Chapter 8- Bonding: General Concepts
Study Guide

Key

8.1 Types of Chemical Bonds

Chemical bonds are "forces that hold groups of atoms together and make them function as a unit." Bonds form because the energy of a system is lower than if the bonds did not form.

The three types of bonds:

1. Ionic Bonding-

An ion pair has energy known as **energy of interaction**. Derived from Coulomb's Law, the energy can be calculated with the formula: $E = 2.31 \times 10^{-19}$ J nm $[(Q_1Q_2)/r]$

Example 8.1 Coulomb's Law

Calculate the energy of interaction (in kJ/mole) between Mg^{2+} and O^{2-} if the distance between the centers of the ions is 0.205 nm.

Solution

$E = 2.31 \times 10^{-19}$ J nm $[(Q_1Q_2)/r] = E = 2.31 \times 10^{-19}$ J nm $[((+2)(-2))/0.205 \text{ nm}]$

$= -4.51 \times 10^{-18}$ J/ion pair $=$ $\mathbf{-4.51 \times 10^{-21}}$ **kJ/ion pair**

2. Covalent Bonding- it involves the "sharing of electrons" between atoms in a molecule
   Polar Covalent bonding is when there is an unequal sharing of electrons.
   Non-polar Covalent bonding is when there is an equal sharing of electrons.

The nature of the bond will depend on the ability of each atom in the bond to attract electrons to itself. This is called **electronegativity**.

Problems:

1. Calculate the energy of interaction for KCl if the internuclear distance is 0.314 nm.

$$E = 2.31 \times 10^{-19}\ \text{J}\cdot\text{nm}\left(\frac{(+1)(-1)}{.314\ \text{nm}}\right) = \boxed{-7.36 \times 10^{-19}\ \frac{\text{J}}{\text{pair}}}$$

2. Calculate the energy of interaction between AgBr if the internuclear distance is 0.120 nm.

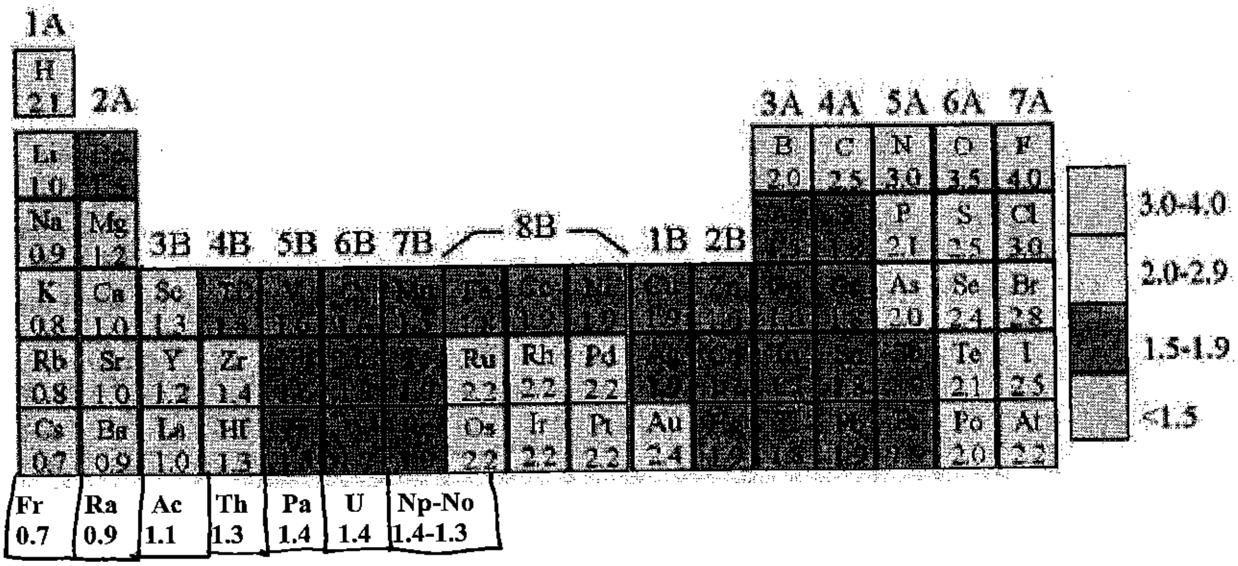
$$E = 2.31 \times 10^{-19}\ \text{J}\cdot\text{nm}\left(\frac{(+1)(-1)}{.120\ \text{nm}}\right) = \boxed{-1.925 \times 10^{-18}\ \text{J}/\text{pair}}$$

