CHAPTER 12

12.1 TYPES OF CHEMICAL BONDS

- Bond→ a force that holds groups of two or more atoms together and makes them function as a unit
- Bond energy→ energy required to break the bond
- Ionic bonding→ the attraction between oppositely charged ions
- Ionic compound→ a compound that results when a metal reacts with a nonmetal to form cations and anions
- Covalent bonding→ a type of bonding in which atoms share electrons
- Polar covalent bonding→ a covalent bond in which the electrons are not shared equally because one atom attracts them more strongly than the other

12.2 ELECTRONEGATIVITY

- Electronegativity→ the relative ability of an atom in a molecule to attract shared electrons to itself
- Generally increases going from left to right on the periodic table
  4.0 Fluorine→0.7 Cesium
- The higher the atom's electronegativity value, the closer the shared electrons tend to be to that atom when it forms a bond
- The polarity of a bond depends on the difference in electronegativities
- The polarity of the bond increases as the difference in electronegativity increases

12.3 BOND POLARITY AND DIPOLE MOMENTS

- Dipole moment→ a molecule with a center of positive charge and a center of negative charge
- The dipole is represented by an arrow pointing toward the negative charge center, and its tail indicating the positive center of charge
- Any diatomic molecule has a dipole moment
- Because water molecules are polar, they can surround and attract both positive and negative ions
- The attractions allow ionic materials to dissolve in water
- The polarity of the water molecule also causes them to attract strongly, meaning much energy is needed to change water from a liquid to gas

12.4 STABLE ELECTRON CONFIGURATIONS AND CHARGES ON IONS

- Group 1 metals always form 1+ cations, group 2 metals always form 2+ cations, and aluminum in group 3 always forms a 3+ cation
- For the nonmetals, group 7 elements always form 1- anions, and the group 6 elements always from 2- anions
- In almost all stable chemical compounds of the representative elements, all of the atoms have achieved a noble gas electron configuration

- When a nonmetal and a group 1,2, or 3 metal react to form a binary ionic compound, the ions form in a way that the valence electron configuration of the nonmetal is completed to achieve the configuration of the next noble gas, and the valence orbital's of the metal are emptied to achieve the configuration of the previous noble gas. In this way both ions achieve noble gas electron configuration.
- When two nonmetals react to form a covalent bond, they share electrons in a way that completes the valence-electron configurations of both atoms. That is, both nonmetals attain noble gas electron configurations by sharing electrons.

  The electronegativity of oxygen is greater than calcium

  Ca: $[Ar]4s^2$

  O: $[He]2s^22p^4$
- Because of the difference, electrons are transferred from calcium to oxygen to form an oxygen anion and a calcium cation
- Oxygen needs two electrons to fill its valence orbitals (2s and 2p) and achieve the configuration of neon ($1s^22s^22p^6$), which is the next noble gas
- By losing two electrons, calcium can achieve the configuration of argon (the previous noble gas)

12.5 BONDING AND STRUCTURES OF IONIC COMPOUNDS

- A cation is always smaller than the parent atom, and an anion is always larger than the parent atom
- When a metal loses all of its valence electrons to form a cation, it gets much smaller
- In forming an anion, a nonmetal gains enough electrons to achieve the next noble gas electron configuration and so it becomes much larger
- Ammonium nitrate contains the NH_4^+ and NO_3^- ions, it also contains covalent bonds in the individual polyatomic ions

12.6 LEWIS STRUCTURES

- Bonding involves just the valence electrons of atoms
- Lewis Structure→ a representation of a molecule that shows how the valence electrons are arranged among the atoms in the molecule

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  No dots are shown on the K^+ ion because it has lost its only valence electron to the Br^- ion (to fill the valence shell)
- Molecules with covalent bonds:
  1. Hydrogen forms stable molecules where it shares two electrons. That is, it follows a **duet rule**. For example, when two hydrogen atoms, each with one electron, combine to form the H_2 molecule, each hydrogen in H_2 has, in effect, two electrons. Each hydrogen has a filled valence shell.

2. Helium does not form bonds because its valence orbital is already filled; it's a noble gas.
3. Second-row nonmetals carbon through fluorine form stable molecules when they are surrounded by enough electrons to fill the valence orbitals. Eight electrons are required to fill the orbitals (2s and 2p), so these elements obey the **octet rule**→ they are surrounded by eight electrons. **Lone pairs** or **unshared pairs** of electrons are pairs that are not involved in bonding.
4. Neon does not form bonds because it already has an octet of valence electrons.

- Rules for writing Lewis structures for a molecule:
    1. Include all the valence electrons from all atoms. The total # of electrons available is the sum of all the valence electrons from all the atoms in the molecule.
    2. Atoms that are bonded to each other share one or more pairs of electrons.
    3. Electrons are arranged so that each atom is surrounded by enough electrons to fill the valence orbitals of that atom. This means two electrons for hydrogen and eight electrons for second-row nonmetals.

12.7 LEWIS STRUCTURES OF MOLECULES WITH MULTIPLE BONDS

- CO_2 --- there are 16 valence electrons in this structure
- Each oxygen has 8 electrons around it, but the carbon has the unshared pair in between too
- A **single bond** involves two atoms sharing one electron pair
- A **double bond** involves two atoms sharing two pairs of electrons
- A **triple bond** has three electron pairs shared
- **Resonance** is a molecule with more than one Lewis structure

12.8 MOLECULAR STRUCTURE

- **Molecular/ geometric structure**s are the 3D arrangements of the atoms in a molecule
- Water is often called "bent" or "V-shaped" the bond angle is 105°
- The CO_2 molecule has a linear structure and a 180° bond angle
- BF_3 is planar (**trigonal planar**) or flat with bond angles of 120°
- CH_4 has a **tetrahedral structure**

12.9 MOLECULAR STRUCTURE: THE VSEPR MODEL

- **VSEPR model**→ valence shell electron pair repulsion; it is useful for predicting the molecular structures of molecules formed from nonmetals
- The bonding and nonbonding electron pairs around a given atom are positioned as far apart as possible

- $BeCl_2$ has two pairs of electrons around the beryllium atom. The best arrangement places the pairs on opposite sides of the beryllium atom at 180° away from each other to minimize repulsion
- Steps for predicting molecular structure using the VSEPR model:
    1. Draw the Lewis structure for the molecule
    2. Count the electron pairs and arrange them in the way that minimizes repulsions (that is, put the pairs as far apart as possible)
    3. Determine the positions of the atoms from the way the electron pairs are shared
    4. Determine the name of the molecular structure from the positions of the atoms

12.10 MOLECULAR STRUCTURE: MOLECULES WITH DOUBLE BONDS

- CO_2 has two double bonds and is known to be linear, so the double bonds must be at 180° from each other
- Each double bond in this molecule must act effectively as one repulsive unit
- This makes sense if we think of a bond in terms of an electron density "cloud" between two atoms
- In $BeCl_2$ the minimum repulsion between these two electron density clouds occurs when they are on opposite sides of the Be atom
- Each double bond in CO_2 involves the sharing of four electrons between the carbon atom and an oxygen atom
- Each double bond should be treated the same as a single bond, although a double bond involves four electrons, they are restricted to the space between a given pair of atoms
- They are "tied together" to form one effective repulsive unit
- A double bond is counted the same as a single electron pair