③ "-ITE" rule: $HNO_2 \rightarrow H^+ + \boxed{NO_2^-}$ ← nitrite

root + "-ous acid" = Nitrous acid

HClO hypochlorite ⟹ hypochlorous acid

Lewis Symbols

- convenient way to show valence electrons
  - dot ("•") represents one e^-
  - placed around the outside of an atom symbol

element	valence e^-	Lewis symbol
carbon	4	·C̈· ← needs 4 more e− for octet
nitrogen	5	·N̈· ← 3 more
oxygen	6	·Ö: ← 2 more
fluorine	7	·F̈: ← 1 more
neon	8	:N̈e: ← super stable with 8 V.E.

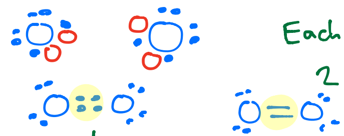
Covalent bonds form so that each atom has 8 V.E.
, we call this an octet

F_2 ← two F atoms covalently bonded to form a stable molecule

1. Draw Lewis symbol for each atom — :F̈· ·F̈: — unpaired = not stable
2. Share e^- to form a pair
   - these shared e^- are "felt" by both nuclei — :F̈:F̈: — shared = covalent bond
3. rewrite molecule showing ALL covalent bonds as a line (":" → "—")
4. check to make sure all atoms have an octet AND correct # of e^- used — :F̈—F̈: (8, 8) — 14 e^- used ✓

Molecular Oxygen O_2

O has 6 valence e^- (v.e.)



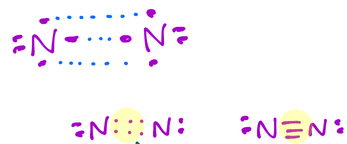
Each atom needs 2 electrons

double bond
2 pairs of shared e^-

* double bonds are shorter and stronger than single bonds *

Molecular Nitrogen N_2

$N \rightarrow 5$ v.e.



* Triple bonds are shorter and stronger than double bonds! *

Triple bond
3 pairs of shared e^-

	pm length	energy (aJ)
C−C	153	0.58
C=C	134	1.02
C≡C	120	1.35

N_2 is VERY unreactive because of the triple bond
- it's very hard to break

Lets move on to more complicated molecules:

$H_2O \rightarrow$ Note that H is written 1st $\rightarrow$ water is an acid

$H_2O \rightarrow H^+ + OH^-$

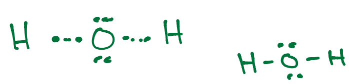
hydroxide

hydroxic acid

Hydrogen is in the 1st shell, so a duplet (1 pair) is STABLE

* NEVER give H more than 1 bond! *

So O must be the central atom because it is the only one that can form multiple bonds



* Always show Lone pairs! * (LP)

# of ve	
7	Cl
7	Cl
6	O
20	

Cl_2O

20 ← Lewis structure needs to have exactly 20 ve

2(10) ✓

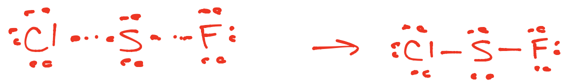
SClF

how are the atoms arranged? The general rule is that the least electronegative atom is the central atom.

E.N $S = 2.5$

$Cl = 3.0$

$F = 4.0$



General strategies for drawing Lewis structures:

← needs to make 1 covalent bond

← needs to form 2 bonds

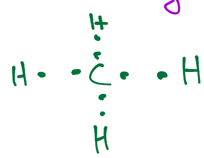
← needs to form 3 bonds

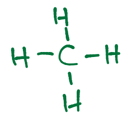
← 4 bonds

① Use valence e- as a guide →

I find this method to be the most useful

CH_4





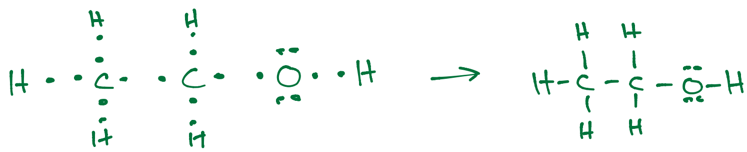
remember H can only form 1 bond, so C must be connected to C

