

Chem 231: Review Sheet for Chapter 1

Electronegativity (en): Ability of bonded atom to attract/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 en \rightarrow electronegative
metals \rightarrow low en \rightarrow electropositive

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

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|-----------------------------------|---|
| If $ \Delta \text{en} \geq 2.0$ | Ionic bonds are present between ions.
Ex. NaF (en : F=4.0; Na=0.9) |
| If $ \Delta \text{en} = 0$ | Covalent bonds are present between atoms.
Ex. F_2 (en : F=4.0) |
| If $0 < \Delta \text{en} < 2.0$ | 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⁻) - (# bonds) - (# nonbonding e⁻)

LEWIS STRUCTURES

Steps for Drawing:

1. 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.
 2. Decide on a skeletal structure (what's bonded to what).
 - a. The central atom is generally written first in the chemical formula.
 - b. Hydrogen is never the central atom but is usually written first in acids.
 3. Place two electrons in each bond.
 4. Surround each of the non-central atoms with an octet of electrons.
NOTE: Hydrogen is an exception and is only surrounded by two electrons.
 5. Determine how many electrons are left. Place any remaining electrons on the central atom in pairs.
 6. 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.
 7. 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.
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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:				
<u>H</u>	<u>IVA</u>	<u>VA</u>	<u>VIA</u>	<u>VIIA</u>
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:

- A. four single bonds
- B. one double and two single bonds
- C. one triple and one single bond
- D. two double bonds