Unit 2: Molecular and Ionic Compound Structure and Properties

Types of Bonds

- **Bonds** form in order to decrease potential energy and increase stability
- **Chemical Bond:**
  - When an attractive force between atoms or ions binds them together as a unit.
- **Covalent Bonding** (AKA Molecular Bonding)
  - When *two nonmetals* share electrons
  - Nonpolar Covalent (Equal Sharing) and Polar Covalent (Unequal Sharing)
- **Metallic Bonding**
  - Metals are held together by sharing their outer electrons.
  - Those e^- can flow around atom to atom, which is represented by the term: **sea of electrons**.
- **Ionic Bonding**
  - Metal and nonmetal transfer electrons
  - Very Polar
  - **Lattice Energy** = the energy that is released when a crystal forms from ions
    - Delta $H_{LE} = (Q_1Q_2 / r)$
      - Q = charge of ions
      - r = distance between nuclei

	Ionic	Covalent	Metallic

Bond Formation	e^- are transferred from metal to nonmetal	e^- are shared between 2 nonmetals	e^- are delocalized among metal atoms
Type of Structure	Crystal lattice	True molecules	"Electron sea"
Physical State	Solid	Liquid or Gas	Solid
Melting Point	High	Low	Very High
Solubility in Water	soluble	Usually insoluble	insoluble
Electrical Conductivity	Only in a solution or liquid form	No	Yes (Any form)
Other Properties	N/A	odorous	Malleable, ductile, lustrous

Ion Formation- Ionic Compounds

- Ions are atoms which gain or lose electrons
- Ions have overall charge
    - **Noble Gas Rule**: Main group elements want to have the same number of electrons as the closest noble gas. They gain or lose electrons in order to do this
        - Losing Electrons = Positive Ions = cations
        - Gaining Electrons = Negative Ions = anions
        - Ex. Sodium: 11 protons and 11 electrons
            - Closest noble gas: Neon ($10e^-$)
            - Sodium loses 1 e^- to have the same number as neon: Na^{+1}, Sodium Ion

Naming Ions

- Positive Ion = Name doesn't change
- Negative Ion = Name's ending changes to -ide

Writing Ionic Formulas

- Positive Ions only bond to negative ions
- Positive Ion is always first
- Positive Charge must balance the negative
- Subscripts are used after each symbol to indicate how many are needed (unless it's a 1)

Naming Ionic Formulas

- Name the Cation (+), Anion (-)
- The name shouldn't include how many atoms of each ion are present
- Examples
    - Na^{+1} Br^{-1} = NaBr, Sodium Bromide

Transition Metals

- Most transition metals can have more than one charge
- Specify the charge in the name with Roman Numerals
    - Fe^{+2} is Iron (ll)
- Exceptions: There are three transition metals that only have one possible charge
    - Zn = +2, Cd = +2, Ag = +1
    - For these ions, we don't use roman numerals in the name

Polyatomic Ions

- Clusters of atoms that have an overall charge
- They are attracted to opposite ions and form compounds
- When using polyatomics in compounds, NEVER alter their formulas
- If more than one needed, use parentheses around the formula followed by the subscript
- List of common polyatomic ions that you should memorized or be familiarized with:

Name	Ion Formula	Name	Ion Formula
ammonium	NH_4^{1+}	sulfite	SO_3^{2-}
hypochlorite	ClO^{1-}	hydrogen sulfite	HSO_3^{1-}
chlorite	ClO_2^{1-}	sulfate	SO_4^{2-}
chlorate	ClO_3^{1-}	hydrogen sulfate	HSO_4^{1-}
perchlorate	ClO_4^{1-}	thiosulfate	$S_2O_3^{2-}$
hypobromite	BrO^{1-}	carbonate	CO_3^{2-}
bromite	BrO_2^{1-}	hydrogen carbonate	HCO_3^{1-}
bromate	BrO_3^{1-}	oxalate	$C_2O_4^{2-}$
perbromate	BrO_4^{1-}	nitrite	NO_2^{1-}
acetate[1]	$C_2H_3O_2^{1-}$	nitrate	NO_3^{1-}
selenate	SeO_4^{2-}	cyanide	CN^{1-}
phosphate	PO_4^{3-}	cyanate	OCN^{1-}
hydrogen phosphate	HPO_4^{2-}	thiocyanate	SCN^{1-}
dihydrogen phosphate	$H_2PO_4^{1-}$	hydroxide	OH^{1-}
chromate	CrO_4^{2-}	permanganate	MnO_4^{1-}
dichromate	$Cr_2O_7^{2-}$	peroxide	O_2^{2-}

