Chapter 4 Homework Key

4.28 How many bonds and lone pairs are typically observed with each element:

- Each element makes bonds to share e⁻ to bring it's valence count up to 8 (2 for hydrogen). Thus the number of bond and lone pairs is determined by looking first at the valence e⁻ and the Lewis dot Structure of the element.

a. O

Lewis Dot	e⁻ required for octet	# bonds required
·Ö· (with 2 lone pairs)	2e⁻	2 bonds

The number of e⁻ required to reach 8 valence is equal to the bonds required as each bond equals the addition of 1e⁻.

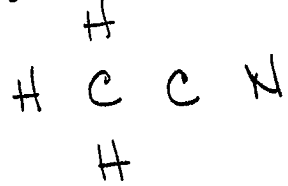
—Ö— (2 lone pairs) 2 bonds 2 lone pairs

	Lewis Dot	e⁻ required	Lewis Struct.	# Bonds	# Lone Pairs
b. Si	·Si· (4 dots)	4e⁻	Si with 4 bonds	4	Ø
c. Ge	·Ge· (4 dots)	4e⁻	Ge with 4 bonds	4	Ø
d. B	·B· (3 dots)	3e⁻	B with 3 bonds	3	Ø

* Boron is an exception. His normal bonding pattern is 3 bonds for 6e⁻ (octet deficient). B can reach an octet with 4 bonds, but then the B will have a negative charge B⊖ (4 bonds) = ·B· (4 dots) one more valence e⁻ than he comes with.

4.34 Draw a valid Lewis Structure for each compound using the given arrangement of atoms.

a. C_2H_3N



* Arrangement given because these are multicenter.

Valence e^- count:

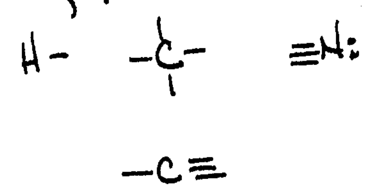
$2\,C \times 4 = 8$

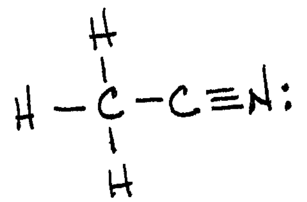
$3\,H \times 1 = 3$

$1\,N \times 5 = +5$

16 valence

Relevant bonding patterns from the chart handed out in class:



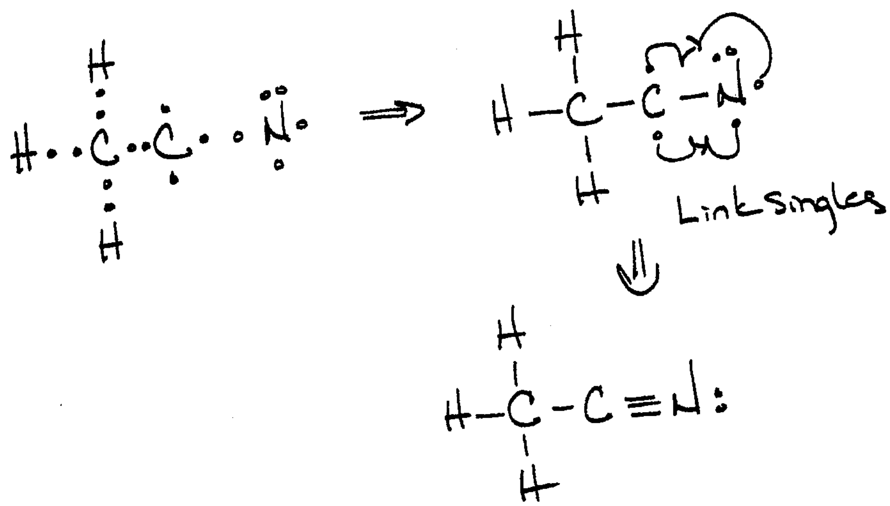


16 valence e^- used ✓

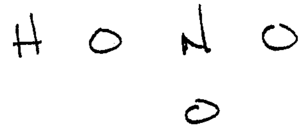
H has duet ✓

C & N have octet ✓

Could also construct from valence e^- as follows:



4.34 b. HNO_3



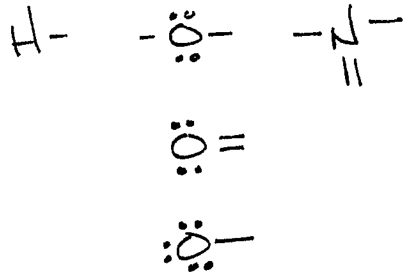
*Arrangement given because these are multicenter.

