Molecular Compounds & Lewis Structures

- In *ionic compounds*, electrons are completely transferred from one atom to another to form ions.
  - The negative and positive ions then attract to each other forming *ionic bonds*
    opposite charges attract (ie. electrostatic attraction)
- In **molecular compounds**, electrons are only shared between nonmetal atoms to satisfy the octet rule for each element in the compound
  - When electrons are only shared between two nonmetal atoms, a **covalent bond** is formed
  - The electrons that are shared by the atoms in a molecular compound are called a **bonding pair**
  - One bonding pair of electrons shared between two atoms creates a **single covalent bond**. Two or three bonding pairs of electrons shared between the same two atoms creates a **double or triple covalent bond**, respectively

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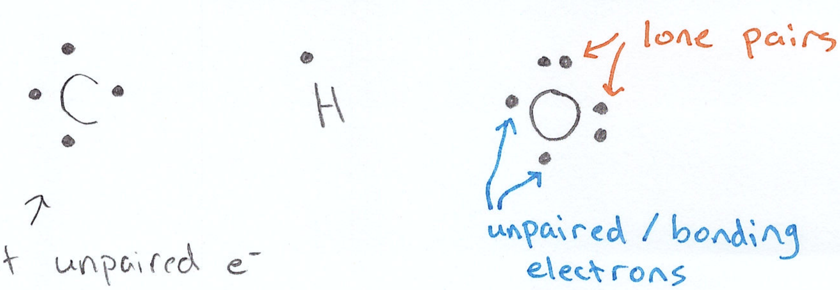
- Electron dot diagrams work good for showing how ionic compounds form due to the attractive force between oppositely charged ions (ie. ionic bonds) because the electron dot diagram illustrate how valence electrons are transferred from one element to another to create ions with satisfied octets
- When electron dot diagrams are used to show *entire molecular molecules*, we call these drawing **Lewis structures**

electron dot diagrams used for individual atoms & ions!

EXAMPLE: Draw the Lewis structure for CH_2O.

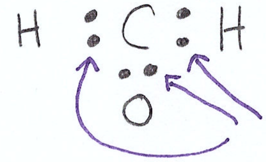
- Steps to drawing Lewis structures for simple molecules
  1. Determine the total number of valence electrons in all the atoms in the molecule. If the molecule is charged, add or subtract electrons to account for the charge.
     - Remember, the number of valence electrons is indicated by the group number (ignoring the "1" in front for all numbers over 10)
  2. Place the symbol of the atom (or atoms if there is more than one) with the most *unpaired/bonding electrons* at the centre.
     - You may need to draw out the electron dot diagram for each atom to determine which atom has the most unpaired/bonding electrons

C = 4 v.e$^-$
H = 1 v.e$^-$ (x2 = 2
O = 6 v.e$^-$
total = 12 v.e$^-$



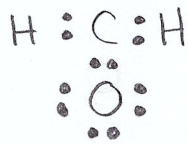
most unpaired e$^-$ ∴ central atom

3. Arrange all other atoms around the central atom as symmetrically as possible and start by drawing one pair of electrons (ie. bonding pair) between each set of atoms to represent a *single covalent bond*.



bonding pairs creating covalent bonds

4. Place *lone pairs* of electrons around the atoms except the central atom to satisfy the octet rule. Remember hydrogen can only have two electrons in its valence/outer shell.



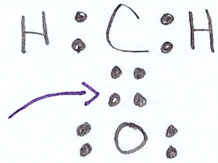
12 valence e^-

5. If all the valence electrons have not been accounted for, add one or more lone pairs around the central atom to complete an octet of electrons.

all 12 valence e^- accounted for

6. If all the valence electrons have been used up but the central atoms still does not have an octet of electrons, move one or more of the lone pairs to form *double or triple bonds* between the central atom and an adjacent atom.

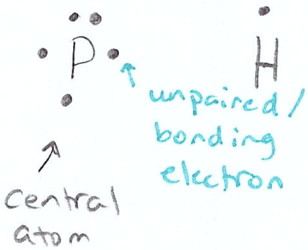
represents a double bond



all elements have their octet satisfied

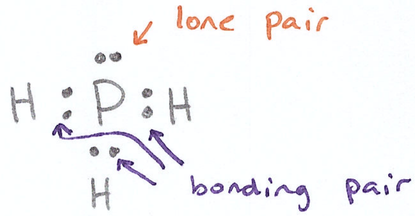
EXAMPLES: Draw the Lewis structure for the following molecules.

