UNIT 5: BONDING

CHEMICAL BONDS

A. Definition: A chemical bond is the force holding two atoms together in a chemical compound.

B. Bonds form from the attraction
   1. Between the positive nucleus of one atom and the negative electrons of another, OR
   2. Between a positive ion (cation) and a negative ion (anion).

TYPES OF BONDS

A. Covalent Bonds

1. Definition: the sharing of electrons between two nonmetals.
2. Electronegativity differences < 1.7 result in covalent bonds.
3. Covalent bonding forms covalent compounds known as molecules. Multiple covalent bonds may form between the same two nonmetal atoms.
   a. One shared pair of electrons forms a single covalent bond.
   b. Two shared pairs of electrons form a double covalent bond.
   c. Three shared pairs of electrons shared form a triple covalent bond.
3. The shared electron pairs complete the outer energy level of both atoms involved; **both** atoms become stable by attaining a noble gas configuration.
4. Diatomic molecules are formed when two atoms of the same element (homonuclear) share electrons. Seven elements exist in nature as diatomic molecules rather than as individual atoms.

$\underline{H}_2$ $\underline{O}_2$ $\underline{Br}_2$ $\underline{F}_2$ $\underline{I}_2$ $\underline{N}_2$ $\underline{Cl}_2$

B. Ionic Bonds

1. Definition: An ionic bond is the electrostatic force holding oppositely charged particles (ions) together in an ionic bond.
2. Electronegativity differences > 1.7 result in ionic bonds.

3. Ionic bonds form because of the attraction between cations and anions, resulting from the transfer of electrons from metals to nonmetals.
4. Ions (charged atoms) may be monatomic or polyatomic.
   a. Monatomic: one atom → one element symbol: Li^{+}, Be^{2+}, Br^{-}, S^{2-}
   b. Polyatomic: more than one atom, two or more nonmetals → two or more symbols; the unit acts as one ion and the charge belongs to the group of atoms. See back side of Periodic Table for examples: NH_4^{+}, OH^{-}, NO_3^{-}, SO_4^{2-}, PO_3^{3-}
5. Ionic compounds may be binary or ternary.
   a. Two monatomic ions – a metallic cation and a nonmetallic anion – combine to form a binary ionic compound.
   b. Ternary ionic compounds are formed when a polyatomic ion is involved either as the cation or, more often, as the anion.
   c. The chemical formula for an ionic compound represents the simplest ratio of ions (the simplest ratio of cations to anions) and is known as a formula unit.

Formation of Ions

1. Metals lose electrons to become positively charged cations. Ionization energy is the amount of energy required to remove most loosely held electron from an atom.
   - $Li + \text{energy (in kJ)} \rightarrow Li^{+} + e^{-}$
   - $Be + \text{energy (in kJ)} \rightarrow Be^{2+} + 2e^{-}$
2. Nonmetals gain electrons to become negatively charged anions.
   - $Br + e^{-} \rightarrow Br^{-} + \text{energy (in kJ)}$
   - $S + 2e^{-} \rightarrow S^{2-} + \text{energy (in kJ)}$
3. Oxidation number equals the number of electrons transferred to or from an atom to form its ion and becomes the charge of the monatomic ion. It is used to help determine the ratio of ions in a compound.
   - Oxidation numbers are predicted by the group and are written above the group numbers on the periodic table.
   - Metals have positive oxidation numbers, while nonmetals have negative oxidation numbers.
4. Transition elements and metals under the stairs have variable oxidation numbers. These elements <u>usually</u> form ions with 2^{+} or 3^{+} charges.

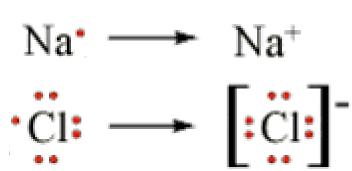
Formation of Ionic Compounds

A. Electron dot (or Lewis dot) diagrams can be used to show the formation of ions and ionic compounds.

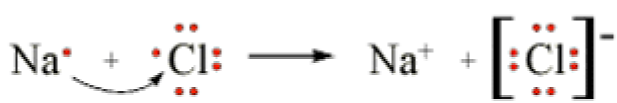
1. Sodium loses one electron to become a cation while chlorine gains one electron to become an anion.

   $Na \rightarrow Na^{+} + e^{-}$

   $Cl + e^{-} \rightarrow Cl^{-}$



2. The electron lost by sodium may be transferred to chlorine, resulting in two charged atoms or ions.



3. The attraction between the opposite charges of the sodium ion and the chloride ion result in the ionic compound sodium chloride.

NaCl

B. Ionic Compounds

1. The overall charge on one formula unit of an ionic compound is zero. The number of electrons lost must equal the number of electrons gained. Oxidation numbers are used to help determine the ratio of cations and anions necessary to form a neutral compound.

