Chapter 26 THE POLYATOMIC UNIT, Part 1

We have completed the basics of bonding. We are now ready to tackle the polyatomic unit as a whole. For this, we need to develop a bigger picture of Lewis. This can be tedious at first but it gets easier with practice. This is essential to do: not only is the Lewis system so important by itself, but it is also the starting point for other systems which we will use in order to do shapes and to do orbitals in polyatomics. In the last Chapter, we introduced dots, we introduced some terminology and we showed some very simple examples. Now and into the next Chapter, we proceed into more details.

Before we proceed, you should understand that there are varying levels of complexity to the Lewis system and all the details are not covered in a typical first-year course. The level of coverage can vary with different instructors, based on their experience with their students and with the needs for their courses and program. I will first present to you a Basic Level of approach for doing Lewis. Later, in Chapter 27, I will add another level which provides a Refinement Step to the Basic approach, and we will also look at some of the other complexities which can arise. The Basic Level and the Refinement will cover many cases of polyatomics, but they won't be perfect for everything. That's OK. It is not my intention nor my desire to cover everything. Your instructor may also present different steps in order to present her/his slant on the method. That's OK, too.

Before we embark, we return to the importance of eight. I made note of this in the last Chapter.

> ❝ Even when we get into polyatomics, eight will still represent the optimum usage of valence *s* and *p* orbitals although some electrons are shared. The bottom line is this: the number eight is really important. In fact, this gives rise to the "octet rule". The octet rule says that there is a distinct preference for an atom in a Lewis structure <u>to be associated with eight electron dots</u>. Notice the wording: I said there is a distinct preference; I didn't say there's an absolute requirement. ❞

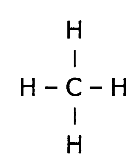
An octet is very favorable but it's not required. In fact, it's not even possible in some cases. At this time, I want to talk a bit about those cases which do not involve an octet. Collectively, they are considered "exceptions to the octet".

26.1 Duets, oddities, deficiencies, and excesses

The most common exception to octet is hydrogen but that's a special case. I'll cover that and then three other cases: odd-electron compounds, electron deficient compounds, and expanded valence.

• HYDROGEN

Hydrogen is NEVER associated with eight electrons in a Lewis structure but there's a good reason. The octet is a result of maximum utilization of valence *s* and *p* orbitals. But H doesn't have *p*! Hydrogen's valence shell is $n = 1$, so there's only 1*s*. With only one valence orbital, H can only handle two electrons max, even when sharing. Consider a simple case: methane, CH_4. The Lewis structure is shown at right. The central C has octet but each H is associated with only two shared electrons. Two is OK for H; that's the way H wants to be. This is sometimes called a "duet" for H. Even though it's only got two, H is actually fulfilling the basis behind octet by fully utilizing its valence orbital. For this reason, we can include H as satisfying the octet rule even when it has two. That's perfectly legit.

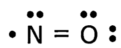


• ODD-ELECTRON COMPOUNDS

In the term "odd-electron compound", odd refers to an odd number of electrons, not how strange the electrons are.

In general, most compounds have an even number of electrons so this category is not real common. I should clarify this statement: our emphasis here lies in compounds composed of Main Group elements and most of these, by far, have an even number of valence electrons. If you get into metal compounds with *d*- or *f*-block elements, then odd-electron things are very common. Examples of Main Group compounds with an odd number of electrons are a bit limited, although they do pop up occasionally. Some of these are part of your world. For example, the simplest one of all of these is toxic but you need it to live.

In an odd-electron compound, one atom will be associated with an odd number of electrons. Typically, this number is seven. One of the most common examples in this category is nitrogen oxide,



NO. Its Lewis structure is shown. Notice that O has a full octet, but N is associated with seven. NO is an interesting compound: deadly but necessary. I mentioned some of this in Chapter 4:

> “ Very harmful. It is also generated in vehicle exhaust and contributes to pollution.
>
> Despite being harmful, NO is another health irony: it is produced in humans and other animals in tiny amounts, and it is absolutely essential for health. ”

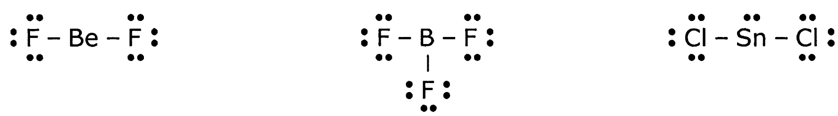
Over recent decades, many biological roles for this little but potent molecule have been realized, ranging from the flash of a firefly all the way up to the regulation of blood pressure in humans.

