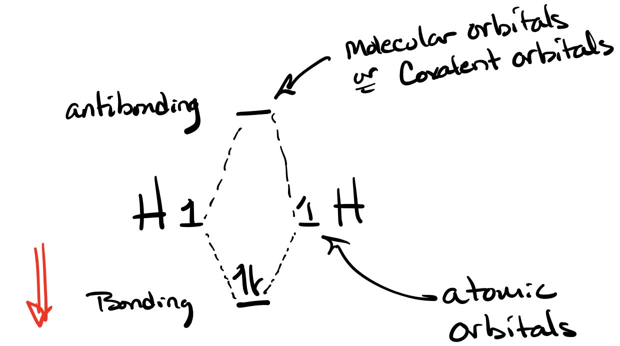
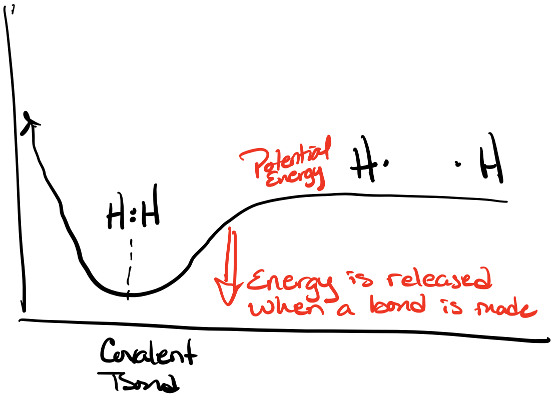
Covalent Bonding