2. **Binary Ionic Compounds**

   a. Determine the charges of the ions involved from their oxidation numbers.
   K^+ Cl^-

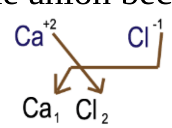
   b. Combine ions so that the overall charge equals zero.
   One 1^+ charge (K^+) are needed to balance one 1^- charge (Cl^-).

   c. Cations are always written first. Use subscripts to show the actual ratio of ions; however, subscripts of one are not written. *No charges appear in the formula*! KCl

   d. Ions do not always form compounds in one-to-one ratios: K^+ and O^{2-}
   Combine ions to form a neutral compound (overall charge of zero).
   Two 1^+ charges (K^+) are needed to balance one 2^- charge (O^{2-}) $\rightarrow K_2O$

   e. Criss-Cross Method

   - The charge of the cation becomes the subscript of the anion, without the positive sign, and the charge of the anion becomes the subscript of the cation, without the negative sign.



$\rightarrow CaCl_2$

The sum of the charges adds to zero
Ca^{2+} Cl^{1-} Cl^{1-} = 0

Practice Problems, p. 221 - #19, #20, #21, #22, #23

3. **Ternary Ionic Compounds**

   a. Determine the monatomic and polyatomic ions involved: Fe^{2+} and NO_3^-

   b. Combine ions so that the overall charge equals zero.
   One 2^+ charge (Fe^{2+}) is balanced by two 1^- charges (NO_3^-).

   c. Write the cation first, and use subscripts to write the formula showing the actual ratio of ions.

   - Parentheses are used when more than one polyatomic ion is needed; parentheses are placed around the polyatomic ion with the subscript outside the closing parenthesis: $Fe(NO_3)_2$
   - Parentheses are <u>never</u> used with monatomic ions or with single polyatomics.

Practice Problems, p. 222 - #24, #25, #26

NAMING IONIC COMPOUNDS

A. Binary Ionic Compounds

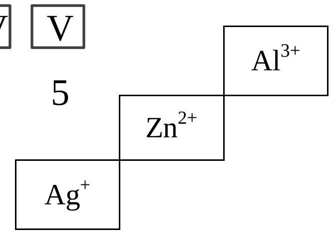
1. Name the cation (metal) using its element name – no changes.
2. Name the anion (nonmetal) with the ending *–ide*.

B. Ternary Ionic Compounds

1. Name the cation.
   a. Monatomic cations: use the element name – no changes.
   b. Polyatomic cations: use the name of the polyatomic ion (on back of PT) – no changes.
2. Name the anion.
   a. Monatomic anions: use the element name with the ending *–ide*.
   b. Polyatomic anions: use the name of the polyatomic ion (on back of PT) – no changes.

C. Metal ions with more than one oxidation number (transition and metals under stairs)

1. Determine the total positive charge needed to form a neutral compound.
2. Stock System: use a Roman numeral in parentheses following the metal name to show the charge for the cation.
3. Name the anion the same way it would be named in any ionic compound.
4. *Examples*
   a. CuCl — total negative charge is 1^- (chlorine has –1 oxidation number)
   in this compound, cation must have a 1^+ charge
   therefore, compound name is copper (I) chloride
   b. $CuCl_2$ — total negative charge is 2^- ($2Cl \times 1^- = 2^-$)
   in this case, cation must have a 2^+ charge
   therefore, compound name is copper (II) chloride
   c. The Roman numeral is the **charge**, <u>not</u> the number of ions involved.
   d. Roman numerals most often used: $\boxed{\text{I}}$ $\boxed{\text{II}}$ $\boxed{\text{III}}$ $\boxed{\text{IV}}$ $\boxed{\text{V}}$
   1 2 3 4 5
   e. *Exceptions*: Ag^+ Zn^{2+}



Practice Problems, p. 223 - #28, #29, #30, #31, #32, #33

PROPERTIES OF IONIC COMPOUNDS

A. Ionic compounds are known as oxides when metals bond to the nonmetal oxygen. Most other ionic compounds are referred to as salts.

B. The formation of ionic bonds is always exothermic, meaning the process releases energy. The ionic compound formed is more stable than the individual ions.

C. Positive and negative ions alternate in a repeating pattern that balances the forces of attraction and repulsion and forms a three-dimensional arrangement of ions called a crystal lattice.

D. Ionic bonds are relatively strong due to the strength of attraction between ions. Since a large amount of energy is necessary to break these compounds apart, they have high melting and boiling points.

E. Ionic compounds are hard, rigid, and brittle solids due to the strong attractive forces between ions. When an external force strong enough to overcome the attraction is applied, the ions in the crystal shift and like-charged ions are repositioned next to each other; resulting repulsive forces break the crystal apart.

F. Many crystals have brilliant colors due to the presence of transition metals in their crystal lattices.

G. Ionic compounds are excellent conductors of electricity when molten (liquid) or aqueous (dissolved in water) but not in the solid state.

- electrolyte: compound whose aqueous solution conducts electricity

H. Lattice energy is defined as the amount of energy required to separate one mole of an ionic compound. The greater the lattice energy, the stronger the attractive force between ions.

- Strength of attraction between ions increases as the distance between ions decreases. Smaller ions produce stronger attractions and greater lattice energies.
- Attraction of ions with larger positive or negative charges is generally associated with greater lattice energy.

METALLIC BONDS

A. In the solid state, metals form lattices in which the outer energy levels overlap.
B. The metal atoms contribute their valence electrons to form a sea of delocalized electrons.
C. A metallic bond is the attraction of a metallic cation for these delocalized electrons.