Lewis Dot Structure Examples: Following the Rules Step-By-Step

Example 1: H_2O

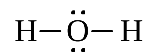
Step 1: Arrange the atoms: H O H

Step 2: Count up total valence electrons: 2(1) + 6 = 8 e^- total

Step 3: Draw single bonds between central atom and surrounding atoms: H – O – H

Step 4: Place remaining electrons, in pairs, around appropriate atoms; start with outer atoms.

8 – 4 = 4 e^- left; therefore 2 lone pairs to add:

In this example, the 2 lone pairs must go around O, since H never gets lone pairs→ 

Step 5: Make sure all atoms that need octets have octets: done in Step 4

Example 2: BCl_3

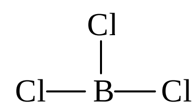
Step 1: Arrange the atoms:

Cl

Cl B Cl

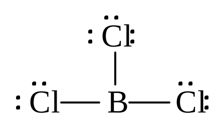
Step 2: Count up total valence electrons: 3(7) + 3 = 24 e^- total

Step 3: Draw single bonds between central atom and surrounding atoms:



Step 4: Place remaining electrons, in pairs, around appropriate atoms; start with outer atoms.

24 – 6 = 18 e^- left; therefore 9 lone pairs to add:



Step 5: Make sure all atoms that need octets have octets: The Cl's have octets, so they're okay; **B is special element that can have an incomplete octet, so B only needs 6.** We're done here.

Example 3: PF_3

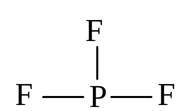
Step 1: Arrange the atoms:

F

F P F

Step 2: Count up total valence electrons: 3(7) + 5 = 26 e^- total

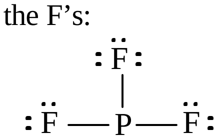
Step 3: Draw single bonds between central atom and surrounding atoms:



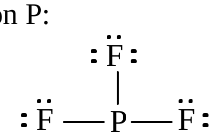
Step 4: Place remaining electrons, in pairs, around appropriate atoms; start with outer atoms.

26 – 6 = 20 e^- left; therefore 10 lone pairs to add:

18 e^- (9 pairs) go around the F's:



The remaining 2 e^- (1 pair) goes on P:



Step 5: Make sure all atoms that need octets have octets: finished in Step 4.

Example 4: CO_2

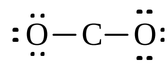
Step 1: Arrange the atoms: O C O

Step 2: Count up total valence electrons: 2(6) + 4 = 16 e^- total

Step 3: Draw single bonds between central atom and surrounding atoms: O – C – O

Step 4: Place remaining electrons, in pairs, around appropriate atoms; start with outer atoms.

16 – 4 = 12 e^- left; therefore 6 lone pairs to add:

Putting electrons on the outer atoms first gives 

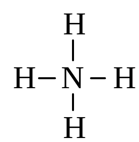
Step 5: Make sure all atoms that need octets have octets: After Step 4, both O's are happy, but C is miserable. So, move a lone pair from each O to C, making 2 double bonds to C. Remember that shared pairs count for both atoms. Done!

Example 5: NH_4^+

Step 1: Arrange the atoms:

```
     H
H    N    H
     H
```

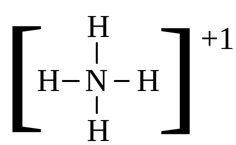
Step 2: Count up total valence electrons: 4(1) + 5 – 1 = 8 e^- total ; **Note that we subtract one electron from the total because NH_4^+ has a +1 charge**

Step 3: Draw single bonds between central atom and surrounding atoms: 

Step 4: Place remaining electrons, in pairs, around appropriate atoms; start with outer atoms.

8 – 8 = 0 e^- left

Step 5: Make sure all atoms that need octets have octets: We're good here, but the structure is not complete until we show the charge on the ion:



Example 3: NO_2^{-1}

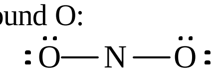
Step 1: Arrange the atoms: O N O

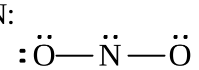
Step 2: Count up total valence electrons: 2(6) + 5 + 1 = 18 e^- total; **Note that we add one electron to the total because NO_2^{-1} has a - 1 charge**

Step 3: Draw single bonds between central atom and surrounding atoms: O – N – O

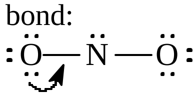
Step 4: Place remaining electrons, in pairs, around appropriate atoms; start with outer atoms.

