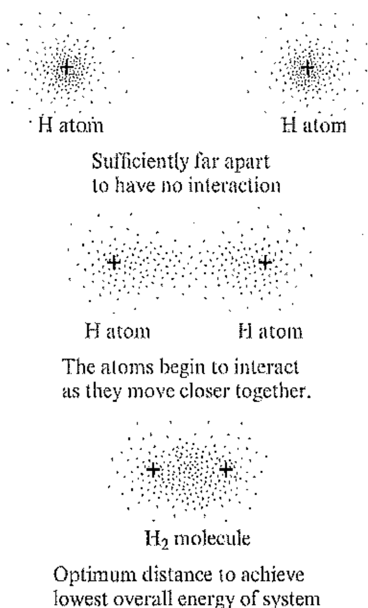
AP Chemistry Unit 5: Bonding

Wkst: Unit 5 Homework Packet

The Interaction of Two Hydrogen Atoms



Key Ideas in Bonding

- Ionic Bonding – electrons are transferred
- Covalent Bonding – electrons are shared equally by nuclei
- What about intermediate cases?

Polar Covalent Bond

- Unequal sharing of electrons between atoms in a molecule.
- Results in a charge separation in the bond (partial positive and partial negative charge).

Electronegativity

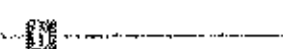
- The ability of an atom in a molecule to attract shared electrons to itself.
- For a molecule HX, the relative electronegativities of the H and X atoms are determined by comparing the measured H–X bond energy with the "expected" H–X bond energy
- On the periodic table, electronegativity generally increases across a period and decreases down a group.
- The range of electronegativity values is from 4.0 for fluorine (the most electronegative) to 0.7 for cesium (the least electronegative).

Table 8.1 | The Relationship Between Electronegativity and Bond Type

Electronegativity Difference in the Bonding Atoms	Bond Type	
Zero	Covalent	Covalent character ↑
↓	↓	
Intermediate	Polar covalent	
↓	↓	↓
Large	Ionic	Ionic character

1. The following electrostatic potential diagrams represent H_2, HCL, or NaCl. Label each and explain your choices.



2. Describe the type of bonding that exists in the $F_2(g)$ molecule. How does this type of bonding differ from that found in the HF(g) molecule? How is it similar?

3. When comparing the size of different ions, the general radii trend is usually not very useful. What do you concentrate on when comparing sizes of ions to each other or when comparing the size of an ion to its neutral atom?

4. Predict which bond in each of the following groups will be post polar.
   a. C-F, Si-F, Ge-F
   b. P-Cl or S-Cl
   c. S-F, S-Cl, S-Br
   d. Ti-Cl, Si-Cl, Ge-Cl

5. Predict the type of bond one would expect to form between the following pairs of elements.
   a. Rb and Cl
   b. S and S
   c. C and F
   d. Ba and S
   e. N and P
   f. B and H

6. Hydrogen has an electronegativity value between boron and carbon and identical to phosphorous. With this is mind, rank the following bonds in order of decreasing polarity: P-H, O-H, N-H, F-H, C-H

Ionic Bonds

- Any compound that conducts an electric current when melted will be classified as ionic.
- Composed of elements with very large differences in electronegativity
- Usually made up of a metal and a non metal OR metal and polytomic ion
- **Lattice energy** – energy that is released when two ions of an ionic compound come together to form a crystal
- Higher the charges the greater the attractive energy
- Greater the distance, the smaller the attractive energy
  - The change in energy that takes place when separated gaseous ions are packed together to form an ionic solid.

$$\text{Lattice energy} = k\left(\frac{Q_1Q_2}{r}\right)$$

k = proportionality constant
Q_1 and Q_2 = charges on the ions
r = shortest distance between the centers of the cations and anion

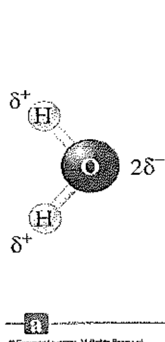
7. Which compound in each of the following pairs of ionic substances has the most exothermic lattice energy? Justify your answer.
   a. NaCl, KCl
   b. LiF, LiCl
   c. $Mg(OH)_2$, MgO
   d. $Fe(OH)_2$, $Fe(OH)_3$
   e. NaCl, Na_2O
   f. MgO, BaS

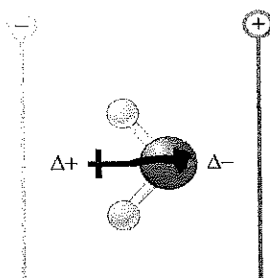
8. The lattice energies of $FeCl_3$, $FeCl_2$, and Fe_2O_3 are (in no particular order) -2631, -5359, -14,774 kJ/mol. Match the appropriate formula to each lattice energy. Explain.

