Atom: The smallest divisible unit of an element
Compound: A substance made of two or more atoms
Ion: A charged atom or molecule
Cation: Positive ion
Anion: Negative ion

NOMENCLATURE
Format for naming chemical compounds using prefixes, suffixes, and other modifications of the names of elements which constitute compounds.

Sometimes things get confusing…

Elements:
H = hydrogen
O = oxygen
C = carbon

Compounds:
H₂ = hydrogen
O₂ = oxygen
H₂O = water
### Ion Charges:

<table>
<thead>
<tr>
<th>Charge</th>
<th>Ion</th>
</tr>
</thead>
<tbody>
<tr>
<td>+1</td>
<td>H+</td>
</tr>
<tr>
<td>+2</td>
<td>Li+</td>
</tr>
<tr>
<td>+3</td>
<td>Na+</td>
</tr>
<tr>
<td>+4</td>
<td>Mg2+</td>
</tr>
<tr>
<td>+5</td>
<td>Al3+</td>
</tr>
<tr>
<td>+6</td>
<td>S6+</td>
</tr>
<tr>
<td>-1</td>
<td>F-</td>
</tr>
<tr>
<td>-2</td>
<td>O2-</td>
</tr>
<tr>
<td>-3</td>
<td>S2-</td>
</tr>
<tr>
<td>-4</td>
<td>Se4-</td>
</tr>
<tr>
<td>-5</td>
<td>Se5-</td>
</tr>
<tr>
<td>-6</td>
<td>Br6-</td>
</tr>
<tr>
<td>-7</td>
<td>I7-</td>
</tr>
<tr>
<td>0</td>
<td>He</td>
</tr>
</tbody>
</table>

### Compounds fall into one of two classes:

**Inorganic Salts**
- **Metal cation**
- **Non-metal or polyatomic anion**

**Molecules**
- **Non-metal**
- **Non-metal**

The two use different formalisms for naming…

### Binary Compounds: Metal & non-Metal

Metal of fixed oxidation (charge) state combined with a non-metal.

**Examples:**
- **K**
- **Ca**
- **Na**
- **Al**

<table>
<thead>
<tr>
<th>Cation</th>
<th>Anion</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>K+</td>
<td>Cl−</td>
<td>KCl</td>
<td>Potassium chloride</td>
</tr>
<tr>
<td>Ca2+</td>
<td>O2−</td>
<td>CaO</td>
<td>Calcium Oxide</td>
</tr>
<tr>
<td>Na+</td>
<td>S2−</td>
<td>Na2S</td>
<td>Sodium sulfide</td>
</tr>
<tr>
<td>Al3+</td>
<td>S2−</td>
<td>Al2S3</td>
<td>Aluminum sulfide</td>
</tr>
</tbody>
</table>

### Metals of variable charge (transition) with a non-metal

**Examples:**
- modify transition metal name with roman numeral

<table>
<thead>
<tr>
<th>Cation</th>
<th>Anion</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Pb2+</td>
<td>Cl−</td>
<td>PbCl2</td>
<td>lead (II) chloride</td>
</tr>
<tr>
<td>Pb4+</td>
<td>Cl−</td>
<td>PbCl4</td>
<td>lead (IV) chloride</td>
</tr>
<tr>
<td>Fe3+</td>
<td>O2−</td>
<td>Fe2O3</td>
<td>Iron (III) oxide</td>
</tr>
</tbody>
</table>

non-metal takes on “ide” suffix

- pronounced: lead - two - chloride
Some common polyatomic ions:

<table>
<thead>
<tr>
<th>Ion</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH₄⁺</td>
<td>ammonium</td>
</tr>
<tr>
<td>H₃O⁺</td>
<td>hydronium</td>
</tr>
<tr>
<td>CO₃²⁻</td>
<td>carbonate</td>
</tr>
<tr>
<td>HCO₃⁻</td>
<td>hydrogen or bicarbonate</td>
</tr>
<tr>
<td>NO₂⁻</td>
<td>nitrite</td>
</tr>
<tr>
<td>NO₃⁻</td>
<td>nitrate</td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>sulfate</td>
</tr>
<tr>
<td>SO₃²⁻</td>
<td>sulfite</td>
</tr>
<tr>
<td>PO₄³⁻</td>
<td>phosphate</td>
</tr>
<tr>
<td>C₂H₃O₂⁻</td>
<td>acetate</td>
</tr>
</tbody>
</table>