orbitals that reside on an atom

Energy is lower

Energy is released when bond is made



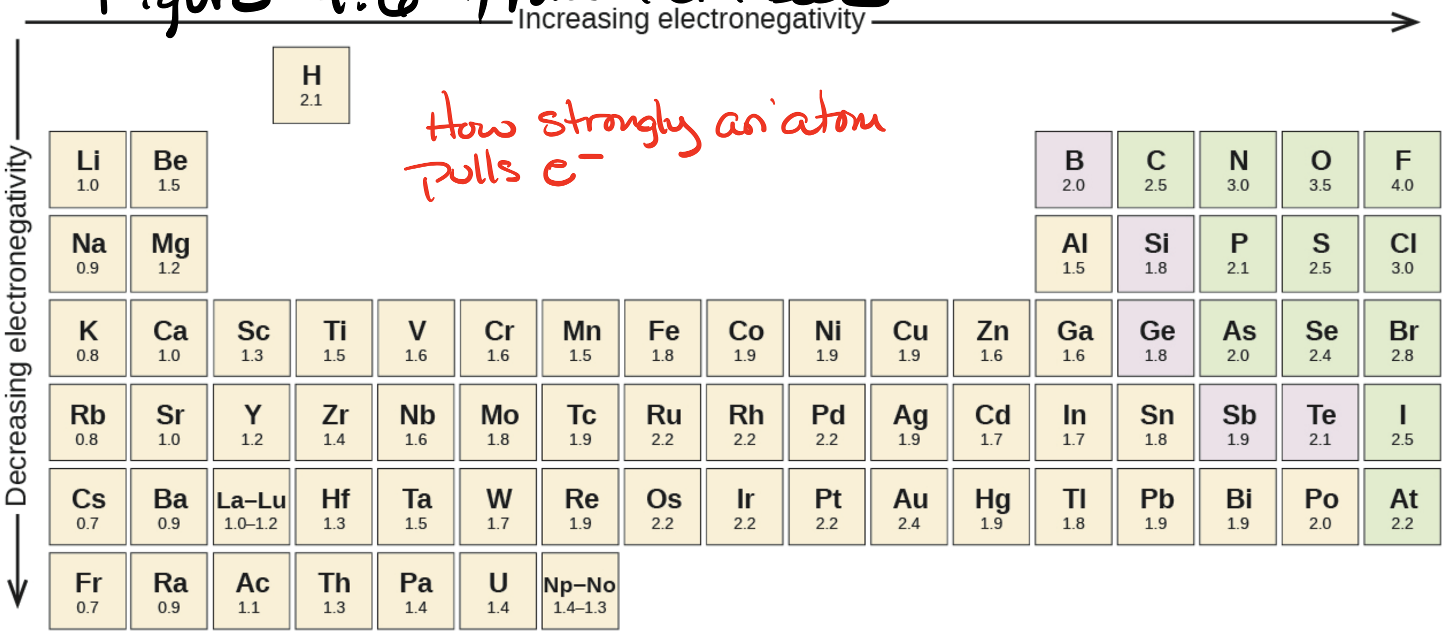
Sharing of $2e^-$

$H:H$ same as $H-H$

Lewis Dots

Lewis Structure

Figure 4.6 from textbook



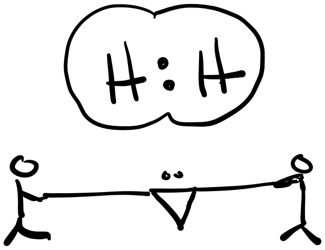
		H 2.1															
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La–Lu 1.0–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac 1.1	Th 1.3	Pa 1.4	U 1.4	Np–No 1.4–1.3										

How strongly an atom pulls e^-

Element 1 — Element 2

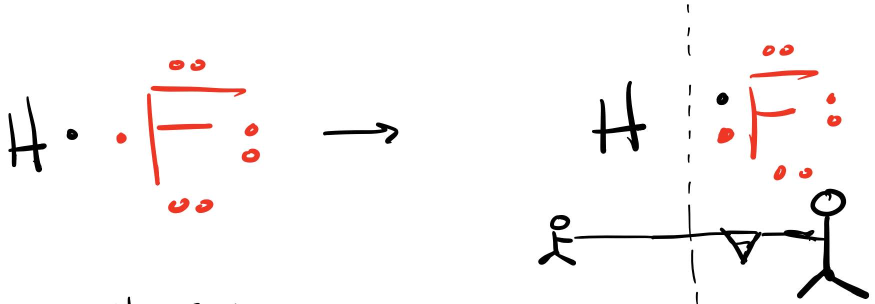
difference in EN $= |EN_1 - EN_2|$

EN of H = 2.1 unitless



$\Delta EN = |2.1-2.1| = |0| = 0$
↑
difference

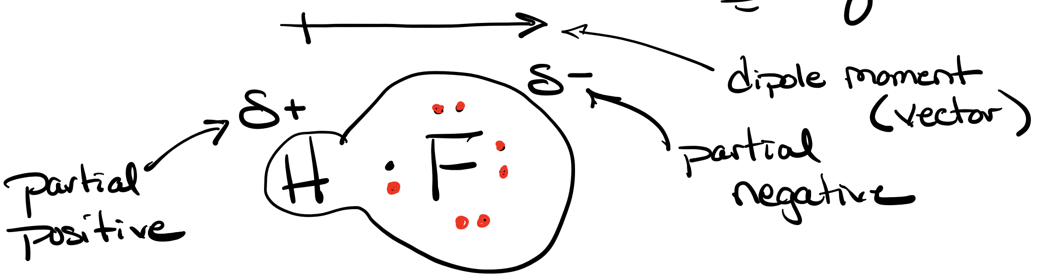
e^- absolutely equally divided between the two atoms.