CH_3CH_2OH

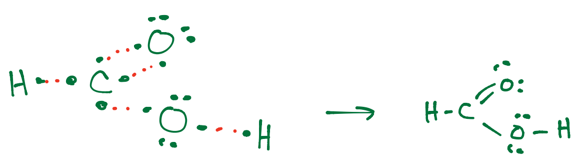
Things to ALWAYS check for:

① Do all atoms have an octet? (except H)

② Did you use the correct # of VE?

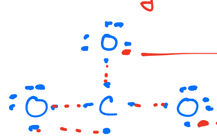


CHO_2H



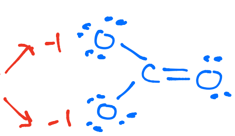
CO_3^{2-} ← we'll be needing 2 extra electrons

EN: 2.5 3.5



2 extra electrons on the molecule because of the -2 charge

extra e- went here!



octets? yes

VE? $C = 4$

$O = 3(6) = 18$

$-2 = 2$

24 → 24 used ✓

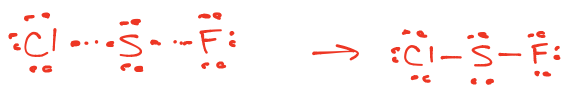
SClF

E.N $S = 2.5$

$Cl = 3.0$

$F = 4.0$

how are the atoms arranged? The general rule is that the least electronegative atom is the central atom.



General strategies for drawing Lewis structures:

① Use valence e- as a guide →

← needs to make 1 covalent bond

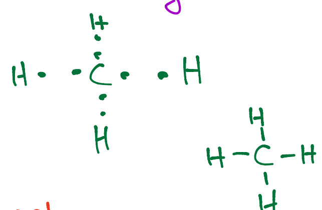
← needs to form 2 bonds

← needs to form 3 bonds

← 4 bonds

I find this method to be the most useful

CH_4



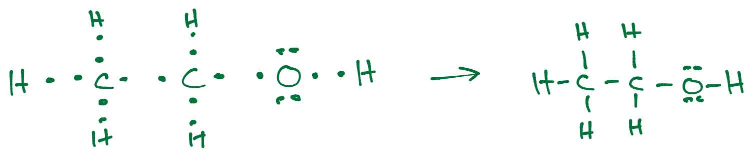
remember H can only form 1 bond, so C must be connected to C

CH_3CH_2OH

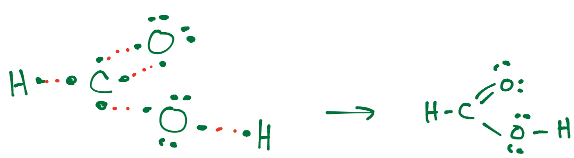
Things to ALWAYS check for:

① Do all atoms have an octet? (except H)

② Did you use the correct # of VE?

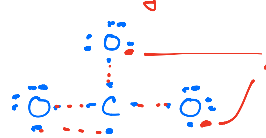


CHO_2H



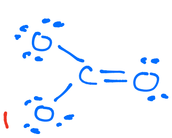
CO_3^{2-} ← we'll be needing 2 extra electrons

EN: 2.5 3.5



2 extra electrons on the molecule because of the -2 charge

extra e- went here!



octets? yes

VE? $C = 4$

$O = 3(6) = 18$

$-2 = 2$

24 → 24 used ✓

How can we figure out which atom(s) carry the charge?

