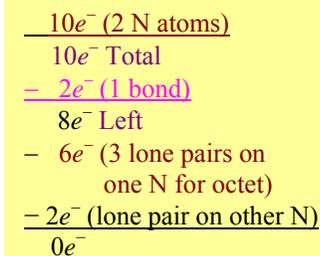
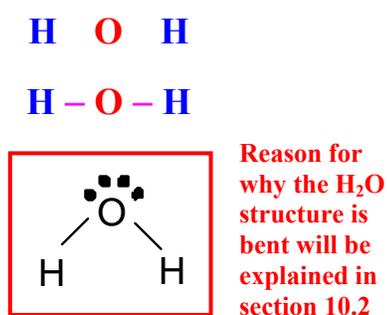
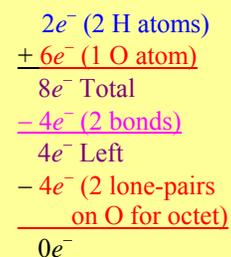
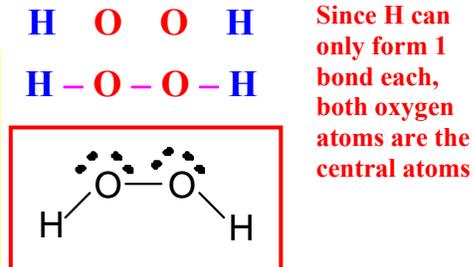
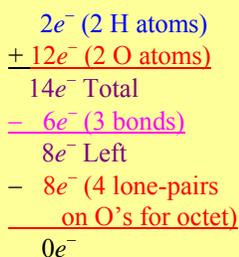
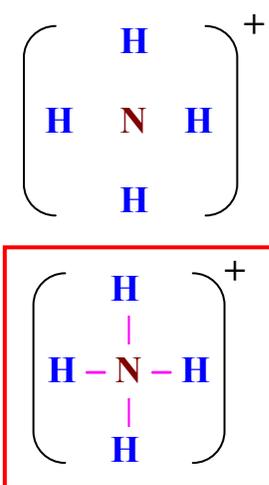
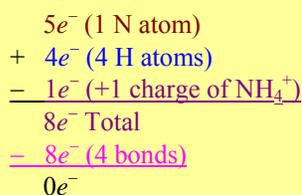
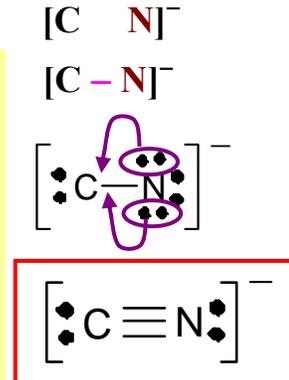
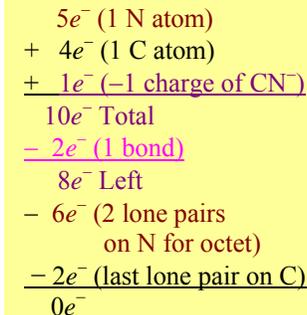
Example 1: Draw the Lewis structure for the following molecules and polyatomic ions.

a. H₂



b. Cl₂



c. O₂d. N₂e. H₂Of. H₂O₂g. NH₄⁺h. CN⁻

Assignment
9.6 pg. 307 #41

9.9: Exceptions to the Octet Rule

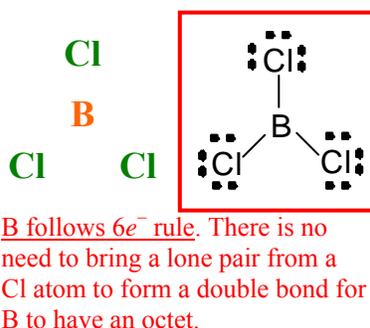
- In the second row, due to the $2s$ and the $2p$ orbitals, C, N, O, and F should always follow the octet rule. Overall, the second period elements will never exceed the octet rule.
- However, because Be is a metal and B is a metalloid, they are most likely to have fewer than eight electrons around them. Thus, these compounds are fairly reactive due to their electron-deficiency. This phenomenon is referred to as the Incomplete Octet.
- In the third period, heavier atoms towards the right of the Table often satisfy the octet rule. This being said, there are times that they can use the empty $3d$ orbitals to exceed the 8 electrons limit. (Examples: P can have $10e^-$ rule; S and Cl can have $10e^-$ or $12e^-$ rule) – Expanded Octet
- Elements in higher row can definitely exceed the octet rule. This sometimes applies to larger noble gases as well. (Examples: Br, I, Xe, Rn can have $10e^-$ or $12e^-$) – Expanded Octet