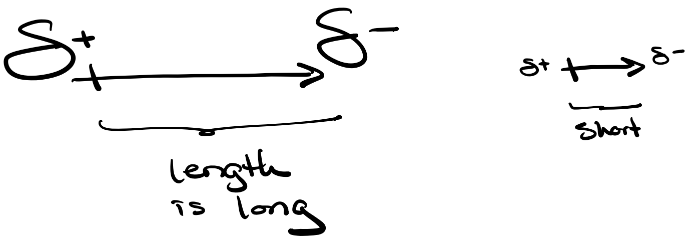
EN H = 2.1
F = 4.0

Fluorine getting more of the e^- than hydrogen

$\Delta EN = |4.0-2.1| = 1.9$ ⇐ the sharing is <u>not</u> equal



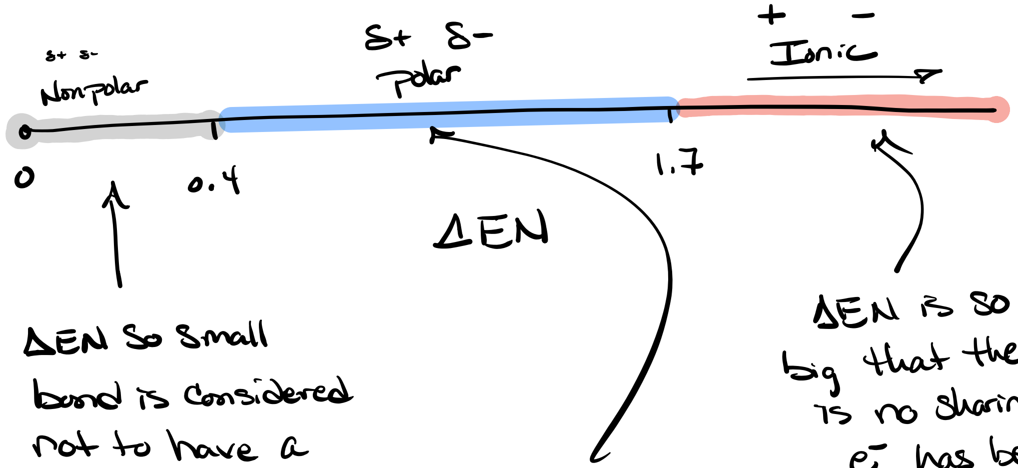
Vector has size & direction



dipole = two poles one $\delta+$ & $\delta-$

Indication of how equal or unequal the sharing of the e^- are.

Bonding Contineum



ΔEN so small bond is considered not to have a dipole

$0 \leq 0.4$

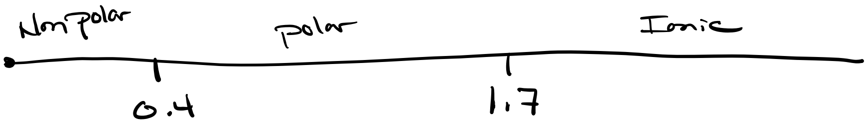
ΔEN is big enough to say the bond is considered polar.

$0.4 \leq 1.7$

ΔEN is so big that there is no sharing e^- has been traded.

$1.7 <$

Bond	ΔEN	non-polar, polar, Ionic ?	
$\overset{0.9}{Na} - \overset{3.0}{Cl}$	$\Delta EN = \|3.0-0.9\|$ $= 2.1$	Ionic	$Na^+ \; Cl^-$
$\overset{3.5}{O} - \overset{2.1}{H}$	$\Delta EN = \|3.5-2.1\|$ $= 1.4$	polar bond	$\delta^- \leftarrow\!\!+ \; \delta^+$ O—H
$\overset{2.5}{C} - \overset{2.1}{H}$	$\Delta EN = \|2.5-2.1\|$ $= 0.4$	non-polar	$\delta^- \leftarrow\!\!+ \; \delta^+$ C—H
$\overset{1.0}{Li} - \overset{2.1}{P}$	$\Delta EN = \|2.1-1.0\|$ $= 1.1$	polar	$\delta^+ \; +\!\!\rightarrow \delta^-$ Li—P
$\overset{1.8}{Fe} - \overset{3.5}{O}$	$\Delta EN = \|3.5-1.8\|$ $= 1.7$	**polar** most would say this is Ionic	$\delta^+ \; +\!\!\rightarrow \delta^-$ Fe—O



0 ≤ 0.4   0.4 ≤ 1.7   1.7 <

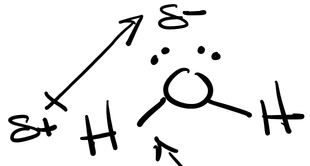
Ionic

Cation & anion comprised of a **metal** cation & nonmetal anion

Metal ⟹ Ionic

Strict Definition   Ionic $\Delta EN > 1.7$

H_2O water



polar, nonpolar or Ionic

larger
↓

$$\overset{3.5}{O} - \overset{2.1}{H}$$

ΔEN

$\Delta EN = |3.5-2.1| = 1.4$ polar bond

$$H - \overset{2.5}{C} \equiv \overset{3.0}{N}\colon$$

$\Delta EN = 0.5$

polar, nonpolar, or ionic

triple bond, but treated the
Same as a single bond.

