Chapter 8

Bonding: General Concepts

8.1 Types of Chemical Bonds –

Bond energies – energy required to break a bond

Ionic bonding – bonding between two ions – bonding forces that result from electrostatic attractions between adjacent oppositely charged atoms

Ionic compound – formed when an atom that loses electrons easily (metals) reacts with an atom that has a high affinity for electrons (non-metals)

Energy of interaction between pairs of ions can be calculated with Coulomb's Law

Bond length – distance of a bond where the energy is minimal

Covalent bonding – electrons are shared between nuclei

Polar covalent bonds – unequal sharing of electrons ex. H-F

Slightly negative (lowercase delta) – used to indicate fractional charge

8.2 Electronegativity –

Electronegativity – the ability of an atom in a molecule to attract shared electrons to itself

Determining electro negativity – compare the measured H-X bond energy with "expected" bond energy, which is an average of H-H, and X-X bond energies

$$\frac{\text{H-H b.e.} - \text{X-X b.e.}}{2} =$$

Electro negativity generally increases from left to right and decreases going down on the periodic table

Sample Exercise 8.1

Order the following bonds according to polarity: H-H, O-H, Cl-H, S-H, and F-H.

Solution:

Refer to figure 8.3. The polarity of the bond increases as the difference in electronegativity increases. The information from the table gives the following results:

H-H<S-H<Cl-H<O-H<F-H

8.3 Bond Polarity and Dipole Moments

Dipolar molecules – Molecules that have a center of positive charge and a center of negative charge

```
-----|-------------->
    8+        8-
```

Refer to figures 8.4 and 8.5

8.4 Ions: Electron configurations and sizes

-two non-metals react –covalent bond- both go to noble gas state

Nonmetal-metal react to form binary ionic compound, metal emptied, nonmetal goes to noble gas state

Isoelectronic ions – ions containing the same number of electrons

Sample Exercise 8.3

Arrange the ions Se^{2-}, Br^{-}, Rb^{+}, and Sr^{2+} in order of decreasing size.

Solution:

This series is isoelectronic with krypton electron configuration. They all have the same number of electrons so their size will depend on the nuclear charge. The Se^{2-} ion is therefore the largest, and the Sr^{2+} is the smallest.

$$Se^{2-} > Br^{-} > Rb^{+} > Sr^{2+}$$

8.5 Formation of binary ionic compounds

Lattice energy – the change in energy that occurs when separated gaseous ions are packed together to form ionic solids

-Bringing cations and anions together is an endothermic process

-This process becomes more exothermic as ionic charges increase

$$\text{Lattice E} = K\left(\frac{Q_1 Q_2}{r}\right)$$

8.6 Partial Ionic character of covalent bonds

-No individual bonds are completely ionic but the larger the difference in electronegativity of the two atoms, the more ionic character they possess

$$\text{Percent ionic character} = \frac{\text{measured dipole moment X-Y}}{\text{calculated dipole moment X-Y}} \times 100\%$$

-If more than 50% ionic character they are considered ionic solids.

-Ionic compound – any compound that conducts electricity in its molten state

8.7 The covalent chemical bond: A model

-a bond represents the quantity of energy obtained from the overall molecular energy of stabilization

-The chemical bond is only a model

-Discrete bonds are relatively independent of the molecular environment, which lends us to expect each to behave the same no matter where they are located

8.8 Covalent bond energies and chemical reactions

Bond energies – the average of individual bond energies

ΔH – sum of the energies required to break old bonds (pos. signs) plus the sum of the energies released in the formation of new bonds (neg. signs)

ΔH = (bonds broken) – (bonds formed)

8.9 The localized electron bonding model

LE model – a molecule is composed of atoms that are bound together by sharing pairs of electrons using the atomic orbitals of the bound atoms

Lone pairs – pairs of electrons localized on an atom
Bonding pairs – pairs of electrons found in the space between the atoms

8.10 Lewis structures
-Shows how valence electrons are arranged among the atoms in a molecule
-Only valence electrons included
-Hydrogen – duet rule – not octet rule
-Octet rule – elements that form stable molecules when they are surrounded by enough electrons to fill in the valence orbitals, that is, the 2s and the three 2p orbitals.
See example 8.6

8.11 Exceptions to the Octet rule
Boron – tends to form compounds where it has less than eight
Expanded octets – atoms in molecules that exceed the octet rule, 3rd period and beyond

Sample exercise 8.7
Write the Lewis structure for PCl_5.

Solution:
Step 1
Sum the valence electrons: 5 + 5(7) = 40 electrons
Step 2
Indicate single bonds between bound atoms.

Step 3
Distribute the remaining electrons. 30 electrons remain in this case because of the ten used in the bonding.

8.12 Resonance
Definition - When more than one valid Lewis structure can be written for a particular molecule
-Bond lengths and strengths in molecules that undergo resonance are in actuality an average or superposition of all resonance structures
-The localized electron model must be amended to accommodate the delocalized electrons present in these molecules
Formal charge – aids in determining the appropriate Lewis structure that describes the actual bonding of a molecules
-Atoms in molecules try to achieve formal charges as close to zero as possible

-Negative formal charges are expected to reside on the most electro negative atoms

8.13 VSEPR Model

Molecular structure – 3-dimensional arrangement of atoms in a molecule

Valence shell electron pair repulsion model (VSEPR) – useful in predicting the geometries of molecules formed

-The molecular structure based on VSEPR is determined primarily by minimizing electron-pair repulsion

Sample Exercise 8.11

Describe the molecular structure of the water molecule.

Solution:

The Lewis structure for water is:

There are four pairs of electrons: two bonding pairs and two nonbonding pairs. To minimize repulsions, these are best arranged in a tetrahedral array, as shown in Fig. 8.16(a). Although H_2O has a tetrahedral arrangement of electron pairs, it is not a tetrahedral molecule. The atoms in the H_2O molecule form a V shape, as shown in Fig. 8.16(b) and (c).