Compounds

Read from on-line textbook: Chap 4.1-4.3, 4.5 and Chap 5.1, 5.2 (first part, but don’t worry about naming), 5.3-5.5, 5.7
Also video tutorials: Compounds, part 1 (don’t worry about naming) and part 2. Shapes and Interactions, part 1 & 2.
http://chemwiki.ucdavis.edu/Wikitexts/Sacramento_City_College/SCC%3A_Chem_309/Chapters

Problems: Lab Manual Appendix III: Part GG and H2

I. Classify different forms of matter.

A. Classification based on purity 
(of sorts)

B. Element: Cannot be separated into simpler substances by chemical means.
   1. Composed of one type of atom
   2. Examples: phosphorous (P) and potassium (K)
      -and all those other substances on the periodic table
   3. Review: What determines the identity of an element?

   Note: With respect to elemental identity, an atom with 6 \( p^+ \) and 6 \( n \) in its nucleus is considered to be the same type of atom as one with 6 \( p^+ \) and 8 \( n \) in its nucleus. These two atoms are said to both be ____________________________ of carbon.

4. Being “in its elemental state.”
   Elemental state means the state in which that element exists in nature.
   The elemental state of:
      a) He, helium, is as a gas,
      b) Fe, iron, is as a solid
      c) Hg, mercury, is as a liquid (Not many elements exist as liquids in nature.)

   Note: Seven elements are “diatomic” (two atoms of an elements bonded covalently) in their elemental state.
   “I never order broiled clams from Hardees.”

C. Compounds: Pure substances formed when at least __________ ________________ elements combine in specific ratios.
   The combining is when the atoms form some kind of_________________________.

   Is Mg a compound?
   Is MgCl\(_2\) a compound?
   Is O\(_2\) a compound?
D. Mixtures: Two or more types of substances mixed together in varying ratios. They can be separated by “physical means”.

1. Homogeneous: Uniform composition. This is also called a ______________. This later term will be used frequently in this class.
   Examples:

   Examples:

II. Formulas and Models of Compounds
   A. Chemical formula, like C₃H₈O or H₂O
      1. Tells what elements are present in a substance.
      2. The subscripts indicate the number of each type of atom.
   B. Structural formula
      1. Tells how the atoms are connected
      2. These connections are called bonds

   C. Condensed structural formula
      CH₃-CH₂-CH₂OH

   D. Line structures
      1. Carbons are located at the ends and junctions of lines.
      2. Hydrogens bonded to carbons are not indicated.

Problem:

<table>
<thead>
<tr>
<th>Structural formula</th>
<th>Structural formula</th>
<th>Structural formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>2-methylbutane</td>
<td>ethylene oxide</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Condensed formula</th>
<th>Condensed formula</th>
<th>Condensed formula</th>
</tr>
</thead>
</table>

<table>
<thead>
<tr>
<th>Line structure</th>
<th>Line structure</th>
<th>Line structure</th>
</tr>
</thead>
</table>

<table>
<thead>
<tr>
<th>Chemical or molecular formula</th>
<th>Chemical or molecular formula</th>
<th>Chemical or molecular formula</th>
</tr>
</thead>
</table>
III. Formation of two types of bonds

covalent bonds & ionic bonds

A. Energy & Stability

Things at higher energy are less stable!!

1. Living things are high energy & unstable.

2. They must obtain energy from other unstable & high energy things….ie. food!

(The compounds in the food you eat must be relatively unstable for you to get useful energy from the food.)

3. Generally speaking, calorically rich food is made up of reduced (hydrogen containing) carbon compounds.

4. The higher the hydrogen content (and lower oxygen content), the more energy the food contains.

(Example: Fat is more caloric than carbohydrate or protein.)

5. One way to consider stability:

*Stable* - if something exists for a relatively long time.

*Unstable* – if something exists for a brief time and changes.

SO

What makes atoms etc. be stable?

It depends on the $e^-$. 


For representative elements, a full valence shell is considered stable. Why?

*Noble or (inert gases) are inert (non-reactive, stable), and they have full valence shells.*

“Unstable atoms gain, lose, or share electrons (forming ions or compounds) to end up with eight valence electrons (and become stable).”

Remember: The valence electrons ($e^-$) are those in the $s$ & $p$ orbitals of the outermost occupied shell.

Which elements have full valence shells?

How many electrons do they each have?

B. Gain, Lose or Share Electrons

1. A *covalent bond is a shared e− pair.* (tug of war)

2. The atoms in molecules are held together by covalent bonds.
3. **An ionic bond is the attraction between oppositely charged ions.**

   a. When atoms **gain** or **lose** electrons, they become ions with a charge (like +1, +2, +3, or -1, -2, -3, etc).

   b. **Metals** tend to _____ to become _____ ions.

   c. **Non-metals** tend to ___ to become ___ ions.

C. As stated above, **covalent bonding** is between two non-metals. **Ionic bonds** are formed between a metal and a non-metal.

Based on these two statements, are the compounds below are ionic or covalent.