$$\overset{2.1}{\underset{}{}}\quad \overset{2.5}{\underset{}{}}$$

$\delta+ \mapsto \delta \quad \delta+ \longmapsto \delta-$

$$H - C \equiv N$$

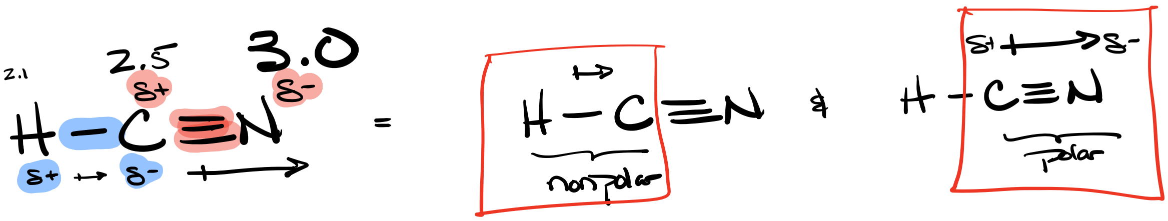
$\Delta EN = |2.5-2.1| = 0.4$

nonpolar bond $\Delta EN = 0.4$

polar bond $\Delta EN = 0.5$

$$H - C \equiv N$$

2.1	2.0	2.5	3.0	3.5	4.0
H	B	C	N	O	F



Ended Chapter 4.2 on Electronegativities

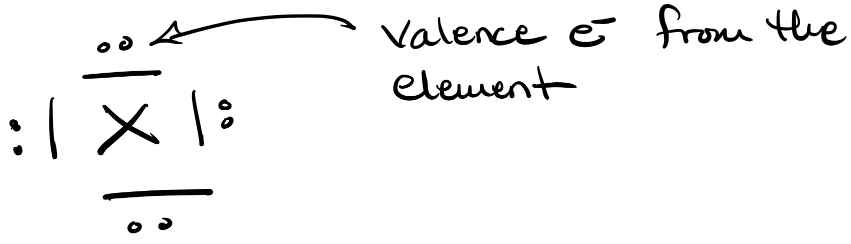
Chapter 4.3 is nomenclature (we covered this earlier)

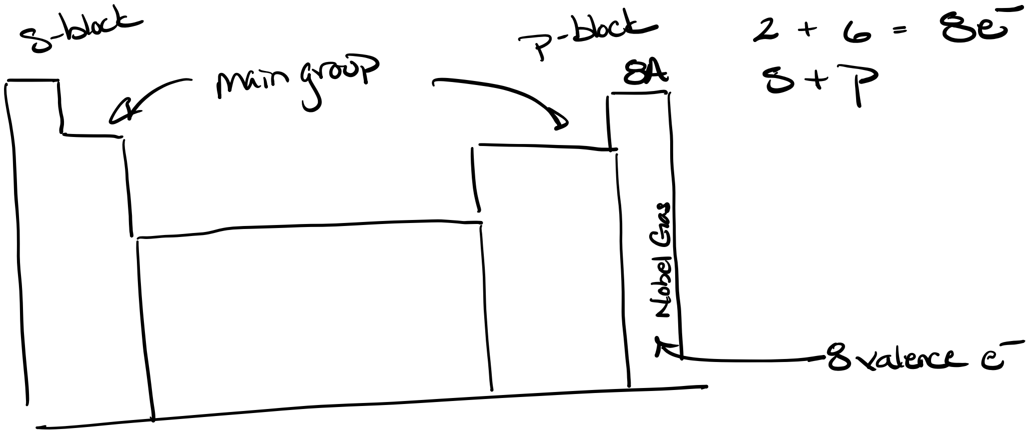
Chapter 4.4 is Lewis Structures

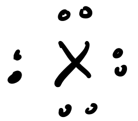
Lewis Structures

- provide or Communicate atom Connectivity
- Model or theory that helps to find or predict the atom Connectivity of a Simple inorganic or organic molecule.
- System used to write the atom Connectivity of Simple & Complicated molecules.