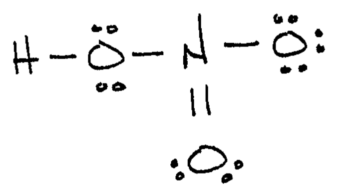
Valence Count

$1 \times H \times 1 = 1$
$1\,N \times 5 = 5$
$3\,O \times 6 = 6$

12 valence

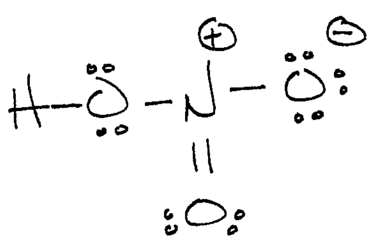
Relavent bonding patterns from the chart handed out in class:





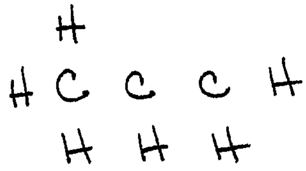
12 valence e^- used ✓
H has duet ✓
O & N have octed ✓

There are formal charges in this structure. Nitrogen has a positive $-\overset{+}{N}-$ and one oxygen has a negative $-\ddot{O}:^-$ but the net charge is zero.



Net = ∅

4.34 c. ~~C_3H_4~~ C_3H_6



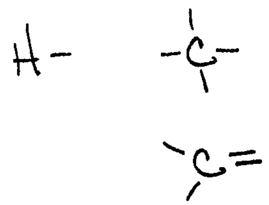
* Arrangement given because these are multicenter.

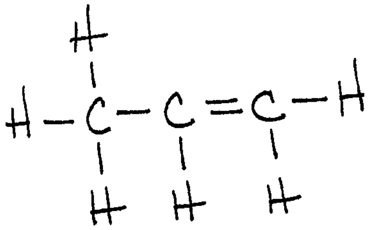
Valence e⁻ count $3C \times 4 = 12$

$6 ~~4~~ H \times 1 = + ~~4~~ 6$

~~16~~ 18 valence e⁻

Relavent bonding patterns from the chart handed out in class:





16 valence e⁻ used X

- Something wrong. Notice how I had to go back to the beginning to find my error.

18 valence e⁻ used ✓

H has duet ✓

C has an octet ✓

4.38 Draw a valid Lewis structure for each ion:

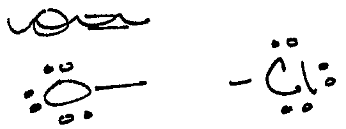
a. OCl^-

Valence e^- count $1O \times 6 = 6$
$1Cl \times 7 = 7$
$+1$ for * 1- charge

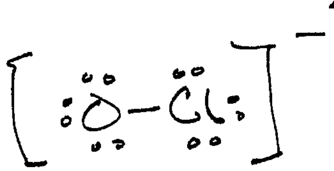
$14e^-$

* e^- are negative
– charge means extra e^-

Bonding patterns:



Charge Required for ions



$14e^-$ used ✓

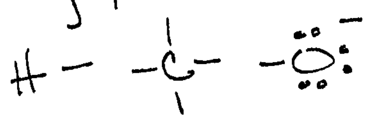
Both oxygen and Chlorine have octet ✓

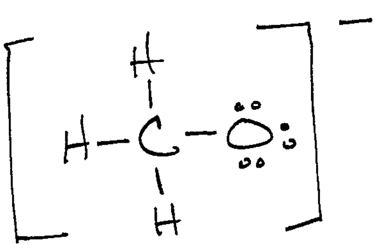
b. CH_3O^-

Valence e^- count $3H \times 1 = 3$
$1C \times 4 = 4$
$1O \times 6 = 6$
$+1$ for 1- charge

$14e^-$

Bonding patterns



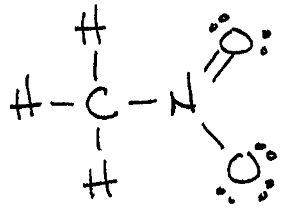


$14e^-$ used ✓

H has duet ✓

Oxygen and Carbon have octet ✓

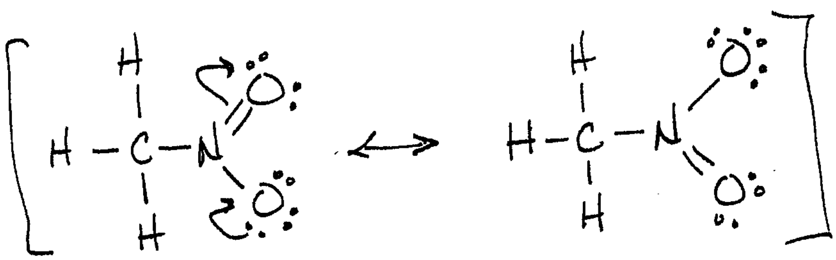
4.44 Draw a second resonance structure for nitromethane, a compound used in drag racing fuel and in the manufacture of pharmaceuticals, pesticides and fibers.