Note: For atoms that deviate from the octet and duet rules, Formal Charges will have to be assigned to each atom in a molecule to assess all possible Lewis structures. This will be discussed in following section.

Example 1: Draw the Lewis structure for the following molecules and polyatomic ions.

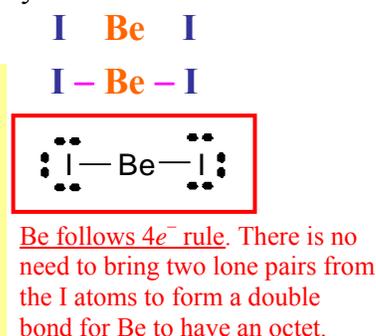
a. BCl_3

$$\begin{array}{l}
 3e^- \text{ (B atom)} \\
 + 21e^- \text{ (3 Cl atoms)} \\
 \hline
 24e^- \text{ Total} \\
 - 6e^- \text{ (3 bonds)} \\
 \hline
 18e^- \text{ Left} \\
 - 18e^- \text{ (lone-pairs on} \\
 \text{Cl's for octet)} \\
 \hline
 0e^-
 \end{array}$$



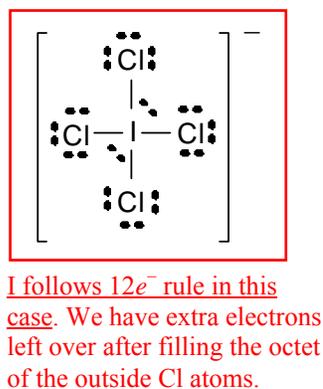
b. BeI_2

$$\begin{array}{l}
 2e^- \text{ (Be atom)} \\
 + 14e^- \text{ (2 I atoms)} \\
 \hline
 16e^- \text{ Total} \\
 - 4e^- \text{ (2 bonds)} \\
 \hline
 12e^- \text{ Left} \\
 - 12e^- \text{ (lone-pairs on} \\
 \text{I's for octet)} \\
 \hline
 0e^-
 \end{array}$$



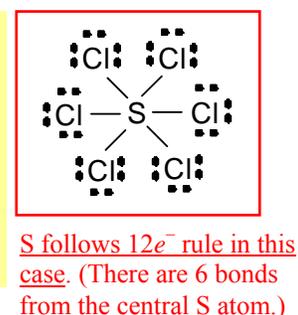
c. ICl_4^-

$$\begin{array}{l}
 7e^- \text{ (I atom)} \\
 + 28e^- \text{ (4 Cl atoms)} \\
 + 1e^- \text{ (-1 charge of } \text{ICl}_4^-) \\
 \hline
 36e^- \text{ Total} \\
 - 8e^- \text{ (4 bonds)} \\
 \hline
 28e^- \text{ Left} \\
 - 24e^- \text{ (lone pairs on} \\
 \text{Cl's for octet)} \\
 - 4e^- \text{ (lone pairs on I)} \\
 \hline
 0e^-
 \end{array}$$



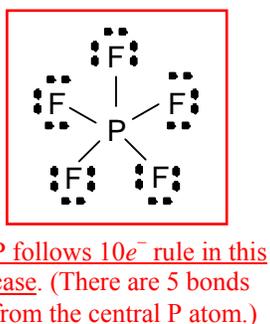
d. SCl_6

$$\begin{array}{l}
 6e^- \text{ (S atom)} \\
 + 42e^- \text{ (6 Cl atoms)} \\
 \hline
 48e^- \text{ Total} \\
 - 12e^- \text{ (6 bonds)} \\
 \hline
 36e^- \text{ Left} \\
 - 36e^- \text{ (lone-pairs on} \\
 \text{Cl's for octet)} \\
 \hline
 0e^-
 \end{array}$$



