Models of Chemical Bonding
- Bonds are forces holding atoms or ions together
- Bonds form as a result of lowering of the total energy (energy of separated species is higher than that of bonded species)

9.1 Types of Bonding
- Bond formation is accompanied by rearrangement of valence electrons
- Complete transfer of electrons between metals (low I, low A) and nonmetals (high I, high A)
  - Formation of ions → ionic bonding
  - Electrostatic attraction between oppositely charged ions
- Sharing of electrons between nonmetals (high I, high A)
  - Formation of molecules → covalent bonding
  - Attraction between the nuclei and the shared electrons
  - The shared electrons are localized between the bonded atoms
- Sharing of electrons between metals (low I, low A)
  - Formation of metallic solids → metallic bonding
  - Attraction between metal cations and a “sea” of shared electrons
  - The shared electrons are delocalized in the entire volume of the metal

Lewis Symbols for Atoms and Ions
- Lewis symbol → chemical symbol + a dot for each valence electron
  - For metals, the # of dots equals the max. # of e−s lost in cation formation
  - For nonmetals, the # of unpaired dots equals the # of e−s gained in anion formation or the # of covalent bonds the element forms
9.2 The Ionic Bonding Model

- **The octet rule** – when atoms bond, they gain, lose, or share electrons in order to attain an **octet** (eight) or a **duplet** (two) configuration of a noble gas.
  - Most *s-* and *p-* block metals form cations by loosing all valence electrons (loosing all dots).
    - *s-* block metals achieve the electron configuration of the previous noble gas; *p-* block metals achieve a pseudo-noble gas electron configuration.
  - Nonmetals form anions by gaining electrons until they reach the configuration of the next noble gas.

**Example:**
Predict the formula of magnesium chloride using Lewis structures.

\[ \text{Mg}^{2+} + 2\text{Cl}^- \rightarrow \text{MgCl}_2 \]

Lattice Energy (Lattice Enthalpy)

- **Lattice enthalpy** \( (\Delta H_L) \) – the enthalpy change for the separation of 1 mol of an ionic compound into isolated gaseous ions.

\[
\text{MX(s)} \rightarrow \text{M}^+(g) + \text{X}^-(g) \quad \Delta H_L > 0
\]

- At constant pressure the lattice enthalpy is numerically equal to the heat of formation of one mol of the ionic compound from gaseous ions.

\[
\text{M}^+(g) + \text{X}^-(g) \rightarrow \text{MX(s)} \quad q_p = \Delta H_L < 0
\]

- The heat released in the formation comes from the potential energy drop due to the attraction between the oppositely charged ions.

\[
E_p \propto \frac{q_1 q_2}{r_{12}}
\]

- Potential energy of interaction between two ions with charges \( q_1 \) and \( q_2 \) separated by a distance \( r_{12} \).

\( \Rightarrow \) The lattice enthalpy increases with increasing the charge of the ions and decreasing the distance between them (decreasing the size of the ions).

- The charge factor is more dominant.
- The size factor becomes important only when comparing ionic compounds with equivalent ionic charges.
Examples:
Lattice enthalpies in kJ/mol:

\[
\begin{align*}
\text{LiCl} & \rightarrow 861 \\
\text{NaCl} & \rightarrow 787 \\
\text{KCl} & \rightarrow 717 \\
\text{LiF} & \rightarrow 1050 \\
\text{MgO} & \rightarrow 3923
\end{align*}
\]

- **Size factor** – $\Delta H_L$ decreases moderately with increasing the size of the ion ($\text{Li}^+ < \text{Na}^+ < \text{K}^+$)
- **Charge factor** – $\Delta H_L$ increases greatly with increasing the charges of the ions ($\text{Li}^+, \text{F}^-, \text{Mg}^{2+}, \text{O}^{2-}$)

$Lattice enthalpies are measured indirectly through Hess’s law using the Born-Haber cycle$

Atomization of elements $\rightarrow \Delta H_{at}$ (formation of gas phase atoms)
Ionization (cations) $\rightarrow I$
Ionization (anions) $\rightarrow A$
Formation of solid (from gaseous ions) $\rightarrow -\Delta H_L$
Formation of solid (from elements) $\rightarrow \Delta H_f$

\[
\Delta H_f = \Delta H_{at} + (I + A) + (-\Delta H_L)
\]

Example: Calculate the lattice enthalpy of KBr

\[
\begin{align*}
\Delta H_f &= \Delta H_{at} + (I + A) + (-\Delta H_L) \\
\Delta H_L &= \Delta H_{at} + I + A - \Delta H_f \\
\Delta H_L &= \Delta H_f (\text{K, g}) + \Delta H_f (\text{Br, g}) + I(\text{K}) + A(\text{Br}) - \Delta H_f (\text{KBr, s})
\end{align*}
\]

Data from Appendix B and Figures 8.12 & 8.14:

\[
\begin{align*}
\Delta H_f &= (89) + (112) + (419) + (-325) - (-394) \text{ kJ/mol} \\
\Delta H_L &= 689 \text{ kJ/mol}
\end{align*}
\]

- The Born-Haber cycle shows that the energy required for atoms to lose or gain electrons is supplied by the lattice energy of ionic solids

The Properties of Ionic Compounds

- Ionic solids are crystalline solids (regular three-dimensional arrays of stacked ions)
  - High melting and boiling points – very strong attractions between the ions (hard to separate)
  - Hard, rigid and brittle
  - Do not conduct electricity in the solid state, but conduct electricity when melted or dissolved (electrolytes)