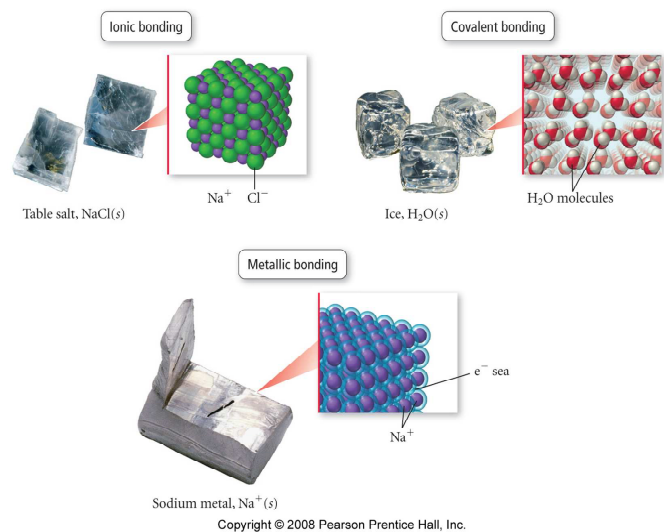


## Chapter 9: Chemical Bonding I: Lewis Theory

### Three types of chemical bonding:

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



**Coulomb's law:**  $E = k \frac{q_1 q_2}{r}$

- 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

**Ionic bond:** 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.



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

Na:

$\text{Na}^+$ :

Cl:

$\text{Cl}^-$ :

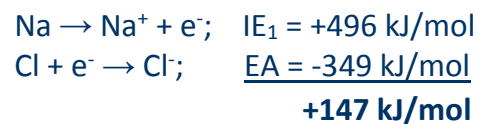


## Lattice energy

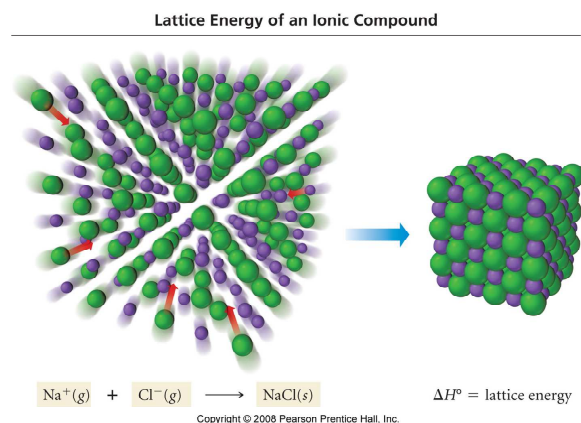
Actual formation of an ionic compound is usually very exothermic:



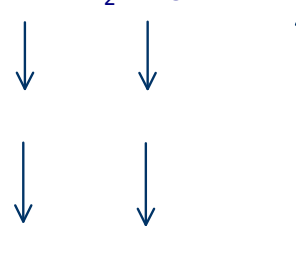
...but when we add up the ionization energy of Na and the electron affinity of Cl, that's actually endothermic!



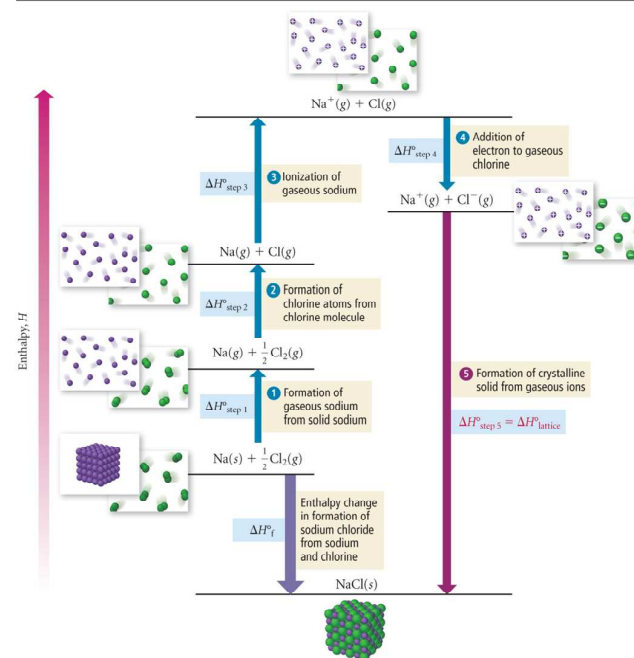
... so the release of energy does not come from formation of the ions! It actually primarily comes from the **lattice energy**, the energy associated with forming a crystalline ionic lattice from separated, gaseous ions.



