

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

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

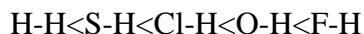
Electro negativity generally increases from left to right and decreases going down on the periodic table

Sample Exercise 8.1

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

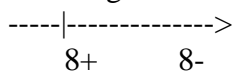
Solution:

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:



8.3 Bond Polarity and Dipole Moments

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



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

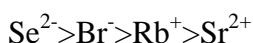
Isoelectronic ions – ions containing the same number of electrons

Sample Exercise 8.3

Arrange the ions  $\text{Se}^{2-}$ ,  $\text{Br}^-$ ,  $\text{Rb}^+$ , and  $\text{Sr}^{2+}$  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  $\text{Se}^{2-}$  ion is therefore the largest, and the  $\text{Sr}^{2+}$  is the smallest.



### 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

$$\text{Lattice } E = \frac{K(Q_1Q_2)}{r}$$

### 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

$\Delta H$  – 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)

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

### 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, 3<sup>rd</sup> period and beyond

#### Sample exercise 8.7

Write the Lewis structure for  $\text{PCl}_5$ .

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  $\text{H}_2\text{O}$  has a tetrahedral arrangement of electron pairs, it is not a tetrahedral molecule. The atoms in the  $\text{H}_2\text{O}$  molecule form a V shape, as shown in Fig. 8.16(b) and (c).