Chapter 4: Structure and Properties of Ionic and Covalent Compounds

- 4.1 Chemical Bonding
  - **Chemical Bond** - the force of attraction between any two atoms in a compound.
  - Interactions involving valence electrons are responsible for the chemical bond.
  - Lewis symbol (Lewis structure) - a way to represent atoms (and their bonds) using the element symbol and valence electrons as dots.
  - **Principal Types of Chemical Bonds: Ionic and Covalent**
    - **Ionic bond** - a transfer of one or more electrons from one atom to another.
    - forms attractions due to the opposite charges of the atoms.
    - **Covalent bond** - attractive force due to the sharing of electrons between atoms.
  - <u>Essential Features of Ionic Bonding</u>
    - Atoms with low I.E. and low E.A. tend to form positive ions.
    - Atoms with high I.E. and high E.A. tend to form negative ions.
    - Ion formation takes place by electron transfer.
    - The ions are held together by the electrostatic force of the opposite charges.
    - <u>Reactions between metals, and metals and nonmetals (representative) tend to be ionic.</u>
  - **COVALENT BONDING**
    - Let's look at the formation of H_2
    - $H + H \rightarrow H_2$
    - Each hydrogen has one electron in it's valance shell.
    - If it were an ionic bond it would look like this:
    - However, both hydrogen atoms have the same tendency to gain or lose electrons.
    - Both gain and loss will not occur.
    - Instead, each atom gets a noble gas figuration by ***sharing*** electrons.
  - <u>Features of Covalent Bonds</u>
    - Covalent bonds tend to form between atoms with similar tendency to gain or lose electrons.
    - The diatomic elements have totally covalent bonds (totally equal sharing.)
  - **Polar Covalent Bonding and Electronegativity**
    - **Polar covalent bonding** - bonds made up of unequally shared electron pairs.
    - A truly covalent bond can only occur when both atoms are identical.
    - Electronegativity is used to determine if a bond is polar and who gets the electrons the most.
    - **Electronegativity** - a measure of the ability of an atom to attract electrons in a chemical bond.
    - The greater the difference in electronegativity between two atoms, the greater the polarity of a bond.
    - Which would be more polar, a H-F bond or a H-Cl bond? The HF bond is more polar than the HCl bond.

- 43.2 Naming Compounds and Writing Formulas of Compounds
  - **Nomenclature** - the assignment of a correct and unambiguous name to each and every chemical compound.
  - We will learn two systems
    - one for naming ionic compounds and
    - one for naming covalent compounds.
  - I. **Ionic Compounds**
  - A. Writing Formulas of Ionic Compounds from the Identities of the Component Ions.
  - B. Writing Names of Ionic Compounds from the Formula of the Compound
  - C. Writing Formulas of Ionic Compounds from the Name of the Compound
  - IV. **Covalent Compounds**
  - A. Naming Covalent Compounds
  - B. Writing Formulas of Covalent Compounds
  - Metals and nonmetals usually react to form ionic compounds.
  - The metals are the cations and the nonmetals are the anions.
  - The cations and anions arrange themselves in a regular three-dimensional repeating array called a crystal lattice.
  - **Formula** - the smallest whole number ratio of ions in the crystal.
  - **<u>A. Writing Formulas of Ionic Compounds</u>**
    - Determine the charge of the ions (usually can be obtained from the group number.)
    - Cations and anions must combine to give a formula with a net charge of zero, it must have the same number of positive charges as negative charges.
  - **<u>B. Writing Names of Ionic Compounds from the Formula</u>**
    - **Stock System:**
    - 1. Name cation followed by the name of anion.
    - 2. Give anion the suffix *-ide*.
    - Examples:
    - NaCl is sodium chloride.
    - $AlBr_3$ is aluminum bromide.
    - If the cation of an element has several ions of different charges (as with Transition metals) use a Roman numeral after the metal name.
    - Roman numeral gives the *charge* of the metal.
    - Examples:
      - $FeCl_3$ is iron(III) chloride
      - $FeCl_2$ is iron(II) chloride
      - CuO is copper(II) oxide
    - **Common Nomenclature System**
    - Use *-ic* to indicate the higher of the charges.
    - Use *-ous* to indicate the lower of the charges.
    - Examples:
    - $FeCl_2$ is ferrous chloride, $FeCl_3$ is ferric chloride
    - Cu_2O is cuprous oxide, CuO is cupric oxide

- Note that the common system requires you to know the common charges and use the Latin names of the metals.
- **Monatomic ions** - ions consisting of a single atom.
- Examples
  - K^+ potassium ion
  - Ba^{2+} barium ion
- Table 4.2: common monatomic cations and anions.
- **Polyatomic ions** - ions composed of 2 or more atoms bonded together. Within the ion, the atoms are bonded using covalent bonds. The ions will be bonded to other ions with ionic bonds.
- Examples:
  - NH4+ ammonium ion
  - SO42- sulfate ion

<u>C. Writing Formulas of Ionic Compounds from the Name</u>

- Determine the charge on the ions.
- Write the formula so the compounds are neutral.
- Example:
- Barium chloride:
- Barium is +2, Chloride is -1
- Formula is $BaCl_2$

