Chemical Bonds

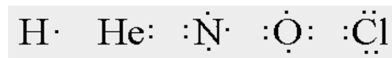
- Forces holding atoms or ions together
- Bonds form as a result of lowering of the total energy (energy of separated species is higher than that of bonded species)
- Bond formation is accompanied by rearrangement of valence electrons
  - complete transfer of electrons – formation of ions (ionic bonding)
  - sharing of electrons – formation of molecules (covalent bonding)

Ionic Bonds

- electrostatic attractions between oppositely charged ions in ionic compounds

8.1 Lewis Symbols for Atoms and Ions

- Lewis symbol – chemical symbol + a dot for each valence electron
  - single dots – single electrons
  - double dots – paired electrons



- 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 (closed shell) **K· → K^+ ↔ [Ar]**
  - **p**-block metals achieve a pseudo-noble gas electron configuration, **[NobleGas](n-1)d^{10}** **:Ga· → Ga^{3+} ↔ [Ar]$3d^{10}$**
- Nonmetals form anions by gaining electrons until they reach the configuration of the next noble gas, ns^2np^6 (closed shell)
- Noble gas configuration – eight valence **e^-** (**octet**), or two valence **e^-** (**duplet**) for **He**

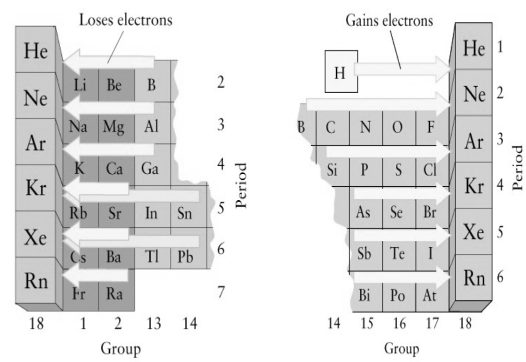


Fig. 8.2, 8.3

- 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^-

$$:\ddot{\underset{..}{Cl}}\cdot + Mg: + :\ddot{\underset{..}{Cl}}\cdot \longrightarrow 2[:\ddot{\underset{..}{Cl}}:]^- + Mg^{2+}$$

8.2 Variable Valence

- **Inert-pair effect** – tendency of the heavier **p**-block metals to form cations two units lower in charge than the expected from their group number
  - the ***np***-electrons have higher energy than the ***ns***-electrons and are lost first
  - the two ***ns***-electrons may or may not be lost

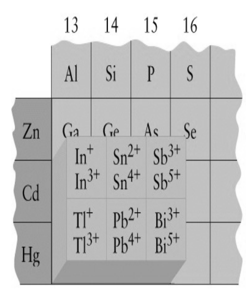


Fig. 8.5

- **Variable valence** – ability to form compounds with different number of bonds (**$InCl_3$, InCl, PbO_2, PbO, $FeCl_3$, $FeCl_2$**)

Example: Use Lewis symbols to write the formulas of **$InCl_3$** and **InCl**.

In → 3 valence electrons (one p-electron and two s-electrons)

In^{3+} or In^{+} and Cl^{-}

$In^{3+}\ 3[:\ddot{\underset{..}{Cl}}:]^{-}$ $\qquad$ $:In^{+}\ [:\ddot{\underset{..}{Cl}}:]^{-}$

8.3 Lattice Enthalpies

- **Lattice enthalpy** (***DH$_L$***)– enthalpy change when a mol of an ionic compound is separated into isolated gaseous ions (positive)

  MX(s) → M^+(g) + X^-(g) $\qquad$ ***DH$_L$*>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) → MX(s) $\qquad$ ***q_p=-DH$_L$*<0**
- The heat released in the formation comes from the potential energy drop due to the attraction between the oppositely charged ions

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

$$E_p \propto \frac{q_1 q_2}{r_{12}}$$

⇒ 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 when comparing ionic compounds with equivalent ionic charges

Table 8.1 Lattice enthalpies at 25°C in kilojoules per mole

Halides							
LiF	1046	LiCl	861	LiBr	818	LiI	759
NaF	929	NaCl	787	NaBr	751	NaI	700.
KF	826	KCl	717	KBr	689	KI	645
AgF	971	AgCl	916	AgBr	903	AgI	887
$BeCl_2$	3017	$MgCl_2$	2524	$CaCl_2$	2260.	$SrCl_2$	2153
		MgF_2	2961	$CaBr_2$	1984		
Oxides							
MgO	3850.	CaO	3461	SrO	3283	BaO	3114
Sulfides							
MgS	3406	CaS	3119	SrS	2974	BaS	2832

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

atomization of elements (formation of gas phase atoms)	***DH$_{at}$***
ionization (cations)	***I***
ionization (anions)	***-A***
formation of solid	***-DH$_L$***
reverse of formation	***-DH$_f$***

***DH$_{at}$* + (*I* - *A*) + (-*DH$_L$*) + (-*DH$_f$*) = 0**

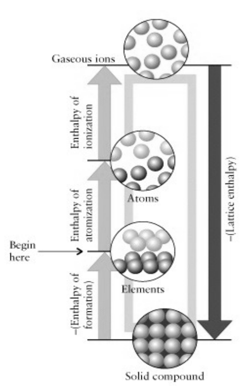


Fig. 8.8

Example: Calculate the lattice enthalpy of KBr.

***DH$_{at}$* + (*I* - *A*) + (-*DH$_L$*) + (-*DH$_f$*) = 0**

***DH$_L$* = *DH$_{at}$* + *I* - *A* - *DH$_f$* =**

= *DH$_f$* (K, g) + *DH$_f$* (Br, g) + *I*(K) - *A*(Br) -
- *DH$_f$*(KBr, s)

Data from Appendix 2A and 2D:

***DH$_L$* = [89 + 112 + 418 - 325 - (-394)] kJ/mol =**
= 688 kJ/mol