Since NO has one electron which is not paired, then NO is paramagnetic. We first used the terms paramagnetic and diamagnetic for monatomic things in Chapter 24 and now we use these terms for polyatomics. All odd-electron compounds are paramagnetic. Most even-electron compounds are diamagnetic, although a few are paramagnetic. Let me add a new term at this time: free radical. A free radical is a paramagnetic chemical unit. Sometimes this just goes by radical, without the free part. These terms are quite common.

• ELECTRON DEFICIENCY (EVEN-NUMBERED)

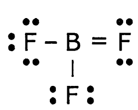
Electron deficient atoms are those with fewer electrons than their valence shell would allow. For most atoms, that means fewer electrons than eight. As described above, hydrogen's valence shell can only handle two, so a hydrogen with two electrons is <u>not</u> electron deficient. Also, an odd-electron atom commonly has fewer than eight electrons but I exclude those cases from this current category because they are in their own category above. That is why I specify even-numbered cases here.

Some elements are more prone to be electron deficient than others. For example, boron is frequently electron deficient. (I said frequently, not always.) Other elements which sometimes do this include beryllium, other elements in Group 13 below boron, and elements in Group 14. Lewis structures for BeF_2, BF_3 and $SnCl_2$ are shown below; all of these are simple covalent molecules in the gas phase. Notice that Be is associated with only four electrons, while B and Sn are associated with six.



Here's a point about terminology. We can refer to the polyatomic unit as a whole as electron deficient or we can refer to the specific atom within it as electron deficient. For example, we can say that BeF_2 is an electron deficient compound or we can say that beryllium in BeF_2 is electron deficient.

In reality, there's actually a way by which we can draw some of the above Lewis structures with octets for everybody. I'll show one example of this for BF_3. By bringing a double bond into the picture,

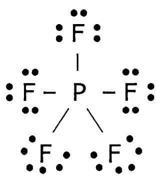


B can achieve an octet. Unfortunately, there are problems with this Lewis structure and the principal Lewis structure is the one shown above with only single bonds. Why? It's a question of balance. It's got to do with "formal charge" but I haven't gotten there yet. It's in Chapter 27. Just forget the Lewis structure at left for now.

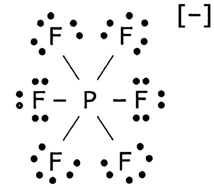
As noted, electron deficiency occurs primarily for some compounds of Be, some of Group 13 and some of Group 14. I will restrict to these cases for this category and I will refer to this grouping of elements as Be/13/14 cases. Furthermore, I will point out that it is rare for carbon to be electron deficient and those instances are energetically-stressful circumstances. I will exclude carbon from this category for our purposes here because this will simplify our steps for Lewis structures. Your instructor may wish to keep C in this category and that's fine but, for the steps presented below, C is excluded from electron deficient cases. For doing Lewis structures, carbon will get octet (unless it has seven, but that's the odd category).

• EXPANDED VALENCE

Historically, this category involved cases where an atom in a Lewis structure was associated with more than eight valence electrons. Besides saying the atom had "expanded valence", the atom was also said to "exceed octet" or the atom was "in excess of octet". The most common numbers beyond eight in a Lewis structure were ten and twelve.



Here are two examples. At left is PF_5. The P atom shows expanded valence, and it is associated with ten electrons. At right, the anion PF_6^-. Again, expanded valence, but now P is associated with twelve electrons.



Before going too far, let me note that the issue of expanded valence can get into troubled water due to various complications which can be involved. This can impact how best to draw a Lewis structure which conveys the best information regarding the distribution of electrons in a particular polyatomic unit. We will adopt a compromise level of coverage here, and I will return to this point near the end of the next Chapter. For now, we will deal with a Basic Level of approach for expanded valence. Furthermore, when we get into orbitals in Chapter 30, we will exclude all coverage of expanded valence. There are simply too many issues involved.

