

SECTION 8 - CHEMICAL BONDING

8.1 -- Types of Chemical Bonds

- Ionic Bonds vs. Covalent Bonds
- Bond Energy
- Coulomb's Law
- Bond Length
- The Potential Energy Diagram for H₂

8-2 -- Electronegativity Trends

- Bond Polarity
- Dipole Moment

8-4 -- Overall Polarity and Dipole Moments

8-6 -- Ions: Electron Configurations and Relative Sizes

8-8 -- Formation of Binary Ionic Compounds

- Ionic Bonds
- Lattice Energy

8-9 -- Energies Contained in Various Covalent Bonds

- Bond Dissociation Energies (BDE)
- Bond Energies of Multiple Bonds

8-10 -- Calculating the Enthalpy Change (ΔH) from Bond Energies

- Calculating the Enthalpy Change from Standard Enthalpies of Formation (ΔH_f°)

8-12 -- Electron Dot Notation

- The Octet Rule
- Non-Bonding Electrons (Lone Pairs of Electrons)

8-13 -- Lewis Structures

- Exceptions to the Octet Rule
- Expanded Octets

8-14 -- Rules for Drawing Lewis Structures

- Drawing Lewis Structures of Molecules

8-21 -- Drawing Lewis Structures of Ions

8-22 -- Resonance Structures

- Nonequivalent resonance structures
- Formal Charges

8-24 -- Molecular Structure and Molecular Geometry

- VSEPR Theory (Valence Shell Electron-Pair Repulsion)
- Linear Geometry
- Trigonal Planar Geometry
- Tetrahedral Geometry
- Trigonal Pyramidal Geometry

- Bent Geometry
- Trigonal Bipyramidal Geometry
- Octahedral Geometry
- Square Planar Geometry

Section 8 = Chemical Bonding

8-1

- everything around us is comprised of a complex mixture of chemical compounds, all of which are bonded together with different kinds of atoms.

↳ the manner in which atoms bind together has a profound effect on chemical and physical properties.

* There are two types of chemical bonds.

↳ first introduced in Section 2.

① Ionic Bonds.

② Covalent Bonds.

→ - ionic bonds are stronger than covalent bonds and have a high bond energy.

↳ = the energy (J) required to break a bond.

- remember, ionic compounds are formed when an atom that loses e^- 's easily (cation - usually a metal) reacts with an atom that gains e^- 's easily (anion - usually a nonmetal).

ex: Find the energy of interaction ("the bond energy") between a pair Na^+ and Cl^- ions, when the distance between them is 2.76 \AA (0.276 nm).

↳ to do this, we must use

Coulomb's Law:

$$E = (2.31 \times 10^{-19} \text{ J}\cdot\text{nm}) \left(\frac{Q_1 Q_2}{r} \right)$$

→ Q_1 = charge on ion₁

→ Q_2 = charge on ion₂

→ r = distance between ion centers

so, we have: $E = (2.31 \times 10^{-19} \text{ J} \cdot \text{nm}) \left(\frac{(+1)(-1)}{0.276 \text{ nm}} \right)$

$$E = -8.37 \times 10^{-19} \text{ J}$$

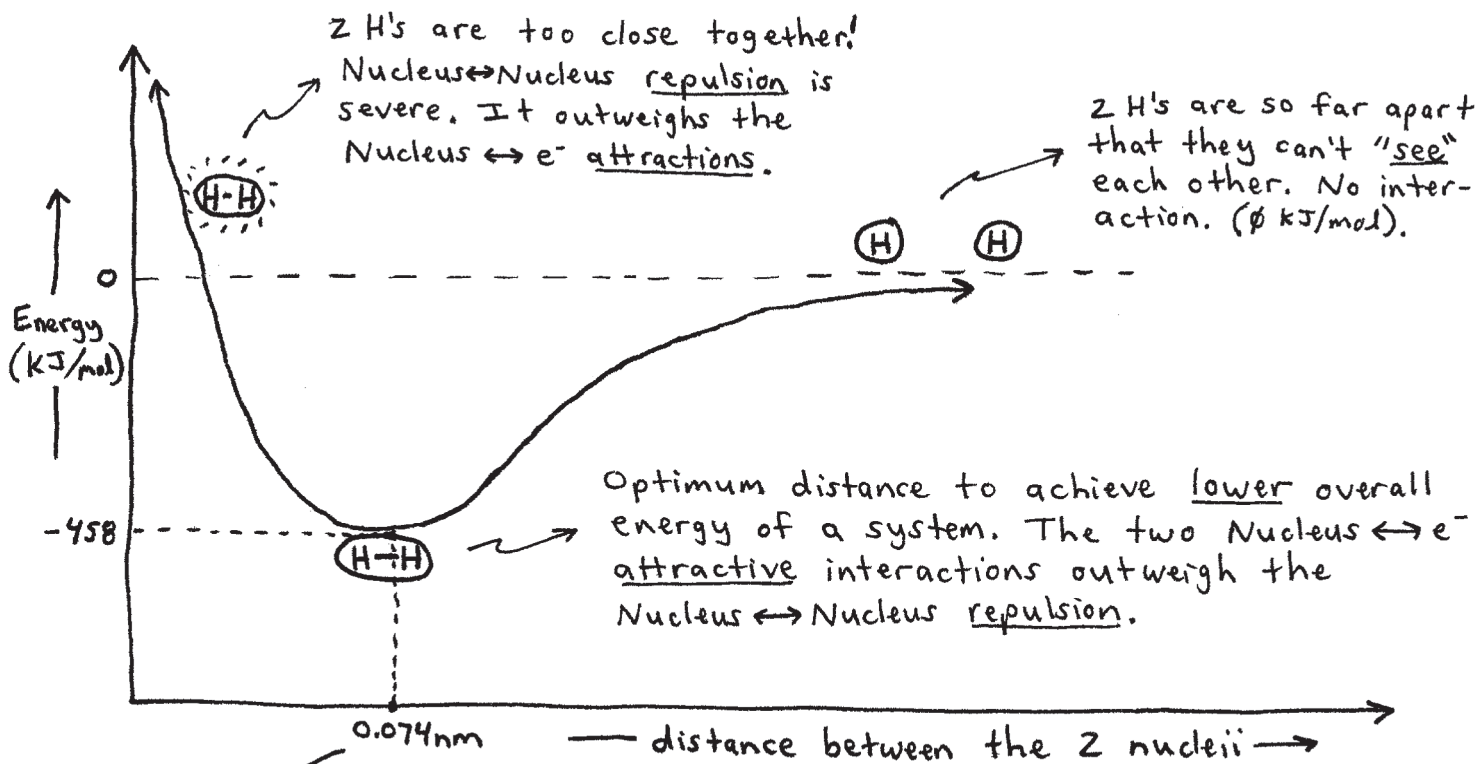
→ negative sign means we have an attractive force.

→ this is the bond energy of NaCl.

* bond length = the distance between 2 atoms (or ions) where the energy is minimal (see plot below).

→ the shorter the bond length, the stronger the bond, and therefore the harder it is to break the bond (higher bond energy).

ex: Consider H_2 's potential energy diagram (H-H is a covalent bond)



bond length for H_2 = 0.074 nm

bond energy for H_2 = -458 kJ/mol

- so, what determines whether a bond is ionic or covalent?

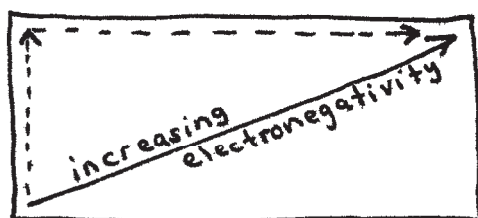
↳ you must consider the 2 atoms' or ions' differences in electronegativity.

↳ = the ability of an atom to attract electrons to itself

* Electronegativity Trend

↳ electronegativity values range from 0.7 to 4.0 on the Periodic Table.

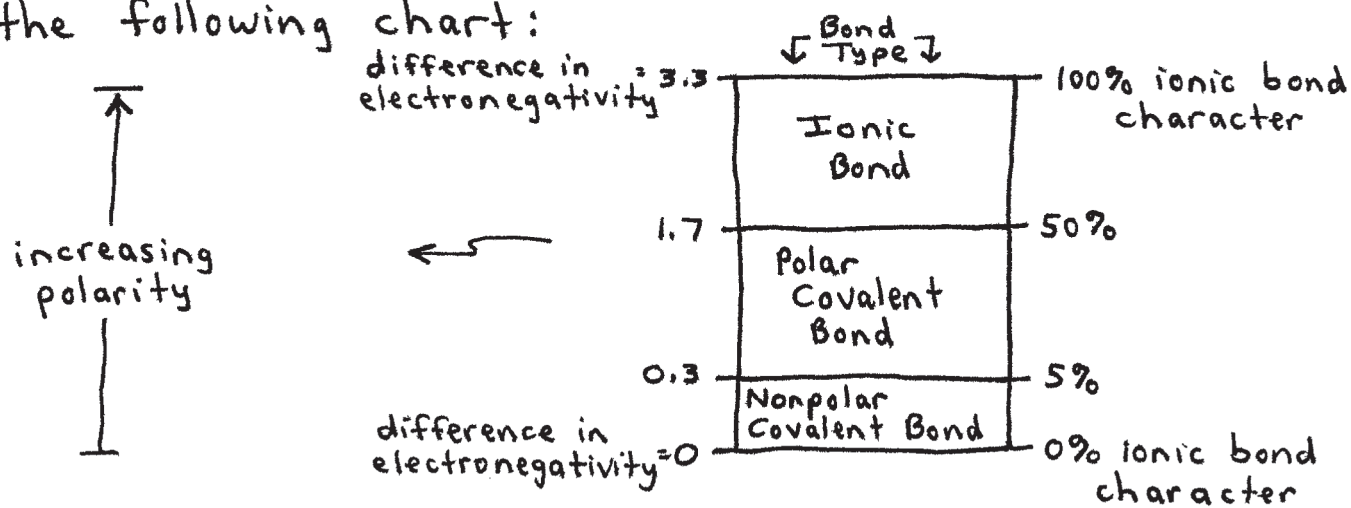
↳ Periodic Trend:



Periodic Table

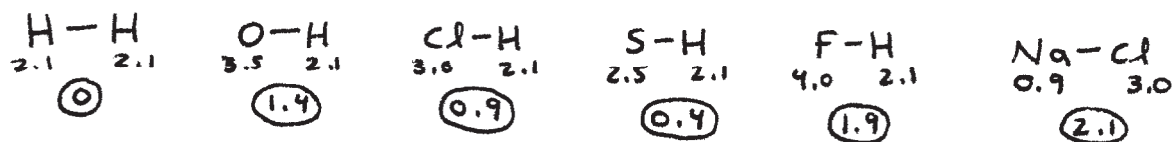
very electronegative

- to decide if a bond is ionic or covalent, subtract the 2 electronegativity values, and compare the result to the following chart:



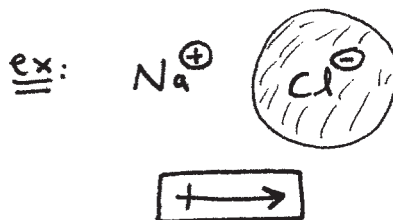
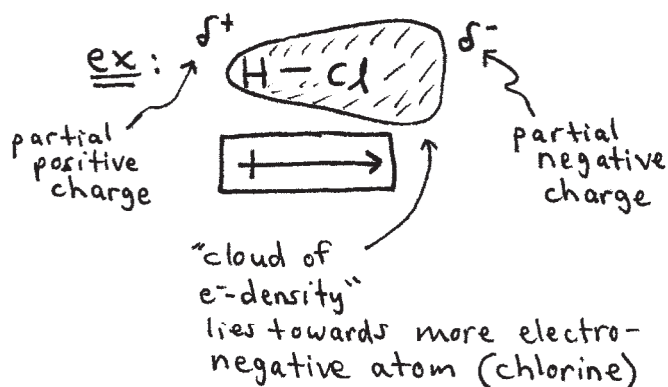
ex: Bond Polarity. Order the following bonds according to polarity: H-H, O-H, Cl-H, S-H, F-H, Na-Cl

→ use your textbook to look up each atom's electronegativity value (0.7-4.0). Then, subtract and compare, using the chart on p. 8-3.



- when bonds are polar (ionic or polar covalent), they have a "dipole moment"

→ = a direction of polarity
→ shown by: $\boxed{+ \rightarrow}$

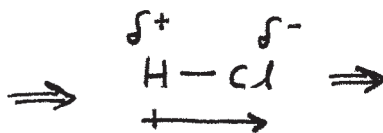
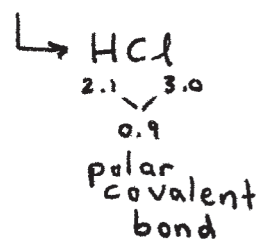


* Polarity and Dipole Moments

→ besides individual bonds, as we have just seen, these two ideas can also be applied to entire molecules.

ex: For each molecule or ionic compound, indicate which ones are overall polar (i.e. they have an overall dipole moment). Also show the direction of the individual bond polarities.

a) HCl

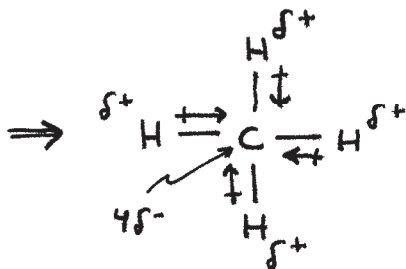
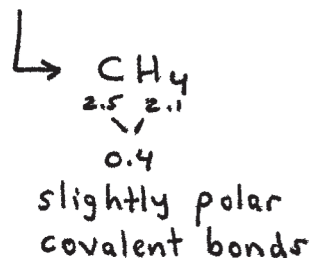


overall:



so,

polar molecule

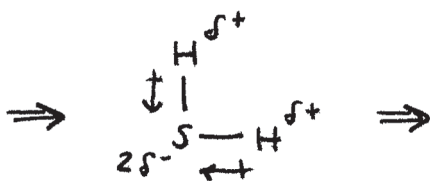
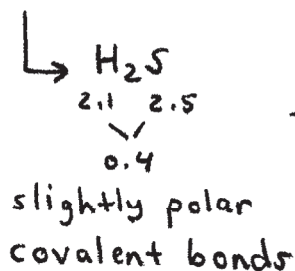
b) CH₄

overall: individual bond dipoles cancel each other out, so

no net dipole moment

so,

nonpolar molecule

c) H₂S

overall:

additive =



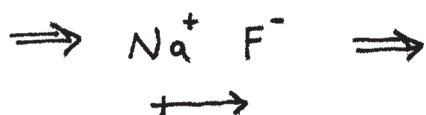
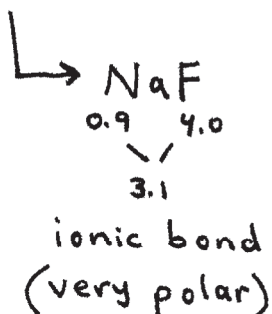
so,

polar molecule

Note: later in this section, we will learn why the following is incorrect:



d) NaF



overall:

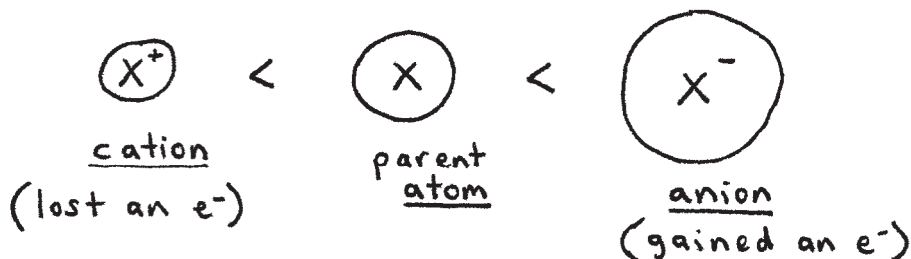


so,

polar ionic compound

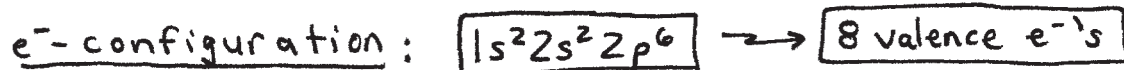
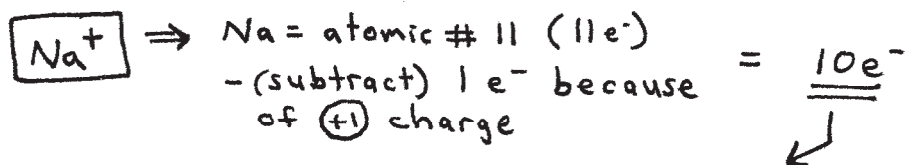
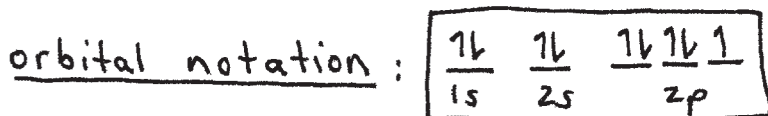
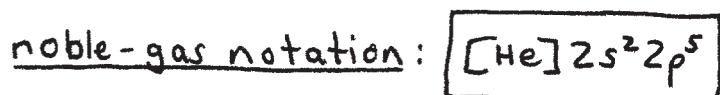
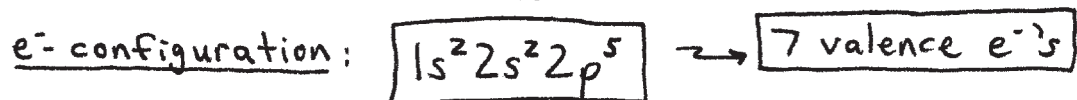
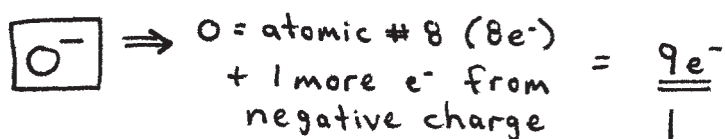
* Ions \Rightarrow Electron Configurations and Relative Sizes.

- generally, a cation (X^+) is smaller than the parent atom (X).
- generally, an anion (X^-) is larger than the parent atom (X).



ex: Write the electron configuration, noble-gas notation, and orbital notation for O^{-1} , Na^+ , and Fe^{2+} .

→ To do this, we really only need to know the number of e^- 's for each ion.



noble-gas notation: $[\text{He}]2s^22p^6$ or $[\text{Ne}]$
 → probably the best answer.

orbital notation: $\begin{array}{ccccc} \uparrow\downarrow & \uparrow\downarrow & \uparrow\downarrow & \uparrow\downarrow & \uparrow\downarrow \\ 1s & 2s & & 2p & \end{array}$

$\text{Fe}^{2+} \Rightarrow \text{Fe atomic \# 26 (26e}^-)$
 but must subtract $2e^- = 24e^-$
 → Note! $24e^-$ is same # of electrons as chromium, which is an exception to the diagonal rule!

e⁻-configuration:

$1s^22s^22p^63s^23p^64s^13d^5$

noble-gas notation: $[\text{Ar}]4s^13d^5$

orbital notation:

$\begin{array}{ccccccccc} \uparrow\downarrow & \uparrow\downarrow & \uparrow\downarrow & \uparrow\downarrow & \uparrow\downarrow & \uparrow\downarrow & \uparrow\downarrow & \uparrow\downarrow & \uparrow\downarrow & \uparrow & \uparrow & \uparrow & \uparrow & \uparrow \\ 1s & 2s & & 2p & & 3s & & 3p & & 4s & & & 3d & \end{array}$

ex: Order the ions from smallest to largest ionic radius.

a) $\text{Li}^+, \text{Na}^+, \text{K}^+, \text{Rb}^+, \text{Cs}^+$

→ all are group 1 cations and size increases down a group:

$\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$

b) $\text{Ba}^{2+}, \text{Cs}^+, \text{I}^-, \text{Te}^{2-}$

→ in this grouping it's not so easy as in (a).

→ all four ions have $54e^-$, so their e^- -clouds would be the same size. HOWEVER, each ion has a →

different amount of protons (+) in its nucleus 8-8 (differing atomic #'s). The ion with the most protons is best at "pulling" the $54e^-$ in, toward the nucleus, thus decreasing the size of the e^- -cloud.

↳ (+) ↔ (-) attraction



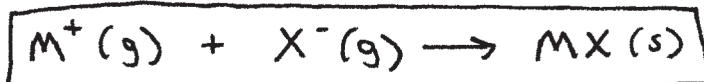
⊛ Formation of Binary Ionic Compounds.

↳ when cations and anions are mutually attracted to each other, they form ionic bonds and ionic compounds.

↳ just how strongly the ions attract each other is indicated by the lattice energy.

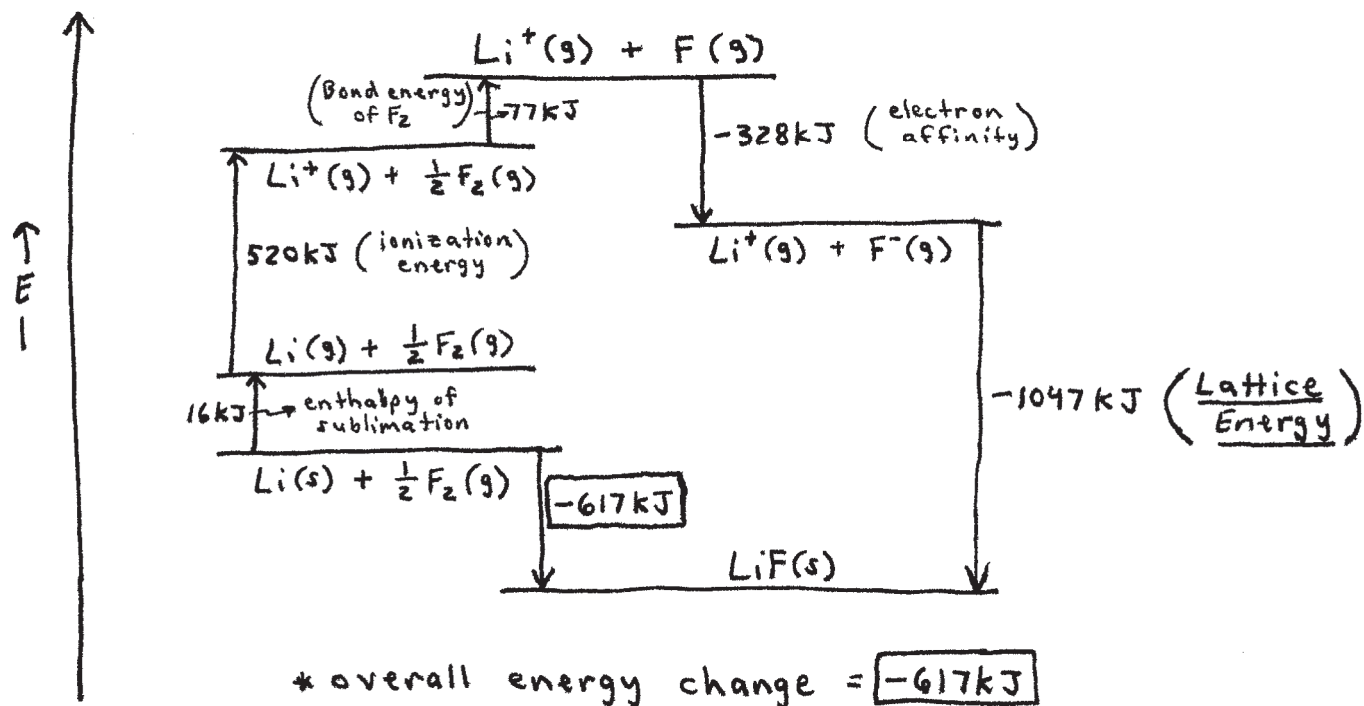
- lattice energy = the change in energy (kJ) that takes place when 2 gaseous ions are packed together to form an ionic solid.

