Lesson #4: Lewis Structures

May 23, 2018 4:27 PM

D. PREDICTING MOLECULAR STRUCTURES

How to Draw Lewis Structures

1. Determine the total number of valence electrons in the molecule

   - if species is a molecular ion, then adjust number accordingly (add for anions; subtract for cations)

most 2. Determine which atoms are bonded together and put two electrons into each bond.

least

Tips: i) The central atom is usually the highest valence, ~~most electropositive species.~~ → least electronegative
ii) Unless necessary (see next section), do not exceed the valence of each atom!

24 − 6e⁻ bond = 18e⁻ left to place

3. Use the remaining valence electrons to complete the octets of the surrounding atom(s). use 2e⁻ unpaired from Oxygen to satisfy C's octet.

4. Place any remaining electrons in pairs on the central atom.

5. If a central atom has less than an octet of electrons, then have an adjacent atom share an additional pair of its non-bonding electrons with the electron deficient atom (results in a double bond). If necessary (and appropriate), assign a triple bond.

6. Replace all bonding electron pairs with dashes.

Carbonate Ion: CO_3^{-2} added 2e⁻

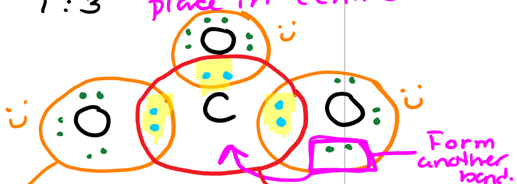
Example

C 3×O ion e⁻

#e⁻: 4 + 3·(6) + 2 = 24e⁻

Ratio C : O 1 : 3

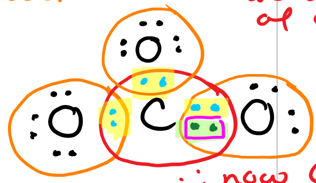
*C is least electroneg. place in centre



all O atoms satisfy octet = 8e⁻

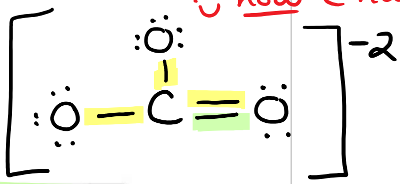
carbon is 2e⁻ short of octet.

Form another bond.



8e⁻

now C has 8e⁻



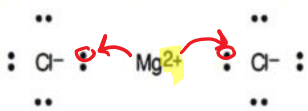
a) The Lewis Structures of Simple Ionic Compounds

The Lewis Structure of an ionic compound is simple to construct. You previously have been shown some example structures for such compounds; this section extends the previous work.

EXAMPLE: Draw the Lewis Structure of $MgCl_2$.

First, determine the charge expected for each ion. In this case, the ions are Mg^{2+} and Cl^-.

For the purposes of Chemistry 11, the nonmetal ions are symmetrically arranged around the metal ion. Remove the two electrons from the Mg atom to form the Mg^{2+} ion. Add one electron to each Cl atom to form the Cl^- ion.



Check back to IONIC bonding lesson

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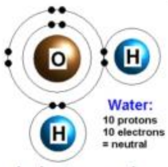
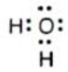
b) The Lewis Structures of Covalent Compounds that Obey the Octet Rule

Lewis Structures show how the ***VALENCE*** electrons are distributed in a molecule. The octet rule states that most atoms, other than hydrogen, tend to attain an octet of electrons as a result of forming covalent bonds.



EXAMPLE: The Lewis structure of H_2O is shown below.



For water, above, H and O contribute one electron each to the covalent bonds between them and share the two electrons in the bond. Each H can then "lay claim" to a ***closed shell*** of 2 electrons. The O atom has 4 electrons which it does not share with the H's in addition to the 4 electrons shared with the H's, for a total of 8 electrons: a closed shell and a "full octet".

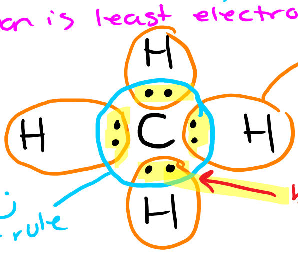
Drawing the Lewis Structures of molecules follows a simple set of rules.

Example 3 Draw the Lewis Structures for the methane, CH_4, and ethene, C_2H_4 .

a) methane : CH_4

valence e⁻ : 4+4·(1) = 8 e⁻ total

* Carbon is least electronegative.

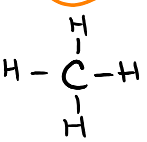


H has 2e⁻ 1s² orbital

bond

8e⁻ ☺ ✓octet rule

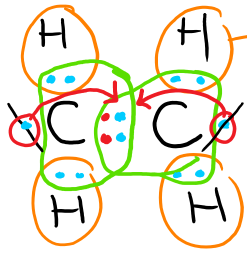
Lewis Structure



b) ethene : C_2H_4

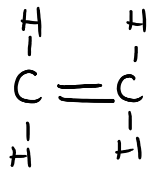
valence e⁻: 2·(4) + 4·(1) = 12e⁻

* Carbon least electroneg.



2e⁻ ☺

Carbon only had 7e⁻ in its orbital, so move 2e⁻ to form another bond.



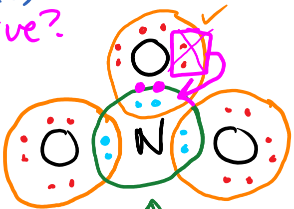
Example 4 Draw the Lewis diagram for the nitrate ion.

valence e⁻ : N 5 + 3×O 3×(6) + e⁻ +1 = 24e⁻ NO_3^-

*Least electronegative?

N or O

centre

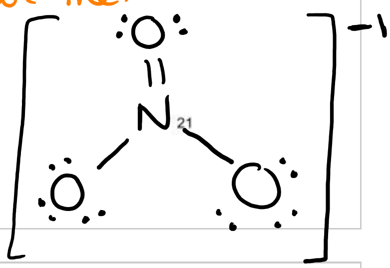


•• 2e⁻ for each bond.

24e⁻ − 6 = 18e⁻

N needs 2 more e⁻

have we met octet rule? 8e⁻



Exceptions to the Octet Rule:

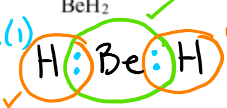
1. Fewer than 8 valence electrons
   Such molecules are called ***electron-deficient molecules***

1s²2s² ⟹ doesn't need to "fill" octet rule
1s² ⟶ " " " " 8 valence e⁻

(is "happy" with 6e⁻ in outer shell)

Examples: BeH_2 BF_3

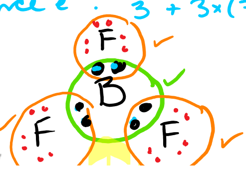
valence e⁻: 2+2(1) = 4

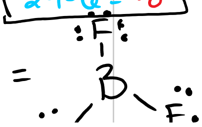


H−Be−H

valence e⁻: B 3 + F 3×(7) = 24e⁻

$\boxed{24-6=18}$





2. Expanded Valence (more than 8 valence e⁻)
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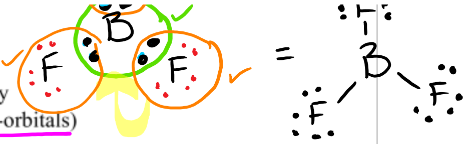
2. Expanded Valence (more than 8 valence e^-)



- central atoms from the third and fourth periods can be surrounded by more than 8 valence e^-s. (The extra e^-s get promoted to low-level d-orbitals)
- Please remember these two: P (up to 10 e^-s) and S (up to 12 e^-s)
- The good news is that the rules given earlier still apply!

Examples: PCl_5 SCl_6

P 5×Cl

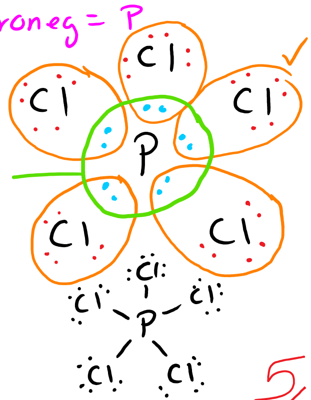
valence e^-: $5+5\cdot(7)=40e^-$

least electroneg = P

$40e^- - 10 = 30e^-$

✓$10e^-$

Cl ✓$8e^-$

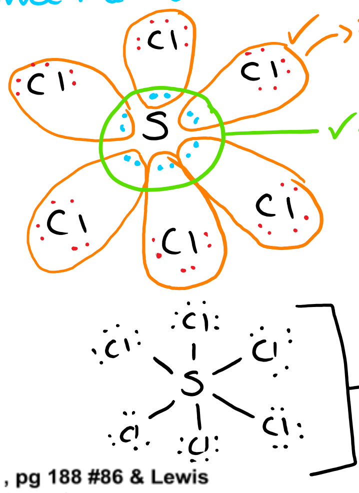


S Cl

valence: $6+6\cdot(7)=48e^-$

✓ stable with $8e^-$ (octet rule)

✓ stable with $12e^-$



Lewis Structure

Homework

Assignment ~~#4~~ 5 Hebden pg. 183 #85 , pg 188 #86 & Lewis Diagrams / Bonding Worksheet *(next page)*

all assignments are to be completed on a separate page with the assignment number & heading

EXERCISE:

85. Draw the Lewis Structure of each of the following ionic compounds.
(a) KBr (b) $AlCl_3$ (c) MgO (d) Li_2S (e) K_3P

86. Assign Lewis structures to the following molecules.

(a) H—Cl (b) I—I (c) I—Cl

(d) $H_3C–CH_3$ (H—C—C—H, each C bonded to two H above and below) (e) C—C with two H on each C (f) H—C—C—H

Answers in textbook

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Lewis Diagrams / Bonding Worksheet

work may be completed in the space provided on this page

Answers Posted online.

1. Draw electron dot diagrams for the following atoms: (1 mark each)

a. C
b. H
c. Na
d. Cl
e. P
f. Kr
g. Fr
h. Mg
i. Al
j. O

2. Draw electron dot diagrams for the following ions: (1 mark each)

a. Na^+
b. Cl^-
c. Fr^+
d. Mg^{2+}
e. Si^{4+}

3. Draw Lewis structures for the following compounds: (1 mark each)

a. CO
b. PBr_3
c. SeF_2
d. KI
e. CH_4
f. C_2H_4
g. N_2
h. H_2O
i. O_2
j. N_2Br_4
k. $SeCl_2$
l. H_2S
m. $CH_3CH_2CH_3$
n. SO_3^{-2}
o. H_2