Termary Compounds: Those with three different elements

**Type A:** metal of fixed charge with a complex ion

<table>
<thead>
<tr>
<th>Cation</th>
<th>Anion</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>K⁺</td>
<td>OH⁻</td>
<td>KOH</td>
<td>Potassium hydroxide</td>
</tr>
<tr>
<td>Ca²⁺</td>
<td>OH⁻</td>
<td>Ca(OH)₂</td>
<td>Calcium hydroxide</td>
</tr>
<tr>
<td>Na⁺</td>
<td>SO₄²⁻</td>
<td>Na₂SO₄</td>
<td>Sodium sulfate</td>
</tr>
<tr>
<td>Al³⁺</td>
<td>SO₄²⁻</td>
<td>Al₂(SO₄)₃</td>
<td>Aluminum sulfate</td>
</tr>
</tbody>
</table>

Metal of variable charge transition) with a complex ion

<table>
<thead>
<tr>
<th>Cation</th>
<th>Anion</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe³⁺</td>
<td>NO₃⁻</td>
<td>Fe(NO₃)₃</td>
<td>Iron (III) nitrate</td>
</tr>
<tr>
<td>Fe²⁺</td>
<td>NO₂⁻</td>
<td>Fe(NO₂)₂</td>
<td>Iron (II) nitrite</td>
</tr>
</tbody>
</table>

Non-metal with a non-metal

When non-metals combine, they form molecules. They may do so in multiple forms:

CO → CO₂

Because of this we need to specify the number of each atom by way of a prefix.

1 = mono    2 = di    3 = tri    4 = tetra
5 = penta   6 = hexa  7 = hepta
Examples:

<table>
<thead>
<tr>
<th>Formula</th>
<th>Name:</th>
</tr>
</thead>
<tbody>
<tr>
<td>BCl₃</td>
<td>boron trichloride</td>
</tr>
<tr>
<td>SO₃</td>
<td>sulfur trioxide</td>
</tr>
<tr>
<td>NO</td>
<td>nitrogen monoxide</td>
</tr>
<tr>
<td>we don't write:</td>
<td>nitrogen monooxide or mononitrogen monoxide</td>
</tr>
<tr>
<td>N₂O₄</td>
<td>dinitrogen tetraoxide</td>
</tr>
</tbody>
</table>

D) Writing formulas for acids and Bases

• An **acid** is a substance that produces H⁺ when dissolved in water.
• Certain gaseous molecules become acids when dissolved in water.

• A **base** produces OH⁻ when dissolved in water.
• Bases often are Group I and Group II hydroxide salts.

**Type I Acids:** Acids derived from –ide anions.

The names for these acids follow the formula:

“hydro” + the root of the ide anion + **ic** “acid”

<table>
<thead>
<tr>
<th>Anion</th>
<th>Acid</th>
<th>Name:</th>
</tr>
</thead>
<tbody>
<tr>
<td>chloride</td>
<td>HCl</td>
<td>hydrochloric acid</td>
</tr>
<tr>
<td>fluoride</td>
<td>HF</td>
<td>hydrofluoric acid</td>
</tr>
</tbody>
</table>

H⁺ and S²⁻

\[
\text{H}_2\text{S} \quad \text{it takes 2} \ H^+ \ \text{to cancel one} \ S^2- \\
\]

 hydro sulfuric acid
**Examples:**

<table>
<thead>
<tr>
<th>Anion:</th>
<th>Acid:</th>
<th>Name:</th>
</tr>
</thead>
<tbody>
<tr>
<td>(nitrate)</td>
<td>NO₃⁻</td>
<td>nitric acid</td>
</tr>
<tr>
<td>(sulfate)</td>
<td>SO₄²⁻</td>
<td>sulfuric acid</td>
</tr>
<tr>
<td>(acetate)</td>
<td>C₂H₃O₂⁻</td>
<td>acetic acid</td>
</tr>
</tbody>
</table>

