

BRIEF SUMMARY OF LEWIS STRUCTURES

Ionic bonds are the result of the **transfer** (giving, taking) of one or more electrons from one atom (or groups of atoms) to another atom (or group). The resultant difference in charge between the ions so formed is responsible for the bonding forces in ionic materials. There is no bond as such, just a net electrostatic attraction between the ions.

Covalent bonds result from atoms **sharing** (loaning, borrowing) of electrons. The electrons become common to the atoms involved, and are usually located between the two atoms involved. This is thus a true bond.

The properties of ionic and covalent compounds are quite different but the underlying driving force to form the bonds is the same. All atoms involved are trying to fill their outer orbital (valence shell); the octet rule.

The rest of this summary pertains to covalent bonding only.

A pair of shared electrons constitutes a **bond**, two pairs, a double bond and three pairs a triple bond. There are no quadruple bonds or higher. Any pairs of electrons left over in the valence shell, not used in bonding, are called **lone pairs**.

Some atoms do not have enough electrons to share and need to borrow more electrons than they give (e.g. BF_3). This sometimes results in double bonds, but in other cases the octet rule is violated resulting in an atom with too few electrons. These are called **electron deficient** compounds. In the case of BF_3 , the introduction of a double bond would result in a positive *formal charge* on the fluorine, which is highly unfavorable.

Note that hydrogen only needs one electron to fill its outer shell so it does not obey the octet rule. One consequence of this is that hydrogen is always terminal... at the end, it never has more than one other atom bonded to it.

For elements in the third row and beyond *d*-orbitals become important. If the *d*-orbital is full (10 electrons), the octet rule can be applied as usual. However, if the *d*-orbital is part filled or empty, the octet rule may again be violated. Extra electrons can be loaded into the *d*-orbital allowing the central atom to bond more atoms than would be allowed if the octet rule was obeyed (review examples given). This phenomena is called **valence shell expansion**.

Beyond the second row, the Lewis theory is not useful for predicting stoichiometry, only rationalizing it. It is still useful for predicting structure though (in conjunction with VESPR).

Many of the ambiguities in structures may be resolved using *formal charges*.

Method of Drawing Structures e.g. for **HNO₂** nitrous acid (an oxy acid)

1. Place the least electronegative atom in the middle. Join the remaining atoms with single bonds, remembering that an oxyacid has an –O-H group.
2. Sum the valence electrons (including or excluding those involved in ionic charge).
e.g. **H** 1, **O** 6, **N** 5, **O** 6 = 18 valence electrons
3. Count the electrons already committed to forming the single bond skeleton.
i.e. in this case there are 3 single bonds or 6 bonding electrons
4. Count the number of electrons that would be required to fulfill the octet rule.
i.e., in this case 14 electrons, but only 12 electrons are available (18–6)
5. Add a multiple bond for each pair of electrons that are deficient. In this case, we have one multiple bond.
6. On each atom, make up the octet of electrons with lone pairs. If more than 1 structure is possible (i.e., *resonance structures*), assign *formal charges* to narrow down the options.

Try the following examples:

- (a) species with only single bonds: **NH₃** **SiF₄** **N₂F₄** **PH₄⁺**
- (b) species with multiple bonds: **SO₂** **SO₃** **HNO₃** **SO₄²⁻** **C₂H₂** **O₃** **N₂O**
- (c) species with less than an octet of electrons: **BH₃** **BeF₂**
- (d) species with an odd number of electrons: **ClO₂** **NO₂**
- (e) species with more than an octet of electrons: **PCl₅** **XeF₄** **GeCl₅⁻** **TeF₅⁺**

Hints for Drawing Lewis Structures

1. Consider only valence electrons (neglect the inner electrons).
2. **H** atoms only form single bonds **H-**
3. **C, N, O, F** nearly always obey the octet rule.

C always has 4 bonds (exception **CO**)

e.g.

N can have 3 or 4 bonds

e.g.

O has 2 or 1 bond

e.g.

F has 1 bond

4. **O-O** bonds are rare, e.g. **H₂O₂** hydrogen peroxide **H-O-O-H**
5. In oxyacids, the **H** atom is normally joined to an **O** atom.
6. Cyclic structures are rare.
7. A few species have less than an octet of electrons.
e.g. **BeCl₂** and **BF₃**
8. Some species have more than an octet of electrons (*valence shell expansion*).
e.g. **BrF₅** and **IF₇** and **SiCl₅²⁻**
9. Molecules with an odd number of electrons (*free radicals*) cannot obey the octet rule.
e.g. **N=O** and **ClO₂**