8.2 Electronegativity

Recall that unequal sharing of electrons in a bond results in a polar covalent bond. Ionic bonds result from the transfer of electrons between atoms. We can get a degree of ionic character of a bond by **measuring bond energies**. The more electrostatic interaction that occurs between 2 atoms, the greater will be the difference in bond energies when compared to the average of perfectly covalent bonds involving the atoms. The difference in bond energy is called Δ and is the relative **electronegativity** difference between the bonding atoms.

For the representative elements, **electronegativity decreases going down a group** and **increases going across a period**. The greater the Δ, the more ionic character a bond has. If Δ = 0, the bond is perfectly covalent. There are no precise cut-offs. All bonds have some ionic and some covalent character.

Electronegativities of Elements



Δ (Electronegativity Diff) (approx)	Type of Bond (In General)	Example
0.0 – 0.4	Covalent (Non-polar)	H-H (0.0)
0.4 – 1.0	Covalent (Mod. Polar)	H +→ Cl (0.9)
1.0 – 1.9	Covalent (Very Polar)	H +→ F (1.9)
> 1.9	Ionic	Na (+) Cl (-) (2.1)

Problems:

3. Using a periodic table, order the following from lowest to highest electronegativity.
   a. Fr, Mg, Rb Fr < Rb < Mg c. P, As, Ga, O Ga < As < P < O
   b. B, Al, C, N Al < B < C < N d. Cl, S, P P < S < Cl

4. Using the periodic chart of the elements, place the following in order from the lowest to the highest electronegativity: F, Nb, N, Si, Rb, Ca, Pt
   4.0 1.6 3.0 1.8 .8 1.0 2.2
   Rb < Ca < Nb < Si < Pt < N < F

5. Using the table above, calculate the difference in electronegativity (Δ) for each of the following bonds:
   a. Cl – Cl Δ = 3-3 = 0 Non-polar covalent
   b. K – Br Δ = 2.8 - .8 = 2.0 ionic
   c. Fe – O Δ = 3.5 – 1.8 = 1.7 very polar covalent
   d. H – O Δ = 3.5 - 2.1 = 1.4 very polar covalent

6. Indicate whether the bonds between the following would be primarily covalent, polar covalent or ionic.
   a. O – H very polar $\Delta = 1.4$ covalent
   c. H – Cl $\Delta = .4$ Non-polar cov.
   b. Br – Br non-polar $\Delta = 0$
   d. Cs – Cl $\Delta = 2.3$ ionic

7. Place the following in order of increasing polarity: NaBr, I_2, H_2O, MnO_2, CN

   $\Delta = 2$ $\Delta = 0$ $\Delta = 1.4$ $\Delta = 2$ $\Delta = .5$

   $I_2 < CN < H_2O < NaBr / MnO_2$

8. Which of the following molecules contain polar covalent bonds? List in order of increasing bond polarity.

Least → Greatest

O_3, P_8, NO, CO_2, CH_4, H_2S

$O_3, P_8 < CH_4, H_2S < NO < CO_2$ [illegible] $\Delta = 0$ 0 .5 1 .4 .4

Polar CO_2 but molecule is non polar

9. How will the charge be distributed on each of the following molecules: HF, NO, CO and HCl?

   H +→ F N +→ O C +→ O H +→ Cl

8.3 Bond Polarity & Dipole Moments

Remember that you can determine the polarity of a bond by the size of the Δ. In the molecule HF, it is polar covalent with fluorine being the more electronegative atom. A partial negative charge (δ-) resides on the fluorine atom and a partial positive charge (δ+) resides on the hydrogen atom.

+⟶
H—F
δ+ δ-

The arrow points to the center of the negative charge while the tail is at the center of the positive charge. A **dipole moment** means that the molecule has two poles.