1. PH_3



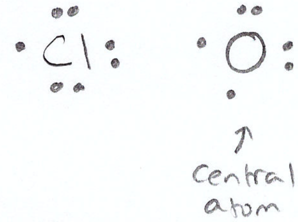
$P = 5$ v.e$^-$
$H = 1$ v.e$^-$ (x3)

total = 8 v.e$^-$



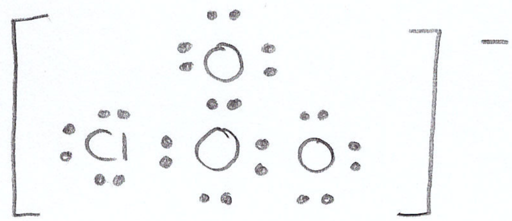
all elements have octet satisfied

3. ClO_3^-



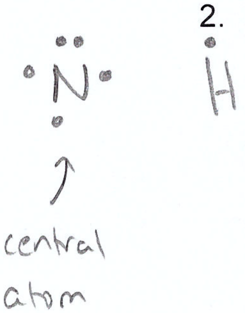
Cl = 7 v.e$^-$
O = 6 v.e$^-$ (x3)
"−" = +1 v.e$^-$

total = 26 v.e$^-$



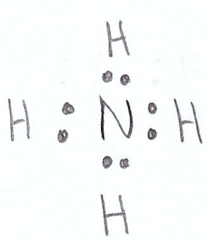
all elements have complete octets

2. NH_4^+

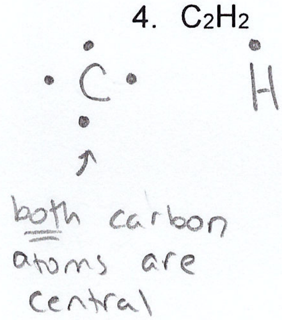


$N = 5$ v.e$^-$
$H = 1$ v.e (x4)
"+" = −1 v.e$^-$

total = 8 v.e$^-$

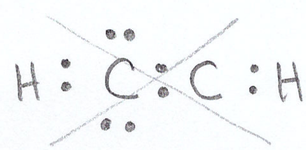


4. C_2H_2

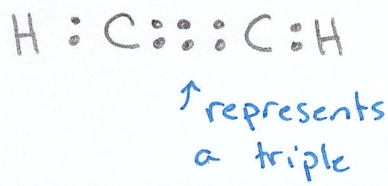


C = 4 v.e$^-$ (x2)
H = 1 v.e$^-$ (x2)

total = 10 v.e$^-$



lone pairs cannot complete all octets ∴ need double or triple bonds



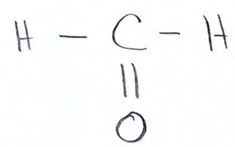
*****Now try Practice Problem #1*****

- Lewis structures can be tedious to draw, so **<u>structural formulas</u>** are another way to draw molecules

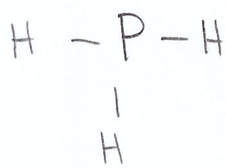
- A structural diagram shows every individual atom in the molecule and shows the bonds between them
    - Structural formulas use a line to represent a bond (a pair of shared electrons or *<u>bonding pair</u>*) but do not show *<u>lone pairs</u>*
    - Two or three lines represent a double or triple bond respectively

- <u>EXAMPLES:</u> Draw structural formulas for the following molecules.

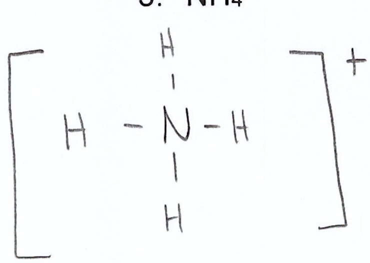
1. CH_2O



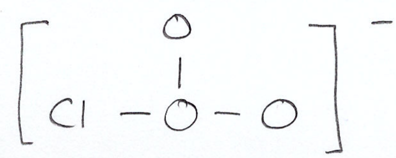
2. PH_3



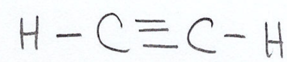
3. NH_4^+



4. ClO_3^-



5. C_2H_2



*****Now try Practice Problem #2*****

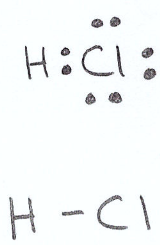
Practice Problems

1. Draw Lewis structures for the following compounds.

a. $HCl_{(g)}$

H = 1 v.e⁻
Cl = 7 v.e⁻

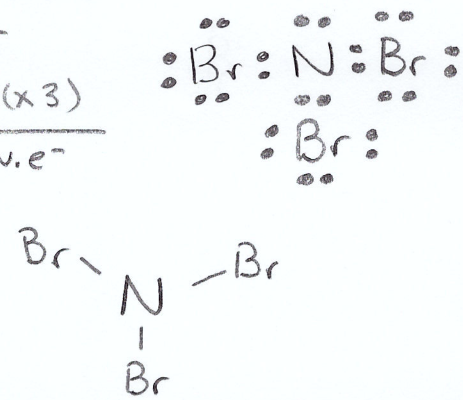
total = 8 v.e⁻



e. $NBr_{3(s)}$

N = 5 v.e⁻
Br = 7 v.e⁻ (x3)

