Chem101 - Lecture 4

Forces Between Particles

Forces between Particles

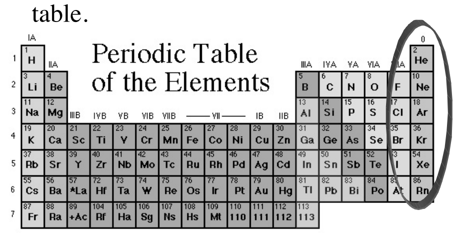
- The chemical and physical properties of matter result from interactions that take place between their constituent particles, *i.e.* their molecules and atoms.
  - Particles interact because there are forces that arise between them.
- In this lecture we will be looking at some of these different forces and will try to understand what produces them.

Noble Gases

- One approach to trying to understand what causes particles to interact with one another is to first compare them to some that do not.
- Among the the 100 or so elements that exist, there are only a handful that have atoms that do not interact with either themselves, or with atoms from other elements.

Noble Gases

- These elements are in a group called the noble gases and are located on the far right-hand side of the periodic table.



Noble Gases

- Noble gases do not interact with themselves.
  - Unlike most other elements, their molecules are monoatomic.
    - Their atoms do not form bonds with other atoms.
  - Unlike most other elements, they exist primarily as gases down to very low temperatures.
    - Their molecules are not attracted to other molecules.

Noble Gas Configuration

- Chemical reactivity involves an atom's electron configuration.
- This lead scientists to ask, what is special about the electron configurations of noble gases that makes them so inert (unreactive).

Element	*Electron Configuration*
He:	$1s^2$
Ne:	$1s^2 2p^6$
Ar:	$1s^2 2p^6 3s^2 3p^6$
Kr:	$1s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
Xe:	$1s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4f^{14} 5p^6$
Rn:	$1s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6$

Noble Gas Configuration

- What they noticed is that all the noble gases have the same number of electrons in their valence shell.
  - Remember, the valence shell is the highest occupied shell (highest *n* number).

Element	*Electron Configuration*	*Valence Electrons*
He:	$\mathbf{1s^2}$	2
Ne:	$1s^2\mathbf{2p^6}$	8
Ar:	$1s^22p^6\mathbf{3s^23p^6}$	8
Kr:	$1s^22p^63s^23p^6\mathbf{4s^2}3d^{10}\mathbf{4p^6}$	8
Xe:	$1s^22p^63s^23p^64s^23d^{10}4p^6\mathbf{5s^2}4f^{14}\mathbf{5p^6}$	8
Rn:	$1s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6\mathbf{6s^2}4f^{14}5d^{10}\mathbf{6p^6}$	8

Chemical Bonding

- The observation that the noble gases have a particularly stable electron configuration, led G. N. Lewis (an American) and Walter Kossel (a German) in 1916 to independently propose a rule that can be used to predict how and why elements combine to form compounds.
- The rule is called the octet rule.
  - *Octet* means a group of 8.

The Octet Rule

- Elements combine to form compounds so that each element in the compound will have 8 electrons in their valence shell.
- Hydrogen is the exception.
  - It combines with other elements so that like helium, it can have 2 electrons in its valence shell.

Lewis Dot Structures

- G. N. Lewis developed a symbolic notation that can be used in applying the octet rule.
- He used the chemical symbol of an element to represent the element's nucleus plus all of its electrons except those in the valence shell.
- Dots are placed around the chemical symbol to represent the valence electrons.

Lewis Dot Structures

- For example:
  - Sodium (Na) has and atomic number of 11.
    - Its electronic configuration is
    - $1s^22s^22p^6\mathbf{3s^1}$
    - It has **1** valence electrons.
    - The Lewis Dot structure for Sodium is Na·
  - Chlorine (Cl) has and atomic number of 17.
    - Its electronic configuration is
    - $1s^22s^22p^6\mathbf{3s^23p^5}$
    - It has **7** valence electrons.
    - The Lewis Dot structure for Chlorine is :C̤̈l·

Exercise 4.1

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

a. lithium
b. sodium
c. chlorine
d. boron

Exercise 4.3

Write abbreviated electronic configurations for the following:

a. iodine
b. element number 38
c. As
d. phosphorous

Ions

- For some interactions, the octet rule is satisfied by transferring one or more electrons from one element to another.
- The resulting atoms will be charged
  - These charged atoms are called simple ions.