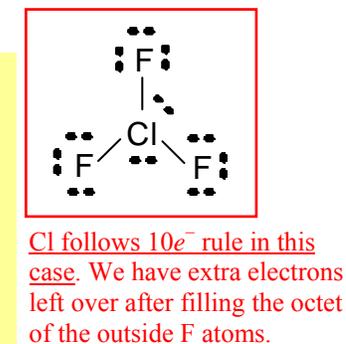
e. PF_5

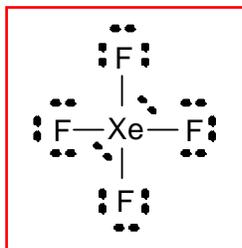
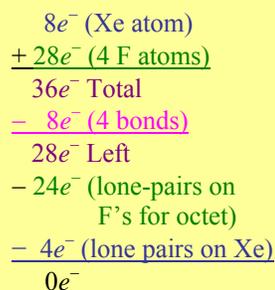
$$\begin{array}{l}
 5e^- \text{ (P atom)} \\
 + 35e^- \text{ (5 F atoms)} \\
 \hline
 40e^- \text{ Total} \\
 - 10e^- \text{ (5 bonds)} \\
 \hline
 30e^- \text{ Left} \\
 - 30e^- \text{ (lone-pairs on} \\
 \text{F's for octet)} \\
 \hline
 0e^-
 \end{array}$$



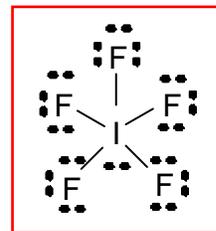
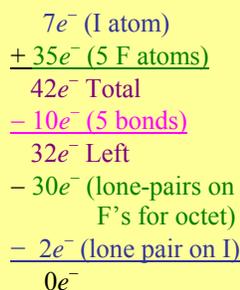
f. ClF_3

$$\begin{array}{l}
 7e^- \text{ (Cl atom)} \\
 + 21e^- \text{ (3 F atoms)} \\
 \hline
 28e^- \text{ Total} \\
 - 6e^- \text{ (3 bonds)} \\
 \hline
 22e^- \text{ Left} \\
 - 18e^- \text{ (lone-pairs on} \\
 \text{F's for octet)} \\
 - 4e^- \text{ (lone pairs on Cl)} \\
 \hline
 0e^-
 \end{array}$$



g. XeF₄

Xe follows $12e^-$ rule. We have extra electrons left over after filling the octet of the outside F atoms.

h. IF₅

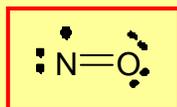
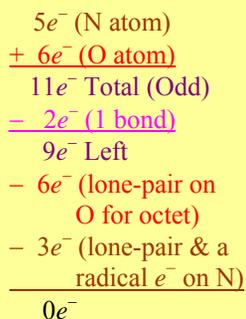
I follows $12e^-$ rule. We have extra electrons left over after filling the octet of the outside F atoms.

Odd Electron Molecules: - when the sum of all valence electrons and any net charges yields an odd number of electrons available.

- odd electron molecules do not fit well with the localized electron model because the model deals with pairs of electrons. Hence, **odd electron molecules are very unstable and very reactive.**

Example 2: Draw all the possible Lewis structures for NO and NO₂. Comment on the stability of each compound.

For NO:

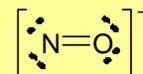


Need to move a lone pair from oxygen to between the atoms because the nitrogen does not have enough e^- for an octet.

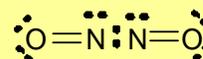
Note that we cannot bring another lone pair in from the oxygen atom because it means nitrogen would exceed the octet rule ($9e^-$ around N).

The overall **NO molecule is unstable** because nitrogen has odd number of electrons ($7e^-$) around itself.

