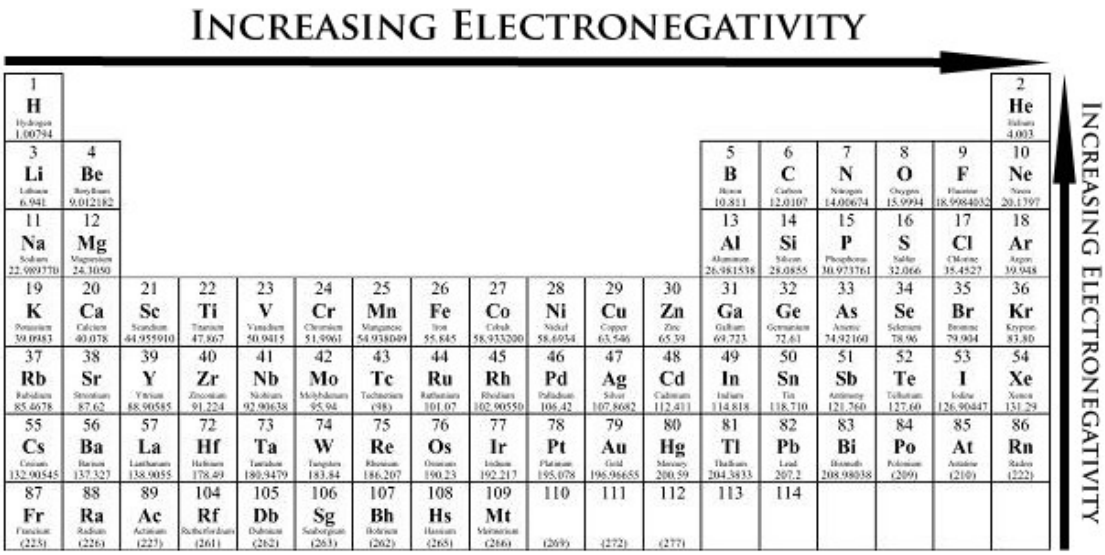
- Electronegativity
  - **Electronegativity** is a measure of how much atoms want electrons.
    - Electronegativity differences between two atoms helps tells you which type of bond they participate in.
      - Nonpolar bond has electronegativity differences less than 0.5.
      - Polar, covalent bonds have electronegativity differences between 0.5 and 2.
      - Ionic bonds generally have electronegativity differences greater than 2.
    - Trend:

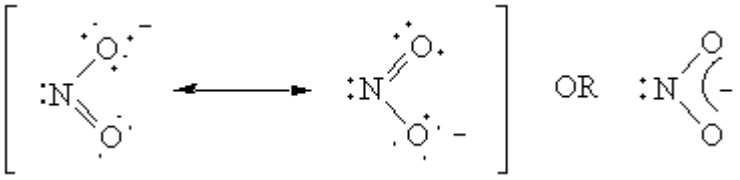


- Lewis structures special notes
  - Elements not part of the transition metal block and are in the 3rd period and below can have expanded octets.
  - Boron is electron deficient, only needed 6 electrons to fill its shell.
  - When adding up the valence electrons in a molecule, if you get an odd number, then it is a radical.
- Formal charge
  - The **formal charge** is the charge of an atom in a molecule. By adding all the formal charges of all the atoms in a molecule, you can determine if the overall molecule is positive, negative, or neutral.
  - Equation:

$$VE = valence\ electrons$$
$$NE = nonbonding\ electrons$$
$$BE = bonding\ electrons$$
$$Formal\ Charge = VE - NE - \frac{BE}{2}$$

- Resonance structures
  - A **resonance structure** is another way to depict the same molecule. Only electrons can move around, not atoms!
  - How to determine importance of resonance structure to the overall structure (most important to least important factors):
    - 1) All atoms have full octets
    - 2) Minimal separation of charge
    - 3) More electronegative atom should have the negative charge whereas the more electropositive atom should have the positive charge.
  - In reality, all resonance structures for the same molecule exist **at the same time, all the time**.
  - Most important resonance contributor best describes the shape of the molecule whereas the second most important resonance contributor best describes the reactivity of the molecule.

- Bond length
  - **Bond length** describes the distance of a bond between two atoms.
  - How do rank bond lengths:
    - Bond length follows the atomic radius trend. Larger atoms that bond together will have larger bond lengths for this reason. (H-Br > H-F)
    - single bond > double bond > triple bond
    - higher % s character ($sp > sp^2 > sp^3$) of the atoms participating in the bond, the shorter the bond
  - NOTE: watch for resonance!



