Chemistry: Form WS4.1.6A　　　　Name ______________________________

BONDING　　　　Date ________________ Period _____

Writing Lewis Structures

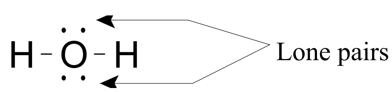
Lewis structures show how valence electrons are arranged among atoms in a compound using dots to represent the valence electrons that are not shared in a covalent bond. To draw Lewis structures showing covalent bonds for elements in periods 1 and 2, do the following: [1] sum the valence electrons; [2] use a pair of electrons to form a bond (bonding electrons); and [3] arrange the remaining electrons to satisfy the duet rule for hydrogen and the octet rule for other elements (lone pairs).

<u>Example</u>: H_2O

Step 1: sum the valence electrons
H =1, O = 6　　1 + 1 + 6 = 8

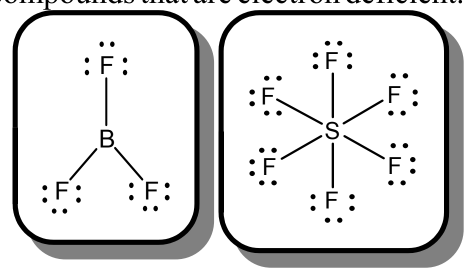
Step 2: use a pair of electrons per bond (each represented as a line)
H–O–H

Step 3: distribute the remaining electrons



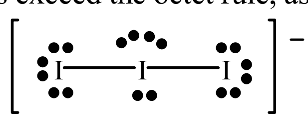
Step 4: check to see that the duet rule is satisfied for hydrogen and the octet rule is satisfied for all other atoms

There are exceptions to the octet rule. Boron and beryllium tend to form compounds that are electron deficient.

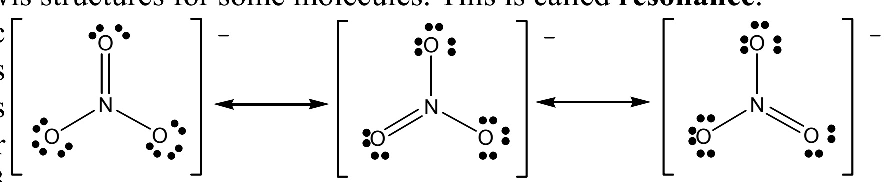


Elements in period 3 and beyond sometimes exceed the octet rule by using space in *d* orbitals to accommodate extra electrons. In SF_6, for example, S exceeds the octet rule.

When writing Lewis structures, satisfy the octet rule first; if extra electrons remain, place them on elements having available d orbitals; if one of several elements exceed the octet rule, assume the electrons are on the central atom.



There may several valid *equivalent* Lewis structures for some molecules. This is called **resonance**. The resonance structures for the polyatomic ion NO^{3-} is pictured to the right. It does NOT have one double and two single bonds nor does it flip from one to the other structure as it appears. Rather it has 3 equivalent bonds. Resonance is represented by showing the possible Lewis structures with a double headed arrow indicating the actual structure is an average of the resonance structures.



When there are several nonequivalent Lewis structures for a molecule, it is possible to choose among them using formal charge. Formal charge is the difference between the number of valence electrons on the free atom and the number of valence electrons assigned to the atom in the molecule.

<u>Calculating Formal Charge</u>

- lone pairs belong entirely to the atom in question
- shared electrons are divided equally among the sharing atoms

∴ (Valence electrons)assigned = (number of lone pair electrons) = ½(number of shared electrons)

<u>Evaluating Lewis Structures with Formal Charge</u>

- atoms in molecules should have formal charges as close to zero as possible
- negative formal charges reside with the most electronegative element

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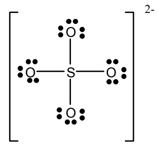
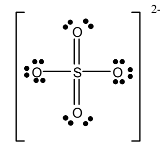
Example: SO_4^{-2}

Structure		 32 valence electrons	 32 valence electrons
Formal Charge	Oxygen with single bonds	Valence electrons assigned: 6 + ½ (2) = 7 Formal charge: 6 – 7 = –1	Valence electrons assigned: 6 + ½ (2) = 7 Formal charge: 6 – 7 = –1
	Oxygen with double bonds	—	Valence electrons assigned: 4 + ½ (4) = 6 Formal charge: 6 – 6 = 0
	Sulfur	Valence electrons assigned: 0 + ½ (8) = 4 Formal charge: 6 - 4 = 2	Valence electrons assigned: 0 + ½ (12) = 6 Formal charge: 6 – 6 = 0
	Total	4(–1) + 2 = –2 ✗	2(–1) + 2(0) + 0 = –2 ✓

The preferred Lewis structure for the sulfate ion, as shown above, has two doubly bonded oxygens and two singly bonded oxygens. As such, it uses all six of sulfur's valence electrons to form bonds. This is preferred, because it results in the lowest formal charge on each atom.

Draw the Lewis structures for each of the following using the procedures described above.

1. SO_3
2. CO_2
3. NH_3
4. H_2S
5. NH_4^+
6. CH_3Cl
7. P_2O_5
8. H_2SO_4