There is one consideration which I can bring in at this time with respect to expanded valence and that has to do with crowding and repulsion. When you have more atoms bonded to a specific atom, and/or more lone pairs associated with that specific atom, then you have more crowding due to the close proximity. This means repulsions and exclusion play a bigger and bigger role. These situations are better accommodated when the atoms are bigger, which means it's better for atoms lower in the Periodic Table. This leads to one very important consequence: we will not give expanded valence to any Second Period element, because they are typically too small to accommodate all of those bonded atoms and lone pairs. Keep this point in mind, since it will help you with this category. Repulsions can have other impacts, too, and we will see more of this in Chapter 28.

OK, the inevitable has become inescapable. It is time to get more fully into Lewis.

26.2 Lewis

We start with our Basic Level, for which there are six Basic Steps. Before going to those Steps, let me preface this with two Notes to be mindful of as we go.

‣ **Note A.** Remember that H takes two. The "octet" for H is satisfied by two.

‣ **Note B.** If you are doing an odd-electron compound, proceed through Step 4 until only one electron is left. Add this last electron to the atom which would next receive a lone pair. This atom will end with seven electrons instead of octet in Steps 5 and 6.

Here are the six **Basic Steps** for now. Just read them at first; I'll explain afterwards.

- **Step 1.** Arrange the atoms according to who's connected to whom.
- **Step 2.** Find the total number of valence electrons in the chemical unit.
- **Step 3.** Draw single bonds between all bonded atoms. If this uses all electrons, then stop here.
- **Step 4.** Distribute the remaining electrons as follows.
  a. Assign electrons as lone pairs on terminal atoms (except H). Do not exceed octet on terminal atoms.
  b. If there are still electrons left, then assign these as lone pairs on the central atom. This can exceed octet.
- **Step 5.** Assess the result so far.
  a. If all atoms have octet or more, stop here.
  b. If the only atom short of octet is a Be/13/14 element, then stop here.
  c. If an atom other than a Be/13/14 element is short of octet, then go to Step 6.
- **Step 6.** For the atom which is short of octet, convert one lone pair from an adjacent atom into a shared bond pair for both atoms. Repeat as needed to achieve octet.

There's a Refinement to this Basic approach in the next Chapter. The six Steps here are good enough to start with.

Before we can execute the Steps, I've got some explaining to do. Most of this involves Step 1, but I must also introduce the terms "central" and "terminal". If you mess up Step 1, the whole thing is kaput. Capisce?

The simple fact is that, when you start to do a Lewis structure, you already need to have an idea of what atoms are connected to what other atoms. In Step 1, you simply sketch this out. You can't sketch it out, however, if you don't know the connections. When I write H_2O, by now you have come to recognize that the molecule has two H's bonded to one O. But if I write Cl_2O or ClO_2 or CH_2Cl_2 or N_2O, do you know who's connected to whom? It's not necessarily obvious. Nonetheless, you must set this up correctly in order to do the Lewis structure.

Frequently, the format of the chemical formula conveys the necessary information for the connections, but that is not always the case and additional information may be required. This is one aspect which different instructors handle differently. Some instructors have their students set up the connections by Group number, some go by electronegativity, some go by combinations of these, etc. Unfortunately, all of those methods have numerous exceptions. I'm just going to stick to the format of the chemical formula when it applies. When it doesn't apply, then I'll say so.

Here's what I mean by this so far. Some formulas show strings of three or more atoms, bonded one after another, wherein all subscripts are ones (which, of course, are not shown). For example, the formula HSCN applies to the molecule with those atoms bonded in that sequence, namely H S C N, so the formula provides the necessary information. Other formulas apply to binary chemical formulas which are of the types AX_n or X_nA. For many of these, the polyatomic unit has a central atom A and terminal atoms X which are bonded to A only and not to other atoms of X. Notice that I just used that term "central" atom again, so let me define it. A central atom is an atom which is bonded to two or more other atoms. WARNING! This definition is based on the number of bonded atoms and it's $\underline{\text{not}}$ based on being physically at the center of something. You must remember this. In fact, you can have more than one central atom in a polyatomic unit. And I also used the term, "terminal" atom. A terminal atom is an atom which is bonded to one other atom only.