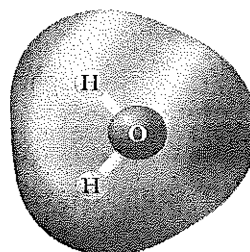
Dipole Moment

- Property of a molecule whose charge distribution can be represented by a center of positive charge and a center of negative charge.
- Use an arrow to represent a dipole moment.
  - Point to the negative charge center with the tail of the arrow indicating the positive center of charge.

Examples





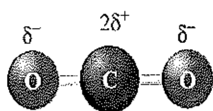


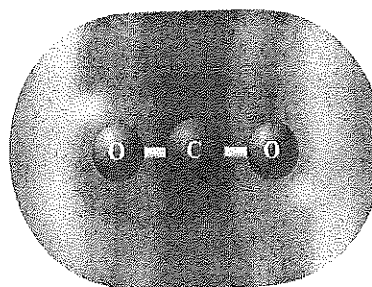
a

Molecular Structure

b
Dipole moment

c
Electrostatic Potential MAP



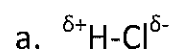


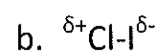
a

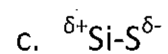
b

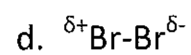
c

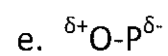
9. Which of the following incorrectly shows the bond polarity? Show the correct bond polarity for those that are incorrect.

   a. 

   b. 

   c. 

   d. 

   e. 

Bond Energies

- To break bonds, energy must be *added* to the system (endothermic, energy term carries a positive sign).
- To form bonds, energy is *released* (exothermic, energy term carries a negative sign).

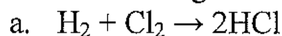
$$\Delta H = \Sigma\ \Delta H\ (\text{bonds broken}) - \Sigma\ \Delta H\ (\text{bonds formed})$$

ΔH in this case represents the BOND ENTHALPY per mole of bonds (always has a positive sign).

Average Bond Energies (kJ/mol)

Bond	Energy	Bond	Energy	Bond	Energy	Bond	Energy
Single Bonds							
H—H	432	N—H	391	Si—H	323	S—H	347
H—F	565	N—N	160	Si—Si	226	S—S	266
H—Cl	427	N—P	209	Si—O	368	S—F	327
H—Br	363	N—O	201	Si—S	226	S—Cl	271
H—I	295	N—F	272	Si—F	565	S—Br	218
		N—Cl	200	Si—Cl	381	S—I	~170
C—H	413	N—Br	243	Si—Br	310		
C—C	347	N—I	159	Si—I	234	F—F	159
C—Si	301					F—Cl	193
C—N	305	O—H	467	P—H	320	F—Br	212
C—O	358	O—P	351	P—Si	213	F—I	263
C—P	264	O—O	204	P—P	200	Cl—Cl	243
C—S	259	O—S	265	P—F	490	Cl—Br	215
C—F	453	O—F	190	P—Cl	331	Cl—I	208
C—Cl	339	O—Cl	203	P—Br	272	Br—Br	193
C—Br	276	O—Br	234	P—I	184	Br—I	175
C—I	216	O—I	234			I—I	151
Multiple Bonds							
C=C	614	N=N	418	C≡C	839	N≡N	945
C=N	615	N=O	607	C≡N	891		
C=O	745 (799 in CO_2)	O_2	498	C≡O	1070		

10. Use the bond energies to estimate ΔH for each of the following reactions in the gas phase:



a. $H_2 + Cl_2 \rightarrow 2HCl$

b. $N_2 + 3H_2 \rightarrow 2NH_3$

11. Use bond energies (In chart above) to predict ΔH for the following reaction:

$H_2S(g) + 3F_2(g) \rightarrow SF_4(g) + 2HF(g)$

12. The major industrial source of hydrogen gas is by the following reaction. Use bond energies to predict the ΔH for this reaction: $CH_4(g) + H_2O(g) \rightarrow CO(g) + 3\ H_2(g)$

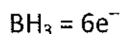
13. Consider the following reaction: $A_2 + B_2 \rightarrow 2AB$ $\Delta H = -285$ kJ
The bond energy for A_2 is one-half the amount of the AB bond energy. The bond energy of B_2 = 432 kJ/mol. What is the bond energy of A_2?

Lewis Structures

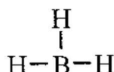
- Shows how valence electrons are arranged among atoms in a molecule.
- Reflects central idea that stability of a compound relates to noble gas electron configuration.
- Used primarily in drawing COVALENT compounds

Exceptions to Octet Rule

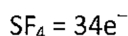
- Boron tends to form compounds in which the boron atom has fewer than eight electrons around it (it does not have a complete octet).

