Chem101: General Chemistry
Lecture 4 - Forces Between Particles

I. Introduction

A. Forces between Particles

1. The chemical and physical properties of matter result from interactions that take place between their constituent particles, *i.e.* their molecules and atoms.
2. For example
   a. Boiling points and melting points are determined by the strength of the forces holding the molecules together in the solid and liquid states.
   b. Electrical conductivity in metals arises from electrons that are loosely held by the metal atoms.
   c. The atoms in molecules are held together by covalent bonds, which involves attractive forces between atoms that arise from the sharing of valence electrons.

II. Noble Gas Configurations

A. Noble Gases

1. One approach to understand what causes particles to interact with one another, is to compare them to some that do not.
2. The noble gas are a family of non-interacting elements.
   a. The far right-hand column of the periodic table represents this group of elements.
3. Noble gases interact only very weakly with themselves, and with few exceptions, do not form compounds with other elements.
4. In Chapter 3 we saw that the chemical and physical properties of the elements, as manifested by their arrangement in the periodic table, are determined by their electron configurations.

B. Noble Gas Configuration

1. With the exception of helium, the noble gases have the same number of electrons in their valence shell.
   a. Remember, the valence shell is the highest energy shell that contains electrons. (highest *n* number).
      i. The electrons in the valence shell are found in the *s* and *p* subshells.
      ii. These are the electrons that are closest to the surface of the atom.
2. For Noble gases, the *s* and *p* subshells in the valence shell are filled.
   a. Consequently there are 8 electrons in the valence shells of the noble gases.

i. The exception is helium, whose valance electrons are in the 1st shell, which has only an *s* subshel and therefore only two electrons instead of eight.

b. The unusual lack of interactions between noble gasses and other elements suggest that there is something particularly stable with having the *s* and *p* subshells in the valence shell filled.

3. This observation led G. N. Lewis (an American) and Walter Kossel (a German) in 1916 to independently propose a rule, called **the octet rule**, which can be used to predict how and why elements combine to form compounds.

The Octet Rule -Elements combine to form compounds so that each element in the compound has 8 electrons in their valence shell.

4. G. N. Lewis developed a symbolic notation that can be used in applying the octet rule.

a. An element is represented by its symbol and is surrounded with zero to eight dots.

i. Each dot representing one of the valence electrons.

b. These notations are called **Lewis dot structure**.

c. The number of valence electrons for the representative elements is equal to the element's group number (IA through VIIA) on the periodic table.

C. Lewis Dot Structures havoc

1. For example:

a. The Lewis dot structure for aluminum is $\cdot\dot{\mathrm{Al}}\cdot$

i. Aluminum is in Group IIIA and has 3 valence electrons.

ii. The abbreviated electronic configuration for aluminum is [Ne]$\mathbf{3s^2 3p^1}$.

b. The Lewis dot structure for chlorine is $:\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}\cdot$

i. Chlorine is in group VIIA and has 7 valence electrons.

ii. The abbreviated electronic configuration for chorine is [Ne]$\mathbf{3s^2 3p^5}$.

c. The Lewis dot structure for selenium is $:\ddot{\underset{\cdot}{\mathrm{Se}}}\cdot$

i. selenium is in group VIA and has 6 valence electrons.

ii. The abbreviated electronic configuration for selenium is [Ar]$\mathbf{4s^2}\,3d^{10}\mathbf{3p^4}$.

Exercise 4.1

4.1 Refer to the group numbers of the periodic table and draw Lewis structures for atoms of the following:

a. Lithium

b. Sodium

c. Chlorine

d. Boron

Exercise 4.3

4.3 Write the abbreviated electronic configurations for the following:

a. iodine

b. element number 38

c. As

d. Boron

III. Ionic bonding

A. On way that elements interact with one another to satisfy the octet rule is to transfer electrons from one atom or group of atoms to another atom or group of atoms.

1. The atoms or groups of atoms that lose electrons become positively charged.
    a. Charged species are called **ions**.
        i. A positively charged ion is called a **cation**.
2. The atoms or groups of atoms that gain electrons become negatively charged.
    a. A negatively charged ion is called an **anion**.
3. Because the number of electrons lost by the cations is equal to the number of electrons gained by the anions, the net charge of the combined ions is zero.