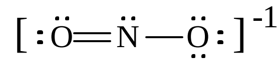
18 – 4 = 14 e^- left; therefore 7 lone pairs to add:

The first 12 e^- (6 pairs) go around O: 

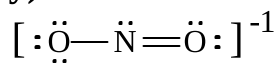
The last 2 e^- (1 pair) must go on N: 

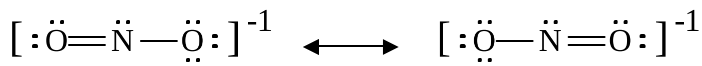
Step 5: Make sure all atoms that need octets have octets: After Step 4, both O's are happy, but N is short 2 electrons. To make N happy, move one lone pair from one of the O's to make one NO double bond:



The correct structure shows the charge of the ion: 

Resonance (super important, especially in organic chemistry)

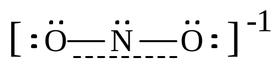
In the above example, we could have moved a lone pair from the other O to get 

The two structures are equivalent: 

These two structures are examples of so-called **equivalent resonance structures**

The Lewis dot structure almost suggests that the double bond "flips" from one O to the other, in an "on-off" kind of way; **this is absolutely NOT the case!**

Rather, the bond that appears to "flip" between oxygens is actually "smeared out" over the 2 N-O bond positions. The electrons in the bond are **delocalized** throughout the molecule. In a Lewis dot structure, you can show this "smearing", or delocalization, as follows:



Because the extra bond is smeared out, you can think of it as being a part of both N-O bonds at once.

*** Each N-O bond is neither a single bond nor a double bond. In the NO_2^{-1}, 3 bonds occupy 2 positions, so the N-O bond is basically a (3/2) = 1.5 bond. ****

Later in general chemistry, you'll see a more elegant description of resonance delocalization when we get to valence bond theory. When you get to organic chemistry, you'll see over and over again how various reactive intermediates are stabilized by resonance.

Expanded Octets

Fact: nonmetal elements in Periods 3 and higher can have more than 8 shared or lone electrons (or both). Why? These elements have empty d orbitals that can hold the extra electrons.

Example 7: I_3^{-1}

Step 1: Arrange the atoms: I I I

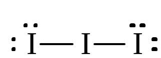
Step 2: Count up total valence electrons: 3(7) + 1 = 22 e^- total

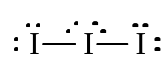
Step 3: Draw single bonds between central atom and surrounding atoms: I – I – I

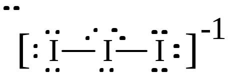
Step 4: Place remaining electrons, in pairs, around appropriate atoms; start with outer atoms.

22 – 4 = 18 e^- left; therefore 9 lone pairs to add:

The first 12 e^- (6 pairs) go around the outer I's: The last 6 e^- (3 pair) go around the central I:

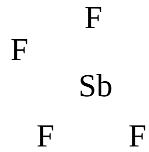




Step 5: The central I has an expanded octet. Show the charge on the ion: 

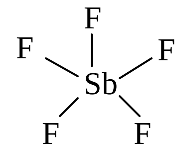
***** Note that "extra" lone pairs will always go on the central atom *****

Example 8: SbF_5



Step 1: Arrange the atoms: **Note that Sb can be surrounded by more than 4 atoms**

Step 2: Count up total valence electrons: 5(7) + 5 = 40 e^- total

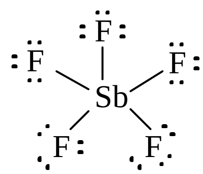


Step 3: Draw single bonds between central atom and surrounding atoms:

Step 4: Place remaining electrons, in pairs, around appropriate atoms; start with outer atoms.