Ions

- An atom that gains one or more electrons becomes negatively charged.
  - It has more electrons than protons.
  - It is called a anion.
- An atom that loses one or more electrons becomes positively charged.
  - It has more protons than electrons.
  - It is called an cation.

Ionic Bonding

- Positive ions are electrostatically attracted to negative ions.
  - The pairing of ions with opposing charges is called an ionic bond.
- In general
  - Metals lose electrons
  - Non-metals gain electrons
- Remember, the first ionization energies of metals are less than those for non-metals.
- Ionic bonds are commonly formed between ions derived from metals, and ions derived from non-metals.

Ionic Bonding

- The number of electrons gained or lost by a single atom rarely exceeds three
- For the representative (main group) elements, Group IA, IIA and IIIA elements become ions by losing the number electrons equal to their group number.
  - They form +1 (IA), +2 (IIA) and +3 (IIIA) ions.

Ionic Bonding

- On the other hand, Groups VA, VIA and VIIA elements become ions by gaining the number electrons equal to 8 minus their group number.
  - They form -3 (VA), -2 (VIA) and -1 (VIIA) ions.

Ionic Bonding

- Hydrogen is different
  - In some cases it loses one electron to become H^+
  - In other cases it gains one electron to become H^-
  - For this reason hydrogen is sometimes grouped with the IA elements and at other times is grouped with the VIIA elements

Exercise 4.9

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

a. iodine
b. element number 38
c. As
d. phosphorous

Chemical and Physical Properties of Ions

- When an atom become an ion, its chemical and physical properties change dramatically from that of the neutral element.
  - Even though they have an electronic configuration like that of a noble gas, their properties also differ dramatically from those of their corresponding noble gas
  - This is because, unlike the noble gases, they are charged.

Ionic Compounds

- Ionic compounds are held together by ionic bonds.
- Any time and atom gains or loses an electron, that electron is transferred from or to another substance.
- The charges that form on the ions balance one another out.

Ionic Compounds

- For example:
- If the metal sodium interacts with the non-metal chlorine,
  - Sodium loses 1 electron to become Na^{+}.
    $Na \longrightarrow Na^+ + 1e^-$
  - Chlorine gains 1 electron to become Cl^-
    $Cl + 1e^- \longrightarrow Cl^-$
  - Together they form the ionic compound, sodium chloride (NaCl).

Ionic Compounds

- If the metal magnesium interacts with the non-metal chlorine,
  - Magnesium loses 2 electrons to become Mg^{+2}
    $Mg \longrightarrow Mg^{+2} + 2e^-$
  - Chlorine gains 1 electron to become Cl^-
    $Cl + 1e^- \longrightarrow Cl^-$
  - It takes two chlorines to accept both of the electrons given up by one magnesium
  - Together they form the ionic compound, magnesium chloride ($MgCl_2$).

Exercise 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

Naming Ionic Compounds

- Ionic compounds that contain only two different types of atoms are called binary compounds.
- The names given to binary compounds follows the convention:
  - Name = metal + nonmetal stem + *-ide*
  - For example:
    - Sodium chloride
    - Magnesium chloride

Naming Ionic Compounds

- When the metal in an ionic compound is a transition metal, instead of a representative metal, it usually has more than one charged state:
- For example:
  - Copper has two ionic forms - Cu^{+} and Cu^{+2}.
  - Combining with the chloride ion it can form either CuCl or $CuCl_2$.

Naming Ionic Compounds

- CuCl and $CuCl_2$ are two different pure substances with different chemical and physical properties.
- To distinguish them we give them two different names:
  - CuCl is Copper(I) chloride
  - $CuCl_2$ is Copper(II) chloride.