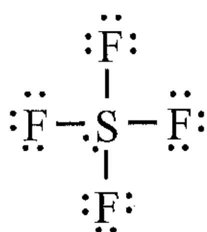


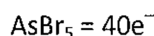
$BH_3 = 6e^-$

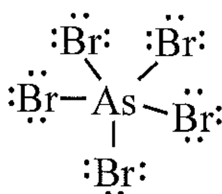


- When it is necessary to exceed the octet rule for one of several third-row (or higher) elements, place the extra electrons on the central atom.


$SF_4 = 34e^-$

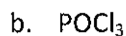


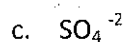

$AsBr_5 = 40e^-$



14. Draw a Lewis structure for each of the following molecules:


a. C_2H_4


b. $POCl_3$


c. SO_4^{-2}


d. SF_4


e. HCN

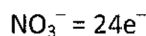

f. NH_4^+

Review

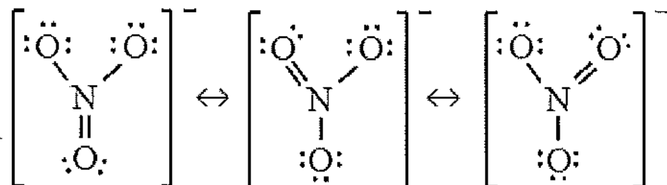
- C, N, O, and F should always be assumed to obey the octet rule.
- B and Be often have fewer than 8 electrons around them in their compounds.
- Second-row elements never exceed the octet rule.
- Third-row and heavier elements often satisfy the octet rule but can exceed the octet rule by using their empty valence *d* orbitals.
- When writing the Lewis structure for a molecule, satisfy the octet rule for the atoms first. If electrons remain after the octet rule has been satisfied, then place them on the elements having available *d* orbitals (elements in Period 3 or beyond).

Resonance

- More than one valid Lewis structure can be written for a particular molecule.
- Actual structure is an average of the resonance structures.
- Electrons are really delocalized – they can move around the entire molecule.



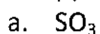
$NO_3^- = 24e^-$



Formal Charge

- Used to evaluate nonequivalent Lewis structures.
- Atoms in molecules try to achieve formal charges as close to zero as possible.
- Any negative formal charges are expected to reside on the most electronegative atoms.
- Formal charge = (# valence e^- on free neutral atom) – (# valence e^- assigned to the atom in the molecule).
    - FC = Valence e^- - (Bonds + Dots)
- Assume:
    - Lone pair electrons belong entirely to the atom in question.
    - Shared electrons are divided equally between the two sharing atoms.
- To calculate the formal charge on an atom: Another formula FC = Valence E - Assigned e
    - Take the sum of the lone pair electrons and one-half the shared electrons.
    - Subtract the number of assigned electrons from the number of valence electrons on the free, neutral atom.

15. Write the Lewis structures for the following. Check with Formal charge, and show all resonance structures where applicable.



a. SO_3

b. N_2O_4

c. OCN^-

16. A toxic cloud covered Bhopal, India in December 1984 when water leaked into a tank of methyl isocyanate and the product escaped into the atmosphere. Methyl isocyanate is used in the production of many pesticides. Draw the Lewis structures for it, CH_3NCO, using Formal charge and including resonance forms if necessary.

17. Order the following species with respect to carbon-oxygen bond length, longest to shortest. Then list the order from the weakest to the strongest bond. CO, CO_2, CO_3^{2-}, CH_3OH

18. Use the formal charge arguments to rationalize why BF_3 would not follow the octet rule.

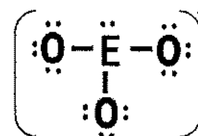
19. Use formal charge moments to explain why CO has a much smaller dipole moment than would be expected on the basis of electronegativity.

20. Oxidation of the cyanide ion produces the stable cyanate ion, OCN^-. The fulminate ion, CNO^-, on the other hand, is very unstable. Fulminate salts explode when struck; $Hg(CNO)_2$ is used in blasting caps. Write the Lewis structure for both ions. Why is fulminate ion so unstable?

21. Complete the chart below:

Molecular Formula	Molecular Structure	Name of Molecular Structure	Bond Angles	Polarity
SeO_3				
SeO_2				
PCl_3				
SiF_4				
$XeCl_2$				
ICl_5				
$XeCl_4$				
$SeCl_6$				

22. Consider the following Lewis structure where E is an unknown element. What are some possible identities for element E? Predict the molecular structure and bond angles for this ion.



23. The molecules BF_3, CF_4, CO_2, PF_5, and SF_6 are all nonpolar, even though they all contain polar bonds. Explain why?