B. Positive ions (cations) are electrostatically attracted to the negative ions (anions).

1. Such interactions, which hold ionic species together, is called an **ionic bond**.

C. The number of electrons gained or lost by a single atom rarely exceeds three

1. Group IA, IIA, and IIIA elements lose 1, 2 or 3 electrons respectively to become positively charged ions.
   a. After losing the electrons they have an electronic configuration that is the same as one of the noble gases.
2. On the other hand, Groups VA, VIA and VIIA elements become ions by gaining 3, 2 or 1 electrons respectively.
   a. the number electrons equal to 8 minus their group number.
   b. They form -3 (VA), -2 (VIA) and -1 (VIIA) ions.
   c. Again, after gaining the electrons they have an electronic configuration that is the same as one of the noble gases.
3. Hydrogen is different
   a. In some cases it loses one electron to become a H^+ ion.
   b. In other cases it gains one electron to become H^- ion.
   c. For this reason hydrogen is sometimes grouped with the IA elements and at other times is grouped with the VIIA elements

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Exercise 4.9

4.3 Indicate both the minimum number of electrons that would have to be added and the minimum number that would have to be removed to change the electronic configuration of the each element listed in Exercise 4.3 to a noble gas configuration```` 0:

a. iodine

b. element number 38

c. As

d. Boron

D. Chemical and Physical Properties of Ions

1. When an atom become an ion, its chemical and physical properties change dramatically from that of the neutral element.
   a. Its electronic structure becomes more stable than its corresponding neutral element so they are less chemically reactive.
   b. Eventhough their electronic configuration is like that of a noble gas, their properties are quite different than that of a noble gas
      i. This is because ions are charge and interact strongly with other ions.

IV. Ionic Compounds

A. Ionic compounds are held together by ionic bonds.

B. Any time and atom gains or loses an electron, that electron is transferred from or to another atom or group of atoms.

1. The charges that form on the ions balance one another out.

C. For example:

1. If the metal sodium interacts with the non-metal chlorine:
    a. sodium loses 1 electron to become Na^+.
        i. $Na \rightarrow Na^+ + 1e^-$
    b. Chlorine gains 1 electron to become Cl^-
        i. $Cl + 1e^- \rightarrow Cl-$
    c. Together they form the ionic compound, sodium chloride (NaCl).
2. If the metal magnesium interacts with the non-metal chlorine:
    a. Magnesium loses 2 electrons to become Mg^{+2}
        i. $Mg \rightarrow Mg^{+2} + 2e^-$
    b. Chlorine gains 1 electron to become Cl^-
        i. $Cl + 1e^- \rightarrow Cl^-$
    c. Since chlorine only requires one electron to attain a noble gas electronic configuration, it takes two chlorines to accept both of the electrons given up by the one magnesium
    d. Together they form the ionic compound, magnesium chloride ($MgCl_2$).
    e. Ionic Compounds

Exercise 4.19

4.19 Write equations to represent positive and negative ion formation for the following pairs of elements. Then write a formula for the ionic compound that results when the ions combine:

a. Mg and S

b. strontium and nitrogen

c. elements number 3 and 34

V. Naming Ionic Compounds

A. Ionic compounds that contain only two different types of atoms are called binary compounds.

1. The names given to binary compounds follows the convention:
    a. Name = metal + nonmetal stem + *-ide*
2. For example:

a. Sodium chloride
b. Magnesium chloride

B. When the metal in an ionic compound is a transition metal, instead of a representative metal, it usually has more than one charged state:

1. The name must indicate which charged state the ion is in

C. Naming Ionic Compounds

1. $CuCl$ and $CuCl_2$ are two different pure substances with different chemical and physical properties.
2. To distinguish them we give them two different names:
   a. $CuCl$ is Copper(I) chloride
   b. $CuCl_2$ is Copper(II) chloride.
3. A Roman numeral is used to indicate the magnitude of the charge on the ion.

Exercise 4.27

4.27 Write equations to represent positive and negative ion formation for the following pairs of elements. Then write a formula for the ionic compound that results when the ions combine:

a. K_2O

b. $SrCl_2$

c. Al_2O_3

d. $LiBr$

e. CaS

Exercise 4.31

4.31 Name the following binary ionic compounds, using a Roman numeral to indicate the charge on the metal ion:

a. $CrCl_2$ and $CrCl_3$

b. CoS and Co_2S_3

c. FeO and Fe_2O_3

d. $PbCl_2$ and $PbCl_4$

Exercise 4.35

4.35 Write formulas for the following binary ionic compounds::

a. manganese(II) chloride

b. iron(III) sulfide

c. chromium(II) oxide

d. iron(II) bromide

E. tin(II) chloride

VI. The Smallest Unit of Ionic Compounds

A. Ionic compounds are not made of true molecules.

1. In the solid state they form a **crystal lattice** in which the ions have a very regular arrangement in which no ion is associated with any other particular ion.
2. The chemical formula for an ionic compound gives the ratio of ionic components.
    a. It is not meant to imply that a particular positive ion is associated with a particular negative ion.
3. For this reason, the chemical formula for ionic compounds usually is called the **formula unit**.
4. We usually speak of an ionic compound's **formula weight** instead of its molecular weight.
    a. Like the molecular weight, the formula weight can be used to determine the number of formula units present in a sample of an ionic compound.