nitromethane

— Answer: resonance structures are easy to draw when you follow a few rules.

1. only move double bonds and lone pairs.
2. Atoms must be connected by ~~all~~ at least one bond.
3. You may not add or remove e⁻ from the total count.
4. Do not move the positions of the atoms.



Resonance structures are drawn with a double arrow ⟷ between structures. All structures are inclosed in square brackets [ ].

4.52 Write a formula that corresponds to each name.

a. Silicon tetrafluoride $\Rightarrow$ SiF_4

b. nitrogen oxide

- This one is an error, there are many nitrogen oxides. I believe the auther intended it to be nitrogen monoxide $\Rightarrow$ NO

c. phosphorus triiodide $\Rightarrow$ PI_3

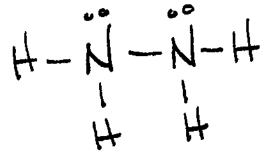
4.64 Draw a Lewis structure for N_2H_4 and explain why the shape around each N atom should be described as trigonal pyramidal.

Valence e⁻ count $2\,N \times 5 = 10$

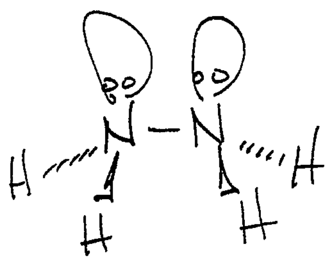
$4\,H \times 1 = +4$

$14\,e^-$

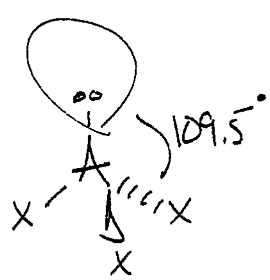
Bonding Patterns: H– $\quad$ –N̈–



14e⁻ used ✓
H has duet ✓
N has octet ✓



The geometry about each nitrogen is AEX_3 corresponding to trigonal pyramidal.



4.72 Label the bond formed between fluorine and each of the following as nonpolar, polar, or ionic.

a. F-H $\Delta EN = 4.0 - 2.1 = 1.9$ **Polar**

b. F-F $\Delta EN = 4.0 - 4.0 = \emptyset$ **nonpolar**

c. F-C $\Delta EN = 4.0 - 2.5 = 1.5$ **polar**

d. F-Li $\Delta EN = 4.0 - 1.0 = 3.0$ **ionic**

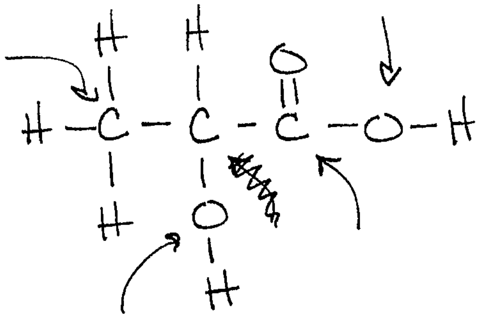
e. F-S $\Delta EN = 4.0 - 2.5 = 1.5$ **polar**

The electronegativity values are taken from ~~pat~~ page 111. The book lists the following values (which are slightly different than the ones I gave in class. Lets agree to <u>use</u> the books values.)

0 —— 0.5 ———————— 1.9 ————→

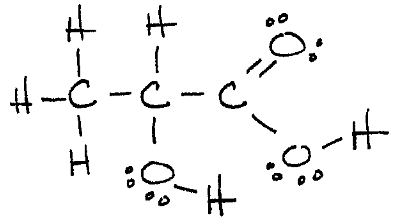
nonpolar   polar   Ionic

4.90 Lactic acid gives sour milk its distinctive taste. Lactic acid is also ~~the~~ a ingredient in several skin care products that purportedly smooth fine lines and improve skin texture.



a. Add lone pairs where needed, and then count the total number of valence e^- in lactic acid.