The situation is clear-cut with HF. It becomes more difficult with 3 or more atoms in a molecule because the individual dipoles can cancel each other out. **Look at Table 8.2 in your book.** This shows how individual bond polarities can cancel each other out to yield a molecule with no dipole moment. Although you will be able to derive the geometries later on, you should memorize them for now.

Example 8.3A Dipole Moment

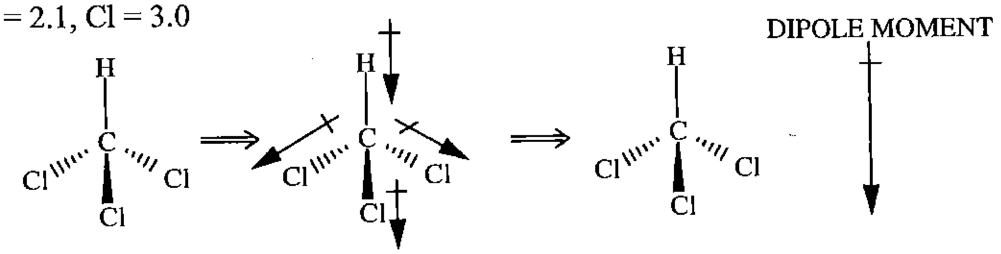
Does $CHCl_3$ (a tetrahedral molecule with carbon at the center) have a dipole moment? If so, show the orientation of the dipole moment.

Strategy

Perform the following steps:
a. Look up electronegativity of each atom.
b. Draw the molecule in a 3-D space.
c. Determine the polarity of each bond and the net polarity on each atom.
d. Draw the dipoles and determine the direction (if any) of the molecule dipole moment.

Solution

C = 2.5, H = 2.1, Cl = 3.0



Example 8.3B Practice with Dipole Moments

For each of the following, determine the orientation of the dipole moment (if any).

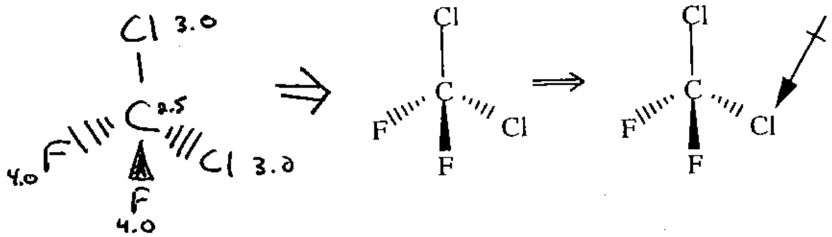
a. HI	b. N_2	c. CCl_2F_2 (carbon is central atom)

Solution

a. H = 2.5, I = 2.1	$\underset{2.1}{H} - \underset{2.5}{I} \Rightarrow \underset{\delta^+}{H} - \underset{\delta^-}{I} \Rightarrow \underset{+\!\!\longrightarrow}{H - I}$

b. N = 3.0	$\underset{3.0}{N} \equiv \underset{3.0}{N} \Rightarrow$ No dipole moment $(\Delta = 0)$

c. C = 2.5, Cl = 3.0, F = 4.0



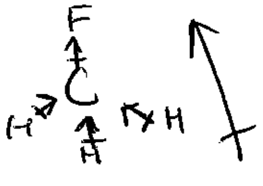
Problems:

10. Determine the orientation of the dipole of the following, if any:

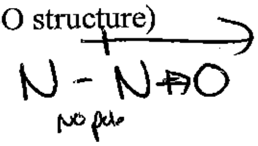
a. $AlCl_3$ (planar w/ aluminum in the center)

No overall dipole.

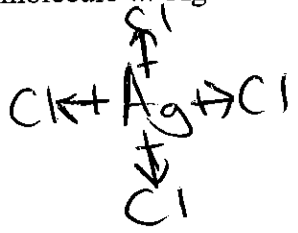
b. CH_3F (tetrahedral w/ carbon at center)



c. N_2O (linear w/ N-N-O structure)



d. $AgCl_4$ (planar molecule w/ Ag at center, Ag-Cl bonds 90° apart)



No dipole moment.

