Chapter 7: How to Draw Lewis Dot Structures

A few guidelines to the uses and restrictions of the Lewis Dot Structure:

- The Lewis structure is produced in 2-dimensions on the plane of a piece of paper.

- The 3-dimensional structure of a molecule is not drawn (wait for VSEPR in Chapter 8.)

- Connectivity, bond order (double and triple bonds) and resonance between atoms is demonstrated.

- The distribution of valence electrons around the atoms is generated.

- The octet rule is commonly applied, which assumes elements achieve an s^2p^6 electronic configuration.

- Exceptions to the octet rule are Be and B, odd number of electrons, and when more than 4 bonds are placed on the central atom. This can occur for atoms in the n=3 shell and larger.

The procedure for Lewis dot structures that satisfy the octet rule:

Step 1: Arrange the atoms on the page to achieve a high degree of symmetry. Surround the more electropositive central atom with electronegative atoms.

Step 2: Determine the following parameters for bonding electrons:

- **B:** The number of bonding sites (regions between atoms that share electrons)
- **N:** The number of electrons needed by each atom (usually 8 except for H, which needs 2; Be, which needs 4; B, which needs 6, and larger atoms which may need 10 or 12.)
- **A:** The number of electrons actually available (the number of valence electrons for each atom)
- **S:** The number of shared electrons (S = N – A)
- Calculate **(S/2) / B** and observe the following interesting features of the bonds:

  If (S/2) / B = 1 then only **single** bonds are present in the molecule
  If (S/2) / B > 1 then **multiple** bonds are present
  If (S/2) / B is not an integer then **resonance** occurs in molecule
  If (S/2) /B < 1 indicates that 5 or 6 electron rich regions are involved (see Chapter 8)

Step 3: Evenly distribute the shared electrons, S, among the bonding sites, B, with a minimum of 2 electrons per site.

Step 4: Distribute the remaining non-bonding electrons trying to achieve the octet rule for non-H peripheral atoms

Examples

An example with a triple bond (S/2)/B = 3

N_2 Step 1: **N N**

Step 2:

B = 1

N = 16 = 8 for each N atom

A = 10 = 5 for each N atom

S = N – A = 6 shared (bonding) electrons

(S/2) / B = 3 so have triple bond between nitrogens

Step 3: **N : : : N**

Step 4: **: N : : : N :**

An example with a single bond and perimeter H atoms

NH_3 Step 1:

H N H
H

Step 2:

B = 3 (three sites for bonds between N and H)
N = 14 (8 electrons needed for N atom and 2 electrons needed for each of 3 H)
A = 8 (5 electrons available for N atom and one electron for each of 3 H)
S = N – A = 6 shared (bonding) electrons
(S/2) / B = 1 therefore single bonds between atoms

Step 3:

H ·· N ·· H
:
H

Step 4:

··
H – N – H
|
H

An example with resonance (S/2)/B not an integer

$CO_3^=$

Step 1:

O C O

O

Step 2:

B = 3

N = 32

A = 24

S = N – A = 8 shared (bonding) electrons

(S/2)/B = 4/3 so is greater than one and have resonance

Step 3:

O :: C : O

|

O

Step 4: three equivalent resonance structures I can't draw here

An example of that does not satisfy the octet rule:

Perform the same procedure as above, but recognize that the S = N-A calculation doesn't work. Instead, create single bonds, fill all the outer atoms to satisfy the octet rule, and add the remaining electrons to the larger central atom. It can now hold 10 and even 12 electrons because of the availability of d and f orbitals.

$XeCl_2$ Step 1: **Cl Xe Cl**

Step 2:

B = 2
A = 22 (7 valence electrons for 2 Cl and 8 valence electrons for Xe)
S = N – A = 2 shared (bonding) electrons
(S/2)/B = 0.5   See, it makes no sense.

Step 3. Form two single bonds using the first 4 four of the 22 available electrons).

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Cl  : Xe : Cl
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Step 4a. Satisfy the octet rule for peripheral Cl (the next 12 of the 22 electrons).

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 . .        . .
:Cl  : Xe : Cl:
 . .        . .
```

Step 4b. Add the last 6 of the 22 available electrons to the central Xe atom.

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. .  . . . .  . .
:Cl  : Xe : Cl:
 . .   . .    . .
```