   NaCl  N₂  H₂O  K₃N  

IV. Covalent Bonding

A. **This type of compound results from combining ___________________________**

B. **Hydrogen gas (H₂(g)) reacting with oxygen (O₂(g)) to form water (H₂O(g)):**

   \[
   2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g)
   \]

   In the rxn. shown above, **two non-metals combine to form a covalent compound.** The product is a molecule.

V. Lewis Structures of Atoms (**e⁻ accounting.**)  

A. Count valence e⁻ (the e⁻ in the outermost shell).

B. Arrange the e⁻ (dots) on the 4 sides of the elemental symbol.

C. Don’t worry transition metals in CHM 109.

D. **Examples:**  

   Xe  C  F  S  He  

   Which are stable?
VI. Seeing patterns in molecules:

A. Atoms share electrons to fill their valence $e^-$ shells.

B. For valence $e^-$ purposes, we count the shared $e^-$ pair as belonging to both atoms.

What would be the stable form of hydrogen gas? Is the stable form of “H”? Why/why not?

Does H have a full valence shell?

Here we show $e^-$ sharing in H$_2$ with pictures.

Lewis Dot  
\[
\begin{align*}
\text{H} & : \text{H} \\
\end{align*}
\]

“Flat” (no 3-D info)

Ball & Stick  
\[
\begin{align*}
\text{H} \quad \text{H} \\
\end{align*}
\]

Space filling (most real?)

VII. Electronegativity and Bond Polarity

(view nuclei as fixed reference points)

A. Covalent bonds involve $e^-$ sharing, but how evenly is the negative charge arranged around the bonded atoms?

How you decide? Electronegativity chart!

*Electronegativity: the relative of ability an element to attract electrons in a bond.*

(“Sharing a blanket” analogy)

Electronegativity Table (Pauling scale) (p. 106 in text)
B. Non-polar covalent bonds have **completely even sharing**. Examples:

\[
\begin{array}{cc}
\text{H} & \text{O} \\
2.1 & 3.5 \\
\end{array}
\]

\[
\begin{array}{cc}
\text{H}_2 & \text{O}_2 \\
\end{array}
\]

C. Polar covalent bonds: when bonding atoms have different electronegativities.

\[
\begin{array}{cc}
\text{H} & \text{I} \\
2.1 & 2.5 \\
\end{array}
\]

\[
\begin{array}{cc}
\text{H} & \text{F} \\
2.1 & 4.0 \\
\end{array}
\]

Continuum: Non-polar covalent bonds  
Polar covalent bonds  
Formation of ions and ionic bonding

**Representing molecules & compound ions w/ Lewis Structures**

Lewis Structures: visual bookkeeping device for valence e⁻.

Steps to Writing Lewis Structures:
1. Add up number of valence e⁻ (group #) for each atom
2. Make adjustments for non-zero net charge
   a) Add one valence e⁻ for each negative charge
   b) Subtract one valence e⁻ for each positive charge
3. Write elemental symbols (which atom links to which?)
4. Form single bonds to connect bonded atoms
5. Fill in non-bonding e⁻ pairs
6. Only if not enough e⁻ to go around, form double or triple bonds

Write Lewis structures for:
- Cl atom
- Cl⁻ ion
- CH₄
- H₂O
- NO₃⁻
- HCN
III. Using Lewis dot structures to determine Electron Pair Geometry, Molecular Shape & Polarity

A. VSEPR stands for Valence Shell Electron Pair Repulsion. We will use this model to determine 3-D shapes of molecules.

1. VSEPR is based on idea that electron pairs are arranged to be as far apart as possible.

2. Three electron pair geometries (discussed in this class)
   
   a) **linear** (2 electron pairs) 
   b) **trigonal planar** (3 electron pairs) 
   c) **tetrahedral** (4 electron pairs)

3. Molecular Shapes (determined by location of atom centers)
   
   a) **linear** electron pair geometry → one possible molecular shape
      
      linear
   
   b) **trigonal planar** electron pair geometry → two possible molecular shapes
      
      trigonal planar  bent
   
   c) **tetrahedral** electron pair geometry → three possible molecular shapes
      
      tetrahedral  trigonal pyramidal  bent

4. Try some examples
   
   a) CH₄
   
   b) CH₃Cl
   
   c) H₂O
   
   d) CO₂
5. Hybrid Atomic Orbitals
The above shapes for molecules are not consistent with the shape and orientation of atomic orbitals. Various models are used to explain this including “orbital hybridization”

Results of this model:
- Tetrahedral: uses four $sp^3$ hybrid orbitals
- Trigonal planar: uses three $sp^2$ hybrid orbitals
- Linear: uses two $sp$ hybrid orbitals

B. Polarity of molecules
1. Polarity: uneven distribution of charge.

2. Since $p^+$ must stay in nucleus, uneven charge distribution refers to $e^-$ distribution around specific nuclei or the molecule as a whole.