Exercise 4.27

Name the following binary ionic compounds:

a. K_2O
b. $SrCl_2$
c. Al_2O_3
d. LiBr
e. CaS

Exercise 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

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

The Structure of Ionic Compounds

- Ionic compounds are not made of true molecules.
- In the solid state they form a crystal lattice in which the ions have a very regular arrangement.
  - The chemical formula for an ionic compound gives the ratio of ionic components.
    - It is not meant to imply that a particular positive ion is associated with a particular negative ion.

The Structure of Ionic Compounds

- For ionic compounds the chemical formula is sometimes called the formula unit.
- We usually speak of an ionic compounds formula weight instead of its molecular weight.
- 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.

The Structure of Ionic Compounds

- When an ionic compound dissolves in water to form a solution, its component ions separate from one another.
  - The formula weight can still be used to determine the number of positive and negative ions present:
  - $NaCl \longrightarrow Na^{+} + Cl^{-}$
  - $MgCl_2 \longrightarrow Mg^{+2} + 2Cl^{-}$

Covalent Bonds

- Forming ionic bonds is not the only way that elements can satisfy each other's desire to have a noble gas electronic configuration
  - Carbon would have to either gain or lose 4 electrons to obtain a noble gas configuration.
  - Many atoms form molecules with themselves where neither wants to give up any electrons:
    - O_2, F_2, N_2, H_2, Cl_2

Covalent Bonds

- 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.
  - This sharing leads to the formation of a covalent bond.

Covalent Bonds

- Figure 4.4 illustrates how this happens:
  - Unfilled atomic orbitals on the atoms involved overlap to form a molecular orbital.
  - The molecular orbital, which contains the shared electrons, lies between the atomic nuclei so that both nuclei is attracted to this negatively charged region.

Covalent Bonds

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

Lewis Dot Structures

1. From the molecular formula determine the atoms involved.
2. Use the given connecting pattern of the atoms to draw an initial structure.
   - Connect the atoms at first by single bond.
3. Add up all the valence electrons of the atoms involved.
4. Subtract from this number, two electrons for each bond.

Lewis Dot Structures

5. Place these remaining electrons on the atoms, attempting to complete an octet for each atom.
   - These electrons are called non-bonding electrons.
   - The bonding electrons, which are being shared, are counted by both of the atoms when determining if they have an octet.

Lewis Dot Structures

6. If there are not enough electrons to complete an octet of all of the atoms, then non-bonding pairs of electrons are moved between the atoms to form double and triple bonds.
   - When counting up the octets, double bonds count as four electrons and triple bonds count as six.

Lewis Structures

- Below is a link to a Web Site at Georgia Southern University
- Among other things, the site provides a nice example of how to draw Lewis dot structures of molecules

Click here to go to Lewis Structures at Georgia South University http://www2.gasou.edu/chemdept/general/molecule/index.htm

- Table 4.2 in your text also provides some examples.

Exercise 4.45

Represent the following reaction using Lewis structures:

$$I + I \longrightarrow I_2$$

Exercise 4.45 - Answer

Represent the following reaction using Lewis structures:

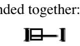
$$I + I \longrightarrow I_2$$

A. Iodine has an atomic number of *53* and an abbreviated electron structure [Kr]$\mathbf{5s^2}4d^{10}\mathbf{5p^5}$, it therefore has 7 valence electrons. Its Lewis dot structure is

Exercise 4.45 - Answer

B. The covalent structure of the diatomic iodine molecule, I_2, can be determined by applying the systematic rules for drawing Lewis structures.

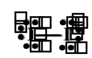
1. There are two iodine atoms in the molecule.
2. Since this is a diatomic molecule there is only one way that the atoms can be bonded together:



3. There are a total of 2 x 7 = 14 valence electrons.
4. Currently 2 of those electrons are being used to make the covalent bond, leaving 14 - 2 = 12 non-bonding electrons.

Exercise 4.45 - Answer

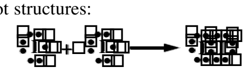
5. Distributing theses non-bonding electrons equally about the two iodine atoms produces the following proposed Lewis dot structure for I_2:



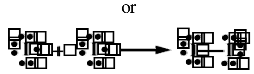
6. This structure contains the requisite 8 electrons around each atom and therefore satisfies the octet rule. No double or single bonds are required.

