Chapter 9, Part 1  
Models of Chemical Bonding

Recall Chapter 2: Chemical bonds hold atoms together in a compound.

- transfer of electrons, forming cations and anions, results in ionic bonding
- sharing of electron pairs results in covalent bonding

One additional type of bonding: pooling of bulk quantity of electrons by many atoms: metallic bonding

All three types of bonding are electrostatic (attraction between + and – charges) at their hearts.

Lewis Dot Symbols  A way to depict valence electrons.
e.g., for S: [Ne]3s\textsuperscript{2}3p\textsuperscript{4} for a total of 6 valence electrons  [S is in Group 6A(16)]

The Octet Rule

**ATOMS TEND TO GAIN, LOSE, or SHARE ELECTRONS to ATTAIN A FILLED OUTER SHELL of 8 ELECTRONS.**

Covalent Bonding

A covalent bond consists of a pair of electrons shared by two atoms.

Fig. 9.12, energy diagram of H\textsubscript{2} bond formation

balance of attractions & repulsions
bond length: distance between nuclei of bonded atoms. Reported as averages just like bond enthalpies.

bond order: number of electron pairs shared between two bonded atoms

BO = 1

BO = 2

BO = 3

For a given pair of atoms:
The HIGHER the bond order, the SHORTER the bond, and the HIGHER the bond energy.

<table>
<thead>
<tr>
<th>Bond</th>
<th>Bond Order</th>
<th>Average Bond Length (pm)</th>
<th>Average Bond Energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C—O</td>
<td>1</td>
<td>143</td>
<td>358</td>
</tr>
<tr>
<td>C=O</td>
<td>2</td>
<td>123</td>
<td>745</td>
</tr>
<tr>
<td>C≡O</td>
<td>3</td>
<td>113</td>
<td>1070</td>
</tr>
<tr>
<td>C—C</td>
<td>1</td>
<td>154</td>
<td>347</td>
</tr>
<tr>
<td>C≡C</td>
<td>2</td>
<td>134</td>
<td>614</td>
</tr>
<tr>
<td>C≡C</td>
<td>3</td>
<td>121</td>
<td>839</td>
</tr>
<tr>
<td>N—N</td>
<td>1</td>
<td>146</td>
<td>160</td>
</tr>
<tr>
<td>N≡N</td>
<td>2</td>
<td>122</td>
<td>418</td>
</tr>
<tr>
<td>N≡N</td>
<td>3</td>
<td>110</td>
<td>945</td>
</tr>
</tbody>
</table>

Properties of Covalent Compounds

Covalent bonds WITHIN molecules: STRONG
Intermolecular forces BETWEEN molecules: WEAK

Covalent molecular substances:
- Poor electrical conductors (even molten or aqueous)
- Relatively low melting points

Quartz, graphite, diamond
- are examples of “network covalent solids”.
- don’t exist as individual molecules.
Electronegativity

Electronegativity expresses the ability of a bonded atom to attract electrons to itself.

![Electronegativity Table](image)

**Figure 9.21**

Pauling scale: F assigned EN = 4.0

**Bond Polarity**

Unequal sharing of electrons in a covalent bond: “polar covalent bond” e.g., H–F

ΔEN tells us how ionic or covalent a bond is.

<table>
<thead>
<tr>
<th>ΔEN</th>
<th>IONIC CHARACTER</th>
</tr>
</thead>
<tbody>
<tr>
<td>&gt;1.7</td>
<td>Mostly ionic</td>
</tr>
<tr>
<td>0.4–1.7</td>
<td>Polar covalent</td>
</tr>
<tr>
<td>&lt;0.4</td>
<td>Mostly covalent</td>
</tr>
<tr>
<td>0</td>
<td>Nonpolar covalent</td>
</tr>
</tbody>
</table>

**PROTIP**

Being the most electronegative, F always has ON of –1 in compounds. Also, we can now see why H has ON +1 with nonmetals and –1 with metals, based on its EN of 2.1 (lower than that of other nonmetals; higher than that of metals and boron). In a bond, the more electronegative atom carries the more negative ON.
Chapter 9, Part 2  Ionic and Metallic Bonding

Ionic Bonding

Let’s start with the formation of a pair of ions (anion + cation), e.g., \( \text{Li}^+ + \text{F}^- \). Figure 9.5

(remember that elemental fluorine exists as \( F_2 \))

Important factors: IE, EA, stabilization of attractions of opposite charges

IE\(_1\) of Li + EA of F is endothermic total. The ionic bond (the attraction of the opposite charges of the cations and anions) is what makes the compound stable.

**Lattice Energy:** energy associated with separating 1 mol of ionic crystal into gaseous ions.

e.g.,

Lattice energy is related to melting point, hardness, solubility of ionic compounds.

Lattice energy is proportional to the charge on the ions and inversely proportional to the size of the ions (i.e., as the ions get smaller, the lattice energy increases).

**Periodic Trends in Lattice Energy**

**Lattice Energies of Alkali Metal Halides** (kJ·mol\(^{-1}\))

<table>
<thead>
<tr>
<th></th>
<th>F(^-)</th>
<th>Cl(^-)</th>
<th>Br(^-)</th>
<th>I(^-)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li(^+)</td>
<td>1050</td>
<td>853</td>
<td>807</td>
<td>757</td>
</tr>
<tr>
<td>Na(^+)</td>
<td>923</td>
<td>787</td>
<td>747</td>
<td>704</td>
</tr>
<tr>
<td>K(^+)</td>
<td>821</td>
<td>715</td>
<td>682</td>
<td>649</td>
</tr>
<tr>
<td>Rb(^+)</td>
<td>785</td>
<td>689</td>
<td>660</td>
<td>630</td>
</tr>
<tr>
<td>Cs(^+)</td>
<td>740</td>
<td>659</td>
<td>631</td>
<td>604</td>
</tr>
</tbody>
</table>

**Lattice Energies of Compounds of the OH\(^-\) and O\(^2-\) Ions** (kJ·mol\(^{-1}\))

<table>
<thead>
<tr>
<th></th>
<th>OH(^-)</th>
<th>O(^2-)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na(^+)</td>
<td>900</td>
<td>2,481</td>
</tr>
<tr>
<td>Mg(^{2+})</td>
<td>3,006</td>
<td>3,791</td>
</tr>
<tr>
<td>Al(^{3+})</td>
<td>5,627</td>
<td>15,916</td>
</tr>
</tbody>
</table>

PROTIP

You would never be able to directly measure lattice energy in a constant-pressure calorimeter in 3150:152 lab.
Properties of Ionic Compounds
- Solids non-conductors of electricity
- Molten/aqueous: conductors
- Hard, rigid, brittle, high melting points

Metallic Bonding
Valence electrons are shared among all of the atoms in the metal—a “sea of electrons”. The array of atoms is regular, but not rigid. The electrons are mobile.

Properties of Metals
- MPs not super-high—need not break attractions between cations and electrons.
- BPs very high—must break attractions between cations and electrons.
- Ductile, malleable—cations able to move though electron sea

 PROTIP
Hammers are not one cation wide in general.

 PROTIP
Same comment about hammers as above.