CHAPTER NOTES – CHAPTER 15

Ionic Bonding and Ionic Compounds

Goals: To gain an understanding of:

1. Valence electron and electron dot notation.
2. Stable electron configurations.
3. Ionic and metallic bonding.

NOTES:

Valence electrons are the electrons in the highest energy level of an atom. For example, in the calcium atom (electron configuration 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)4s\(^2\)) the 4s\(^2\) electrons are the valence electrons. In the titanium atom (electron configuration 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)4s\(^2\)3d\(^2\)) The 4s\(^2\) electrons are still the valence electrons - they are in the highest energy level. In the phosphorus atom (electron configuration 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^3\)) the 3s\(^2\)3p\(^3\) are the valence electrons.

The valence electron numbers of the representative elements are:

- Group 1 - 1 valence electron
- Group 2 - 2 valence electrons
- Group 13 - 3 valence electrons
- Group 14 - 4 valence electrons
- Group 15 - 5 valence electrons
- Group 16 - 6 valence electrons
- Group 17 - 7 valence electrons
- Group 18 - 8 valence electrons

Dot formulas use dots surrounding the symbol of the element to represent the valence electrons. The dot formulas for period 2 and 3 would appear as follows.

<table>
<thead>
<tr>
<th>Period</th>
<th>1A</th>
<th>2A</th>
<th>3A</th>
<th>4A</th>
<th>5A</th>
<th>6A</th>
<th>7A</th>
<th>0</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Li</td>
<td>Be</td>
<td>B</td>
<td>C</td>
<td>N</td>
<td>O</td>
<td>F</td>
<td>Ne</td>
</tr>
<tr>
<td>3</td>
<td>Na</td>
<td>Mg</td>
<td>Al</td>
<td>Si</td>
<td>P</td>
<td>S</td>
<td>Cl</td>
<td>Ar</td>
</tr>
</tbody>
</table>

Note electrons are usually shown as far apart as possible - they have the same charge and therefore repel each other.

The Noble gases (Group 0) have a stable electron configuration (s\(^2\)p\(^6\)) with 8 electrons filling the outer s and p orbitals. This stability comes from the low energy state of this configuration and also accounts for the low reactivity of these elements (most elements react with other elements to get to a lower, more stable energy state). For example the halogens (Group 7A) have 7 valence electrons (s\(^2\)p\(^5\)) and want to gain one electron to get the low energy, stable electron configuration of the noble gases. The elements in group 6 (s\(^2\)p\(^3\)) want to gain 2 electrons to get the low energy, stable electron configuration of the noble gases. The Group 1A elements (s\(^1\)) want to lose their outer electron to empty their outer shell and get a stable electron configuration. For example if sodium (1s\(^2\)2s\(^2\)2p\(^6\)3s\(^1\)) loses its 3s\(^1\) electron it will have filled s and p orbitals in its outer energy level.

Gilbert Lewis, in 1916, proposed the octet rule: Atoms react by changing their number of electrons so as to acquire the stable electron configuration of a noble gas (s\(^2\)p\(^6\)).

An exception to the octet rule is the electron configuration of helium. Helium(1s\(^2\)) is a noble gas, only it has only one orbital, the s orbital. It is filled and therefore stable and elements close to it (lithium, beryllium and sometimes hydrogen) try to acquire its electron configuration by losing or gaining electrons.)
The pseudo noble-gas electron configuration has the outer three orbitals filled, the s, p and d- \( s^2 p^6 d^{10} \) (18 electrons total) and so is fairly stable. Elements that attain this electron configuration are at the right side of the transition metals (d-block).

An ion is a charged atom that is formed by the gaining or losing of electrons. When an atom gains or loses electrons it is no longer neutral because the number of electrons (negative charges) and protons (positive charges) are not equal. For example when a sodium atom loses an electron to get the noble gas electron configuration of neon it gets a charge of +1 and becomes the sodium ion, because it now has 11 protons (+) and only 10 electrons (-). When a chlorine atom gains one electron to get the noble gas electron configuration of argon it gets a charge of -1 and becomes the chloride ion, because it has 17 protons (+) and 18 electrons (-).

Ionic compounds are compounds formed when the elements bond together with ionic bonds - bonds formed by the electrostatic charges (+ and -) formed when there is a transfer of electrons from one element to another.

Atoms bond together in order to achieve a lower, more stable electron configuration (noble gas or pseudo noble-gas electron configurations). They can do this by transferring electrons from one element to another, creating ions which then bond together due to their electrostatic charges. This type of bond is called an ionic bond because it involves ions. This type of bonds occurs between metals (have few electrons in their outer energy level so they want to get rid of those electrons) and nonmetals (have many electrons in their outer energy level so they want to take in more electrons to get the noble gas electron configuration).

A common example of an ionic bond is between sodium (a metal) and chlorine (a nonmetal). Sodium (1s\(^2\)2s\(^2\)2p\(^6\)3s\(^1\)) needs to lose one electron (the 3s\(^1\)) to achieve the noble gas electron configuration of neon. Chlorine (1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^5\)) needs to gain one electron to get the noble gas electron configuration of argon. A Bohr model of this bond would look like:

![Bohr model of NaCl bond]

An electron dot representation of this electron transfer and bond would look like:

![Electron dot representation of NaCl]

A quantum mechanical representation of the electron transfer would look like:

![Quantum mechanical representation of NaCl]

A common name for ionic compounds is "salt." We tend to think of salt as table salt, or sodium chloride. This is just one example of a very large class of compounds called salts. While all salts are an ionic compound not all ionic compounds are salts. Acids and bases are also ionic compounds. (The more technical definition of a salt is an ionic compound formed from the negative ion of an acid and the positive ion of a base).
Common characteristics of ionic compounds are:

- bond is formed by electrostatic attraction of opposite charges
- crystalline solids at room temperature
- brittle
- electrolytes - conduct electricity when dissolved or melted

Metal atoms tend to lose electrons to become cations. Metallic bonds are bonds between metallic cations. These bonds are formed by the mutual attraction the cations have for free floating electrons. It may be pictured as follows.