II. Covalent Compounds

- <u>Covalent compounds are usually formed from nonmetals.</u>
- **Molecular Compounds** - compounds characterized by covalent bonding.
- not a part of a massive three dimensional crystal structure.
- **<u>A. Naming Covalent Compounds</u>**
  - 1. The names of the elements are written in the order in which they appear in the formula.
  - 2. A prefix indicates the number of each kind of atom

mono-	1	hexa-	6
di-	2	hepta-	7
tri-	3	octa-	8
tetra-	4	nona-	9
penta-	5	deca-	10

  - If only one atom of a particular kind is present in the molecule, the prefix mono- is usually omitted from the first element.
  - Example: CO is carbon monoxide
  - The stem of the name of the last element is used with the suffix –ide
  - The final vowel in a prefix is often dropped before a vowel in the stem name.
- **<u>B. Writing Formulas of Covalent Compounds</u>**
  - Use the prefixes in the names to determine the subscripts for the elements.
  - Example:
  - diphosphorus pentoxide
  - P_2O_5

        - Some common names are used.
- 4.3 Properties of Ionic and Covalent Compounds
    - Physical State
    - Ionic compounds are solids at room temperature
    - Covalent compounds are solids, liquids and gases
    - Melting and Boiling Points
    - **melting point** - the temperature at which a solid is converted to a liquid
    - **boiling point** - the temperature at which a liquid is converted to a gas
    - Melting and Boiling Points
        - Ionic compounds have much higher melting points and boiling points than covalent compounds due to the large amount of energy required to break the attractions between ions.
    - Structure of Compounds in the Solid State
        - Ionic compounds are crystalline
        - Covalent compounds are crystalline or **amorphous** - have no regular structure.
    - Solutions of Ionic and Covalent Compounds
        - Ionic compounds often dissolve in water, when they do they **dissociate** - form positive and negative ions in solution.
        - **Electrolytes** -ions present in solution allow the solution to conduct electricity.
        - Covalent solids usually do not dissociate and do not conduct electricity - **nonelectrolytes**
- 4.4 Drawing Lewis Structures on Molecules and Polyatomic Ions
    - 1. Use chemical symbols for the various elements to write the skeletal structure of the compound.
        - the least electronegative atom will be placed in the central position
    - 2. Determine the total number of valence electrons associated with each atom in the compound.
        - for polyatomic cations, subtract one electron for every positive charge;
        - for polyatomic anions, add one electron for every negative charge.
    - 3. Connect the central atom to each of the surrounding atoms using electron pairs. Then give each atom an octet.
        - Remember, hydrogen needs only two electrons
    - 4. Count the number of electrons you have and compare to the number you used.
        - If they are the **same**, you are finished.
        - If you **used more electrons than you have** add a bond for every two too many you used. Then give every atom an octet.
        - If you **used less electrons than you have**….(see later when discuss exceptions to the octet rule)
    - 5. Check that all atoms have the octet rule satisfied and that the total number of valance electrons are used.
    - **Lewis Structure, Stability, Multiple Bonds, and Bond Energies**
        - **Single bond** - one pair of electrons are shared between two atoms
        - **Double bond** - two pairs of electrons are shared between two atoms
        - **Triple bond** - three pairs of electrons are shared between two atoms

            - **Bond energy** - the amount of energy required to break a bond holding two atoms together.
            - triple bond > double bond > single bond
            - **Bond length** - the distance separating the nuclei of two adjacent atoms.
            - single bond > double bond > triple bond
            - **Resonance** - two or more Lewis structures that contribute to the real structure.
    - **Lewis Structures and Exceptions to the Octet Rule**
        - Incomplete Octet - less then eight electrons around an atom other than H.
            - Let's look at BF_3
        - Odd Electron - if there is an odd number of valence electrons it isn't possible to give every atom eight electrons.
            - Let's look at NO
        - Expanded Octet - elements in 3rd period and below may have 10 and 12 electrons around it. Expanded octet is the most common exception.
            - Write the Lewis structure of SF6
    - **Lewis Structures and Polarity**
        - Polar molecules are molecules that are polar or behave as a dipole (two "poles"). One end is positively charged the other is negatively charged. Typically the result of polar bonds or geometry.
        - Nonpolar molecules do not have dipoles or react in an electric field.
- 4.5 Properties Based On Molecular Geometry
    - **Intramolecular forces** - attractive forces *within* molecules. (Chemical bonds)
    - **Intermolecular forces** - attractive forces *between* molecules.
    - We will look at how intermolecular forces affect:
        - I. Solubility.
        - II. Boiling points and melting points.
    - **Solubility** - the maximum amount of solute that dissolves in a given amount of solvent at a specific temperature.
    - I. "Like dissolves like"
        - polar molecules are most soluble in polar solvents
        - nonpolar molecules are most soluble in nonpolar solvents.
            - Example: ammonia (NH_3) in water.
            - Example: water is polar and oil is nonpolar.
    - II. Boiling Points and Melting Points
        - The greater the intermolecular force the higher the melting point and boiling point.
        - Two factors to consider:
            - Larger molecules have higher m.p. and b.p. than smaller molecules.
            - Polar molecules have higher m.p. and b.p. than nonpolar molecules (of similar molecular mass.)