8.4 Ions: Electron Configurations & Sizes

Your textbook deals only w/ nonmetals, representative metals and ionic bonds in this discussion. It has been observed that atoms that form bonds in stable compounds have a **noble gas configuration**. (Each is **isoelectric** with the closest noble gas).

To determine the formula of a binary ionic compound, remember that chemical compounds are electrically neutral. For examples, the formula of the ionic compound formed from magnesium (Mg^{2+}) and chlorine (Cl^-) to form $MgCl_2$.

We discussed in chapter 7, as we go down a group, ion size increases (higher energy levels have a larger average distance). Because the ratio of protons to electrons becomes greater, **cations are smaller than their neutral atoms**. Because of electron-electron repulsion and less effective shielding **anions are always larger than their neutral atoms**. The larger the charge, the more pronounced the effect. Basically, **the higher the nuclear charge, the smaller the ion.**

Problems

11. Determine the most stable ion for each of the following atoms and indicate which element they would be isoelectric with if they lost or gained electrons.
    - a. O O^{-2} [Ne]
    - b. Be Be^{+2} [He]
    - c. I I^{-1} [Xe]
    - d. Te Te^{-2} [Xe]
    - e. Na Na^{+1} [Ne]

12. List the four ions that are isoelectric with argon and have charges from -2 to +2. Arrange those in order of increasing atomic radius.

    S^{-2}, Cl^-, K^+, Ca^{+2} $Ca^{+2} < K^+ < Cl^- < S^{-2}$

13. Determine the formula of the binary compound from the following sets of atoms.
    - a. Ca and O → CaO
    - b. K and Cl → KCl
    - c. Rb and S → Rb_2S
    - d. Ba and P → Ba_3P_2

14. Using shorthand notation, list the core electron configurations for the ions in the compounds in problem 13.

a) Ca^{+2} [Ar]
O^{-2} [Ne]

b) K^+ [Ar]
Cl^- [Ar]

c) Rb^{+1} [Kr]
S^{-2} [Ar]

d) Ba^{+2} [Xe]
P^{-3} [Ar]

15. Place the following in an order of increasing atomic size.

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

b. Cs^{+}, S^{2-}, O^{2-}, K^{+} $K^{+} < Cs^{+} < O^{-2} < S^{-2}$

8.5 Formation of Binary Ionic Compounds

The theme of this section is that there are many separate processes that go into forming an ionic solid. As we have said before, the ionic compound forms because its energy is lower than if its elements remained separated. However, not every part of the process is energetically favorable. Your book points out that **lattice energy,** the energy released when an ionic solid is formed from its gaseous ions, that is the most favorable, and more than makes up for some parts of the process that are energetically unfavorable. Lets examine the formation of the **KCl** ionic solid from K(s) and Cl_2(g).

Processes that Must Occur

1. K(s) must form K(g): Energy of Sublimation = + 64 KJ
2. K(g) must form K+ (g): First Ionization Energy = +419 KJ
3. $1/2Cl_2$ must form Cl (g): Bond Energy x ½ = 120 KJ
4. Cl(g) must form Cl^-(g); Electron Affinity = -349 KJ
5. K^+(g) must combine with Cl^-(g) to form $KCl_{(s)}$: Lattice Energy = -690 KJ

The net energy of formation equals the sum of the energy changes, -436 kJ. So you see that the value for any of the processes that make up salt formation can be obtained if you understand the processes involved, and are given suitable data.

Problem

16. Given the following data, determine the ΔH_f for LiBr from the equation

$$Li(s) + \frac{1}{2} Br_2(g) \rightarrow LiBr(s)$$

1. Li(s) must form Li(g): Energy of Sublimation = + 161 KJ
2. Li(g) must form Li+ (g): First Ionization Energy = +520 KJ
3. Br_2 must form Br (g): Bond Energy = 193 KJ
4. Br(g) must form Br^-(g); Electron Affinity = -324 KJ
5. Li^+(g) must combine with Br^-(g) to form $LiBr_{(s)}$: Lattice Energy = -787 KJ

① +161
② +520
③ ½(+193)
④ -324
⑤ -787

$\boxed{\Delta H = -333.5 \text{ KJ}}$

The remainder of the section makes the point that the higher the charge on each ion, the greater the lattice energy will be. This value counteracts the higher endothermic ionization energies, thus resulting in a more energetically stable crystal.