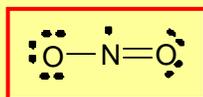
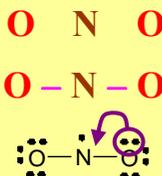
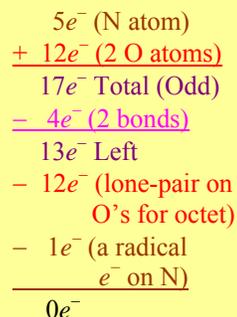
Remedy #1: Pick up an electron and form a polyatomic ion – hyponitrite (NO^-)



Remedy #2: Bond with another NO molecule to form a dimer – dinitrogen dioxide (N_2O_2)



Remedy #3: Decomposition back into N_2 and O_2 using a catalyst.

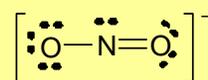
For NO₂:

Need to move a lone pair from oxygen to between the atoms because the nitrogen does not have enough e^- for an octet.

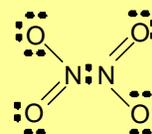
Note that we cannot bring another lone pair in from the oxygen atom because it means nitrogen would exceed the octet rule ($9e^-$ around N).

The overall **NO₂ molecule is unstable** because nitrogen has odd number of electrons ($7e^-$) around itself.

Remedy #1: Pick up an electron and form a polyatomic ion – nitrite (NO_2^-)



Remedy #2: Bond with another NO₂ molecule to form a dimer – dinitrogen tetroxide (N_2O_4)



Remedy #3: Decomposition back into N_2 and O_2 using a catalyst.



Assignment

9.9 pg. 308 #55, 57, 61 to 64

9.7 & 9.8: Formal Charges and the Concept of Resonance

Resonance: - occurs when more than one Lewis structure exists for a molecule.

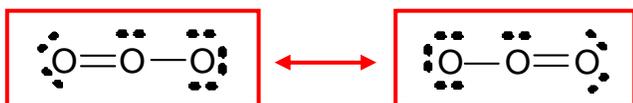
Resonance Structure: - different variations of a resonance molecule can be drawn using a double headed arrow, \leftrightarrow , between the diagrams.

Example 1: Draw the Lewis structure for the following molecules or polyatomic ions.

a. O_3

$18e^-$ (3 O atoms)
 $18e^-$ Total
 $- 4e^-$ (2 bonds)
 $14e^-$ Left
 $- 12e^-$ (6 lone pairs on outside O's for octet)
 $- 2e^-$ (lone pair on central O)
 $0e^-$

There are only $6e^-$ around the central O atom. An outside O atom needs to contribute a lone pair to form a double bond.

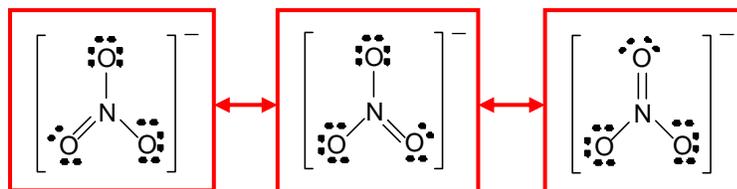


The double bond can be on either the left or right side of the central oxygen atom (resonance).

b. NO_3^-

$5e^-$ (1 N atom)
 $+ 18e^-$ (3 O atoms)
 $+ 1e^-$ (-1 charge of NO_3^-)
 $24e^-$ Total
 $- 6e^-$ (3 bonds)
 $18e^-$ Left
 $- 18e^-$ (lone pairs on O's for octet)
 $0e^-$

There are only $6e^-$ around the central N atom. An outside O atom needs to contribute a lone pair to form a double bond.



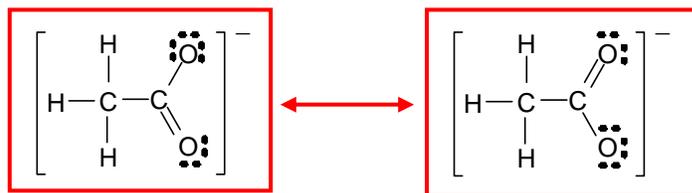
The double bond can come from any one of the O atoms (resonance).

c. CH_3COO^-

$8e^-$ (2 C atoms)
 $+ 3e^-$ (3 H atoms)
 $+ 12e^-$ (2 O atoms)
 $+ 1e^-$ (-1 charge)
 $24e^-$ Total
 $- 12e^-$ (6 bonds)
 $12e^-$ Left
 $- 12e^-$ (lone pairs on O's for octet)
 $0e^-$

C atoms are the central atoms – most bonding capacity (can make 4 single bonds). Note the structure is like how the formula is written.