**Practice:**

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>N₂SO₄</td>
<td>sodium sulfate</td>
<td></td>
</tr>
<tr>
<td>barium carbonate</td>
<td>BaCO₃</td>
<td></td>
</tr>
<tr>
<td>FeO</td>
<td>Iron (II) oxide</td>
<td></td>
</tr>
<tr>
<td>zinc phosphide</td>
<td>Zn₃P₂</td>
<td></td>
</tr>
<tr>
<td>NiBr₂</td>
<td>nickel (II) bromide</td>
<td></td>
</tr>
</tbody>
</table>

**Common Names:**

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>CO₂</td>
<td>carbon dioxide</td>
<td></td>
</tr>
<tr>
<td>carbon monoxide</td>
<td>CO</td>
<td></td>
</tr>
<tr>
<td>P₂O₅</td>
<td>diphosphorous pentaoxide</td>
<td></td>
</tr>
<tr>
<td>nitrogen trihydride</td>
<td>NH₃</td>
<td>(ammonia)</td>
</tr>
<tr>
<td>H₂O</td>
<td>water</td>
<td></td>
</tr>
<tr>
<td>ammonia</td>
<td>NH₃</td>
<td></td>
</tr>
<tr>
<td>CH₄</td>
<td>methane</td>
<td></td>
</tr>
<tr>
<td>NO</td>
<td>nitric oxide</td>
<td></td>
</tr>
<tr>
<td>N₂O</td>
<td>nitrous oxide</td>
<td></td>
</tr>
</tbody>
</table>
please go to my chem. 1A web site and download a nomenclature practice worksheet if you need more review.
http://www.csus.edu/indiv/m/mackj/chem1a/chem1A_lab.html
The link to the worksheet is at the top of the page.

Modern Atomic Theory:

• Atoms are made of protons, neutrons and electrons.
• The nucleus of the atom carries most of the mass.
• It consists of the protons and neutrons surrounded by a cloud of electrons.
  The charge on the electron is −1
  The charge on the proton is +1
  There is no charge on the neutron

The Atomic Number or number of protons in the nucleus defines an element.

Isotopes, Atomic Numbers, and Mass Numbers

• Atomic number (Z) = number of protons in the nucleus.
• Mass number (A) = total number of nucleons in the nucleus (i.e., protons and neutrons).
• One nucleon has a mass of 1 amu
  (Atomic Mass Unit) a.k.a “Dalton”
• Isotopes have the same Z but different A.
• The elements are arranged by Z on the periodic table.

By convention, for element X, we write \[ A^Z_X \]

The isotope \[ _{34}^{75}\text{Se} \] is used medically for diagnosis of pancreatic disorders. How many protons, neutrons, and electrons does an atom of \[ _{34}^{75}\text{Se} \] have?

75 protons + 34 neutrons

protons = 34 electrons = 34
neutrons = 75-34 = 41
**Avagadro’s Number**

Since one mole of $^{12}$C has a mass of 12g (exactly), 12g of $^{12}$C contains $6.022142 \times 10^{23}$ $^{12}$C-atoms.

But carbon exists as 3 isotopes: C-12, C-13 & C-14

The average atomic mass of carbon is 12.011 u.

From this we conclude that 12.011g of carbon contains $6.022142 \times 10^{23}$ C-atoms

*Is this a valid assumption?* Yes, since $N_A$ is so large, the statistics hold.

---

**Molar Masses (Molecular Weights) of Compounds:**

The molar mass of a *molecular compound* is the sum of the molar masses of its atoms.

**Example:**

The molar mass of CO$_2$ is:

\[
1 \times (12.01 \text{ g/mol}) + 2 \times (16.00 \text{ g/mol}) = 44.01 \text{ g/mol}
\]

---

How many oxygen atoms are there in 25.1g of chromium (III) acetate?

