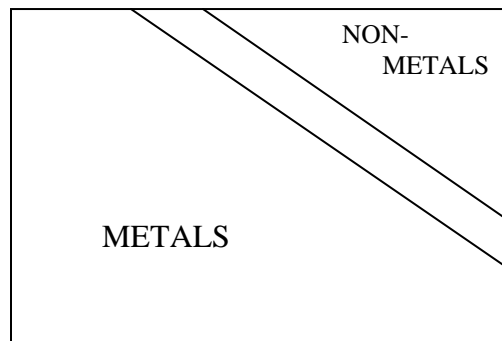
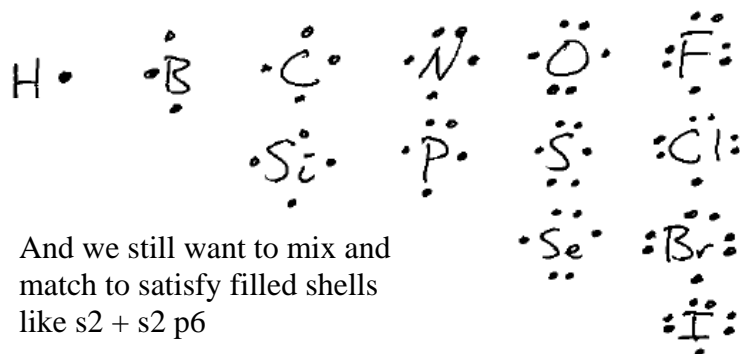


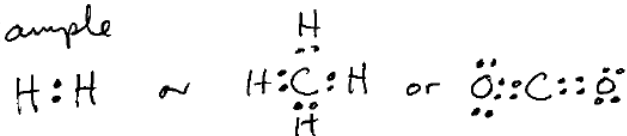
LECTURE 8. INTRODUCTION TO COVALENT LEWIS DOT STRUCTURES

In the upper right corner of the periodic table are the non metals. These relatively few atoms have the following Lewis structures.



And we still want to mix and match to satisfy filled shells like $s^2 + s^2 p^6$

Example



In all cases, covalent bonds involve the sharing of electrons. NO ions are formed.

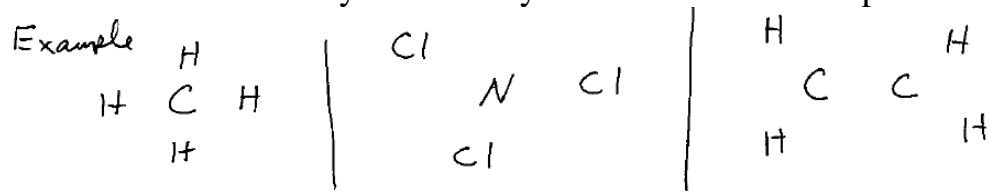


Note that even though the e^- pair is shared, it is found closer to one atom than another. Electronegativity will explain this.

Systematic Approach to Drawing Lewis Structures

Is there a systematic way to build Lewis structures of molecules? Try this:

1. Write out the atoms symmetrically with the most electropositive in the center.



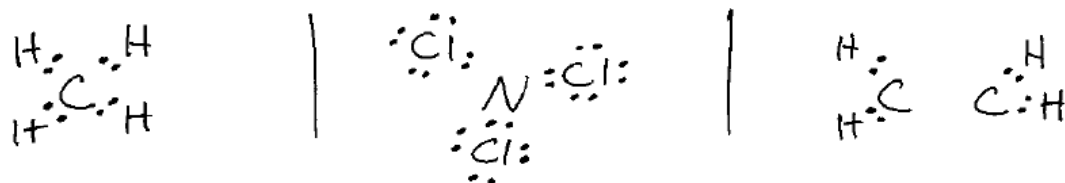
2. Count up all the valence electrons (H=1, C=4, N=5, etc)

