Chapter 5: Atoms, Bonding, and the Periodic Table

Valence electrons and bonding
Valence electrons – electrons in the highest energy level (outermost electron shell) and are held most loosely
The number of valence electrons in an atom determines its chemical and physical properties
Lewis Electron Dot (LED) diagram (or formula) – a method used to depict the number of valence electrons in an atom which includes the symbol for the electron and a number of dots which represent electrons

Chemical bonds and stability
Atoms of most elements are more stable when they have eight valence electrons
H and He are most stable with only two valence electrons
Atoms tend to react in ways that make them more stable
Metals tend to lose valence electrons
Nonmetals tend to gain electrons until their valence shell contains eight electrons
(except H and He which are most stable with two valence electrons)
Chemical bond – the force of attraction that holds two atoms together as a result of a rearrangement of electrons between them
Covalent bond – results from sharing electrons to form a chemical bond
Covalent bonds form when two nonmetals share electrons to form a chemical bond
Ionic bond – results from transferring electrons to create a chemical bond
Ionic bonds form when a metal and a nonmetal transfer electrons to form a chemical bond

How the Periodic Table Works
The periodic table gives information about the arrangement of electrons in atoms.
The period number (along the left edge) tells the number of energy levels (electron shells) the atoms of elements in that period will contain
The older American chemistry group numbers tell the number of valence electrons in an atom of an element
The elements within a group have similar properties because they all have the same number of valence electrons in their atoms

<table>
<thead>
<tr>
<th>Group</th>
<th>alkali metals</th>
<th>alkaline earth metals</th>
<th>boron</th>
<th>carbon</th>
<th>nitrogen</th>
<th>oxygen</th>
<th>halogen</th>
<th>noble gases</th>
</tr>
</thead>
<tbody>
<tr>
<td>New</td>
<td>1</td>
<td>2</td>
<td>3</td>
<td>4</td>
<td>5</td>
<td>6</td>
<td>7</td>
<td>8</td>
</tr>
<tr>
<td>Old</td>
<td>IA</td>
<td>IIA</td>
<td>IIIA</td>
<td>IVA</td>
<td>VA</td>
<td>VIA</td>
<td>VIIA</td>
<td>VIIIA</td>
</tr>
<tr>
<td>LED</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>Name</td>
<td>lithium</td>
<td>beryllium</td>
<td>boron</td>
<td>carbon</td>
<td>nitrogen</td>
<td>oxygen</td>
<td>fluorine</td>
<td>neon</td>
</tr>
</tbody>
</table>

| LED   | Na            | Mg                    | Al    | Si     | P        | S      | Cl      | Ne         |
| Name  | sodium        | magnesium             | aluminum | silicon | phosphorus | sulfur | chlorine | argon      |

Noble gases – atoms of these elements already have 8 valence electrons (except for H which has 1 and He which has 2 valence electrons) and it is very difficult to cause these elements to react
Reactive metals and nonmetals

Metals tend to lose electrons and become positive ions and become more reactive to the left and toward the bottom of the periodic table.
- Alkali metals – the most reactive metals on the periodic table are in group IA (1)
- Alkaline earth metals – group IIA (2) metals are also very reactive

Nonmetals tend to gain electrons and become negative ions and become more reactive to the right and toward the top of the periodic table.
- Noble gases – group VIIIA (18) the noble gases which tend not to react are exceptions
- The halogens – the most reactive nonmetals on the periodic table are in group VIIA (17)
  - Fluorine (F) is the most reactive element on the periodic table and gains 1 electron
  - Oxygen (O) is the next most reactive nonmetal and tends to gain 2 electrons
- Hydrogen (H) belongs in group IA because it has only 1 electron in its valence shell but because the first valence shell can hold only 2 electrons total, H acts as a nonmetal

If nonmetals react with other nonmetals, they tend to share electrons instead of gaining or losing electrons.

Metalloids

The elements that lie along the zigzag line on the periodic table (except Al which is a metal) are metalloids and can gain, lose, or share electrons with other elements behaving like either metals or nonmetals depending on what other element with which they react.

Chemical bonds

All chemical bonds occur with the positive nuclei of two atoms are attracted to electrons (usually a pair of electrons) between them (see to the right).