Let's apply these terms to some examples which we've already covered in this Chapter. Flip back several pages and look at those examples as we go through here. In CH_4, C is central and all H's are terminal. In BeF_2, Be is central and the F's are terminal. In BF_3, B is central and the F's are terminal. In $SnCl_2$, Sn is central and the Cl's are terminal. In PF_5 and PF_6^-, the P is central and all F's are terminal. Notice that I skipped the example of NO from earlier. Diatomics do not have a central atom; they have two terminal atoms. Except for NO, all of those prior examples follow the AX_n format. For Basic Step 1, if you are given a formula in this format, then you sketch this out accordingly. Unfortunately, this formula interpretation does not hold true in all cases. For example, N_2O (laughing gas) looks like it would be connected as N O N but it's really N N O. So, yes, there are exceptions to this format and I will point out the exceptions when they arise.

For polyatomic units with three different elements and one central atom, I will indicate the central atom by an underline in the formula, when the compound is first given. An example is shown at right. Thus, if there is an underline, then that atom is the only central atom and the other atoms are automatically terminal. I'll follow this practice for several Chapters here to give you some experience with setting these up, but I'll stop using the underline further down the road.

$\underline{C}H_2Cl_2$:

$$\begin{array}{ccc} & H & \\ H & C & Cl \\ & Cl & \end{array}$$

There are many other types of formulas which can be encountered, and more information will be given with those cases. For example, consider C_2H_2, acetylene. This could be represented as a simple string, H C C H, or I can describe it as follows: each H is terminal; both C's are central atoms; and, each C is bonded to the other C and to one H. For more complicated cases, more information will be provided.

One final point: we will limit H to terminal positions. Don't make H a central atom. In reality, H can bond to more than one atom but that's beyond our Lewis coverage here. I'm leaving those out. (So far in this book, there's a picture of one example which has hydrogen as a central atom. Think you can find it? It's in the Aqueous chapters.)

Well, I just spent all that time explaining Step 1. Now let me add a bit about Step 2 and then we can mosey on.

In Step 2, the job is to find the total number of valence electrons in the chemical unit. If the unit is neutral, all you do is add up the numbers of valence electrons for all the atoms. If the unit is a polyatomic ion, then you adjust the total for the ion charge. For polyatomic $\underline{\text{anions}}$, you $\underline{\text{add}}$ the value of the negative charge. For polyatomic $\underline{\text{cations}}$, you $\underline{\text{subtract}}$ the value of the positive charge. Let's do a few examples in order to show how the total number of valence electrons is obtained.

BeF_2	Be (Group 2) brings in two valence electrons:	2
	Each F (Group 17) brings in seven:	14
	The total number of valence electrons is:	16

CH_2Cl_2	C (Group 14) brings in four valence electrons:	4
	Each H (Group 1) brings in one:	2
	Each Cl (Group 17) brings in seven:	14
	The total number of valence electrons is:	20
SO_3^{2-}	S (Group 16) brings in six valence electrons:	6
	Each O (Group 16) brings in six:	18
	The 2– charge adds two more:	2
	The total number of valence electrons is:	26

See if you're getting the hang of it:

NH_4^+	N (Group 15) brings in:	_____
	Each H (Group 1) brings in _____:	_____
	The 1+ charge subtracts one:	–1
	The total number of valence electrons is:	8

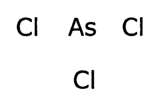
As we go, you'll see how this is applied.

26.3 Examples

Let's work through a number of Examples. Refer back to Basic Steps 1 - 6 as we go. If you tire of these, just take a break. Most of the action is highly repetitive, but I'm trying to point out many of the little details.

Example 1. Do the Lewis structure for arsenic trichloride, $AsCl_3$.

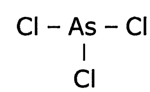
‣ Step 1. Based on the formula, we take As to be central and the three Cl's to be terminal, each bonded to the As. We can sketch this as shown on the right.



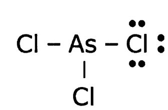
‣ Step 2. Find the total number of valence electrons:

$AsCl_3$	As (Group 15) brings in five valence electrons:	5
	Each Cl (Group 17) brings in seven:	21
	The total number of valence electrons is:	26

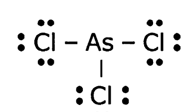
‣ Step 3. We draw single bonds between As and each Cl as shown on the right. Each of the three single bonds involves two electrons, so this Step uses six electrons. We subtract this from the original 26 (Step 2) to get 20 electrons left.



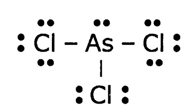
‣ Step 4, Part a. We assign electrons as lone pairs to the terminal Cl atoms, giving each Cl an octet. This means adding six electrons (as three lone pairs) to each Cl. One of the Cl's is shown with its octet at right.