Exercise 4.45 - Answer

C. Representing the chemical reaction using Lewis dot structures:



or



Exercise 4.47

Represent the following moelceuls by Lewis dot structures:

a. HF
b. IB
c. PH_3 (each H atom is bonded to the P atom)
d. $HClO_2$ (the O atoms are each bonded to the Cl, and the H is bonded to one of the O atoms

Polyatmomic Ions

- Earlier we looked at *simple ions*, which comprise single atoms that have either gained or lost electrons.
- There are also many ions that contain more than one atom.
  - These atoms are held together by covalent bonds and as a group have an electrical charge and therefore are ions.
  - These species are called polyatomic ions.
  - With the exception of the ammonium ion, NH_4^+, most are negative ions.

Polyatmomic Ions

- Lewis dot structures can be used to represent the structures of polyatomic ions
  - In the third step the number of valence electrons is modified to account for the charge on the ion.

Table 4.6-Some common polyatomic ions

Very common		Common	
NH_4^+	ammonium	CrO_4^{2-}	chromate
$C_2H_3O_2^-$	acetate	$Cr_2O_7^{2-}$	dichromate
CO_3^{2-}	carbonate	NO_2^-	nitrite
ClO_3^-	chlorate	MnO_4^-	permanganate
CN^-	cyanide	SO_3^{2-}	sulfite
HCO_3^-	hydrogen carbonate (bicarbonate)	ClO^-	hypochlorite
OH^-	hydroxide	HPO_4^{2-}	hydrogen phosphate
NO_3^-	nitrate	$H_2PO_4^-$	dihydrogen phosphate
PO_4^{3-}	phosphate	HSO_4^-	hydrogen sulfate (bisulfate)
SO_4^{2-}	sulfate	HSO_3^-	hydrogen sulfite (bisulfite)

Exercise 4.49 c

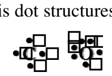
Draw Lewis structures for the following polyatomic ion:

c. CO_3^{2-} (each O atom is bonded to the C atom)

Exercise 4.49 c - Answer

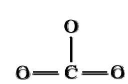
1. This is the carbonate ion.

- It contains 1 carbon, 3 oxygens and has a net charge of -2.
- Carbon has an atomic number of 6 and an abbreviated electrons structure [He]$\boldsymbol{2s^2 2p^2}$. It has 4 valence electrons.
- Oxygen has an atomic number of 8 and an abbreviated electronic structure of [He]$\boldsymbol{2s^2 2p^4}$. It has 6 valence electrons.
- Their Lewis dot structures are.



Exercise 4.49 c - Answer

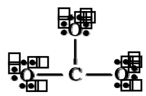
2. The arrangement of the atoms is given to you (each O atom is bonded to the C atom)



3. There are a total of 4 + (3 x 6) = 22 valence electrons coming from the carbon and oxygens, *plus* 2 valence electrons for the -2 charge, for a total of 24 valence electrons.
4. Currently 6 of the valence electrons are being used to make the 3 covalent bonds, leaving 24 - 6 = 18 non-bonding electrons.

Exercise 4.49 c - Answer

5. Distributing theses non-bonding electrons equally about the three oxygen atoms produces the following proposed Lewis dot structure for CO_3^{2-}:

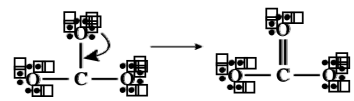


6. This structure contains the requisite 8 electrons about each oxygen atom, but only 6 electrons about the carbon atom.

Exercise 4.49 c - Answer

6. con'd

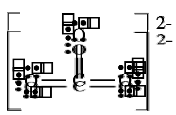
To satisfy the octet rule for the carbon atom, a pair of non-bonding electrons is used to produce a second bond between one of the oxygens and the carbon.



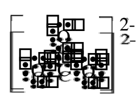
Exercise 4.49 c - Answer

6. con'd

To indicate that this is an ion with a -2 charge, brackets are drawn on about the ion and the charge designated with a 2- subscript:



or



Shapes of Molecules and Polyatomic Ions

- Molecules and polyatomic ions have three dimensional shapes
  - Predicting the shape is important to understanding the properties of molecules and polyatomic ions.
- 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.