Ionic bonds – occur when electrons are transferred from a metallic to a nonmetallic element:

\[
\text{Na} .\quad +\quad \text{Cl} :\quad \rightarrow\quad \text{Na}^{1+}\quad :\quad \text{Cl}^{1–}
\]

\text{sodium} \quad \text{chlorine} \quad \text{sodium ion} \quad \text{chloride ion}

Covalent bonds – occur when electrons are shared between two nonmetallic elements:

\[
\text{Cl} .\quad +\quad \text{Cl} :\quad \rightarrow\quad \text{Cl} :\quad \text{Cl} :
\]

\text{chlorine atom} \quad \text{chlorine atom} \quad \text{chlorine molecule}

Ionic bonds – form when atoms transfer electrons.

Atoms that lose electrons (1– charge) will form positive ions called cations.
Atoms that gain electrons will form negative ions called anions.

Polyatomic ions – ions that are made up of more than one atom.

Polyatomic ions are a group of atoms held together by covalent bonds that as a group gain or lose electrons.
- Positively charged polyatomic ions are rare
  - Example: ammonium ion \( \text{NH}_4^{1+} \)
- Negatively charged polyatomic ions are more common
  - Examples:
    - bicarbonate ion \( \text{HCO}_3^{1–} \)
    - carbonate ion \( \text{CO}_3^{2–} \)
    - nitrate ion \( \text{NO}_3^{1–} \)
    - phosphate ion \( \text{PO}_4^{3–} \)
    - sulfate ion \( \text{SO}_4^{2–} \)
Ionic bonds occur when electrons are transferred from a metallic to a nonmetallic element. 

\[
\text{Na}^+ + \text{Cl}^- \rightarrow \text{Na}^1+ : \text{Cl}^- 1^- \\
\text{sodium} \quad \text{chlorine} \quad \text{sodium ion} \quad \text{chloride ion}
\]

Formulas of ionic compounds
Comounds like NaCl which are held together by ionic bonds are called **ionic compounds**

The charges of the ions in an ionic compound must add up to a neutral charge

**Example:** \(\text{Na}^1+ + \text{Cl}^- = \text{NaCl}, \text{a neutral compound}\)

But what about ions with charges other than 1+ or 1–?

From the periodic table, we know that Mg atoms have 2 valence electrons

We also know that Mg is an alkaline earth metal and will form a 2+ ion

But, that means that the charges of one \(\text{Mg}^2+\) cation will not add up to a neutral charge when combined with one \(\text{Cl}^1–\) anion

The solution to this problem is to combine one one \(\text{Mg}^2+\) cation with two \(\text{Cl}^1–\) anions

**Example:** \(\text{Mg}^2+ + 2 \text{Cl}^- = \text{MgCl}_2, \text{a neutral compound}\)

Notice that the chloride anion required a subscript of 2 to make the formula represent a neutral compound.

Atoms in formulas without any subscripts have an understood but unwritten subscript of 1

**Naming ionic compounds**
The name of the positive ion comes first

For monatomic ions, this is just the element name

**Example:** \(\text{Mg}^2+ + 2 \text{Cl}^- = \text{MgCl}_2, \text{a neutral compound}\)

For polyatomic ions, the name must be learned or looked up on a chart

**Properties of ionic compounds**
Physical properties of ionic compounds

Brittle and hard

Nonconductors in the solid state

The ions cannot move in an ionic crystal therefore they do not conduct electricity

Conductors if melted or dissolved in water

The ions are free to move in the liquid phase or in solution so they can conduct electricity

High melting points and boiling points

The ions are firmly surrounded (see to the right) by ions of the opposite charge and much energy is needed to break all the bonds to melt the crystal

Covalent bonds – form when atoms share electrons

One pair of shared electron forms a single bond (see the \(\text{F}_2\) molecule at the right)

Oxygen has six valence electrons leaving two unpaired electrons and therefore can form two single bonds with two hydrogen atoms (see the \(\text{H}_2\text{O}\) molecule at the far right)

One pair of shared electron forms a single bond (see the \(\text{F}_2\) molecule at the right)
Valence electrons and the number of covalent bonds that can form

<table>
<thead>
<tr>
<th>Name</th>
<th>Hydrogen</th>
<th>Beryllium</th>
<th>Boron</th>
<th>Carbon</th>
<th>Nitrogen</th>
<th>Oxygen</th>
<th>Fluorine</th>
<th>Neon</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bonds</td>
<td>1</td>
<td>2</td>
<td>3</td>
<td>4</td>
<td>3</td>
<td>2</td>
<td>1</td>
<td>0</td>
</tr>
</tbody>
</table>

**Multiple covalent bonds**

The element oxygen can also bond two atoms to form a molecule (see below)

(Note the double bond with two shared pairs of electrons.)

