

Chem 233: Review Sheet for Chapters 1 & 2

LEWIS STRUCTURES

Steps for Drawing:

- Count up the total number of valence electrons in the structure.
NOTE: For A-Group elements, the number of valence electrons is equal to the Group number.
- Decide on a skeletal structure (what's bonded to what).
 - The central atom is generally written first in the chemical formula.
 - Hydrogen is never the central atom but is usually written first in acids.
- Place two electrons in each bond.
- Surround each of the non-central atoms with an octet of electrons.
NOTE: Hydrogen is an exception and is only surrounded by two electrons.
- Determine how many electrons are left. Place any remaining electrons on the central atom in pairs.
- If the central atom does not have an octet of electrons, form double or triple bonds to give it an octet of electrons.
NOTE: H, B, Be, F, and metals do not generally form multiple bonds.
- If the central atom in Step #6 does not form multiple bonds, leave it short of electrons.
NOTE: H is generally surrounded by 2 electrons; Be, by 4 electrons and B, by 6 electrons.

Valence: Number of bonds that an atom of an element likes to form in covalent compounds. This is also equal to the number of e⁻ needed to complete the octet. Use valences to easily and quickly draw Lewis structures for organic substances.

Common Valences:

Non-metals in Groups:				
H	IVA	VA	VIA	VIIA
1	4	3	2	1

Example: Carbon in Group IVA has a valence of 4 and consequently likes to form 4 bonds. These four bonds can consist of:

- four single bonds
- one double and two single bonds
- one triple and one single bond
- two double bonds

Electronegativity (\underline{en}): Ability of bonded atom to attract (i.e. inductively pull) electrons in bond toward itself.

1. Increases in value as proceed across the periodic table within a row from Group IA to Group VIIA. Note: Electronegativity of Group VIIIA elements is zero.
2. Decreases in value as proceed down the periodic table within a column.
3. In general: non-metals \rightarrow high \underline{en} \rightarrow electronegative
metals \rightarrow low \underline{en} \rightarrow electropositive

Difference in electronegativity ($|\Delta \underline{en}|$): Value of $|\Delta \underline{en}|$ gives information on type of bond and compound present.

- | | |
|--|--|
| If $ \Delta \underline{en} \geq 2.0$ | Ionic bonds are present between ions.
Ex. NaF (en: F=4.0; Na=0.9) |
| If $ \Delta \underline{en} = 0$ | Covalent bonds are present between atoms.
Ex. F_2 (en: F=4.0) |
| If $0 < \Delta \underline{en} < 2.0$ | Polar covalent bonds are present between atoms. Ex. HCl (en: Cl=3.2; H=2.2) |

- Polar covalent bonds:**
1. Have a character intermediate between covalent and ionic bonds.
 2. In a polar covalent bond, electron pairs are shared but not equally. The electron pair spends more time near the more electronegative atom giving this atom a partial negative charge.

Formal Charge: Charge on atom in given Lewis structure. A formal charge of +1 means that the atom has one less valence e⁻ surrounding it than the neutral atom. A formal charge of -2 means that the atom has two more valence e⁻ surrounding it than the neutral atom.

To calculate:

Formal Charge = (# Valence e⁻) - ($\frac{1}{2} \times$ # bonding e⁻) - (# nonbonding e⁻)

OR

Formal Charge = (# Valence e⁻) - (# bonds) - (# nonbonding e⁻)