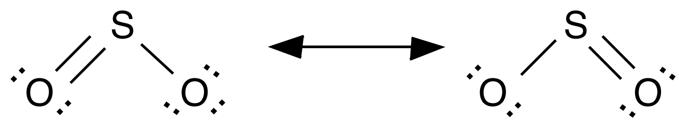
Student Handout 2 of 3: Chemical Bonding

Resonance: Comes up when there are more than one are equivalent Lewis structures.



Nothing tricky, when you draw a Lewis diagram just ask yourself if you could get an equivalent structure by moving a π-bond.

Hit Parade: SO_2, SO_3, NO_3^-, HNO_3, CO_3^{2-}

Expanded Octet: Once you get to period 3 and below, Group V, VI, VII atoms can use inner atomic p and d orbitals for bonding, so the octet rule can be violated. They are either brutally obvious, SF_6, or if they are subtle, I_3^-, problems with VSEPR will clue you in. Like for I_3^-, step 1 says that you need 12 pairs of electrons, step 2 says that the total number of covalent bonds possible is 11, step 3 says that the actual number of covalent bonds is 12-11 = 1 <---- obviously VESPR is screwed up and this almost always indicates an expanded octet molecule.

Formal Charge: This is a rarely used calculation, but you kind of have to know how to answer the question "what is the formal charge of a particular atom in a molecule or of the molecule as a whole." It's also supposed to be a way of determining which is the most probable Lewis structure out of all different possibilities. The basic idea is that the lower the formal charge of the molecule, the more stable the structure.

For an atom in a molecule: F.C. = group number - [(# of bonds) + (# of unshared e^-'s)]

In a neutral molecule the sum of all the formal charges should equal zero, in an ion it should equal the ionic charge

Bond Polarity/Electronegativity

Electronegativity is an idea of Linus Pauling's. He assigned a number between 0.8 and 4 to each element according to the element's ability to attract an additional electron (.8 to C and 4.0 to F). In a covalent bond, you have a shared electron that sort of "figure-eights" between the two nuclei. By comparing the electronegativity values of the two elements in the bond, you can get a relative idea of how much time the shared electron spends with each nucleus. Since electrons are negatively charged. This will induce a negative/positive polarity in the bonds between elements of different electronegativities. Generally an electronegativity difference greater than 2.0 is an ionic bond, between .4 and 2.0 is a polar covalent bond and less than .4 is a nonpolar covalent bond. <-- Once you know the polarity of bonds, you can determine the molecule's dipole. To look at the overall molecule you need to use representative vectors. For each bond in a molecule just draw arrows from positive to negative polarity and then add them all up vectorially to get an idea of the overall molecular polarity.

Electronegativity

H 2.2						
Li 1.0	Be 1.6	B 2	C 2.6	N 3.0	O 3.4	F 4.0
Na .9	Mg 1.3	Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2
K .8	Ca 1.0	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0
Rb .8	Sr .9	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.7
Cs .8	Ba .9	Tl 2	Pb 2.3	Bi 2.0	Po 2.0	At 2.2

Polarity