**step 1:** write the correct chemical formula...

\[
\text{Cr}^{3+} \text{ & } \text{C}_2\text{H}_3\text{O}_2^- \rightarrow \text{Cr(C}_2\text{H}_3\text{O}_2)_3
\]

**step 2:** calculate the molar mass...

229.13 g/mol

**step 3:** use dimensional analysis to solve the problem...

\[
25.1 \text{ g Cr(C}_2\text{H}_3\text{O}_2)_3 \times \frac{1 \text{ mol Cr(C}_2\text{H}_3\text{O}_2)_3}{229.13 \text{ g Cr(C}_2\text{H}_3\text{O}_2)_3} \times \frac{6 \text{ mol O}}{1 \text{ mol Cr(C}_2\text{H}_3\text{O}_2)_3} \\
\times \frac{6.022 \times 10^{23} \text{ O-atoms}}{1 \text{ mol O}} = 3.96 \times 10^{23} \text{ O-atoms}
\]

---

**Percent Composition:**

The relative amounts of each atom in a molecule or compound can be represented fraction of the whole.

**Question:** What is the weight % of each element in C$_2$H$_6$?

**First** determine the molar mass of C$_2$H$_6$:

1 mol of C$_2$H$_6$ has a mass of 30.07 g

\[
(2 \times 12.01 + 6 \times 1.008) \text{ g/mol}
\]

**Next** determine the mass of hydrogen in 1 mol of the compound:

\[
1 \text{ mol C}_2\text{H}_6 \times \frac{6 \text{ mol H}}{1 \text{ mol C}_2\text{H}_6} \times \frac{1.0079 \text{ g H}}{1 \text{ mol H}} = 6.047 \text{ g H}
\]
Now relate the mass of H in one mol of the compound to the molar mass of the compound

\[
6.047 \text{ g H} \times \frac{1}{30.07 \text{ g C}_2\text{H}_6} \times 100 = 20.11\% \text{ H}
\]

Since there is only C as the remaining element:

\[
\% \text{ C} = 100 - \% \text{ H} = 79.89 \% \text{ C}
\]

The compound C\(_2\)H\(_6\) is **20.11\% H & 79.89\% C**

---

**Determining a Formula from Percent Composition:**

Given the relative percentages of each element in a compound,

10 \% X, 20 \% Y, 30 \% Z etc…

one can find the **empirical formula** of the compound.

The **empirical formula** of a compound or molecule represents the simplest ratio of each element in 1 mol of the compound or molecule.

---

**Example:** A compound is found to be 64.82 \% carbon, 21.59 \% oxygen and 13.59 \% hydrogen. What is the empirical formula for this compound?

**Solution:** determine X, Y & Z in (C\(_x\)H\(_y\)O\(_z\))

1. Since the percentages for each element sum to 100\%, if one equates \% to grams (g), the sum of the masses must be 100g.
   (i.e. one can assume 100g of the compound)

   \[
   \begin{align*}
   64.82 \text{ g C} & \quad 21.59 \text{ g O} & \quad 13.59 \text{ g H}
   \end{align*}
   \]

2. Convert the grams of each element to moles.
   (g element \(\rightarrow\) mole X etc…)

   \[
   \begin{align*}
   64.82 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} &= 5.397 \text{ mol C} \\
   21.59 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} &= 1.349 \text{ mol O} \\
   13.59 \text{ g H} \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} &= 13.48 \text{ mol H}
   \end{align*}
   \]
3. Divide each of the individual moles by the smallest number of moles to gain the molar ratios for each element in the compound. These are the formula subscripts. \((X_2Y_3\text{ etc…})\)

\[
\begin{align*}
\text{Subscript for C} & \quad \text{Subscript for H} & \quad \text{Subscript for O} \\
\frac{5.397}{1.349} = 4.001 & \quad \frac{13.48}{1.349} = 9.992 & \quad \frac{1.349}{1.349} = 1.000
\end{align*}
\]