$$\text{CH}_4=8$$

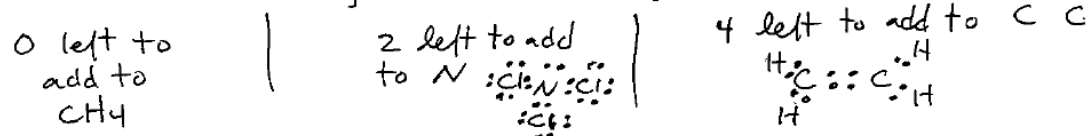
$$\text{NCl}_3=26$$

$$\text{C}_2\text{H}_4=12$$

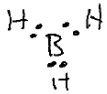
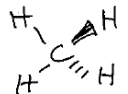
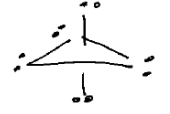
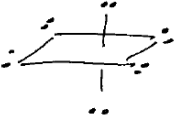
3. Create octets and duplets for perimeter atoms:



4. Fill in the remaining e^- s on the central atoms:



Any other helpful hints? There are a limited number of shapes based upon spreading out e^- pairs surrounding the central atom.

2 e ⁻ rich regions	O=C=O	linear shape
3 e ⁻ rich regions		trigonal shape
4 e ⁻ rich regions		tetrahedral shape
5 e ⁻ rich regions		trigonal bipyramidal
6 e ⁻ rich regions		square bipyramidal (octahedral)

So know that you need to have an open mind about possible expanded sizes to the Lewis structures.

Any other help? Yes, there is a simple calculation involving the number of shared electrons and number of bonding sites that tells you what kind of bonding is going on.

Let $S = N - A$ where $N = \text{needed } e^-$ (either 2 or 8)

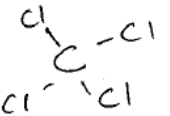
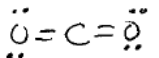
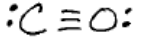
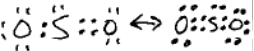
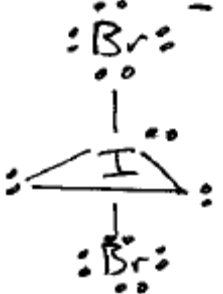
$A = \text{available valence } e^- \text{ s}$

$S = \# \text{ of } e^- \text{ s that need to be shared}$

$B = \text{regions with } e^- \text{ s}$

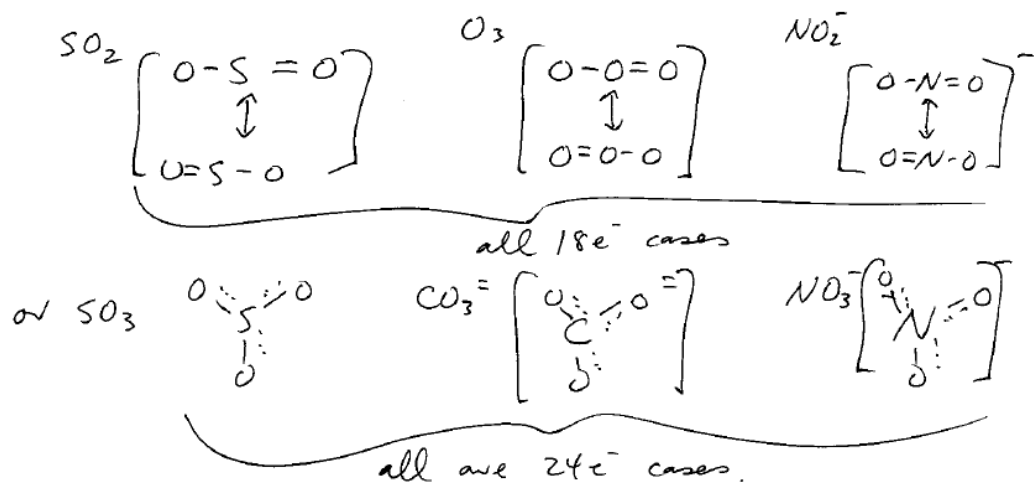
Then $S/2/B = \text{kind of bonding}$


Note this can be very helpful when it works (when the octet rule is satisfied) and really misleading when it doesn't (when you have hypo or hypervalent structures that don't satisfy the octet rule. I personally never use it but then again, I am a chemistry professor. You might find it a good way to get started on learning this material when you are in doubt about resonance or number of double and triple bonds.