- well, we know that we put 1 extra e^- on Oxygen 1 + Oxygen 2, so it would make sense that they would be −1 each (so −2 overall charge), but Formal Charge gives us a way to confirm

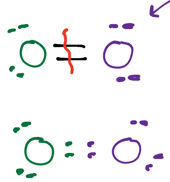
F.C. = V.E. − "owned" electrons

this is the # of V.E. that the atom SHOULD have

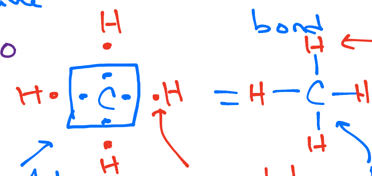
an atom owns its lone pairs AND one electron from each bond

FC = 6−6=0

FC = 6−6=0



6 owned 6 owned



FC = 1−1=0

= 4 owned

owned by carbon

owned by hydrogen

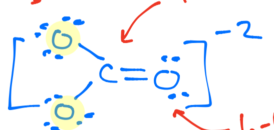
FC = 4−4=0

Formal charges MUST add up to the charge on the molecule!

6−7=−1

4−4=0

these two atoms carry a −1 charge

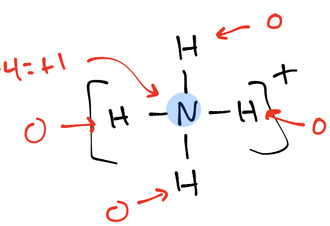


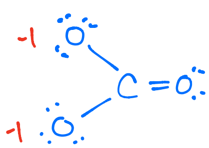
6−6=0

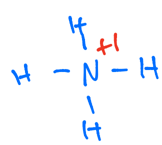
6−7=−1

−1 + −1 + ∅ + ∅ = −2

5−4=+1







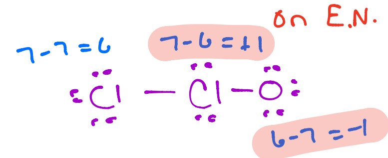
Central Atoms are NOT always the least electronegative.

Minimizing Formal charge is more important

OCl_2 ← this is a neutral molecule. If we can draw it with no formal charge, that would be most ideal

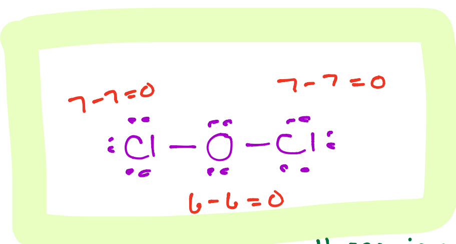
EN → 3.5 3.0

We would predict that Cl would be a central atom based on E.N. values



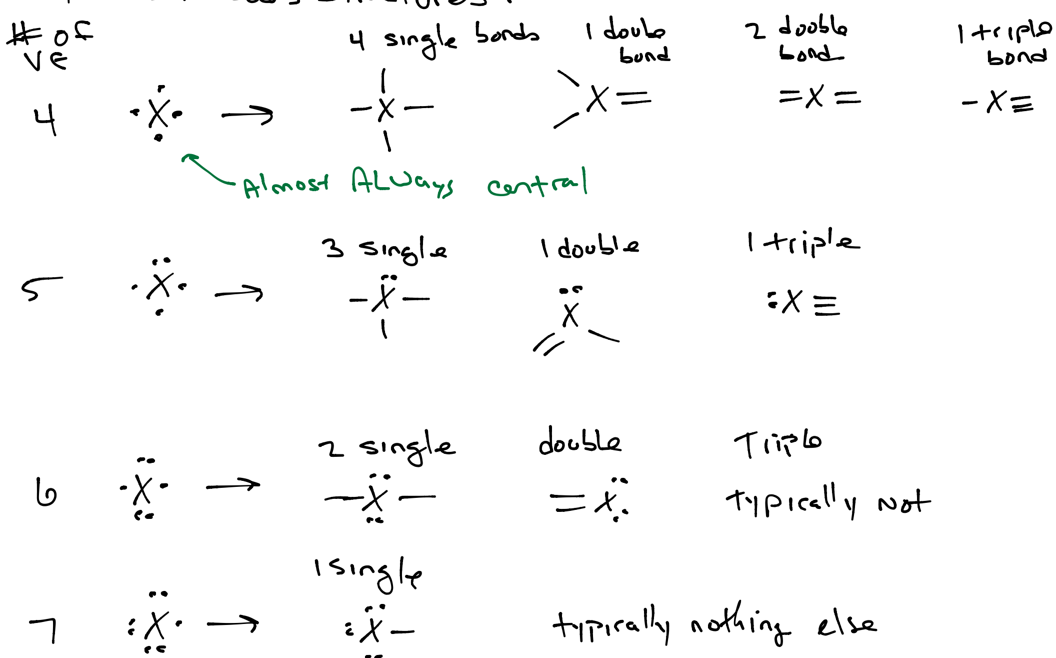
- formal charge adds up to ∅? yes
- correct # of ve? yes
- all octets? yes