3. You must look at $e^-$ pair geometry (molecular symmetry) to determine molecular polarity.

4. Symmetry, vectors or a tractor pull(?)

C. The Double Bond
One bond is a normal bond formed from the direct overlap of orbitals. Called a sigma ($\sigma$) bond. The other is formed from sideways overlap of orbitals. Called a pi ($\pi$) bond.

A sigma bond is stronger than a pi bond.

When a double bond is formed, rotation about that bond is eliminated.

Contrast this w/ free rotation about C–C single bond in ethane. The limited rotation in double bonds is responsible for another group of cis/trans isomers. The cis and trans isomers of 2-butene are shown below.

(The “-ene” ending (The “-ene” ending indicates that there is a double bond, and the “2” indicates the position of the double bond.)

A compound is cis if its higher ranking groups are on the same side of a line running through the double bonded C atoms and trans if they are on opposite sides.
Biological relevance of cis-/trans isomers?

IX. Ionic Bonds
Elemental sodium (Na\(_{(s)}\)) reacting with chlorine (Cl\(_{2(g)}\)) to form table salt (NaCl\(_{(s)}\)), an __________

\[
2\text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s)
\]

Not safe to do, but see: [http://www.youtube.com/watch?v=VBReOjo3ri8](http://www.youtube.com/watch?v=VBReOjo3ri8) & [http://www.youtube.com/watch?v=Ftw7a5ccubs&feature=related](http://www.youtube.com/watch?v=Ftw7a5ccubs&feature=related)

In the rxns shown above, a **metal combined with a non-metal to form an ionic compound** as the product a.k.a. a salt. The ions in the salt are held together by **electrostatic attraction** of oppositely charged ions (+ and -).

X. Ionic bonds, ionic compounds, ionic compound formulas

A. In an ion: \# of \(p^+\) \(\neq\) \# of \(e^-\).

B. Ions can be simple or polyatomic.

1. A simple ion consists of one charged atom (for example Cl\(^-\)).

2. A polyatomic ion has multiple atoms held together by covalent bonds and carries a charge. (for example NO\(_3^\))

3. positively charged ions are called __________________________

   negatively charged ions are called __________________________

Remember: **Ionic bond** refers to the **electrostatic attraction** between oppositely charged ions.
C. Ionic compound formulas must have zero net charge.  
\( \text{(No shock when you touch an ionic compound.)} \)
1. Determine charge of simple ions in ionic components from Periodic Table.  How?

2. A few “polyatomic” ions like sulfate ion, \( \text{SO}_4^{2-} \), are listed in the lab manual appendix.

3. An ionic compound formula consists of the minimum number of each ionic component that results in equal amounts of (+) and (-) charge.

**Problem.** Write ionic compound formulas for:

1. potassium chloride
2. calcium iodide
3. sodium phosphide
4. sodium phosphate
5. magnesium oxide
6. iron (II) nitrate

XI. Ionic Bond  Formation

A. Example of ionic compound formation: \( 2\text{Na}_\text{(s)} + \text{Cl}_2\text{(g)} \rightarrow 2\text{NaCl}_\text{(s)} \)

*Important:* 2 distinct things occur in these rxns. You must be able to recognize them as distinct.

B.1. First, \( e^- \) are being transferred to convert atoms (or molecules) into ions:

\[
\text{Na} \rightarrow \text{Na}^+ + e^- \quad \text{(ox)}
\]
\[
\text{Cl}_2 + 2 e^- \rightarrow 2 \text{Cl}^- \quad \text{(red)}
\]

*metals* tend to lose \( e^- \) (give \( e^- \) away), *non-metals* tend to gain \( e^- \).

2. Second, positively charged ions bind to negatively charged ions to form ionic compounds:

\[
\text{Na}^+ + \text{Cl}^- \rightarrow \text{NaCl}
\]
C. Another example:
Elemental magnesium (Mg\textsubscript{(s)}) reacting with oxygen (O\textsubscript{2}(g)) to form magnesium oxide (MgO\textsubscript{(s)}):

\[ 2 \text{Mg(s)} + O_2(g) \rightarrow 2 \text{MgO(s)} \]

Two steps:
1. \( \text{Mg(s)} \rightarrow \text{Mg}^{2+} + 2e^- \), \( \text{O}_2 + 4e^- \rightarrow 2 \text{O}^{2-} \)
2. \( \text{Mg}^{2+} + \text{O}^{2-} \rightarrow \text{MgO} \)

D. Remember: Ionic compounds are not molecules!

Although the ionic compound formula, NaCl, shows a 1 to 1 ratio of the ions, it is incorrect to think of a specific, directionally-oriented interaction between one Na\textsuperscript{+} and one Cl\textsuperscript{-}.

You can see from the figure that each Na\textsuperscript{+} is actually interacting equally with 6 Cl\textsuperscript{-} ions (and vice versa).

E. An interesting thing happens to ionic compounds that dissolve in water.
The lowest energy state occurs when the ions separate and each is surrounded by water.

The H\textsubscript{2}O molecules reduce the strength of attractive or repulsion between ions significantly.*