    - This is NO_2^-, or nitrite. When looking at the N-O bond lengths, one might be tempted in saying that one N-O bond is strictly a double bond while the other is a single bond.. However, this is not true! Note that the molecule can resonate! Because of this, the N-O bond length is actually less than 2, it is around 1.5.
    - **When given a bond length problem, look for resonance because that can affect the way you answer the question.**
- Bond strength
  - **Bond strength** indicates how much energy is needed to break a bond into two equal atoms.
  - General rule: the shorter the bond is, the stronger the bond is
  - How to rank bond strength:
    - triple bond > double bond > single bond
    - Higher % s character ($sp > sp^2 > sp^3$) of the atoms participating in the bond, the stronger the bond strength. This is because the atoms participating in the bonds are in lower energy orbitals and are more tightly bound by the nucleus.
  - NOTE: watch for resonance (just like bond length)!
- Molecular geometry
  - Determining hybridization:
    - 1) count every single, double, or triple bond around a given atom as "1"
    - 2) count every lone pair around the same atom as "1"
    - 3) add up the two counts and then use the table to determine hybridization

Count	Hybridization
2	sp
3	sp^2
4	sp^3
5	sp^3d
6	sp^3d^2

    - NOTE: watch for resonance! The resonance structure with the smallest observed hybridization best describes the overall hybridization!
  - sp^3 = ~109° bond angles and a tetrahedral arrangement
    - 4 bond sand 0 lone pairs = tetrahedral
    - 3 bonds and 1 lone pair = trigonal pyramidal
    - 2 bonds and 2 lone pairs = bent
    - 1 bond and 3 lone pairs = linear
  - sp^2 = ~120° bond angles and trigonal planar arrangement

        - 3 bonds and 0 lone pairs = trigonal planar
        - 2 bonds and 1 lone pair = bent
        - 1 bond and 2 lone pairs = linear
    - sp = ~180^{o} bond angles and linear arrangement
    - **lone pair electrons** take up extra space and crunches the bond angles. The more lone pairs that exist around an atom, the smaller the bond angles are.
- Intermolecular forces
    - **Intermolecular forces** are forces that occur between neighboring particles. They have big effects on physical properties.
    - In terms of strength: Ion-dipole > hydrogen-bonding > dipole-dipole > van-der-waals
    - **Ion-dipole** is an interaction between an ion and an oppositely charged dipole.
    - **Hydrogen bond** is the interaction between a lone pair on fluorine, oxygen, or nitrogen and a hydrogen on another molecule that is directly bound to fluorine, oxygen, or nitrogen.
        - NOTE: hydrogen bonding can occur between atoms within the same molecule. Intramolecular H-bonds prevent intermolecular H-bonds from forming and thus weakens its effect on physical properties.
    - **Dipole-dipole** occurs from polar covalent bonds.
        - NOTE: dipoles within the same molecule can cancel each other out if the dipoles are in the opposite directions and are of the same magnitude. Factor in the NET molecule dipole.
    - **Van-der-waals** occurs between all atoms. It is the brief attraction between neighboring molecules due to the random movement of electrons.
        - The larger the molecular weight, the stronger the VDW forces.
        - The bigger the molecule is (more surface area), the stronger the VDW forces.
        - The more “branched” the molecule is, the weaker the VDW forces.
- Melting points, boiling points, and solubility
    - Boiling point ranking:
        - 1) The stronger the IMFs, the higher the boiling point
            - Note that for hydrogen bonding, oxygen > nitrogen (you will probably never see a fluorine molecule).
        - 2) The larger the surface area of the molecule, the higher the BP
        - 3) The more polarizable the atom is, the higher the BP
            - **Polarizability** is described as the ease of distortion of the electron cloud of a molecule by an electric field. Polarizability increases down the <u>column</u> of the periodic table. (I > Br > Cl)
            - NOTE: HF > HI > HBr > HCl for boiling point (note that HF can do hydrogen bonding whereas the other 3 acids cannot)
    - Melting point ranking:
        - 1) The stronger the IMFs, the higher the melting point
        - 2) The more branched the molecule is, the higher the melting point
    - Solubility ranking:
        - 1) similar IMFs between solvent and solute
        - 2) the larger the surface area of the solute is, the less dissolving there is