B. When an ionic compound dissolves in water to form a solution, its component ions separate from one another.

1. The formula unit can still be used to determine the number of positive and negative ions present:
2. $NaCl \rightarrow Na^{+} + Cl^{-}$
3. $MgCl_2 \rightarrow Mg^{+2} + 2Cl-$
    a. In the second example, the 2 subscript indicates that there are 2 chlorides in the formula unit.
        i. When $MgCl_2$ dissolves in water, the two choride ions and the magnesium ion separate from one another, therefore the 2 is placed in front of the Cl^{-} ot indicate that there are 2 separate Cl^{-} formed.

VII. Covalent Bonding

A. Forming ionic bonds is not the only way that elements can satisfy each other's desire to have a noble gas electronic configuration

B. Another way for atoms to achieve the electronic configuration of a noble gas, at least part of the time, is to share electrons with other atoms.

Figure 4.4 - Formation of covalent bond between hydrogen atoms

1. Covalent bonding is observed between non-metals
   a. Non metals have a stronger affinity for electrons than metals
      i. They typically need one to three elecrons to fill their valence shell.
   b. In forming covalent bonds they achieve their octet by sharing electrons with other nonmetals

C. Covalent bonds a re formed by overlapping half filled atomic orbitals from to atoms to form a filled **molecular orbital**.

D. There is a systematic way of using *Lewis dot structures* of atoms to predict the structures of the covalently bonded molecules that they form.

1. Use the molecular formula to determine how many atoms of each kind are in the molecule.
2. Us the given connecting pattern of the atoms to draw an initial structure for the molecule with the atoms arranges properly.
3. Determine the total number of valence-shell electrons contained in the atoms of the molecule.
4. Place one pair of electrons between each bonded pair of atoms,
   a. These electrons are located the molecular orbitals that were formed by the overlapping of half-filled atomic orbitals.
5. Place the remaining valence electrons on the atoms to bring each up to an octet.
   a. In determining the number of valence electrons surrounding each atom, The bonding electrons are counted by both of the atoms involved in a bond.
      i. Hydrogens require only 2 electrons, not 8, like the other elements.
   b. These other electrons are called non-bonding electrons
6. If you run out of electrons before satisfying the octet rule for each atom in the molecule, then use non-bonding electrons to form double or triple bonds.
   a. Again, the electrons participating in double and triple bonds are counted by both of the elements participating in the bond.

Exercise - Draw the Lewis dot structure for SO_3 and H_2C_2.

VIII. Polyatomic Ions

A. Earlier we looked at *simple ions*, which comprise single atoms that have either gained or lost electrons.

B. Poly atomic ions are ions that contain more than one atoms

1. This group of atoms are covalently bonded, and the group as a whole has an electrical charge.

C. Most polyatomic ions are negatively charged

1. The ammonium ion (NH_4^+) is an exception

D. As with covalent molecules, Lewis dot structures can be used to represent the structures of polyatomic ions.

1. The steps used are the same, except when counting up the valence electrons, electrons are added, or removed, from the count to produce the correct electronic charge.

Exercise - Draw the Lewis dot structure for the sulfate ion (SO_4^{2-}), and the carbonate ion (CO_3^{2-})

IX. Shapes of molecules and polyatomic ions

A. Molecules and polyatomic ions have three dimensional shapes

1. Predicting the shape is important to understanding the properties of molecules and polyatomic ions.

B. The **V**alence **S**hell **E**lectron **P**air **R**epulsion (**VSEPR**) theory is a simple theory that can be applied to predict the shapes of molecules and polyatomic ions.

1. To determine the shape of a molecule using the VSEPR theory, count the number of regions of electron density surrounding the central atom of a molecule or polyatomic ion.
   a. Count each bonding pair and each non-bonding pair of electrons as one region of electron density.
   b. Count double and triple bonds as a single region of electron density.

