Chapter 9: Chemical Bonding I: Lewis Theory

Three types of chemical bonding:

<u>Bond</u>	<u>Atoms</u>	Electron behavior
lonic	Metal + nonmetal	Electrons transferred
Covalent	Nonmetal + nonmetal	Electrons shared
Metallic	Metal + metal	Electrons pooled



Sodium metal, $Na^+(s)$ Copyright © 2008 Pearson Prentice Hall, Inc.

<u>Coulomb's law</u>: $E = k \frac{q_1 q_2}{r}$

- when **opposite** charges are brought closer, potential energy _____
- when like charges are brought closer, potential energy _____

Dot structures and ionic bonding

Recall that we can use dots to show valence electrons - these are called **Lewis electron-dot structures**:

Li Be B C N O F Ne

<u>Ionic bond</u>: attraction of two oppositely-charged ions (recall Coulomb's law)

Lewis structures can be used as a simple way to show formation of ionic bonds.

Na + Cl \rightarrow

Octet rule: main group (s or p block) atoms or ions tend to be stable when they have 8 valence electrons



Lattice energy

Actual formation of an ionic compound is usually very exothermic:

 $Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s); \Delta H_f^o = -410 \text{ kJ/mol}$

...but when we add up the ionization energy of Na and the electron affinity of Cl, that's actually endothermic!

Na \rightarrow Na⁺ + e⁻; IE₁ = +496 kJ/mol Cl + e⁻ \rightarrow Cl⁻; <u>EA = -349 kJ/mol</u> +147 kJ/mol

... so the release of energy does not come from formation of the ions! It actually primarily comes from the <u>lattice energy</u>, the energy associated with forming a crystalline ionic lattice from separated, gaseous ions.







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Periodic trends in lattice energy

<u>lon size</u>:

Compound	Lattice energy
LiCl	-834 kJ/mol
NaCl	
KCI	
CsCl	



Ion charge:

<u>Compound</u>	Lattice energy	Melting point
NaF	-910 kJ/mol	993 °C
CaO		

<u>Coulomb's law</u>: $E = k \frac{q_1 q_2}{r}$

As lattice energy becomes more negative, the ions become more ______ to separate, and melting point _____.

Covalent bonding

Covalent bond:

- pair of shared electrons between two nonmetal atoms, drawn as a line
- what holds the atoms together in a molecule

Formation of H₂ from 2 H atoms:

Formation of H₂O from 2 H atoms and 1 O atom:

Formation of O_2 from 2 O atoms:

Formation of N₂ from 2 N atoms:

Notice the **octet rule** still applies to the main group elements (except H and He - they are stable with only 2 electrons - **duet rule.**)

Always be sure that all the atoms' original valence electrons are represented in the Lewis structure

 C_2H_4 :

Bond polarity

Lewis theory oversimplifies the behavior of shared electrons in many cases.

In H—F, the pair of electrons is <u>**not**</u> equally shared between H and F.



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HF contains a **polar covalent bond**, where the fluorine has more electron density than the hydrogen.



The polar covalent bond is an **intermediate** between:

- pure covalent bonds:
- ionic bonds:

Electronegativity: the ability of an atom to attract electrons to itself in a chemical bond

 Pauling scale: 4 is most electronegative (F) and 0 is the least electronegative



Electronegativity is another periodic property, opposite of atomic size:

- Going across a period, EN increases
- Going down a column, EN decreases

Electronegativity difference determines the polarity of the bond:



Cl—Cl Δ EN = 0 - The bond is pure (nonpolar) covalent.

HCI: $\Delta EN = 0.9$ - The bond is polar covalent.

NaCl: Δ EN = 2.1 - The bond is ionic.

Which molecule is more polar, HF or CIF?

Lewis structures of molecular compounds

 Draw the correct skeleton structure, connecting atoms with single bonds. (H's are always terminal, more EN atoms tend to be terminal)

CH₂O: HCN:

- Calculate the total number of valence electrons.
 CH₂O: HCN:
- 3. Distribute remaining electrons, filling octets.

4. Make double or triple bonds only if any atoms lack an octet.

Lewis structures of polyatomic ions

When calculating the total number of valence electrons, account for the charge by adding or removing electrons from the total.

 PO_4^{3-} : CO_3^{2-}

<u>**Resonance structures</u>**: Lewis structures that are different only by the **location** of electrons in the structure. (Same # e⁻, same atom positions)</u> Formal charge

 $COCl_2$ has 3 resonance structures, but they are not equivalent: (24 ve⁻ total)

<u>Formal charge</u>: charge that each atom would have if it got half of its bonded electrons.

Start with an atom's original valence electrons,

- subtract its nonbonding electrons
- subtract half of its bonded electrons.

Formal charge: choosing the major resonance structure

The major (best) resonance structure has the fewest formal charges.

If choosing between 2 with the same charges, the better structure has the - on the more EN atom.

COCl₂ major resonance structure:

COCl₂ resonance hybrid:

N₂O (16 ve total)

Nonstandard octets

Boron normally has an incomplete octet in its compounds (6 electrons)

Elements in period 3 and beyond can be bonded with more than 8 valence electrons (usually 10 or 12) especially if it reduces formal charges.

SO₂ (18 ve total)

 SO_4^{2-} (32 ve total)

Bond length and strength

As more electrons are shared between atoms, the bond becomes **<u>stronger</u>** and **<u>shorter</u>**.

<u>Bond</u>	<u>Length</u>	<u>Strength</u>
C–C	154 pm	347 kJ/mol
C=C	134 pm	611 kJ/mol
C≡C	120 pm	837 kJ/mol

Which structure has the longest CO bond, CH₃CO₂⁻, CH₃OH, or CH₃COCH₃?