## Born-Haber Cycle: Calculating lattice energy



Born-Haber Cycle for Production of NaCl from Na(s) and Cl<sub>2</sub>(g)



Periodic trends in lattice energy

**Ion size:**

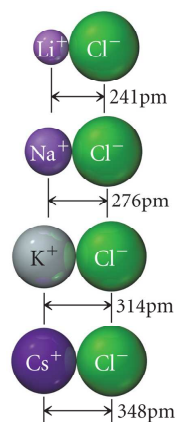
**Compound**    **Lattice energy**

LiCl            -834 kJ/mol

NaCl

KCl

CsCl



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**Ion charge:**

**Compound**    **Lattice energy**    **Melting point**

NaF            -910 kJ/mol    993 °C

CaO

**Coulomb's law:**  $E = k \frac{q_1 q_2}{r}$

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 H<sub>2</sub> from 2 H atoms:

Formation of H<sub>2</sub>O from 2 H atoms and 1 O atom:

Formation of O<sub>2</sub> from 2 O atoms:

Formation of N<sub>2</sub> 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.**)

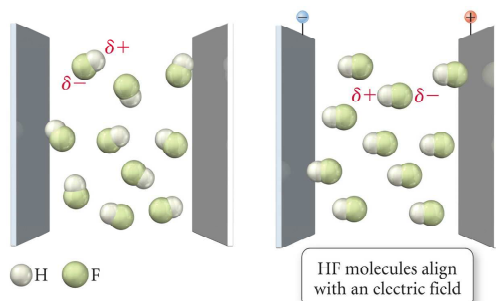
Always be sure that all the atoms' original valence electrons are represented in the Lewis structure

C<sub>2</sub>H<sub>4</sub>:

## Bond polarity

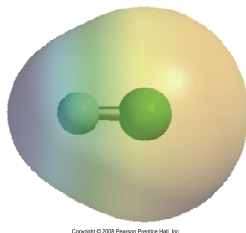
Lewis theory oversimplifies the behavior of shared electrons in many cases.

In H—F, the pair of electrons is **not** 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.



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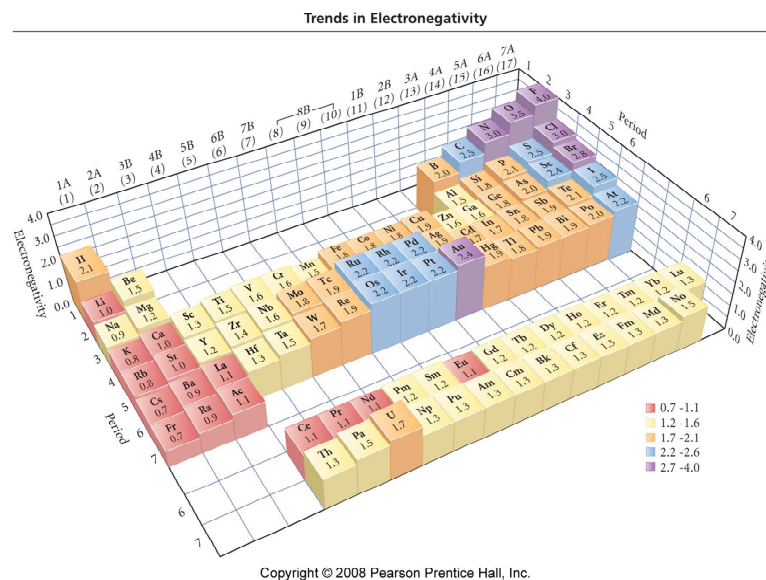
The polar covalent bond is an **intermediate** between:

- pure covalent bonds:
- ionic bonds:

## Electronegativity

**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



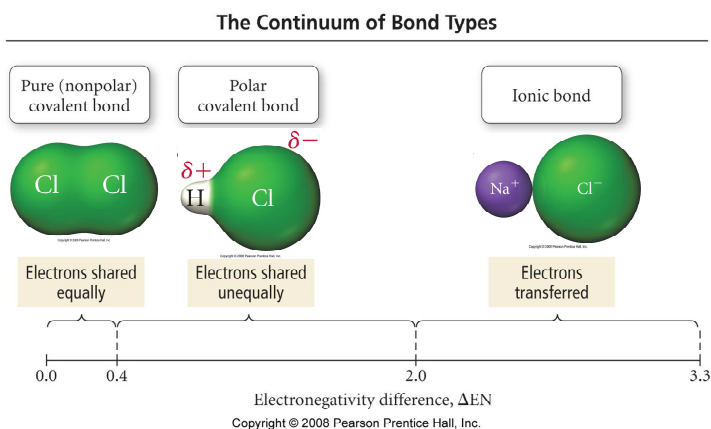
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Electronegativity is another periodic property, opposite of atomic size:

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

## Bond polarity and dipole moment

Electronegativity difference determines the polarity of the bond:



$\text{Cl}-\text{Cl}$   $\Delta EN = 0$  - The bond is pure (nonpolar) covalent.

$\text{HCl}$ :  $\Delta EN = 0.9$  - The bond is polar covalent.

$\text{NaCl}$ :  $\Delta EN = 2.1$  - The bond is ionic.

Which molecule is more polar, HF or ClF?

## Lewis structures of molecular compounds

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



2. Calculate the total number of valence electrons.



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.



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

Formal charge

COCl<sub>2</sub> has 3 resonance structures, but they are not equivalent: (24 ve<sup>-</sup> total)

**Formal charge:** 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.

$\text{COCl}_2$  major resonance structure:

$\text{COCl}_2$  resonance hybrid:

$\text{N}_2\text{O}$  (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.

$\text{SO}_2$  (18 ve total)

$\text{SO}_4^{2-}$  (32 ve total)

Bond length and strength

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

**Bond** **Length** **Strength**

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,  
CH<sub>3</sub>CO<sub>2</sub><sup>-</sup>, CH<sub>3</sub>OH, or CH<sub>3</sub>COCH<sub>3</sub>?