

Lewis Structures

Lewis structures can be drawn for:

Ionic Compounds
Molecular Compounds
Polyatomic Compounds

Ionic Compounds

Contains metallic and non-metallic ions.

Difference in electronegativity is greater than 1.7.

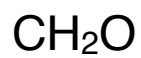
Sodium Chloride NaCl

Aluminum Fluoride AlF₃

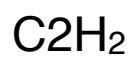
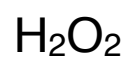
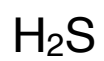
Molecular Compounds

Contains two or more non-metals held together by covalent bonds.

Difference in electronegativity is equal to or less than 1.7.



Draw Lewis structures for:



Co-ordinate Covalent Bonds

A covalent bond involves the sharing of electrons between two atoms, each atom contributing an electron into the pair.

In some cases (example hydronium H_3O^{1+}), one atom contributes both of the electrons to the shared pair.

In cases such as this, the covalent bond is referred to as a ***co-ordinate covalent bond***.

According to the quantum mechanical model of the atom, a co-ordinate covalent bond forms when a filled atomic orbital overlaps an empty atomic orbital. Once formed, a co-ordinate covalent bond behaves in the same manner as every other covalent bond.

Drawing Lewis Structures with co-ordinate covalent bonds

The following steps can be used to help sketch a Lewis Structure for a molecular or a polyatomic ion containing co-ordinate covalent bonds

- 1) Position the least electronegative atom at the centre of the molecule or polyatomic ion. Position the remaining atoms around the central atom, with each atom bonded to the central atom by a single bond. *Always position a hydrogen or halogen at the end
- 2) Determine the total number of valence electrons available. *Be sure to take into consideration that charge of the polyatomic ion.
- 3) Determine the total number of valence electrons required for each atom to have the electron configuration of an inert noble gas.
- 4) Determine the number of valence electrons that must be shared. Subtract the number of valence electrons available from the number of valence electrons required. Determine the number of bonding pairs by dividing by two.

- 5) Determine the number of unbound electron by subtracting the number of valence electrons shared from the number of valence electrons available. Dividing by two will tell you the number of lone or non-bonding electron pairs.
- 6) Arrange the bonding and lone pairs around the atoms present so that each has an electron configuration of an inert noble gas.

Example: ClO_3^{1-}

Example: CO_3^{2-}

Expected Observation: We expect the analysis of this compound to show 2 different bond lengths for the carbon and oxygen bonds, one C=O and one C-O.

Analysis showed only one kind of carbon oxygen bond with a bond length that is longer than a C=O bond but shorter than a C-O.

In this case **resonance structures** exist. In resonance structures the multiple bond constantly shifts from one position to another.

Some Lewis Diagram Exceptions

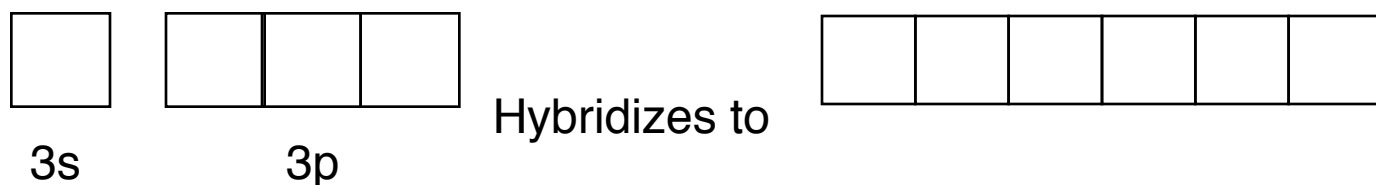
Expanded Octets

Some elements in the periodic table are able to form up to six covalent bonds. This is considered to be an **expanded valence**, where the valence energy level of the central atom has more than eight electrons.

This is achieved through **hybridization** of orbitals. Two or more orbitals combine to form a new orbital.

Example: Let's explain SF₆

Sulfur [Ne]3s²3p⁴



All s and p electrons are involved in bonding, so the central atom has no lone pairs.

Hints you may be dealing with an expanded valence:

- A compound has more than 4 of one type of atom (ex PF_5)
- A halogen has formed more than one bond (ex IF_4)
- A noble gas is involved in bonding

To draw molecules involving expanding valence

- 1) Identify the central atom (lowest electronegativity)
- 2) Draw the skeleton structure (all other atoms are bonded to central atom, assume normal number of bonds)
ONLY CENTRAL ATOM HAS EXPANDED VALENCE!
- 3) Determine the total number of valence electrons.
- 4) Determine the number of bonded electrons.
- 5) Determine number of lone pairs by subtracting the number of bonded electrons from the number of electrons we have.
- 6) Distribute lone pairs. Any leftover lone pairs appear around the central atom.
- 7) Use square brackets to indicate an ion if necessary.

Example: ICl_4^{-1}

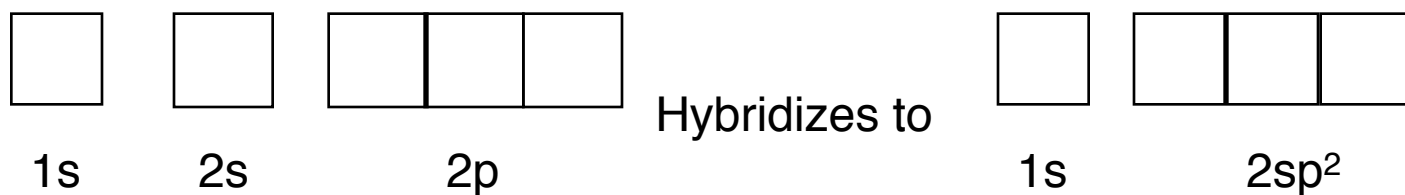
Incomplete Octets

The metals boron and beryllium are able to form unique covalent bonds. When this happens it is considered to be an ***incomplete valence***, where the valence energy level of the central atom has less than eight electrons.

This is achieved through ***hybridization*** of orbitals. Two or more orbitals combine to form a new orbital.

Example: Let's explain BCl_3

Boron $1s^2 2s^2 2p^1$



Try these for homework:

