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 H_2

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 ($\Delta Hf°$)

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

- 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 Å (0.276nm).

  ↳ to do this, we must use Coulomb's Law:

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

  → Q_1 = charge on ion₁
  → Q_2 = charge on ion₂
  → r = distance between ion centers

ChemistryNotes.com

so, we have: $E = (2.31 \times 10^{-19} J \cdot nm)\left(\dfrac{(+1)(-1)}{0.276 nm}\right)$
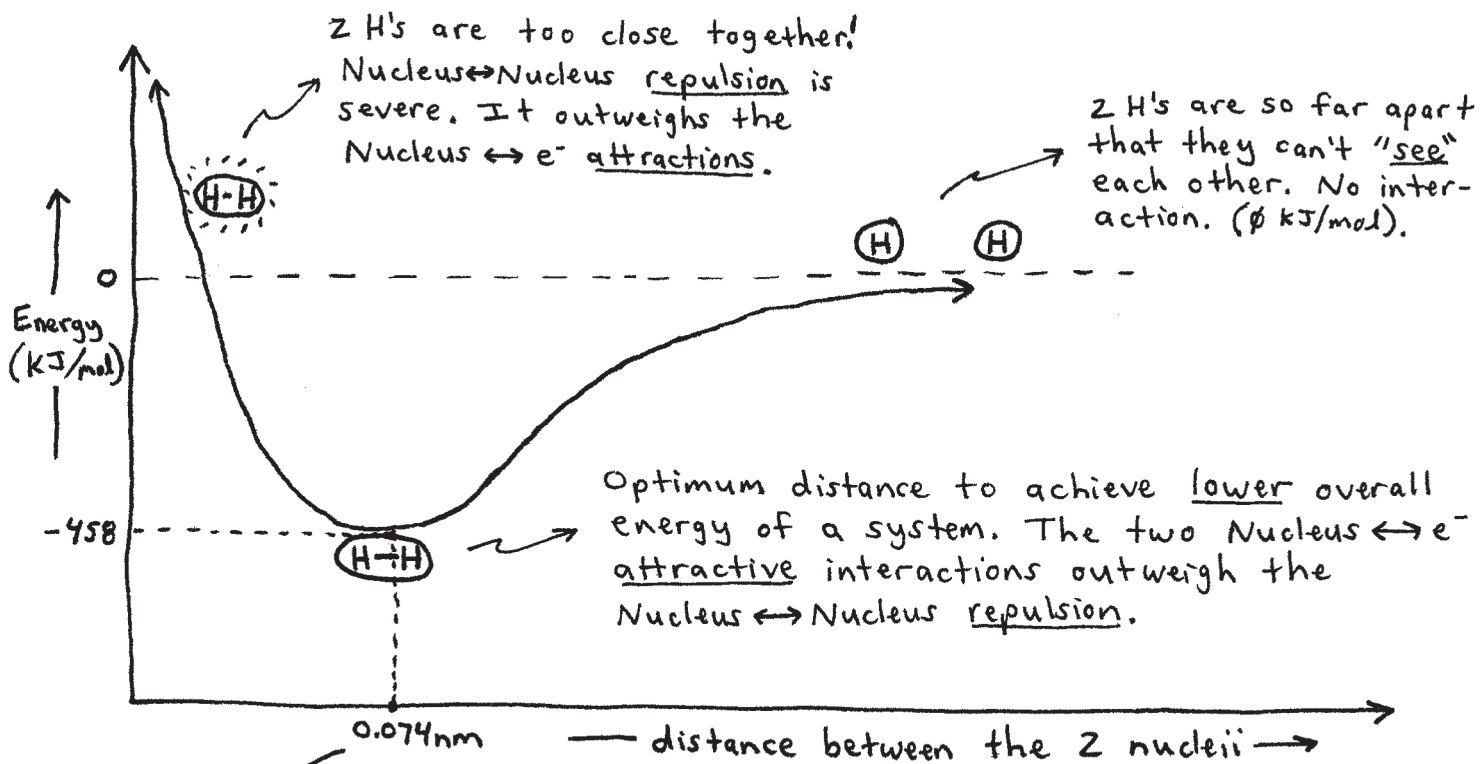
$$\boxed{E = -8.37 \times 10^{-19} J}$$

→ 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)

2 H's are too close together! Nucleus⟷Nucleus **repulsion** is severe. It outweighs the Nucleus ⟷ e⁻ **attractions**.

2 H's are so far apart that they can't "**see**" each other. No inter-action. (∅ kJ/mol).

Optimum distance to achieve **lower** overall energy of a system. The two Nucleus⟷e⁻ **attractive** interactions outweigh the Nucleus ⟷ Nucleus **repulsion**.

Energy (kJ/mol)

0

−458

0.074nm   — distance between the 2 nucleii →

bond length for H_2 = $\boxed{0.074 nm}$
bond energy for H_2 = $\boxed{-458 \text{ kJ/mol}}$

- so, what determines whether a <u>bond</u> is <u>ionic</u> or <u>covalent</u>?
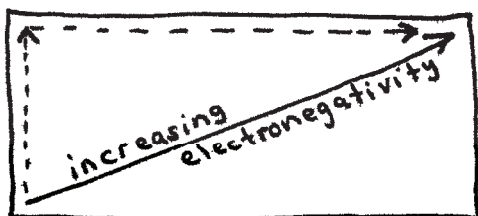
↳ you must consider the 2 atoms' or ions' <u>differences</u> in <u>electronegativity</u>.

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

✱ <u>Electronegativity Trend</u>

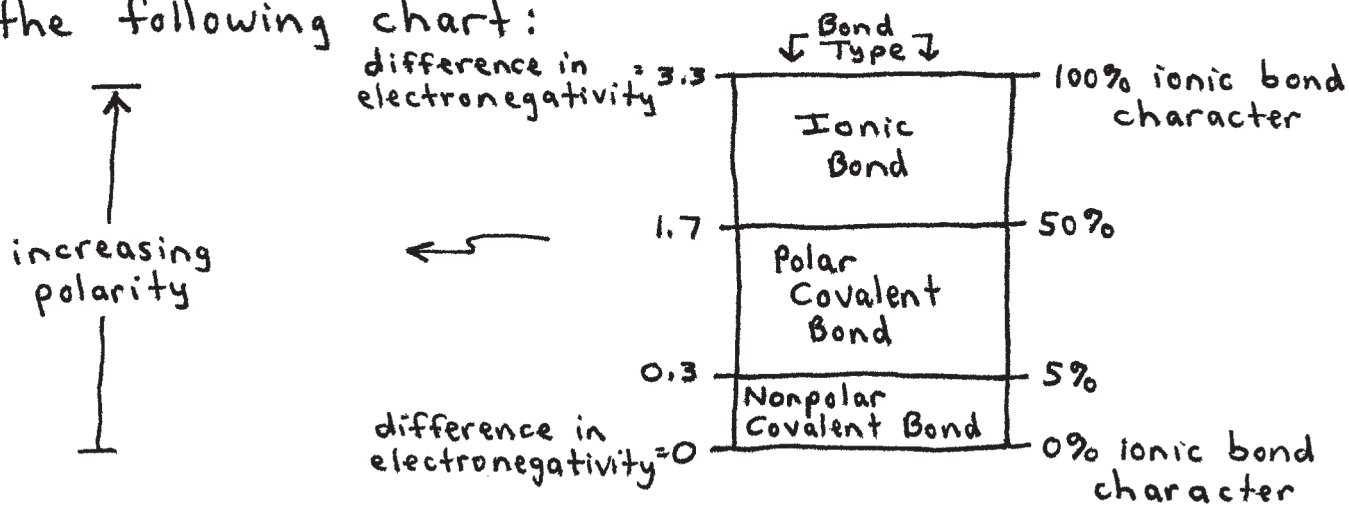→ electronegativity values range from 0.7 to 4.0 on the Periodic Table.

very electronegative

↳ Periodic Trend:

increasing electronegativity

Periodic Table
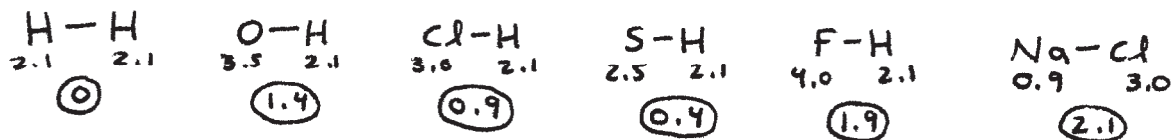
- to decide if a bond is <u>ionic</u> or <u>covalent</u>, subtract the 2 electronegativity values, and compare the result to the following chart:

increasing polarity

↓ Bond Type ↓

difference in electronegativity ≥ 3.3 ——— 100% ionic bond character

Ionic Bond

1.7 ——— 50%

Polar Covalent Bond

0.3 ——— 5%

Nonpolar Covalent Bond

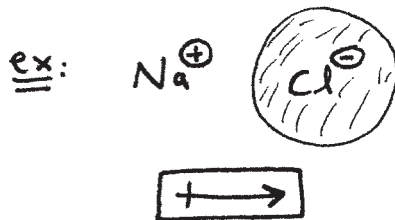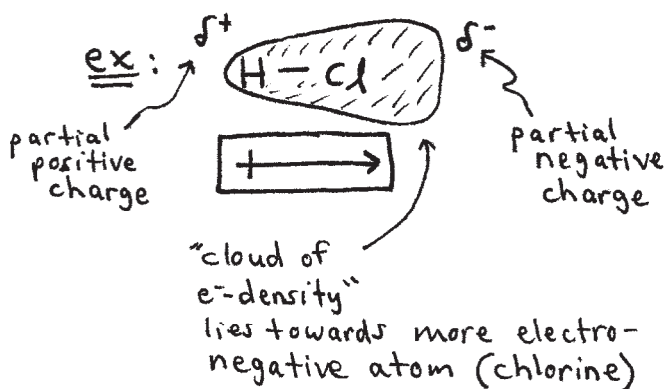difference in electronegativity = 0 ——— 0% ionic bond character

<u>ex</u>: <u>Bond Polarity</u>. 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 elec- tronegativity value (0.7 - 4.0). Then, subtract and compare, using the chart on p. 8-3.

H — H   O — H   Cl — H   S — H   F — H   Na — Cl
2.1   2.1   3.5   2.1   3.0   2.1   2.5   2.1   4.0   2.1   0.9   3.0
(0)   (1.4)   (0.9)   (0.4)   (1.9)   (2.1)

$$H-H < S-H < Cl-H < O-H < H^+F^- < Na^+Cl^-$$

═══

- when bonds are __polar__ (ionic __or__ polar covalent), they have a "__dipole moment__"

     ↳ = a direction of polarity
     ↳ shown by: [+——→]

<u>ex</u>: δ^+ H — Cl δ^-

partial positive charge

[+——→]

"cloud of e⁻ density" lies towards more electro-negative atom (chlorine)

partial negative charge

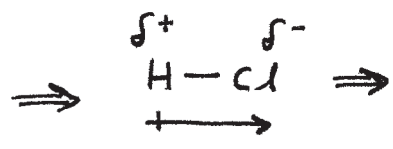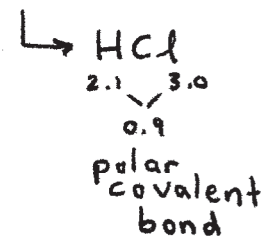<u>ex</u>:   Na⊕   Cl⊖

[+——→]

═══

⊛ <u>Polarity and Dipole Moments</u>

     ↳ besides individual bonds, as we have just seen, these two ideas can also be applied to __entire molecules__.
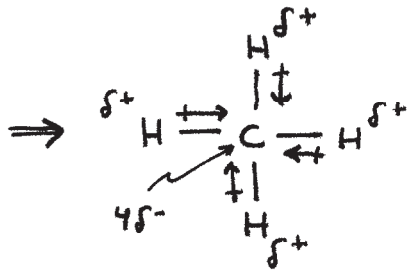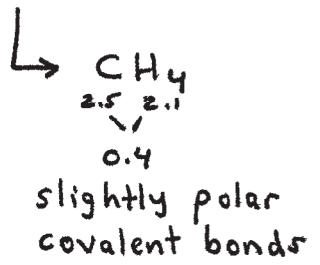
<u>ex</u>: For each molecule or ionic compound, indicate which ones are overall __polar__ (i.e. they have an __overall__ dipole moment). Also show the direction of the individual bond polarities.
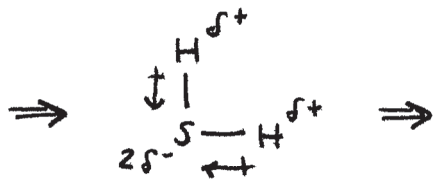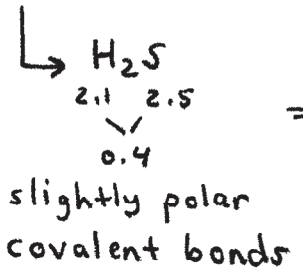
a) HCl

$\hookrightarrow$ HCl
2.1 $\diagdown$ 3.0
0.9
polar
covalent
bond

$\Rightarrow$

$\overset{\delta^+}{H} - \overset{\delta^-}{Cl}$
$\longmapsto$

$\Rightarrow$

overall : $\boxed{\longmapsto}$

$\hookrightarrow$ so, $\boxed{\text{polar molecule}}$

b) CH_4

$\hookrightarrow$ CH_4
2.5 2.1
$\diagdown$
0.4
slightly polar
covalent bonds

$\Rightarrow$

$\overset{\delta^+}{H} \rightleftharpoons \underset{\underset{\underset{H}{\updownarrow}}{H_{\delta^+}}}{\overset{\overset{H^{\delta^+}}{\updownarrow}}{C}} \rightleftharpoons H^{\delta^+}$
$4\delta^-$

$\Rightarrow$

overall : individual
bond dipoles cancel
each other out, so

$\boxed{\text{no net dipole moment}}$

$\hookrightarrow$ so, $\boxed{\begin{array}{c}\text{nonpolar} \\ \text{molecule}\end{array}}$

c) H_2S

$\hookrightarrow$ H_2S
2.1 2.5
$\diagdown$
0.4
slightly polar
covalent bonds

$\Rightarrow$

$\overset{\delta^+}{H}$
$\updownarrow$
$2\delta^- \underset{\rightleftharpoons}{S} - H^{\delta^+}$

$\Rightarrow$

overall :
additive = $\boxed{\swarrow}$

so, $\boxed{\text{polar molecule}}$

$\hookrightarrow$ Note : later in this section, we will learn
why the following is <u>incorrect</u> :

$\boxed{\overset{\longmapsto}{H} - S - \overset{\longleftarrow}{H}}$

d) NaF

$\hookrightarrow$ NaF
0.9 4.0
$\diagdown \diagup$
3.1
ionic bond
(very polar)

$\Rightarrow$

$Na^+ F^-$
$\longmapsto$

$\Rightarrow$

overall :

$\boxed{\longmapsto}$
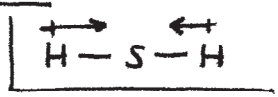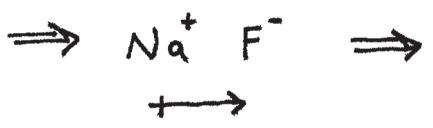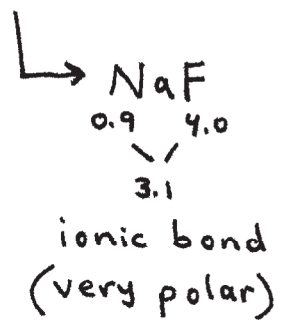
$\hookrightarrow$ so, $\boxed{\begin{array}{c}\text{polar ionic} \\ \text{compound}\end{array}}$

✱ <u>Ions</u> ⇒ Electron Configurations and Relative Sizes.

    → generally, a <u>cation</u> (X^+) is smaller than the parent atom (X).

    → generally, an <u>anion</u> (X^-) is larger than the parent atom (X).

$$\widehat{X^+} \quad < \quad \widehat{X} \quad < \quad \widehat{X^-}$$

     <u>cation</u>   <u>parent</u>   <u>anion</u>

   (lost an e^-)   <u>atom</u>   (gained an e^-)
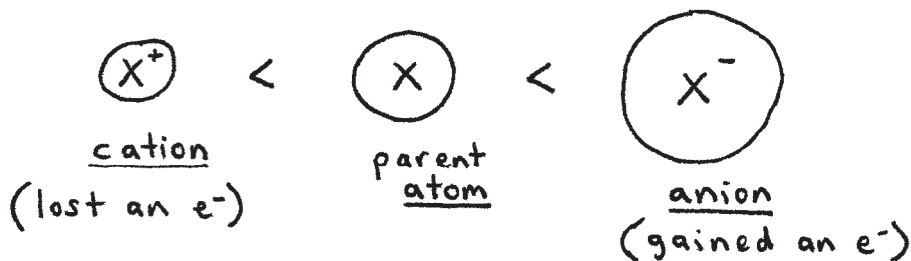
<u>ex</u>: Write the electron configuration, noble-gas notation, and orbital notation for O^{-1}, Na^+, and Fe^{2+}.

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

  $\boxed{O^-}$ ⇒ O = atomic # 8 ($8e^-$)
             + 1 more e^- from   = <u><u>$9e^-$</u></u>
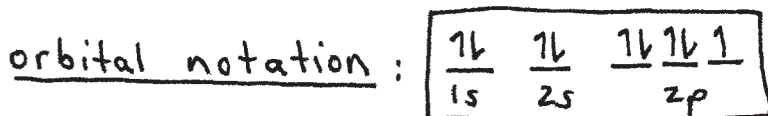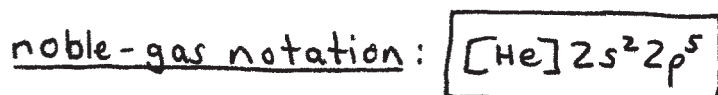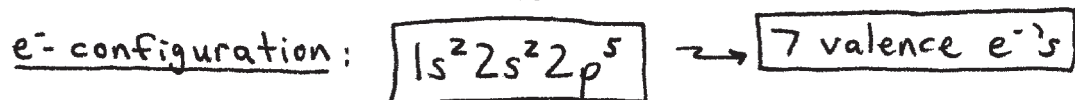   negative charge

    <u>e^- configuration</u>: $\boxed{1s^2 2s^2 2p^5}$ ⟿ $\boxed{7\text{ valence } e^-\text{'s}}$

    <u>noble-gas notation</u>: $\boxed{[He]2s^2 2p^5}$

    <u>orbital notation</u>: $\boxed{\begin{array}{ccc} \underline{\uparrow\downarrow} & \underline{\uparrow\downarrow} & \underline{\uparrow\downarrow\ \uparrow\downarrow\ \uparrow} \\ 1s & 2s & 2p \end{array}}$

  $\boxed{Na^+}$ ⇒ Na = atomic # 11 ($11e^-$)
             − (subtract) 1 e^- because   = <u><u>$10e^-$</u></u>
               of ⊕ charge

    <u>e^- configuration</u>: $\boxed{1s^2 2s^2 2p^6}$ ⟿ $\boxed{8\text{ valence } e^-\text{'s}}$
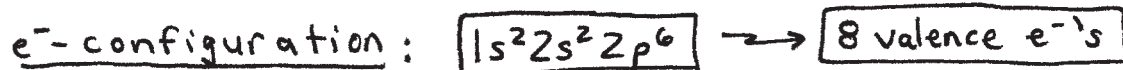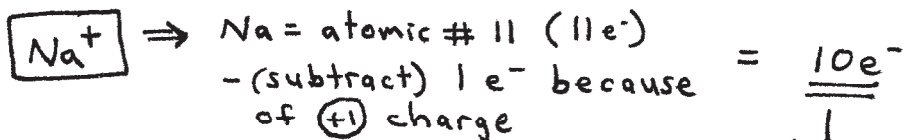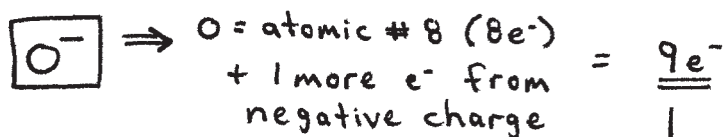
noble-gas notation: $[He]2s^2 2p^6$ or $[Ne]$