Molecular Compounds

- Made of all nonmetals that share electrons
- Use prefixes to indicate how many atoms of each element are in the middle
    - Mono - (1)
    - Di - (2)
    - Tri - (3)
    - Tetra - (4)
    - Penta - (5)
    - Hexa - (6)
    - Hepta - (7)
    - Octa - (8)
    - Nona - (9)
    - Deca - (10)
- When there is only one atom of the first element, leave the “mono” off
- Second element ends with -ide

Polarity

- Polar bonds occur when e- are shared unequally
- A molecule is polar if it has polar bonds that are not symmetrical
- A polar molecule has one side that is positive and the other side is negative

Diatomic Molecules

- Composed of 2 nonmetal atoms
- The H-7 Club
- There are 7 and they make a 7 on the periodic table, which starts at #7
    - H2, N2, O2, F2, Cl2, Br2, L2
- Examples
    - CCl4 = Tetrachloride
    - N2O3 = Dinitrogen Trioxide
    - P2F5 = Diphosphorus Pentafluoride

Hydrates

- A solid compound that containing or linked to water molecules
- Name the ionic compound followed by the numerical prefix and the suffix -hydrate
    - CuSo4 * 5 H20 = Copper (ll) Sulfate Penta Hydrate

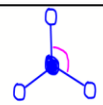
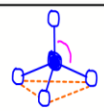
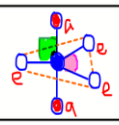
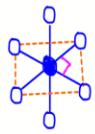
Lewis Diagrams

- Representation of the **valence electrons** in a molecule
- Show how electrons are arranged around individual atoms in a molecule
- The dots represent the valence electrons and the line is shown for a bonding of electrons between two atoms
    - 1 line = Single Bond = 2 electrons
    - 2 lines = Double Bond = 4 electrons
    - 3 lines = Triple Bond = 6 electrons
    - 2 dots = lone pair of electrons
- Steps for drawing a Lewis Diagram

    - Count all the valence electrons
    - Determine the central atom (the element there is only one of)
    - Put all the remaining valence electrons on atoms as lone pairs
    - Turn lone pairs into double/triple bonds to give every atom an octet (or duet)
- Alternate method to drawing a Lewis Structure
    - Figure out how to many bonds each atom wants
    - Put the atom that wants the most bonds in the middle
    - Give each atom the number of bonds it wants
    - Add lone pairs to give every atom an octet

Valence Electron Shell Pair Repulsion (VSEPR)

- VSEPR predicts shapes of molecules
- Can be found by using the number of electron pairs
    - To predict the shape, first draw out the Lewis Structure of the compound and look at its central atom
    - Count its valence electrons
    - Add one electron for each bonding atom
    - Add or subtract electrons for charge
    - Divide the total of these by 2 to find the total number of electron pairs
    - Use this final number to predict the shape

Total no. of electron pairs = bond pairs + lone pairs	Basic Shape	Diagram	Bond angle
2	linear		180°
3	trigonal planar		120°
4	tetrahedral		109.5°
5	trigonal bipyramidal	 axial equatorial	120° 90°
6	octahedral		90°

- Example
  - (PF6)-
    - Central Atom = Phosphorus
    - Valence Electrons on Central Atom = 5
    - 6 (F) Atoms = 6
    - +1 Electron for negative charge on P = 1
    - Total = 12, divide by 2 to become 6
    - Phosphorus is an octahedral

Hybrid Orbitals

- The central atom's coordination number gives the number of hybrid orbitals
- If any element has a coordination number of 4, it will hybridize all four valence orbitals (sp3)
- If an element has a coordination number of 3, it only needs to hybridize three of its orbitals (1 remains as a normal "p" orbital), it is (sp2)
- Hybrid orbitals make sigma bonds, which is the direct overlap of orbitals between two atoms
- Non-hybridized orbitals can still bond but they are different than the hybrid orbital ones

- Non-hybrid orbitals make pi bonds
    - Sharing of electrons between parallel p orbitals

Websites that I used for reference:

- https://www.chem.fsu.edu/chemlab/chm1045/e_config.html
- https://chem.libretexts.org/Bookshelves/Organic_Chemistry/Supplemental_Modules_(Organic_Chemistry)/Fundamentals/Ionic_and_Covalent_Bonds

Pictures:

- https://chemistryguru.com.sg/valence-shell-electron-pair-repulsion-theory
- https://www.pinterest.com/pin/472948398346860399/