Shapes of Molecules and Polyatomic Ions

- 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.
  - Count each bonding pair and each non-bonding pair of electrons as one region of electron density.
  - Count double and triple bonds as a single region of electron density.

Shapes of Molecules and Polyatomic Ions

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

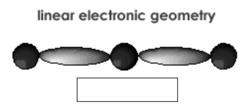
Shapes of Molecules and Polyatomic Ions

- The geometry is determined by the number of regions of electron density.

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

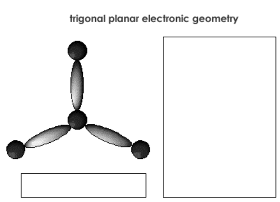
Shapes of Molecules and Polyatomic Ions

VESPR - Linear geometry



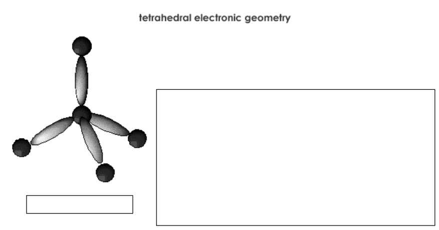
Shapes of Molecules and Polyatomic Ions

VESPR - Trigonal planar geometry



Shapes of Molecules and Polyatomic Ions

VESPR - Tetrahedral geometry



Shapes of Molecules and Polyatomic Ions

- A VSEPR Tutorial available at Georgia Southern University Web site

Click here to go to Lewis Structures at Georgia South University
http://www2.gasou.edu/chemdept/general/molecule/index.htm

Shapes of Molecules and Polyatomic Ions

- The Shape of the molecule is determined by the location of the atoms and the bonds
  - ignoring nonbonding pairs of electrons

Shapes of Molecules and Polyatomic Ions

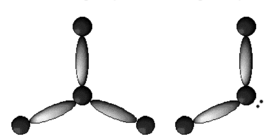
- For a *linear* electron geometry:
  - There are no non-bonding electron pairs, so the shape is also *linear*.

Shapes of Molecules and Polyatomic Ions

- For a *trigonal planar* electron geometry:
  - With no non-bonded electron pairs, the shape is also *trigonal planar*.
  - With one non-bonded electron pair, the shape is a *bent* molecule with a bond angle of 120°

Shapes of Molecules and Polyatomic Ions

- For a *trigonal planar* electron geometry:

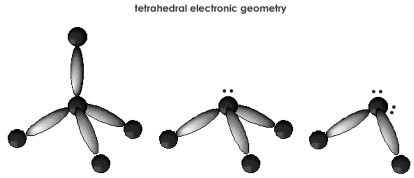


Shapes of Molecules and Polyatomic Ions

- For a tetrahedral geometry:
  - no non-bonded electron pairs, the shape is also *tetrahedral*.
  - With one non-bonded electron pair, the shape is a *triangular pyramidal* molecule with bond angles of 109.5°.
  - With two non-bonding electron pairs, the shape is a *bent* molecule with a bond angle of 109.5°.

Shapes of Molecules and Polyatomic Ions

- For a tetrahedral geometry:



Exercises 4.52, 4.53 & 4.55

Predict the shape of each of the following molecules by first drawing a Lewis structure, then applying the VSEPR theory:

a. O_3 (The O atoms are bonded together, llike beads on a string.)

b. PCl_3 (Each Cl atom is bonded to the P atom.)

c. OF_2 (Each F atom is bonded to the O atom.)

d. ClO_3^- (Each O atom is bonded to the Cl atom.)

e. ClO_4^- (Each O atom is bonded to the Cl atom.)

Polarity of Covalent Molecules

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

Polarity of Covalent Molecules

- Electronegativity
  - Not all atoms have the same affinity for electrons.
  - 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.