↳ probably the best answer.

orbital notation:

$\underline{1\downarrow}$	$\underline{1\downarrow}$	$\underline{1\downarrow}$ $\underline{1\downarrow}$ $\underline{1\downarrow}$
1s	2s	2p

Fe^{2+} ⇒ Fe atomic # 26 (26e⁻) but must subtract 2e⁻ = $\underline{24e^-}$

↳ Note! 24e⁻ is same # of electrons as <u>chromium</u>, which is an exception to ⇐ the diagonal rule!

e⁻-configuration:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$

noble-gas notation: $[Ar]4s^1 3d^5$

orbital notation:

$\underline{1\downarrow}$	$\underline{1\downarrow}$	$\underline{1\downarrow}$ $\underline{1\downarrow}$ $\underline{1\downarrow}$	$\underline{1\downarrow}$	$\underline{1\downarrow}$ $\underline{1\downarrow}$ $\underline{1\downarrow}$	$\underline{1}$	$\underline{1}$ $\underline{1}$ $\underline{1}$ $\underline{1}$ $\underline{1}$
1s	2s	2p	3s	3p	4s	3d

<u>ex</u>: Order the ions from smallest to largest ionic radius.

a) Li^+, Na^+, K^+, Rb^+, Cs^+

↳ all are group 1 cations and <u>size</u> increases <u>down</u> a group:

$Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$

b) Ba^{2+}, Cs^+, I^-, 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. <u>HOWEVER</u>, each ion has a →

different amount of <u>protons (+)</u> in its nucleus θ-θ (differing atomic #'s). The ion with the most protons is best at "<u>pulling</u>" the 54e⁻ in, toward the nucleus, thus decreasing the size of the e⁻-cloud.

$\longrightarrow$ (+)$\leftrightarrow$(-) attraction

$$Ba^{2+} < Cs^+ < I^- < Te^{2-}$$

⊛ <u>Formation of Binary Ionic Compounds.</u>
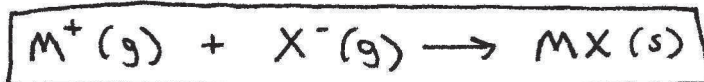
↳ when <u>cations</u> and <u>anions</u> are mutually attracted to each other, they form <u>ionic bonds</u> and <u>ionic compounds</u>.

↳ just how strongly the ions attract each other is indicated by the <u>lattice energy.</u>

– <u>lattice energy</u> = the change in energy (kJ) that takes place when 2 <u>gaseous ions</u> are packed together to form an <u>ionic solid</u>.
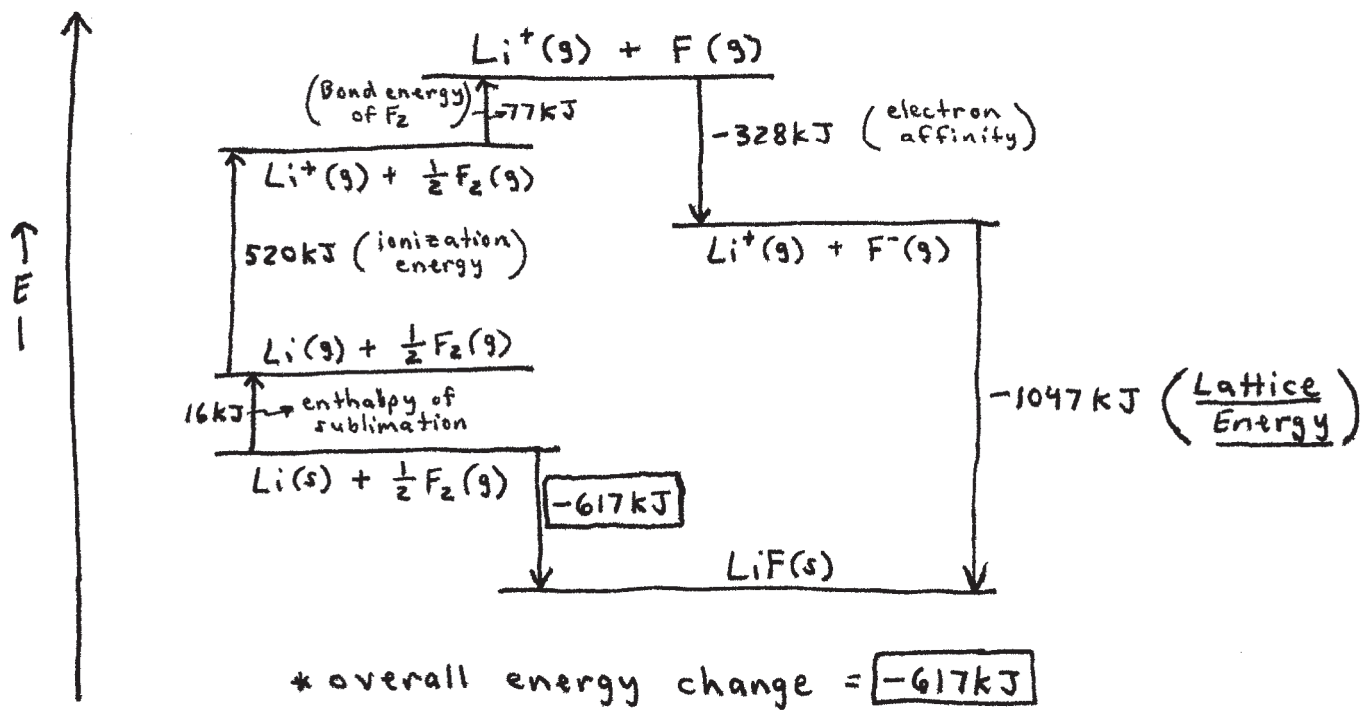
↳ it's the <u>energy released</u> when an ionic solid forms:

$$M^+(g) + X^-(g) \longrightarrow MX(s)$$

↳ sample problem next page...

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

$Li^+(g) + F(g)$

(Bond energy of F_2) $\uparrow$ $-77kJ$ $-328kJ$ (electron affinity)

$Li^+(g) + \frac{1}{2}F_2(g)$

$520kJ$ (ionization energy) $Li^+(g) + F^-(g)$

$Li(g) + \frac{1}{2}F_2(g)$

$16kJ \rightarrow$ enthalpy of sublimation $-1047 kJ$ $\left(\underline{\text{Lattice Energy}}\right)$

$Li(s) + \frac{1}{2}F_2(g)$ $\boxed{-617kJ}$

LiF(s)

$\uparrow E$

* overall energy change = $\boxed{-617kJ}$

⊛ <u>Energies Contained in Various Covalent Bonds.</u>

↳ these are <u>bond energies.</u>

↳ sometimes called "<u>bond dissociation energies</u>" <u>or</u> BDE.

ex: Consider the following processes:

Process	Energy Required (kJ/mol)
$CH_4 \rightarrow CH_3 + H$	435
$CH_3 \rightarrow CH_2 + H$	453
$CH_2 \rightarrow CH + H$	425
$CH \rightarrow C + H$	339
	Total = 1652

$\underline{Avg.} = \dfrac{1652}{4}$

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

- this is the <u>average bond energy</u>
  for $\boxed{C-H}$ ↳ "bond dissociation energy" (BDE)

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

ChemistryNotes.com

✻ Bond energies of <u>multiple bonds</u>
  ↳ double bonds.
  ↳ triple bonds.

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

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

✻ <u>Bond Energies can be used to calculate the enthalpy of a reaction (ΔH)</u>

<u>ex</u>: Consider the following reaction: $\boxed{H_2(g) + F_2(g) \rightarrow 2HF(g)}$

  – we must <u>break</u> 1 mol H–H ↝ +432 kJ/mol
  – we must <u>break</u> 1 mol F–F ↝ +154 kJ/mol
  – we must <u>form</u> 2 mol H–F ↝ (–565 kJ/mol)×2 = –1130 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 <u>method</u> we've already discussed.
    ↳ using standard enthalpies of formation (ΔH_f°) gives –542 kJ/mol
      ↳ "products – reactants"
      ↳

SECTION 8 – CHEMICAL BONDING

Were the FREE Section 8 Notes Useful?

Want the FULL VERSION of the Section 8 Notes?

Download Them Instantly at:

ChemistryNotes.com