The element nitrogen (5 valence e⁻) can bond two atoms to form a molecule (see below)

(Note the triple bond.)

Molecule – a neutral group of atoms joined by covalent bonds

Molecular compound – a compound containing atoms that are covalently bonded

Molecules of compounds can also form multiple covalent bonds (see below)

(Note there are two double bonds in CO₂.)

**Polar and nonpolar covalent bonds**

Polar covalent bonds result when electrons are shared unequally
Nonpolar covalent bonds result when electrons are shared evenly

**Polar and nonpolar molecules**

Typically, nonpolar covalent bonds will form nonpolar molecules
Polar covalent bonds that are arranged asymmetrically form polar molecules
Polar covalent bonds that are arranged symmetrically form nonpolar molecules

**Attraction between molecules**

Two polar molecules can attract by arranging such that the positive and negative ends alternate (opposites attract)
Nonpolar molecules can form a much weaker van der Waals attraction caused by a temporary uneven distribution of e⁻

**Physical properties of molecular compounds**

Brittle and not typically as hard as ionic compounds
Nonconductors in the solid state
The molecules cannot move in the solid therefore they do not conduct electricity
Nonconductors if melted or dissolved in water
The charges in molecules are not free to move in separate directions even in the liquid phase or in solution so they cannot conduct electricity
Low melting points and boiling points
The bonds between molecules are much weaker than ionic, covalent, and metallic bonds therefore much less heat is needed to break the bonds to melt the crystal
Network solids – a solid composed of a three dimensional network of covalent bonds

Physical properties of network solids
Brittle but unusually hard
Nonconductors in the solid or liquid states
High melting points and boiling points
Sometimes decompose instead of melting

Metallic bonds – form when the valence electrons become free to move throughout the lattice of positive metal ions

The properties of metals are due to this ‘sea’ of freely moving electrons
This sea of electrons is the defining characteristic of metallic bonds in both pure metals and in mixtures of metals called alloys
Alloys can retain many of the properties of the original metals, but they can also have both chemical and physical properties that vary greatly from the alloyed metals

Physical properties of metals
Malleable and ductile – the positive ions can be pushed out of position but the sea of freely moving electrons can quickly move to re-form the metallic bond between the positive ions
Nonconductors in both the solid and liquid state
The electrons can freely move throughout the positive ions to conduct heat and electricity
Very high melting points and boiling points
The positive ions are closely surrounded (see to the right) by ions of the very small electrons (the shorter a chemical bond, the stronger it will be)
Very shiny with reflective luster

<table>
<thead>
<tr>
<th>Property</th>
<th>Metals</th>
<th>Ionic Compounds</th>
<th>Molecular Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>State (room T)</td>
<td>Solids</td>
<td>Solids</td>
<td>Solids to gases</td>
</tr>
<tr>
<td>Melting points</td>
<td>Very high</td>
<td>High</td>
<td>Very low</td>
</tr>
<tr>
<td>Boiling points</td>
<td>Extremely high</td>
<td>Very high</td>
<td>Low</td>
</tr>
<tr>
<td>Conducts electricity</td>
<td>Solid – yes</td>
<td>Solid – no</td>
<td>Solid – no</td>
</tr>
<tr>
<td></td>
<td>Liquid – yes</td>
<td>Liquid – yes</td>
<td>Liquid – no</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Aqueous – yes</td>
<td>Aqueous – no</td>
</tr>
<tr>
<td>Luster</td>
<td>Shiny and reflective</td>
<td>Shiny, not reflective</td>
<td>Shiny, not reflective</td>
</tr>
<tr>
<td>Malleable / ductile</td>
<td>Yes / yes</td>
<td>No, hard and brittle</td>
<td>No, brittle to soft</td>
</tr>
</tbody>
</table>