Start with Lewis Dot Structures







8e⁻ ⇒ an octet

filled s & p orbitals
isoelectronic w/ nobel gas

Octet Rule

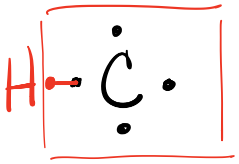
Every element (except H) is trying to obtain an octet.

$H \longrightarrow He$ $1s^2$ ← duet

Group "A" values are the # of valence e^-

not usually in Lewis structures because they make ions

The goal is to share e^- in order to obtain an octet



Each bond adds $1e^-$ to the count for each atom

First of two Systems - Traditional

- Locate central atom ✓
  - ⇒ often the least EN atom
  - ⇒ often communicated by an underline

$H\underline{C}N$ $\leftarrow$ central $\qquad H-C\equiv N:$

$\underline{C}O_2 \qquad :\ddot{O}=C=\ddot{O}:$

- Bond all elements to the central with a single bond. ✓
- place lone pairs of e^- on outter most (most EN) elements to give them an octet ($8e^-$)
- Do an octet count
- use tripple or double bonds to give central atom an octet if needed.
- Check valence e^- used
- Check octets

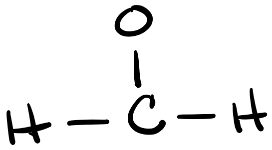
- Check formal charges. Later

$\underline{C}H_2O$ * central element underlined

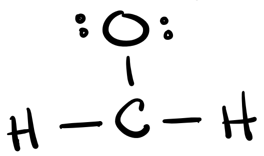
Valence e^- Count

			#ve⁻		
C	1	×	$4e^-$	=	$4e^-$
H	2	×	$1e^-$	=	$2e^-$
O	1	×	$6e^-$	=	$+\ 6e^-$
					$12e^-$ total

①



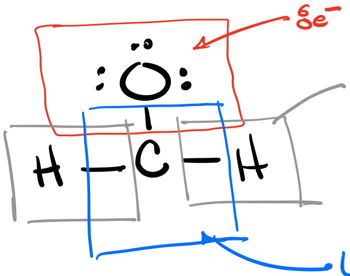
②



more EN
Give octet

Ignore H (they want a duet)