We continue this for the other Cl's until all three chlorines have an octet, as shown on the right. This ends Part a of Step 4. This Part used 18 electrons. We subtract 18 from the 20 left after Step 3 to get 2 electrons remaining. These two electrons carry into Part b.



Part b. We assign the two remaining electrons as a lone pair to the central As.



‣ Step 5. Assess: everyone has octet. Stop here.

Alright, let's see what we've got. The Lewis structure tells us that arsenic trichloride has three AsCl single bonds; the arsenic has one lone pair; and, each Cl has three lone pairs. This is the kind of information obtainable from the Lewis structure.

Next.

. . . .

Example 2. Do the Lewis structure for ethylene, C_2H_4. This does not follow the typical formula format, so I will tell you that both C's are central atoms and each is bonded to two terminal H's and to the other C.

. . . .

‣ Step 1. Based on the description, we set up the arrangement as shown on the right.

```
    H   H

H   C   C   H
```

‣ Step 2. Find the total number of valence electrons:

C_2H_4	Each C (Group 14) brings in four electrons:	8
	Each H (Group 1) brings in one:	4
	The total number of valence electrons is:	12

‣ Step 3. We draw single bonds between each H and its C and between C and C as shown on the right. Each single bond is two electrons, so this Step uses ten electrons. We subtract this from the original 12 electrons (Step 2) to get 2 electrons left.

```
    H   H
    |   |
H – C – C – H
```

‣ Step 4, Part a. We are supposed to assign electrons as lone pairs to the terminal atoms, but the terminals are H's. H's only get two and each H is already associated with two from the bond pair. H's are done. Go to Part b.

Part b. We assign the two remaining electrons as a lone pair to one of the central C's. At this point in the process, it doesn't matter which C gets the two, so I entered them on the left C.

```
    H   H
    |   |
H – C – C – H
    ••
```

‣ Step 5. Assess: H's are OK at two. We've got one C with octet and that's good, but we've got one C short of octet and that's not good. I mentioned upstairs that we won't allow C to be short of octet (unless it has seven) in our coverage. So, we need Step 6.

‣ Step 6. We take the lone pair from the left C and share it with the right C to make a second bond pair. Now both C's have octet and everybody's happy.

```
    H   H
    |   |
H – C = C – H
```

Let's summarize what we've got. The Lewis structure tells us that ethylene has four CH single bonds and one CC double bond. There are no lone pairs in the molecule at all.

Next.

. . . .

Example 3. Hydrogen cyanide, HCN.

. . . .

‣ Step 1. We set up the arrangement as shown on the right.

H C N

‣ Step 2. Find the total number of valence electrons:

HCN	H brings in one valence electron:	1
	C brings in four:	4
	N brings in five:	5
	The total number of valence electrons is:	10

‣ Step 3. We draw one single bond between H and C and one between C and N. Each single bond is two electrons, so this Step uses four electrons. We subtract these 4 electrons from the original 10 (Step 2) to get 6 electrons left.

H – C – N

‣ Step 4, Part a. We now assign electrons as lone pairs to the terminal atoms. The terminal H has two and gets no more. We turn to the N, and assign all 6 remaining electrons as lone pairs.

```
        ••
H – C – N :
        ••
```

Part b. There are no unassigned electrons remaining, so there's nothing to do here. Go to Step 5.

‣ Step 5. Assess: H is fine and N has octet, but our central C is seriously short of octet. Presently, it is associated with only four. We need to fix that; go to Step 6.

‣ Step 6. We take one lone pair from N and share it with the C to make a second bond pair, as shown on the right. This is better since C is now associated with six, but it's still not octet and it's still not right. (By the way, notice that N keeps an octet in this process.)

H – C = N:

Repeat this procedure: take another lone pair from N and share it with C to make a third bond pair. This gets us to where we need to be. Now, C and N have octet.

H – C ≡ N:

There you have it. We see from the Lewis structure that hydrogen cyanide has one CH single bond and one CN triple bond. In addition, there is one lone pair on N.

Let's keep plugging away here. What I want to do now is show you an electron deficient case. I'll use one which we highlighted earlier, BF_3, and show how the Lewis structure is derived from our six Basic Steps.

Example 4. Boron trifluoride, BF_3.