– ~~Add~~ to complete octets –

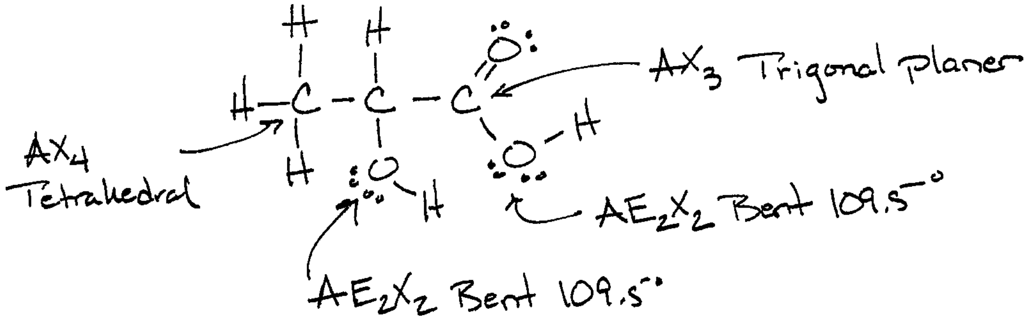


36 e^- total

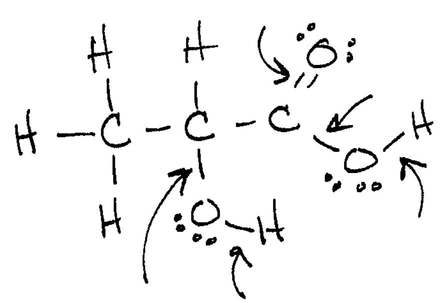
$3\,O \times 6 = 18$
$3\,C \times 4 = 12$
$6\,H \times 1 = +6$

$36\,e^-$ ✓

b. Determine the shape around the four indicated atoms.



4.90 c. Label all of the polar bonds.

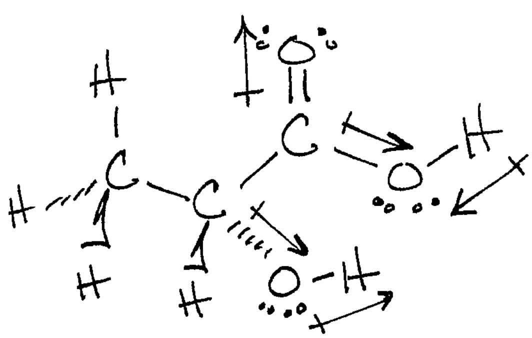


Bonds to consider:

C – H	ΔEN = 0.4	nonpolar
C – O, C = O	ΔEN = 1.0	polar
O – H	ΔEN = 1.4	polar
C – C	ΔEN = ∅	nonpolar

d. Is Lactic acid a polar or nonpolar molecule? Explain.

– First draw the molecule in a way that allows you to see the 3-D shape



– next draw each of the dipoles corresponding to each polar bond

$\delta+ \longrightarrow \delta-$

– Note how complicated the picture is. There is no simple symmetry which allows the dipoles to cancel out. If the dipoles do not cancel out then the molecule is polar.

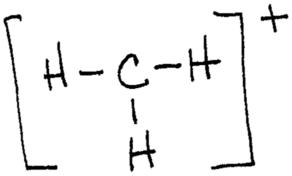
Polar!

4.94 Although carbon has four bonds in stable molecules, sometimes reactive carbon intermediates that contain carbon atoms without four bonds are formed for very short time periods. Examples of these unstable intermediates include the methyl carbocation CH_3^+ and the methyl carbanion CH_3^-. Draw Lewis structures for both unstable ions and predict the shape around each carbon.

Carbocation $\underline{C}H_3^+$

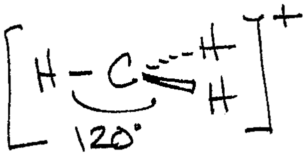
valence e^- count

$1C \times 4 = 4$
$3H \times 1 = 3$ (+)
$7e^-$
-1 for +
6 valence e^-



$6e^-$ used ✓

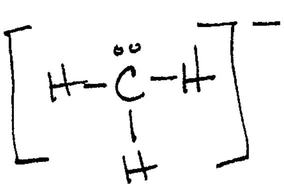
AX_3 = Trigonal planer



Carbanion CH_3^-

valence e^- count

$1C \times 4 = 4$
$3H \times 1 = +3$
$7e^-$
$+1$ for −
8 valence e^-



$8e^-$ used

AEX_3 = Trigonal pyramidal