2. The geometry is determined by the number of regions of electron density.
    a. The geometry used is the one that places the regions of electron density as far apart from one another as possible.

Regions of Electron Density	**Electronic Geometry**	**Bond Angles**
2	Linear	180°
3	Trigonal Planar	120°
4	Tetrahedral	109.5°

    b. Balloons can be used to demonstrate this.

C. A VSEPR Tutorial available at Georgia Southern University Web site

"http://www2.gasou.edu/chemdept/general/molecule/index.htm"

1. The Shape of the molecule is determined by the location of the atoms and the bonds.
    a. For a *linear* electron geometry:
        i. There are no non-bonding electron pairs, so the shape is also *linear*.
    b. For a *trigonal planar* electron geometry:
        i. With no non-bonded electron pairs, the shape is also *trigonal planar*.
        ii. With one non-bonded electron pair, the shape is a *bent* molecule with a bond angle of 120°
    c. For a tetrahedral geometry:
        i. With no non-bonded electron pairs, the shape is also *tetrahedral*.
        ii. With one non-bonded electron pair, the shape is a *triangular pyramidal* molecule with bond angles of 109.5°.
        iii. With two non-bonding electron pairs, the shape is a *bent* molecule with a bond angle of 109.5°.

X. Polarity of Covalent Molecules

A. Nonpolar covalent bonds

1. When covalent bonds form between atoms of the same element, the electrons are shared equally.

B. Electronegativity

1. Not all atoms have the same affinity for electrons.
2. With the exception of the Noble gases, the elements that are in the upper right-hand region of the periodic table have the highest affinity for

electrons, whereas those in the lower left-hand corner have the lowest affinity.

Table 4.3 - Electronegativites for the common representative elements

C. Bond polarization

1. When bonds form between atoms having different electronegativities, the electrons spend more of their time in the vicinity of the atom with the higher electronegativity.
    a. This results in a partial positive charge at one end of the bond (δ^+) and a partial negative charge at the other end (δ^-).
    b. This separation of positive and negative charge is called a dipole.
        i. This is designated with an arrow that points from the positive to the negative charge.
2. The net charge of the dipole zero.
3. The electronegativity scale ranges from 0 to 4 and gives the relative affinity of each atom for its electrons
4. The electronegativity scale is used to determine the polarity of covalent bonds
    a. If the difference in electronegativities for the two atoms is greater than 2.1 ($\Delta EN > 2.1$), the bond is considered ionic.
    b. If the difference in electronegativities for the two atoms is less thatn 2.1 ($\Delta EN < 2.1$), but greater than 0, the bond is considered polar covalent.

Figure 4.7 - The ionic/polar bond continuum

Exercise - Polar bonds in I-Cl, Br-Br and C≡O

D. Polar molecules

1. Have polar bonds
2. The arrangement of the polar bonds is nonsymmetric, so that the molecule as a whole has a net dipole.

XI. More about Naming Compounds

A. Covalent binary compounds that form between nonmetals often combine in varying ratios.

1. For example
    a. SO_2, SO_3
    b. NO, N_2O, NO_2, N_2O_3
2. For this reason, when naming these compounds the numbers of each atom has to be included in the name.
    a. The prefixes, *mono-, di-, tri-, tetra-, penta-, hexa-, hepta-, octa-, nona-*, and *deca-* are used to indicate the numbers.

3. The name of less electronegative element is given first
4. Use *-ide* suffix for the second element.
5. For example
    a. SO_2 sulfur dioxide
    b. SO_3 sulfur trioxide
    c. NO nitrogen oxide
    d. N_2O dinitrogen oxide
    e. NO_2 nitrogen dioxide
    f. N_2O_3 dinitrogen trioxide

B. Polyatomic ions.

1. These are usually anions that are made up of nonmetals that are covalently bonded to one another and which as a group have a net negative charge.

Table 4.6 - Some common polyatomic ions

XII. Other inter particle forces

A. Most pure substances in their pure state exist in a crystal lattice similar to the one we looked at before for the ionic compound NaCl (Figure 4.3)

1. In some cases it is ions that inhabit the lattice sites, while in others it is molelcules or atoms.

Table 4.7 - Some characteristics of selected pure substances.

B. If the temperature of a solid increased, it will at some point convert to a liquid, and upon further heating convert to a gas.