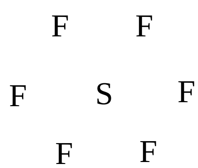
40 – 10 = 30 e^- left; therefore 15 lone pairs to add:

First, give the outer F's octets:



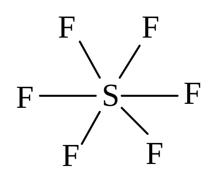
Step 5: Make sure all atoms that need octets have octets: finished in Step 4.

Example 8: SbF_5



Step 1: Arrange the atoms: **Note that S can be surrounded by more than 4 atoms**

Step 2: Count up total valence electrons: 6(7) + 6 = 48 e^- total

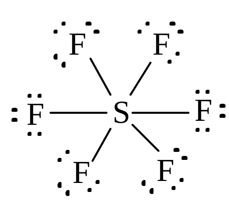


Step 3: Draw single bonds between central atom and surrounding atoms:

Step 4: Place remaining electrons, in pairs, around appropriate atoms; start with outer atoms.

48 – 12 = 36 e^- left; therefore 18 lone pairs to add:

First, give the outer F's octets:

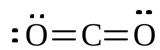
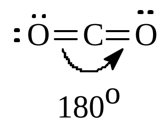
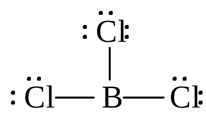
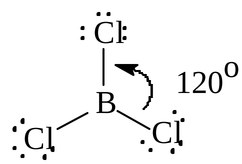
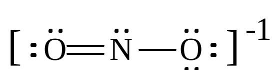
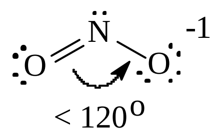
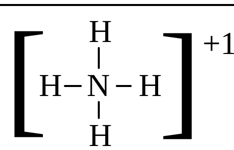
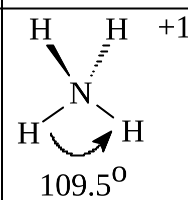
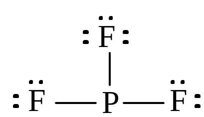
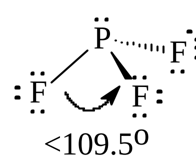
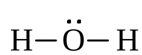
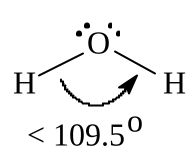


Step 5: Make sure all atoms that need octets have octets: finished in Step 4.

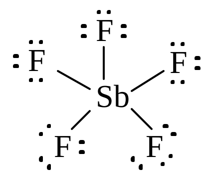
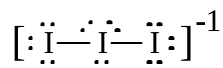
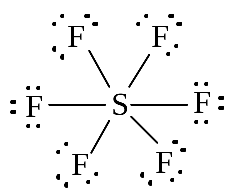
Molecular Shapes by the VSEPR/MXE Method

M – the atom around which you want the shape
X – another <u>atom</u> (or group of atoms) bonded to M
E – a lone pair on M

****** The basic idea of VSEPR theory is that the molecule will always take a shape that keeps the X's and the E's as far apart from each other as possible. ************

Designation	**X + E**	**"Base" geometry (shape) that keeps X's and E's as far apart as possible**	**Actual molecular geometry (i.e. the shape made by the X's only)**	**Example**	**Geometric drawing showing bond angle**
MX_2	2	linear	linear		
MX_3	3	trigonal planar	trigonal planar		
MX_2E	3	trigonal planar	Bent or angular		 **the lone pair on N is bulky, so it "pushes" the bond angle down**
MX_4	4	tetrahedral	tetrahedral		 **a "wedge and dash" structure**
MX_3E	4	tetrahedral	trigonal pyramidal		
MX_2E_2	4	tetrahedral	bent or angular		

The VSEPR/MXE Method for Expanded Octet Molecules and Ions

Designation	**X + E**	**"Base" geometry (shape) that keeps X's and E's as far apart as possible**	**Actual molecular geometry (i.e. the shape made by the X's only)**	**Example**	**Geometric drawing showing bond angle**
MX_5	5	trigonal bipyramidal	trigonal bipyramidal		
MX_2E_3	5	trigonal bipyramidal	linear		
MX_6	6	octahedral	octahedral		
MX_4E_2	6	octahedral	square planar