so this is an acceptable structure, but it is NOT the most stable because we can draw OCl_2 so that there is ∅ formal charge

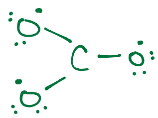


- all requirements (FC, # of e⁻, + octets) are met AND there is no F.C., so this is more stable

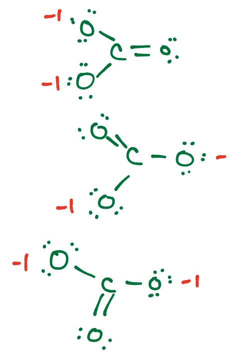
Themes in Lewis Structures:



Lets think about carbonate: CO_3^{2-}



there are multiple ways to connect the dots to satisfy our rules



Carbonate has $\underline{\underline{3}}$ resonance forms/ structures

All three of the structures are perfectly equivalent

Resonance Structures: two or more stable conformations of the same molecule

Since each of these structures are equally stable, each will exist. The physical properties of each bond/atom is an average of all structures

Are the CO bonds single or double? single bond in 2/3 + double in 1/3

$$\frac{2}{3}(1) + \frac{1}{3}(2) = \frac{4}{3}$$

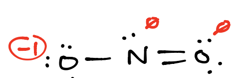
each CO bond is partially single (2/3) and partially double (1/3)

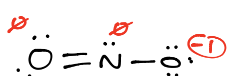
- so each CO bond will be shorter than a single bond but longer than a double bond. Also, weaker than a double bond + stronger than a single

What is the charge on each oxygen? $2/3(-1) + 1/3(\emptyset) = -2/3$

Consider nitrite (NO_2^-). Determine the charge on each atom.

Two resonance forms are possible:





$N \rightarrow \emptyset(1/2) + \emptyset(1/2) = 0$

$O \rightarrow -1(1/2) + \emptyset(1/2) = -1/2$

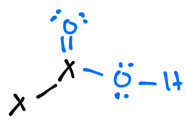
Each oxygen has an average charge of $-1/2$

Resonance Structure are indicated by showing $\longleftrightarrow$ between the L.S.

this unambiguously identifies these as Resonance Structures

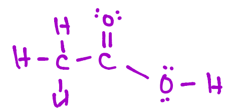
Common themes in L.S.

Acids :

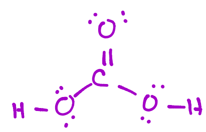


Acids commonly have H written at the beginning of molecular formulas

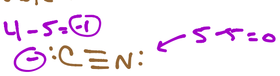
example: CH_3COOH acetic acid



H_2CO_3 carbonic acid

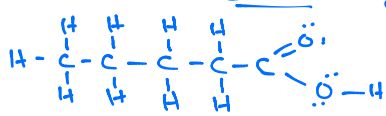


Cyanate: what atom has the (-) ?



commonly see things like: $CH_3CH_2CH_2CH_2COOH$

· You should be comfortable quickly recognizing patters! H only gets one bond, so C-C

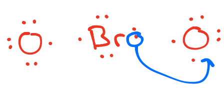


Expanded Octets

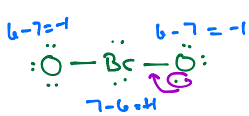
BrO^-

central

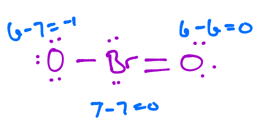
BrO_2^-



feel free to shuffle electrons. After all, the bond MUST be formed



Lots of formal charge! what if we do this



Less formal charge. Is this ok based on our rules? No!