Polarity of Covalent Molecules

- **Table 4.3** - Electronegativites for the common representative elements

■ **TABLE 4.3** Electronegativities for the common representative elements

Increasing electronegativity →

Decreasing electronegativity ↓

		H 2.1				
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9					

Polarity of Covalent Molecules

- Bond polarization
  - 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 VSEPR Tutorial available at Georgia Southern University Web site

Click here to go to Lewis Structures at Georgia South University
http://www2.gasou.edu/chemdept/general/molecule/index.htm

Polarity of Covalent Molecules

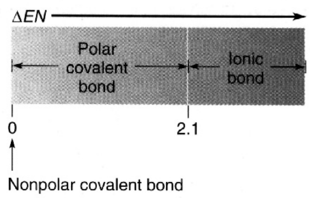
- Bond polarization
  - This results in a partial positive charge at one end of the bond (δ^+) and a partial negative charge at the other end (δ^-).
  - This separation of positive and negative charge is called a dipole.
    - This is designated with an arrow that points from the positive to the negative charge.
  - The net charge of the dipole zero.

Polarity of Covalent Molecules

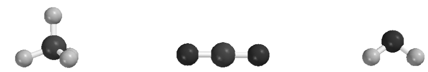
- Bond polarization
  - The electronegativity scale ranges from 0 to 4 and gives the relative affinity of each atom for its electrons
  - The electronegativity scale is used to determine the polarity of covalent bonds
    - $\Delta EN > 2.1$, the bond is considered ionic.
    - $\Delta EN < 2.1$, but greater than 0, the bond is considered polar covalent.

Polarity of Covalent Molecules

- **Figure 4.7** - The ionic/polar bond continuum

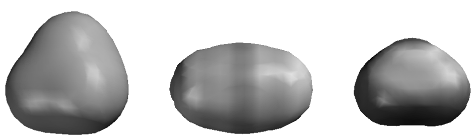


Polarity of Covalent Molecules



Methane Carbon dioxide Water

Polarity of Covalent Molecules



Methane Carbon dioxide Water

Non Polar	Non Polar	Polar
No polar bonds	Polar bonds	Polar bonds
Symmetrical bonds	Symmetrical bonds	Nonsymmetrical bonds

More about Naming Compounds

- Covalent binary compounds that form between nonmetals often combine in varying ratios.
  - For example
    - SO_2, SO_3
    - NO, N_2O, NO_2, N_2O_3

More about Naming Compounds

- For this reason, when naming these compounds the numbers of each atom has to be included in the name.
  - The prefixes, mono-, di-, tri-, tetra-, penta-, hexa-, hepta-, octa-, nona-, and deca- are used to indicate the numbers.
- The name of less electronegative element is given first
- Use -ide suffix for the second element.

More about Naming Compounds

- For example

SO_2	sulfur dioxide
SO_3	sulfur trioxide
NO	nitrogen oxide
N_2O	dinitrogen oxide
NO_2	nitrogen dioxide
N_2O_3	dinitrogen trioxide

More about Naming Compounds

- Polyatomic ions.

Very common		Common	
NH_4^+	ammonium	CrO_4^{2-}	chromate
$C_2H_3O_2^-$	acetate	$Cr_2O_7^{2-}$	dichromate
CO_3^{2-}	carbonate	NO_2^-	nitrite
ClO_3^-	chlorate	MnO_4^-	permanganate
CN^-	cyanide	SO_3^{2-}	sulfite
HCO_3^-	hydrogen carbonate (bicarbonate)	ClO^-	hypochlorite
OH^-	hydroxide	HPO_4^{2-}	hydrogen phosphate
NO_3^-	nitrate	$H_2PO_4^-$	dihydrogen phosphate
PO_4^{3-}	phosphate	HSO_4^-	hydrogen sulfate (bisulfate)
SO_4^{2-}	sulfate	HSO_3^-	hydrogen sulfite (bisulfite)

Other Inter particle forces

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

Other Inter particle forces

- **Table 4.7** - Some characteristics of selected pure substances

■ **TABLE 4.7** Some characteristics of selected pure substances

Substance	Formula or symbol	Classification	Particles occupying lattice sites
Sodium chloride	NaCl	compound	Na^+ and Cl^- ions
Water	H_2O	compound	H_2O molecules
Carbon monoxide	CO	compound	CO molecules
Quartz (pure sand)	SiO_2	compound	Si and O atoms
Copper metal	Cu	element	Cu atoms
Oxygen	O_2	element	O_2 molecules