③

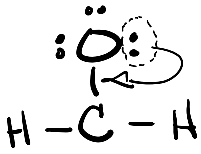


$8e^-$ ✓ octet

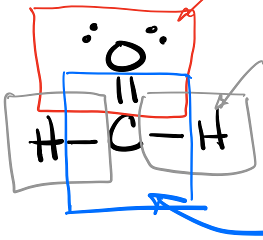
$2e^-$ ✓ Duet

$6e^-$ X need $8e^-$

④



make a double bond →



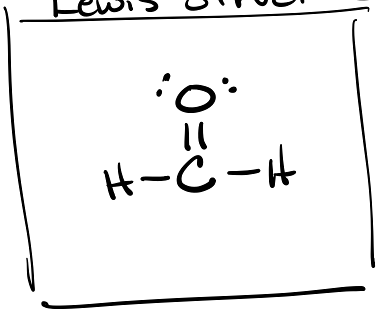
$8e^-$ ✓

$2e^-$ ✓ duet

$8e^-$ ✓

all octets satisfied ✓

Lewis Structure



Valence e^- used $12e^-$ ✓

octets C ✓ O ✓ H ✓

Formal Charge — later

CH_2O

Formaldehyde

4A = $4e^-$

6A = $6e^-$

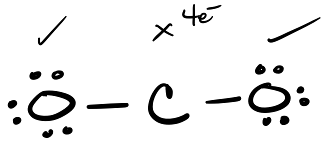
CO_2

Valence e^- Count

C $1 \times 4e^- = 4e^-$

O $2 \times 6e^- = 12e^-$

$16e^-$

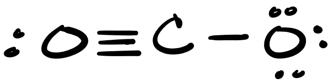


Two choices to fix this

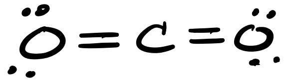
$4e^-$ from one oxygen

or

$2e^-$ from each oxygen



or



Valence e^- ✓

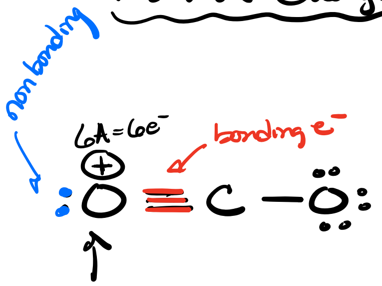
octets ✓

Valence e^- ✓

octets ✓

Formal Charge —

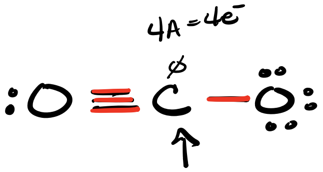
Formal Charge = Valence e^- − ½ bonding e^- − nonbonding e^-



$6e^- - \frac{1}{2}(6) - 2$

$6 - 3 - 2$

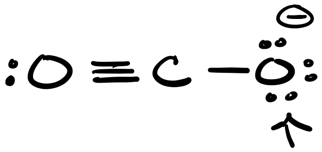
$= +1$



$4e^- - \frac{1}{2}(8) - 0$

$4 - 4$

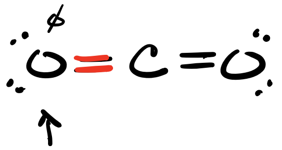
$= 0$



$6e^- - \frac{1}{2}(2) - 6$

$6 - 1 - 6$

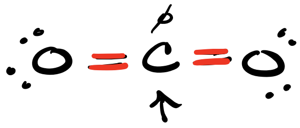
$= -1$



$6e^- - \frac{1}{2}(4) - 4$

$6 - 2 - 4$

$= 0$

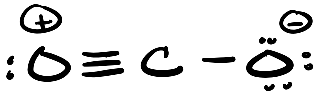


No formal charges

$4e^- - \frac{1}{2}(8) - 0$

$4 - 4$

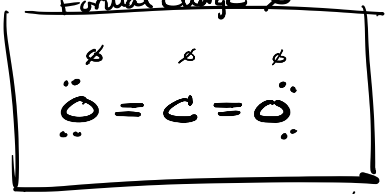
$= 0$



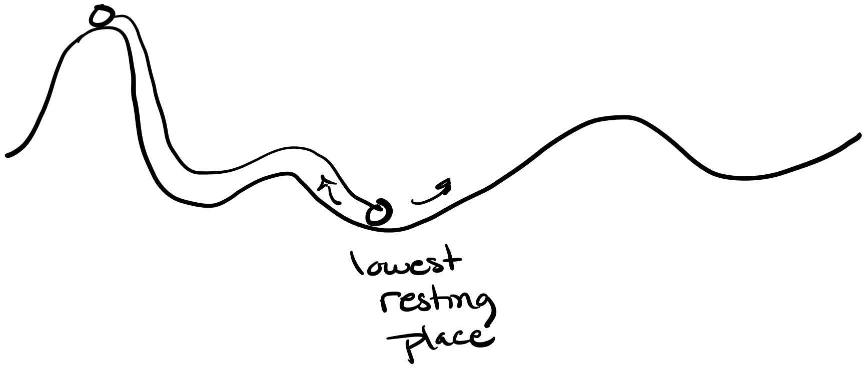
octets ✓

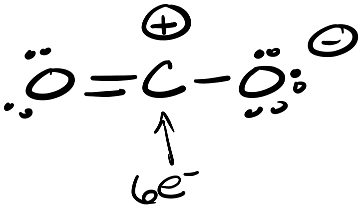
16 valence ✓

Formal Charge 0 ✓



The structure with the lower number of formal charges is the better (lower energy) structure.


lowest
resting
place


+
−
O=C−O
6e⁻