Table 4.8 - Behavior of selected pure substances on heating

C. There is an array of forces that hold matter together,

1. The strongest ones include two forces that we have already introduced:
    a. Covalent bonding
    b. Ionic bonding

D. Network solids

1. The crystal lattice sites are occupied by atoms that are covalently bonded to each other.
2. These produce the strongest solids with the highest melting points.
3. For example
    a. Diamonds (C)
    b. Quarts (SiO_2)

E. Metallic bonding

1. Metals readily lose their valence electrons.
2. In the solid state these valence electrons move freely from one atom to the next.
   a. This is why metals have high electrical conductivity.
3. The passing off of the valence electrons from one metal atom to the next is what holds that atoms together in solid and liquid states.

F. Ionic bonding

1. As discussed earlier, ionic bonding is due the attraction that positive (cations) and negative (anions) have for each other.
   a. The cations are usually metals, which have lost electrons to achieve the electron configuration of a noble gas.
   b. The anions are usually non-metals, which have gained electrons to achieve the electron configuratin of a noble gas.
2. This type of interaction is called a **charge/charge interaction**.
3. Ionic bonds are quite strong in air but can be greatly weaken by the presence of polar substances such as water.
   a. This is why many ionic compounds (salts) dissolve in water.
      i. The water weakens the ionic interactions sufficiently that the ions separate from one another.

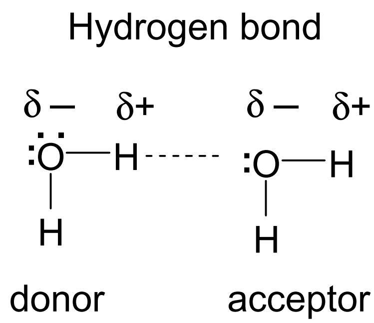
G. Dipolar forces

1. This is the attraction that arises between polar molecules.
   a. Even though polar molecules do not have a net charge like ions do, each molecule contains regions of positive and negative charge.
2. Polar molecules will arrange themselves so that the positive regions of charge on one molecule are close to the negative regions of charge on a neighboring molecule.
3. This type of interaction is referred to as a **dipole/dipole interaction**.
4. Dipole forces are moderately strong in air, but like ionic bonds they can be greatly weakend by the presence of polar substances such as water.

H. Hydrogen bonding

1. A special case of the dipole force is the hydrogen bond.
2. Hydrogens form polar bonds with electronegative elements such as oxygen and nitrogen.
   a. The hydrogen will have positive partial charge (δ^+) relative to the oxygen or nitrogen.
   b. The hydrogen is then attracted to electronegative atoms on other molecules.
      i. The electronegative atom to which the hydrogen is covalent bonded is called the hydrogen bond donor,
      ii. The electronegative atom to which the hydrogen is attracted to is called the hydrogen bond acceptor.
3. Hydrogen bonding is what holds water together.

a. It is what makes water more cohesive than other substances of comparable composition and molecular weight.



b. Comparing the boiling points of the compounds formed by bonding two hydrogens to a group VI element:

   i. Water does not fit the trend established by the other elements in family VI
   ii. The unusally high boiling point for water is due to its ability to form extensive networks of hydrogen bonds.

Compound	Boiling Point
H_2O	100°C
H_2S	-60.3°C
H_2Se	-41.°C
H_2Te	-2.2°C

c. Water molecules are capable of forming 4 hydrogen bonds
   i. Two as an acceptor.
   ii. Two as a donor.

d. In liquid water not all of the possible hydrogens are made.
   i. The hydrogen bonds are made and broken many times a second.
      1. This is what allows the water molecules to slide past one another in liquid water.

e. When the ice forms , all 4 hydrogen bonds are made.
   i. The water must expand in order to accommodate the forth hydrogen bond.
      1. This is why ice floats; it is less dense than liquid water.

I. Dispersion forces

1. The weakest interparticle force is called a dispersive interaction or dispersive force.
    a. It is also called the London dispersion force.
2. Even in nonpolar molecules there are transient dipoles formed as distribution of electrons changes with time.
    a. On average there is no net dipole, but on a very short time scaled there are transient ones.
    b. When two nonpolar molecules are brought close to one another, their transient dipoles will synchronize in order to always attract one another.
3. This interaction is also referred to as and **induced dipole/induced dipole interaction**.
    a. The transient dipole of one molecule induces an attractive transient dipole in its neighboring molecules.
4. The dispersive interaction allows molecules which lack any of the other attractive interactions to form liquids and solids.
5. All matter exhibits dispersive interactions, though the other stronger forces mask the effect if they exist.

Figure 4.12 - Relative strengths of interparticle forces.