Other Inter particle forces

- **Table 4.8** - Behavior of selected pure substances on heating

■ **TABLE 4.8** Behavior of selected pure substances on heating

Temperature (°C)	Behavior or state of substance					
	Oxygen (O_2)	Carbon monoxide (CO)	Water (H_2O)	Salt (NaCl)	Copper (Cu)	Quartz (SiO_2)
−220	Solid	Solid	Solid	Solid	Solid	Solid
−218	**Melts**	Solid	Solid	Solid	Solid	Solid
−199	Liquid	**Melts**	Solid	Solid	Solid	Solid
−192	Liquid	**Boils**	Solid	Solid	Solid	Solid
−183	**Boils**	Gas	Solid	Solid	Solid	Solid
0	Gas	Gas	**Melts**	Solid	Solid	Solid
100	Gas	Gas	**Boils**	Solid	Solid	Solid
801	Gas	Gas	Gas	**Melts**	Solid	Solid
1083	Gas	Gas	Gas	Liquid	**Melts**	Solid
1413	Gas	Gas	Gas	**Boils**	Liquid	Solid
1610	Gas	Gas	Gas	Gas	Liquid	**Melts**
2230	Gas	Gas	Gas	Gas	Liquid	**Boils**
2595	Gas	Gas	Gas	Gas	**Boils**	Gas
2600	Gas	Gas	Gas	Gas	Gas	Gas

Other Inter particle forces

- There is an array of forces that hold matter together
- The strongest ones include two forces that we have already introduced:
  - Covalent bonding
  - Ionic bonding

Other Inter particle forces

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

Other Inter particle forces

- Metallic bonding
  - Metals readily lose their valence electrons.
  - In the solid state these valence electrons move freely from one atom to the next.
    - This is why metals have high electrical conductivity.
  - 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.

Other Inter particle forces

- Ionic bonding
  - As discussed earlier, ionic bonding is due the attraction that positive (cations) and negative (anions) have for each other.
    - The cations are usually metals, which have lost electrons to achieve the electron configuration of a noble gas.
    - The anions are usually non-metals, which have gained electrons to achieve the electron configuratin of a noble gas.
  - This type of interaction is called a **charge/charge interaction**.

Other Inter particle forces

- Ionic bonds are quite strong in air but can be greatly weaken by the presence of polar substances such as water.
  - This is why many ionic compounds (salts) dissolve in water.
  - The water weakens the ionic interactions sufficiently that the ions separate from one another.

Other Inter particle forces

- Dipolar forces
  - This is the attraction that arises between polar molecules.
    - Even though polar molecules do not have a net charge like ions do, each molecule contains regions of positive and negative charge.
  - 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.
  - This type of interaction is referred to as a **dipole/dipole interaction**

Other Inter particle forces

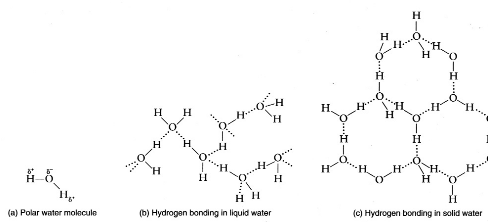
- Dipolar forces
  - Dipole forces are moderately strong in air, but like ionic bonds they can be greatly weakend by the presence of polar substances such as water.

Other Inter particle forces

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

Other Inter particle forces

- Figure 4.10-Hydrogen bonding in water



Other Inter particle forces

- Hydrogen bonding
  - Hydrogen bonding is what holds water together.
    It is what makes water more cohesive than other substances of comparable composition and molecular weight.

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

Other Inter particle forces

- Comparing the boiling points of the compounds formed by bonding two hydrogens to a group VI element:
  - Water does not fit the trend established by the other elements in family VI
  - The unusually high boiling point for water is due to its ability to form extensive networks of hydrogen bonds.

Other Inter particle forces

- Figure 4.12-Interparticle Forces