↳ it's the energy released when an ionic solid forms:



↳ sample problem next page...

ex: Show the energy changes involved in the formation of solid lithium fluoride from its elements.



* Energies Contained in Various Covalent Bonds.

→ these are bond energies.

→ sometimes called "bond dissociation energies" or BDE.

ex: Consider the following processes:

<u>Process</u>	<u>Energy Required (kJ/mol)</u>
$\text{CH}_4 \rightarrow \text{CH}_3 + \text{H}$	435
$\text{CH}_3 \rightarrow \text{CH}_2 + \text{H}$	453
$\text{CH}_2 \rightarrow \text{CH} + \text{H}$	425
$\text{CH} \rightarrow \text{C} + \text{H}$	339
	<u>1652</u>
	Total = 1652

$$\text{Avg.} = \frac{1652}{4}$$

$$= 413 \text{ kJ/mol}$$

- this is the average bond energy
for C-H

→ "bond dissociation energy" (BDE)

→ see your textbook for a table of all the bond energies for other bonds.

* Bond energies of multiple bonds

- ↳ double bonds.
- ↳ triple bonds.

ex: C—C single bond (bond E = 347 kJ/mol) - weakest
 C=C double bond (bond E = 614 kJ/mol)
 C≡C triple bond (bond E = 839 kJ/mol) - strongest

↳ the shorter the bond length, the higher the bond energy.

* Bond Energies can be used to calculate the enthalpy of a reaction (ΔH)

ex: Consider the following reaction: $\boxed{\text{H}_2(\text{g}) + \text{F}_2(\text{g}) \rightarrow 2\text{HF}(\text{g})}$

- we must break 1 mol H—H $\rightarrow +432 \text{ kJ/mol}$
- we must break 1 mol F—F $\rightarrow +154 \text{ kJ/mol}$
- we must form 2 mol H—F $\rightarrow (-565 \text{ kJ/mol}) \times 2 = -1130 \text{ kJ/mol}$

Total: $\boxed{\Delta H = -544 \text{ kJ/mol}}$

↙
 - bond energy values can be found in your textbook.

- by the way, the value of -544 kJ/mol corresponds well with calculating ΔH via a method we've already discussed.

↙ using standard enthalpies of formation (ΔH_f°) gives -542 kJ/mol

↳ "products - reactants"
 ↳ see p. 6-15 through 6-17.

SECTION 8 - CHEMICAL BONDING

Were the FREE Section 8 Notes Useful?
Want the FULL VERSION of the Section 8 Notes?

Download Them Instantly for Only \$5.99 at:

CollegeChemistryNotes.com