- But — this $\boxed{\text{IS}}$ the correct structure of BrO_2^-.

The octet rule works because covalent bonds are formed using valence atomic orbitals.

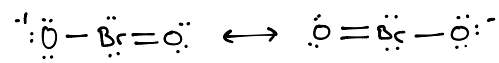
Hydrogen can only form 1 bond because it only has 1 valence orbital.

Atoms in shell 2 have 4 valence orbitals (2s 2px 2py 2pz), so they are restricted to 4 bonds (or 4 Lone pairs)

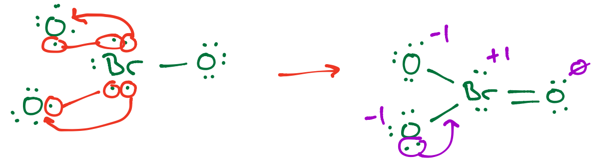
Bromine is in the 4th shell, so it is not restricted to 4 bonds / Lone pairs. It can access the d orbitals to form a bond or occupy the LP.

More on this Later this week!

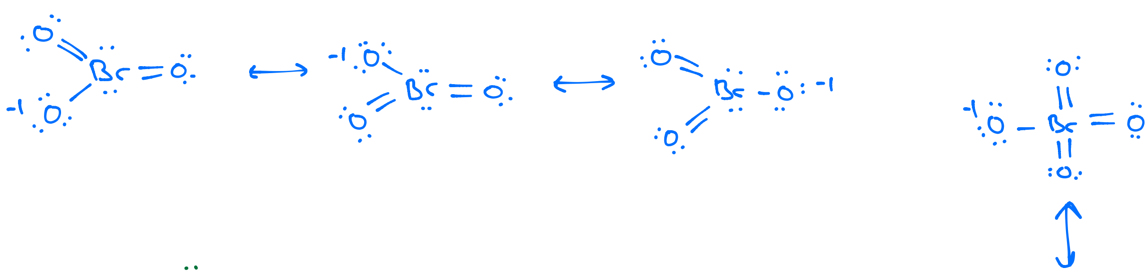
We call this an expanded octet.



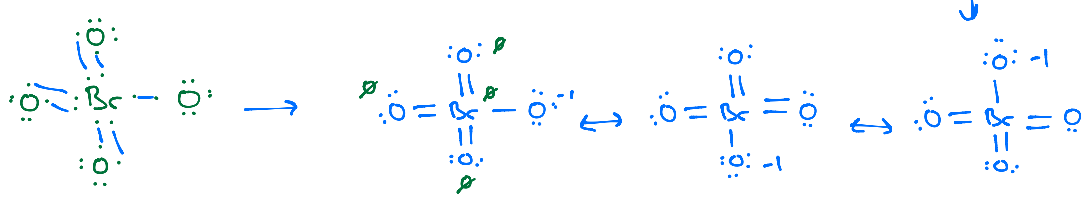
BrO_3^-



Minimize F.C.!



BrO_4^-



General approach to drawing Lewis structures:

1. Draw each atom with correct number of VE
2. Connect the dots
3. Add / subtract e- for anions/cations
4. Check formal charge
5. Can any atom expand the octet to minimize FC? If so, do it.

It's generally true that the least electronegative atom is going to be central.

Always show Lone Pairs!

Your book says "put unique ato in the center"

Formal charge = VE - Electrons owned

2 (LP) + number of bonds