There are only $6e^-$ around the one of the C atoms. An outside O atom needs to contribute a lone pair to form a double



The double bond can be on either the top or bottom side of the second carbon atom (resonance).

Formal Charge: - the difference between the number of valence electron of the original atom and the number of valence electron actually assigned to the atom in the Lewis structure.

- in cases where molecules have **more than one possible Lewis structure, the one with the Formal Charges CLOSEST to ZERO has the Lowest Energy. Hence, it is more stable than the other ones.**

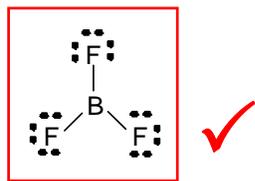
For **EACH atom in a molecule**, its **Formal Charge** is calculated by

$$\text{Formal Charge} = \# \text{ of } e^-_{\text{valence}} - \# \text{ of } e^-_{\text{unshared}} - \frac{1}{2} (\# \text{ of } e^-_{\text{bonding}})$$

Example 2: Draw all the possible Lewis structures for each of the following molecule or polyatomic ion. Determine the formal charge of the central atom and decide which structure is more favourable.

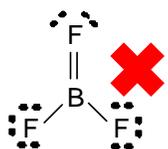
a. BF_3

$$\begin{array}{r} 3e^- \text{ (B atom)} \\ + 21e^- \text{ (3 F atoms)} \\ \hline 24e^- \text{ Total} \\ - 6e^- \text{ (3 bonds)} \\ \hline 18e^- \text{ Left} \\ - 18e^- \text{ (lone-pairs on} \\ \text{F's for octet)} \\ \hline 0e^- \end{array}$$



$$\begin{array}{l} \text{Formal Charge (B)} \\ = 3e^- - 0e^- - \frac{1}{2}(6e^-) \\ \text{(val.) (lone pair) (bonding)} \end{array}$$

Formal Charge (B) = 0
(Correct Structure)

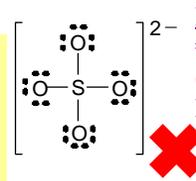


$$\begin{array}{l} \text{Formal Charge (B)} \\ = 3e^- - 0e^- - \frac{1}{2}(8e^-) \\ \text{(val.) (lone pair) (bonding)} \\ \text{Formal Charge (B) = -1} \\ \text{(incorrect structure)} \end{array}$$

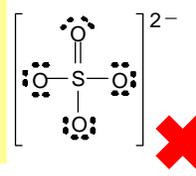
This is why B follows $6e^-$ rule. There is no need to bring a lone pair from a F atom to form a double bond for B to have an octet.

b. SO_4^{2-}

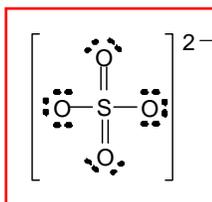
$$\begin{array}{r} 6e^- \text{ (S atom)} \\ + 24e^- \text{ (4 O atoms)} \\ + 2e^- \text{ (-2 charge)} \\ \hline 32e^- \text{ Total} \\ - 8e^- \text{ (4 bonds)} \\ \hline 24e^- \text{ Left} \\ - 24e^- \text{ (lone-pairs on} \\ \text{O's for octet)} \\ \hline 0e^- \end{array}$$



$$\begin{array}{l} \text{Formal Charge (S)} \\ = 6e^- - 0e^- - \frac{1}{2}(8e^-) \\ \text{(val.) (lone pair) (bonding)} \\ \text{Formal Charge (S) = 2} \\ \text{(incorrect structure)} \end{array}$$



$$\begin{array}{l} \text{Formal Charge (S)} \\ = 6e^- - 0e^- - \frac{1}{2}(10e^-) \\ \text{(val.) (lone pair) (bonding)} \\ \text{Formal Charge (S) = 1} \\ \text{(incorrect structure)} \end{array}$$