‣ Step 1. As suggested by the formula, B is central and the F's are terminal. This gives the setup on the right.

F B F

F

‣ Step 2. Find the total number of valence electrons:

BF_3	B brings in three valence electrons:	3
	Each F brings in seven:	21
	The total number of valence electrons is:	24

‣ Step 3. Draw a single bond between B and each F. Each single bond is two electrons, so this Step uses six electrons. We subtract these 6 electrons from the original 24 (Step 2) to get 18 electrons left.

F – B – F
|
F

‣ Step 4, Part a. Start assigning electrons as lone pairs to the F's. Each F will take three LPs (six electrons). This will use up the 18 electrons left from Step 3.

Part b. There are no unassigned electrons left, so there's nothing to do here. Go to Step 5.

:F – B – F:
|
:F:

‣ Step 5. Assess: F's have octets; the central B is short of octet but that is OK since it is one of the Be/13/14 elements. Stop here.

This is exactly the same Lewis structure presented earlier in the Chapter. We can describe the molecule as having three BF single bonds and each F in the molecule has three lone pairs. The boron is electron deficient.

Your turn.

Example 5. CH_2Cl_2.

‣ Step 1. I showed the setup earlier for this compound. If you need to go back and look, then do so.

‣ Step 2. Find the total number of valence electrons. We did this earlier also. Do it below first and then go back and check your answer.

CH_2Cl_2 C brings in four electrons: ______

Each H brings in one electron: ______

Each Cl brings in seven electrons: ______

The total number of valence electrons is: ______

‣ Step 3. Draw single bonds between C and each H and between C and each Cl. Each single bond is two electrons, so this Step uses _______ electrons.

We subtract these electrons from the original total of ______ (Step 2) to get ______ electrons left.

‣ Step 4, Part a. Assign electrons as lone pairs to the terminal atoms. Not to H! Do the Cl's. If you did everything OK up to here, this should use up all the electrons left from Step 3.

Part b. There's nothing to do here. Go to Step 5.

‣ Step 5. What do you think? Are H's OK? Are Cl's OK? Is C OK? Do you stop here? Do you proceed to Step 6? Do you pass GO and collect $200?

Let's do something odd.

Example 6. NF_2.

‣ Step 1. Based on the formula, we take N to be central and the two F's to be terminal. We sketch this out as shown on the right.

F N F

‣ Step 2. Find the total number of valence electrons:

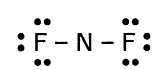
NF_2	N:	5
	F's:	14
	The total number of valence electrons is:	19

Right off the bat, be aware that you have an odd number. We will need Note B which was given with the Basic Steps.

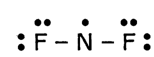
‣ Step 3. Draw a single bond between N and each F as shown on the right. This Step uses four electrons. We subtract these from the 19 to get 15 electrons left.

F – N – F

‣ Step 4, Part a. Assign electrons as lone pairs to the terminal F atoms, giving each an octet. This means adding three lone pairs to each F, as shown on the right. This ends Part a of Step 4. This Part used 12 electrons. We subtract 12 from the 15 (Step 3) to get three electrons remaining. These three electrons carry into Part b.



Part b. We assign a lone pair to the central N. That leaves us with the final odd electron, which we also assign to the N.



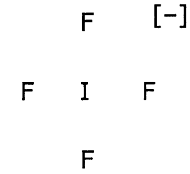
‣ Step 5. The fluorines have octet. Nitrogen has seven; as given by Note B, that's where we leave it. We're done.

In summary, we see that NF_2 has two NF single bonds. Each F has three lone pairs; the N has one LP and one unpaired electron. NF_2 is paramagnetic.

That ends our odd case. Next, let's do an example of expanded valence.

Example 7. IF_4^-.

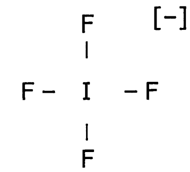
‣ Step 1. The I is central and the F's are terminal.



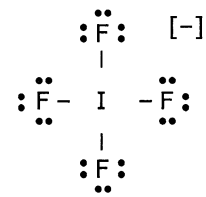
‣ Step 2. Find the total number of valence electrons:

IF_4^-	I:	7
	F's:	28
	Add for anion charge:	1
	The total number of valence electrons is:	36

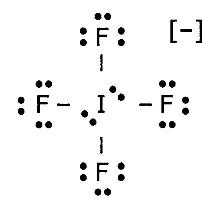
‣ Step 3. Draw a single bond between I and each F. This Step uses eight electrons. We subtract this from the 36 to get 28 electrons left.



