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***) and nonmetals (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

Period	1A(1)	2A(2)	3A(13)	4A(14)	5A(15)	6A(16)	7A(17)	8A(18)
	ns^1	ns^2	ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6
2	Li	Be	B	C	N	O	F	Ne
3	Na	Mg	Al	Si	P	S	Cl	Ar

  - 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

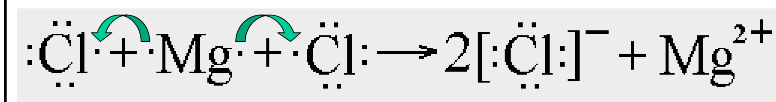
- Electrons lost by the metal are gained by the nonmetal
- Both positive and negative ions reach **octet** (or **duplet**) electron configurations

Example:

Predict the formula of magnesium chloride using Lewis structures.

Mg – group 2 → 2 valence e$^-$ → loss of 2 e$^-$

Cl – group 17 → 7 valence e$^-$ → gain of 1 e$^-$



Formula: $MgCl_2$

Lattice Energy (Lattice Enthalpy)

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

$$MX(s) \rightarrow M^+(g) + X^-(g) \qquad \Delta H_L > 0$$

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

$$M^+(g) + X^-(g) \rightarrow MX(s) \qquad q_p = -\Delta H_L < 0$$

- 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}

⇒ 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:

LiCl	→	**861**
NaCl	→	**787**
KCl	→	**717**

LiF	→	**1050**
MgO	→	**3923**

~x4

Size factor – ΔH_L decreases moderately with increasing the size of the ion ($Li^+ < Na^+ < K^+$)

Charge factor – ΔH_L increases greatly with increasing the charges of the ions (Li^+, F^- Mg^{2+}, O^{2-})

- Lattice enthalpies are measured indirectly through Hess's law using the **Born-Haber cycle**

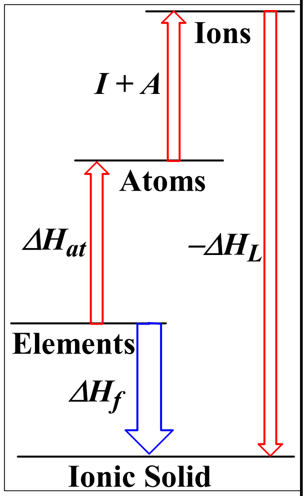
Atomization of elements → ΔH_{at}
(formation of gas phase atoms)

Ionization (cations) → I

Ionization (anions) → A

Formation of solid → $-\Delta H_L$
(from gaseous ions)

Formation of solid → ΔH_f
(from elements)



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

Example: Calculate the lattice enthalpy of KBr

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

$$\Delta H_L = \Delta H_{at} + I + A - \Delta H_f$$

$$\Delta H_L = \Delta H_f(\text{K, g}) + \Delta H_f(\text{Br, g}) + I(\text{K}) + A(\text{Br}) - \Delta H_f(\text{KBr, s})$$

Data from Appendix B and Figures 8.12 & 8.14:

$$\Delta H_L = (89) + (112) + (419) + (-325) - (-394)\ \text{kJ/mol}$$

$$\Delta H_L = 689\ \text{kJ/mol}$$

- 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)