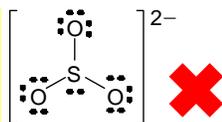
$$\begin{array}{l} \text{Formal Charge (S)} \\ = 6e^- - 0e^- - \frac{1}{2}(12e^-) \\ \text{(val.) (lone pair) (bonding)} \end{array}$$

Formal Charge (S) = 0
(Correct structure)

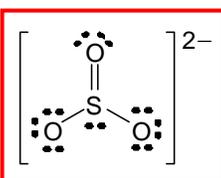
This is why S sometimes follows $12e^-$ rule. Both S=O bonds are opposite to each other to maximize symmetry of the polyatomic ion.

c. SO_3^{2-}

$$\begin{array}{r} 6e^- \text{ (S atom)} \\ + 18e^- \text{ (3 O atoms)} \\ + 2e^- \text{ (-2 charge)} \\ \hline 26e^- \text{ Total} \\ - 6e^- \text{ (3 bonds)} \\ \hline 20e^- \text{ Left} \\ - 18e^- \text{ (lone-pairs on} \\ \text{O's for octet)} \\ - 2e^- \text{ (lone pair} \\ \text{on central S)} \\ \hline 0e^- \end{array}$$



$$\begin{array}{l} \text{Formal Charge (S)} \\ = 6e^- - 2e^- - \frac{1}{2}(6e^-) \\ \text{(val.) (lone pair) (bonding)} \\ \text{Formal Charge (S) = 1} \\ \text{(incorrect structure)} \end{array}$$



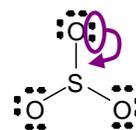
$$\begin{array}{l} \text{Formal Charge (S)} \\ = 6e^- - 2e^- - \frac{1}{2}(8e^-) \\ \text{(val.) (lone pair) (bonding)} \end{array}$$

Formal Charge (S) = 0
(Correct structure)

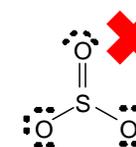
This is why S sometimes follows $10e^-$ rule (as the formal charge is now minimized to zero).

d. SO_3

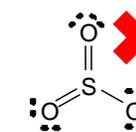
$$\begin{array}{r} 6e^- \text{ (S atom)} \\ + 18e^- \text{ (3 O atoms)} \\ \hline 24e^- \text{ Total} \\ - 6e^- \text{ (3 bonds)} \\ \hline 18e^- \text{ Left} \\ - 18e^- \text{ (lone-pairs on} \\ \text{O's for octet)} \\ \hline 0e^- \end{array}$$



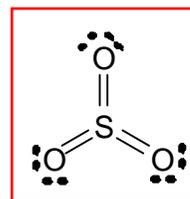
There are only $6e^-$ around the central S atom. An outside O atom needs to contribute a lone pair to form a double bond.



$$\begin{array}{l} \text{Formal Charge (S)} \\ = 6e^- - 0e^- - \frac{1}{2}(8e^-) \\ \text{(val.) (lone pair) (bonding)} \\ \text{Formal Charge (S) = 2} \\ \text{(incorrect structure)} \end{array}$$



$$\begin{array}{l} \text{Formal Charge (S)} \\ = 6e^- - 0e^- - \frac{1}{2}(10e^-) \\ \text{(val.) (lone pair) (bonding)} \\ \text{Formal Charge (S) = 1} \\ \text{(incorrect structure)} \end{array}$$



$$\begin{array}{l} \text{Formal Charge (S)} \\ = 6e^- - 0e^- - \frac{1}{2}(12e^-) \\ \text{(val.) (lone pair) (bonding)} \end{array}$$

Formal Charge (S) = 0
(Correct structure)

This is why S sometimes follows $12e^-$ rule (as the formal charge is now minimized to zero).

Assignment

9.7 & 9.8 pg. 307–308 #40, 42 to 44, 49 to 54