‣ Step 4, Part a. Assign electrons as lone pairs to the terminal F's. Each F will take three lone pairs. This Part uses 24 electrons. We subtract 24 from the 28 following Step 3 to get four electrons remaining.



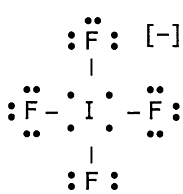
Part b. We assign the four remaining electrons to the central iodine as two lone pairs. The I is now assigned twelve electrons, so it has expanded valence.



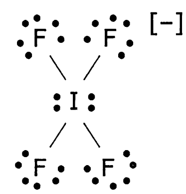
‣ Step 5. Stop here.

In summary, the Lewis structure displays four IF single bonds, three lone pairs on each F, and two lone pairs on the iodine.

This would be a good time to make a little cautionary note about lone pairs. When you do a lone pair in a Lewis structure, be sure to convey the electrons as paired. Don't break them apart. For example, take a look at the Lewis structure on the left. It's wrong. Why? The lone pairs on the central I are not correctly portrayed as paired. Who cares? The notion of paired and unpaired electrons is important, so this should be properly conveyed in the Lewis structure. Except for this detail, it doesn't really matter how you draw all the bond pairs and lone pairs for atoms. For example, another Lewis structure for IF_4^- is on the right. Compared to the first one above, this one just moves the bonds and LPs around on the central atom. That's OK.



WRONG



RIGHT

Next Example: it's your turn again.

. . . .

Example 8. H_2NOH. This molecule has two central atoms, so let me describe the setup. Both the N and the O are central and are bonded to each other. N also has two terminal H's while O has one terminal H.

. . . .

‣ Step 1. Based on the description, set up the arrangement as shown on the right.

H N O H

  H

‣ Step 2. Find the total number of valence electrons:

H_2NOH	H's:	______
	N:	______
	O:	______
	The total number of valence electrons is:	______

‣ Step 3. Draw a single bond between N and its two H's, between O and its H, and between N and O. This Step uses ______ electrons. We subtract this number from the original total of ______ to get ______ electrons left.

‣ Step 4, Part a. Do you do anything here?

Part b. Now assign the remaining electrons as lone pair(s) to central atoms.

‣ Step 5. Now what?

‣ Step 6. Do you need Step 6?

Put your summary here.

Bonds	NH:	How many?	______	What bond order?	______
	OH:	How many?	______	What bond order?	______
	NO:	How many?	______	What bond order?	______
LPs	on N:	How many?	______		
	on O:	How many?	______		

Done?

OK, we've completed eight examples illustrating various aspects of actually doing Lewis structures. Hopefully, you are catching on to this. Hopefully, you can also see some of the importance of the Lewis structure in terms of the information which it provides about bond orders and about lone pairs for all of the atoms in the polyatomic unit. I present this schematically as shown at right. At this stage of the game, there's really not much to this but I am going to develop this further as we go, all the way through Chapter 31. We'll call this Stage 1 for now, and we will eventually take this through Stage 4. There is more yet to do and more yet to be gleaned.

lone pairs

↖

bond pairs (order) ← Lewis structure

But before you go there, you must be able to get here. Go over the examples above. Be able to do them with one hemisphere tied behind your back. Then go on.

Problems

1. True or false.
   a. AlH_3 is electron deficient.
   b. NF_4^+ displays expanded valence.

c. O_3 is an odd-electron molecule.

d. In every Lewis structure, every atom except H must have octet.

2. Do the Lewis structure for each of the following.

a. $Cl\underline{N}H_2$ b. C_2H_2 (HCCH) c. H_2S d. NO_2^+ e. $F_3\underline{C}O^-$

For each, how many bonds are there between what atoms and of what order? How many lone pairs are there on what atoms?

3. Do the Lewis structure for each of the following.

a. $H_2\underline{C}O$ b. PH_2 c. BO_3^{3-} d. SiF_6^{2-} e. CH_3CN:

$$\begin{array}{cccc} & H & & \\ H & C & C & N \\ & H & & \end{array}$$

Which of these (if any) have an odd number of electrons? Which of these (if any) are electron deficient? Which of these (if any) have an atom which displays expanded valence?