total = 26 v.e⁻



b. $F_{2(g)}$

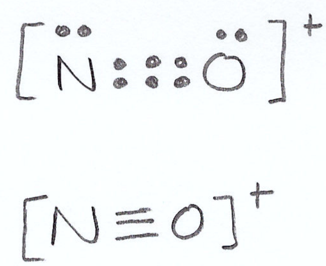
F = 7 v.e⁻ (x2)

total = 14 v.e⁻

f. NO^+

N = 5 v.e⁻
O = 6 v.e⁻
"+" = −1 v.e⁻

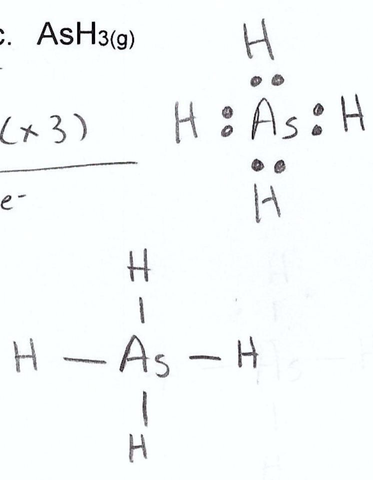
total = 10 v.e⁻



c. $AsH_{3(g)}$

As = 5 v.e⁻
H = 1 v.e⁻ (x3)

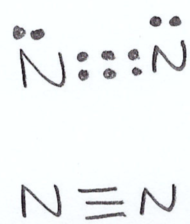
total = 8 v.e⁻



g. $N_{2(g)}$

N = 5 v.e⁻ (x2)

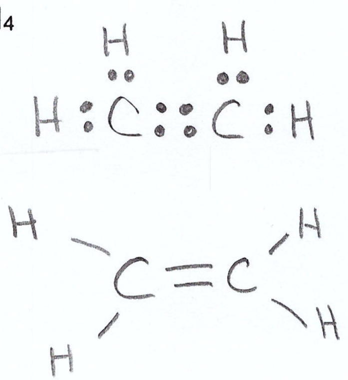
total = 10 v.e⁻



d. C_2H_4

C = 4 v.e⁻ (x2)
H = 1 v.e⁻ (x4)

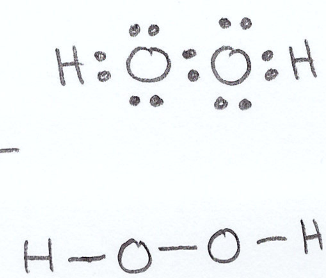
total = 12 v.e⁻



h. $H_2O_{2(l)}$

H = 1 v.e⁻ (x2)
O = 6 v.e⁻ (x2)

total = 14 v.e⁻



i. $HOCl_{(g)}$

H = 1 v.e⁻
O = 6 v.e⁻
Cl = 7 v.e⁻

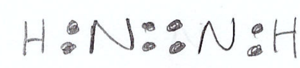
total = 14 v.e⁻

H – O – Cl

k. $N_2H_{2(g)}$

N = 5 v.e⁻ (x2)
H = 1 v.e⁻ (x2)

total = 12 v.e⁻



H – N = N – H

j. $ClNO_{(g)}$

Cl = 7 v.e⁻
N = 5 v.e⁻
O = 6 v.e⁻

total = 18 v.e⁻

Cl – N = O

l. $CO_{2(g)}$

C = 4 v.e⁻
O = 6 v.e⁻ (x2)

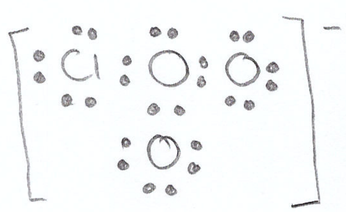
total = 14 v.e⁻

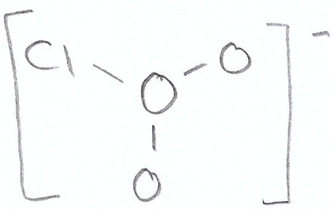
O = C = O

m. $ClO_3^-{}_{(aq)}$

Cl = 7 v.e⁻
O = 6 v.e⁻ (x3)
"–" = +1 v.e⁻

total = 26 v.e⁻

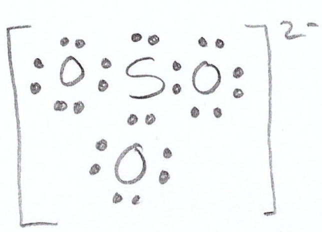


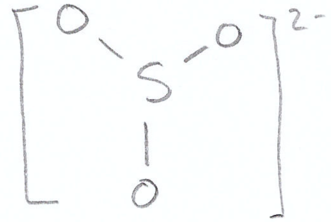


n. SO_3^{2-}

S = 6 v.e⁻
O = 6 v.e (x3)
"2–" = +2 v.e⁻

total = 26 v.e⁻





2. Draw structural formulas for the molecules in #2.