**If the ratios are fractional (0.5, 1.5 or 0.333) multiply each ratio by a whole number to get even number formula subscripts.**

Examples: \(0.5 \times 2 = 1\) \(0.25 \times 4 = 1\) \(0.333 \times 3 = 1\)

Rounding to the nearest whole numbers:

\[
\begin{align*}
X &= 4.001 = 4 \\
Y &= 9.992 = 10 \\
Z &= 1.000 = 1
\end{align*}
\]

The empirical formula is: \(C_4H_{10}O\)

The results of this calculation tells us only about the empirical formula of the compound.

To determine the molecular formula, we need more information. This will be shown in a later example.

---

**Chemical Equations:**

*Mass is conserved in a chemical reaction.*

\[
\text{Total mass of reactants} = \text{Total mass of products}
\]

Chemical equations must therefore be balance for mass.

The number and type of atoms on either side of the equation must be equal!
Reduction and Oxidation Reactions: RedOx

**Oxidation** involves an atom or compound losing electrons

**Reduction** involves an atom or substance gaining electrons

Neither process can occur alone… that is, there must be an exchange of electrons in the process.

The substance that is **oxidized** is the reducing agent

The substance that is **reduced** is the oxidizing agent

Chemists use oxidation numbers to account for the transfer of electrons in a RedOx reaction.

---

Electrochemistry: Oxidation numbers

In the compound potassium bromate (KBrO₃), the oxidation number of bromine (Br) is?

The compound is neutral so the sum of the oxidation numbers should be zero.

\[ KBrO_3 \]

\[ +1 \quad 5 \quad 3 \times (-2) = -6 \]

\[ K^+ BrO_3^- \]

\[ 1 + ?? + (-6) = 0 \quad ?? = 5 \]

---

Balancing REDOX reactions:

\[ Fe + O_2 \rightarrow Fe_2O_3 \]

oxidation states: \[ 0 \quad 0 \quad +3 \quad -2 \]

oxidation half reaction: \[ \{ Fe \rightarrow Fe^{3+} \quad +3e^- \} \times 4 \]

reduction half reaction: \[ \{ O_2 + 4e^- \rightarrow 2O^{2-} \} \times 3 \]

Balance electrons transferred then sum the half RXN’s:

\[ 4Fe + 3O_2 + 12e^- \rightarrow 2Fe_2O_3 + 12e^- \]

\[ 4Fe + 3O_2 \rightarrow 2Fe_2O_3 \]
Consider the following reaction:

\[
\text{HCl (aq)} + \text{Ba(OH)}_2 (\text{aq}) \rightarrow \text{H}_2\text{O (l)} + \text{BaCl}_2(\text{aq})
\]

Balancing:

\[
2 \text{HCl (aq)} + \text{Ba(OH)}_2 (\text{aq}) \rightarrow 2\text{H}_2\text{O (l)} + \text{BaCl}_2(\text{aq})
\]

How many moles of HCl are consumed if 1.50 g of BaCl2 are produced assuming that Ba(OH)2 is in excess?

**Solution:**

\[
\frac{g \text{ BaCl}_2}{1.50 \text{ g} \text{ BaCl}_2} \rightarrow \frac{\text{mol BaCl}_2}{\frac{1 \text{ mol BaCl}_2}{208.24 \text{ g BaCl}_2}} \rightarrow \frac{\text{mol HCl}}{\frac{2 \text{ mol HCl}}{1 \text{ mol BaCl}_2}}
\]

Conversions between masses & moles in chemical reactions.

**Stoichiometry:** Conversions between masses & moles in chemical reactions.

**Limiting Reactant:**

⇒ When one reactant is present in an amount such that it is completely consumed before all other reactants, we say that it limits the reaction.

⇒ The other reactants are said to be in excess.

⇒ The *Theoretical Yield* is determined by the stoichiometry of the limiting reactant.

⇒ The limiting reactant can only be determined through molar ratios. It cannot be identified by mass.