8.8 Covalent Bond Energies and Chemical Reactions

In the beginning of this section, it is calculated that the average bond energy for a C – H bond for methane (CH_4). Assumptions made were:

1. that the energy needed to break each C – H bond was the same (413 kJ), and
2. that each of the bonds was not sensitive to its environment.

Table 8.4 in your book gives **average bond energies** for many covalent bonds. Notice that **multiple bonds** (bonds that involve sharing more than2 electrons between atoms) require more energy to break than single bonds.

You can determine the heat of a reaction by knowing the individual bond energies. It is simply:

$$\Delta H = \Sigma D \text{ (bonds broken)} - ED \text{ (bonds formed)}$$

where ΔH = heat of reaction and D = bond energy per mole of bonds

The key to calculating the heat of reaction from the average bond energy is to **carefully list all bonds broken and all bonds formed.** Also, recognize that it takes energy to break bonds (endothermic, ΔH = +) while energy is released when bonds are formed (exothermic, ΔH = -).

Example 8.8 Heat of Reaction from Bond Energy

Using data from Table 8.4 in your book, calculate ΔH for the following reaction.

$$CH_4(g) + 4F_2(g) \rightarrow CF_4(g) + 4HF(g)$$

Solution:

First, make a list of bonds broken and bonds formed.

Bonds Broken	Energy per bond (kJ)		Total Energy (kJ)
4 C – H	413 x 4	=	1652
4 F – F	154 x 4	=	616
Total Energy to break bonds		=	**2268 kJ**

Bonds Formed	Energy per bond (kJ)		Total Energy (kJ)
4 C – F	485 x 4	=	1940
4 H – F	565 x 4	=	2260
Total Energy to form bonds		=	**4200 kJ**

ΔH = (Energy to break bonds) – (Energy to form bonds) = 2268 kJ – 4200 kJ = -1932 kJ (exothermic)

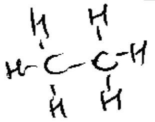
Problems

17. Using bond energy values from Table 8.4 in your book, calculate the ΔH for the following reactions:
    a. $2H_2(g) + O_2(g) \rightarrow 2H_2O$

a) $(2(H\text{-}H) + (O\text{-}O)) = 2(432) + 146$

$-(4(H\text{-}O)) = -4(467)$

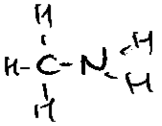
$$\boxed{\Delta H = -858 \text{ kJ/mol}}$$



b. $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$

Broken		Formed	
(2) C-C	2(347)	(8) C-O	8(358)
(12) C-H	12(413)	(12) H-O	12(467)
(7) O-O	7(146)		
	6672		8468

$6672 - 8468 = \boxed{-1796 \text{ KJ}}$



c. $HCN(g) + 2H_2(g) \rightarrow CH_3NH_2(g)$

Broken		Formed	
H-C	413	(3) C-H	3(413)
C-N	305	C-N	305
(2) H-H	2(432)	(2) N-H	2(391)
	1582		2326

$1582 - 2326 = \boxed{-744 \text{ KJ}}$

d. $H\text{-}C \equiv C\text{-}H(g) + H_2(g) \rightarrow CH_2{=}CH_2(g)$

Broken		Formed	
(1) C≡C	1(839)	(1) C=C	1(614)
(2) C-H	2(413)	(4) C-H	4(413)
(1) H-H	1(432)		
	2097		2266

$2097 - 2266 = \boxed{-169 \text{ KJ}}$

18. Compare the values obtained in part "d" of problem 17 to ΔH values calculated from ΔH_f° data in Appendix 4 of your book.

$\Delta H = H_f(\text{prod}) - H_f(\text{react})$

$\Delta H = 52 - (227 + 0) = \boxed{-175 \text{ KJ}}$

8.9 The Localized Electron Bonding Model

If you can answer the following review questions, you understand the important ideas in this brief section.

1. What are the two covalent bonding models?