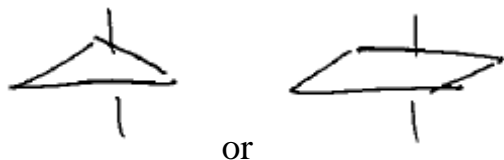
$S/2/B=1$ single bonds  $S=40-32=8$ $S/2/B=8/2/4=1$ Single Bonds around C	$S/2/B=2$ double bonds  $S=24-16=8$ $S/2/B=8/2/2=2$ double bonds around C	$S/2/B=3$ Triple bond  $S=16-10=6$ $S/2/B=6/2/1=3$ triple bond between C and O	$S/2/B=\text{fraction greater than 1}$  $S=24-18=6$ $S/2/B=6/2/2=3/2$ 3 e ⁻ pairs shared between two bonding regions	$S/2/B < 1$  $S=24-22=2$ $S/2/B=2/1/2=1/2$ So need expanded for 5 e ⁻ pairs
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Explaining Resonance

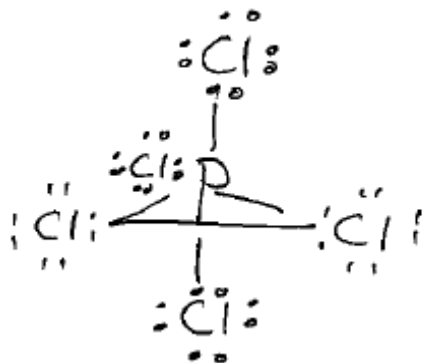
Can you explain resonance? Sure, when we draw certain structures, there are equivalent ways to present them, when this happens, we can draw multiple Lewis structures. The ones we see most often are when 3 e⁻ are placed in two e⁻ regions



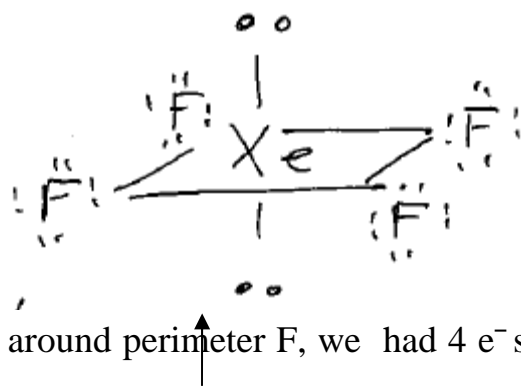
Note that when we get lazy, we draw  instead of all those different structures.
 What about more examples of hypervalent compounds. (Ones with 5 or 6 e⁻ rich regions)
 Sure. Just remember no to panic. The answer is either:



Example, $\text{PCl}_5 = 40 \text{ e}^- \text{ s}$:



*Note: P must accommodate 5 bonds



Example, $\text{XeF}_4 = 36 \text{ e}^- \text{ s}$:

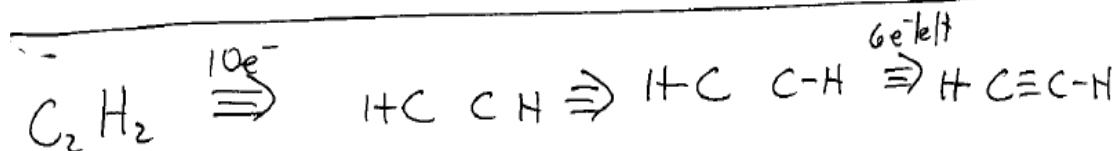
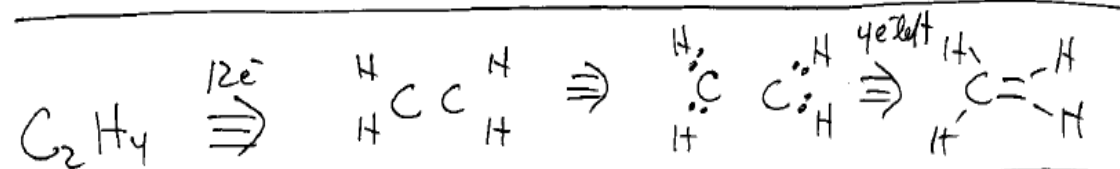
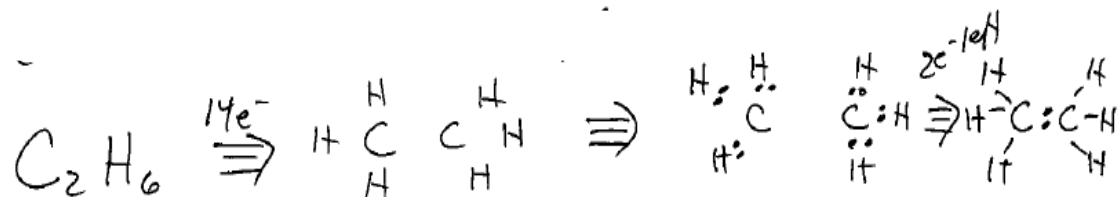
*Note after putting 32 e⁻ s around perimeter F, we had 4 e⁻ s left over to add to 6 e⁻ s rich regions, 4 bonding to Xe and 2 non-bonding.

Multiple Central Atoms

I'm bored by all these single central atoms. What about multiple central atoms like in the real world. How do I do those?
Just like single central atoms:

1. Layout Atoms
2. Count Valence
3. Fill in Perimeter
4. Fill Center

Here are 3 famous examples



Do this for all organics. Almost always you end up putting 8 e⁻ s around C and 2 on hydrogen.

Is the octet + duplet rule always followed?

Nope. There are 3 exceptions.

1. Too many electrons around the central atom, like SF_6 , XeF_4 , IBr_2^- , PCl_5 , etc=happens with $n=3$ and higher situations with d orbitals available.
2. Too few electrons around the central atom like B or Be
3. Odd number of electrons in valence shells means you can't form a
•• electron pair

— empty orbit is an acid

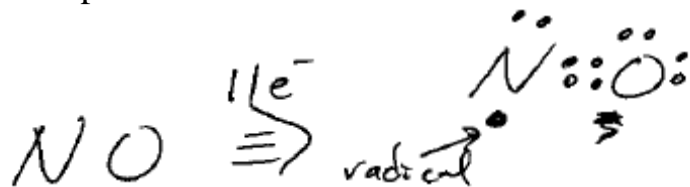
7 one e^- is a radical (causes cancer)

76 is a bond as a nonbonding e^- pair (Lewis base)

Odd case.

So if you ever count an odd number of total valence electrons, you have to form a radical.

Example:



Too few electrons means not enough to form octet.

Example of Lewis atom structures with insufficient e^- s:

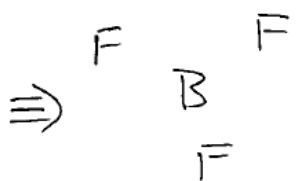
Too few electrons means you don't have enough electrons to form four bonds.

H •
↑
one e⁻
to make one
bond

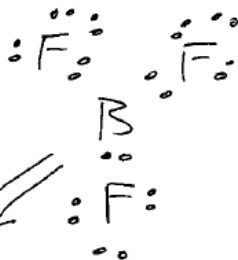
• Be
↑
two e⁻
to make
two bonds

B
↑
3 e⁻ to
make 3 bonds
so can't have octet.

Example
BF₃



24e⁻
⇒



6e⁻
make 3
bonds, no octet

Hint: If you ever see a Boron compound, think octet rule exception.