This model explains the following properties of metals:

- metals are malleable - can be hammered into a shape as the cations slide past each other
- metals are ductile - can be drawn out into a wire as cations slide past each other
- metals are good conductors of electricity - the weakly held electrons are able to flow

When a hammer strikes a malleable metal the cations will slide past each other and still be bonded together by their mutual attractions for the free floating electrons. When a hammer strikes an ionic crystal it forces the like charged ions to come into close proximity. This creates a repulsive force which may shatter the crystal.

An alloy is a solid solution formed by dissolving one metal in another metal. Three types of alloys are:

- substitutational alloys - this is an alloy between two metals whose cations are similar in size an example here is sterling silver - an alloy of copper and silver. It may be pictured as follows (actual ratio is 7.5% Cu and 92.5% Ag):

- interstitial alloys - this is an alloy between two metals which have very different sized cations which can fit into the interstitial spaces of another metal. An example here is steel, an alloy of iron and carbon. It can be pictured as (actual ratio is 99% Fe : 1% C):

- amalgams - alloys of mercury, for example dental amalgam (tooth fillings) is an alloy of mercury, zinc and silver
Ionic Compounds

Introduction to Chemical Bonding
A. Ionic Bonding
1. Chemical bonding that results from the electrical attraction between cations and anions
2. Electrons are transferred in pure ionic bonding and ions are held together by electrostatic forces of attraction
B. Covalent Bonding
1. Results from the sharing of electron pairs between two atoms (we will cover this later)

Ionic Bonding and Ionic Compounds

I. Introduction
A. Ionic Compounds
1. A compound composed of positive and negative ions held together by electrostatic forces of attractions and are combined so that the numbers of positive and negative charges are equal
   a. Most are crystalline solids
   b. Examples include NaCl, MgBr₂, Na₂O  (a metal ion + a nonmetal ion)
   c. Any chemical formula that begins with a metal atom is an ionic compound
B. Formula Unit
1. The simplest ratio of atoms from which an ionic compound's formula can be established (CaO  NOT Ca₂O₂)

II. Formation of Ionic Compounds
A. Electron Configuration Changes
1. Electrons are transferred from the highest energy level of the metal atom to the highest energy level of the nonmetal atom, creating noble gas configurations in all atoms involved
2. Formation of sodium chloride
   a. Na = 3s¹ and Cl = 3s²3p⁵
   b. Na⁺ = 2s²2p⁶ and Cl⁻ = 3s²3p⁶
B. Characteristics of Ionic Bonding
1. Lattice Energy
   a. The energy released when one mole of an ionic crystalline compound is formed from gaseous ions
      \[ \text{Na}(g) + \text{Cl}(g) \rightarrow \text{NaCl}(s) + 787.5 \text{ kJ} \]
   b. Formation of ionic compounds is ALWAYS exothermic (releases heat energy) and forms a stable compound
2. Lattice Structure
   a. Three dimensional arrangements vary depending upon the sizes and charges of the ions

NaCl

Ionic compounds are brittle solids that are usually crystalline in nature. They conduct electricity when the ions are free to move, when they are dissolved or melted. Ionic bonds are very strong and the melting points of ionic compounds are high.

C. A Comparison of Ionic and Molecular Compounds

<table>
<thead>
<tr>
<th></th>
<th>Ionic Compounds</th>
<th>Molecular Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting point</td>
<td>Generally high</td>
<td>Generally low</td>
</tr>
<tr>
<td>Boiling Point</td>
<td>Generally high</td>
<td>Generally low</td>
</tr>
<tr>
<td>Electrical Conductivity</td>
<td>Excellent conductors, molten and aqueous</td>
<td>Poor conductors, except aqueous acids</td>
</tr>
<tr>
<td>Solubility in water</td>
<td>Generally soluble</td>
<td>Polar covalent molecules are soluble</td>
</tr>
</tbody>
</table>
Metallic Bonding

I. The Metallic Bond Model
   A. Metallic Bonding
      1. The chemical bonding that results from the attraction between metal atoms and the surrounding sea of electrons (the electron sea model)
   B. Electron Delocalization in Metals
      1. Vacant $p$ and $d$ orbitals in metal’s outer energy levels allow outer electrons to move freely throughout the metal
      2. Valence electrons do not belong to any one atom, they are free to move (movement of electrons is electrical current)

II. Metallic Properties
   A. Metals are good conductors of heat and light (electrons are free to move)
   B. Metals are shiny
      1. Narrow range of energy differences between orbitals allows electrons to be easily excited, and emit light upon returning to a lower energy level
   C. Metals are Malleable
      1. Can be hammered into thin sheets (atoms can flow over and around each other and not break)
   D. Metals are ductile
      1. Ability to be drawn into wire
         a. Metallic bonding is the same in all directions, so metals tend not to be brittle