CHM 101 GENERAL CHEMISTRY

FALL QUARTER 2008

Section 2

Lecture Notes – 10/13/2008

(last revised: 10/13/08)

2.8 Naming Simple Chemical Compounds (concluded)

- **Naming Binary Covalent Compounds** (The text calls these also Type III binary compounds)

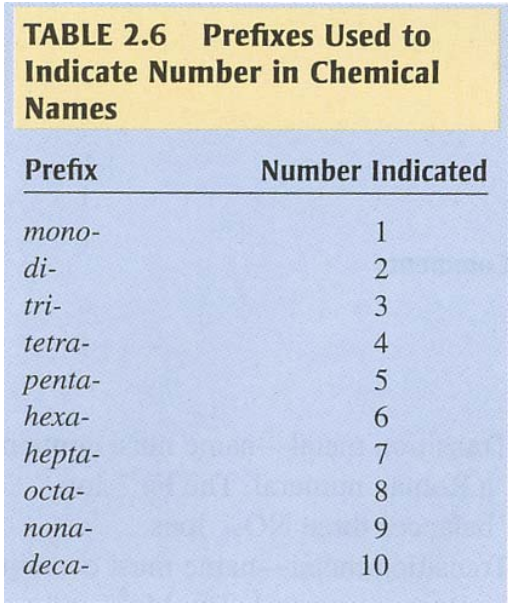


TABLE 2.6 Prefixes Used to Indicate Number in Chemical Names

Prefix	Number Indicated
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

  - Binary covalent compounds are formed between two non-metals. The rules for naming them are very similar to the rules for binary ionic compounds, even though their bonds are covalent rather than ionic.
    - The first element in the formula is named first, using the full name of the element.
    - The second element is named as if it were an anion.
    - Prefixes mono-, di-, tri-, etc., are used to indicate the numbers of each type of atom. See Table 2.6 in your text for a full list.
    - The prefix, mono-, is never applied to the first element in the name. For example, CO is carbon monoxide, and CO_2 is carbon dioxide.
    - The letters, o, and a, on the ends of prefixes can be omitted to avoid awkward pronunciations if the name of the following element begins with a vowel. Examples: carbon monoxide and phosphorous pentoxide.

- Water and ammonia are always referred to by their common names, never by their systematic names based on the formulas, H_2O and NH_3.
- Some examples for class discussion:

Formula	**Name**
H_2O	Water
NH_3	Ammonia
N_2O	Dinitrogen monoxide
NO	Nitrogen monoxide
NO_2	Nitrogen dioxide
N_2O_3	Dinitrogen trioxide
N_2O_4	Dinitrogen tetroxide
N_2O_5	Dinitrogen pentoxide
PCl_5	
PCl_3	
SO_2	
	Sulfur hexafluoride
	Sulfur trioxide
	Carbon dioxide

- **Summary: How to Name Binary Compounds**
  - Here is the flowchart from your text (Figure 2.23). (In the text, binary covalent compounds are also called Type III binary compounds):

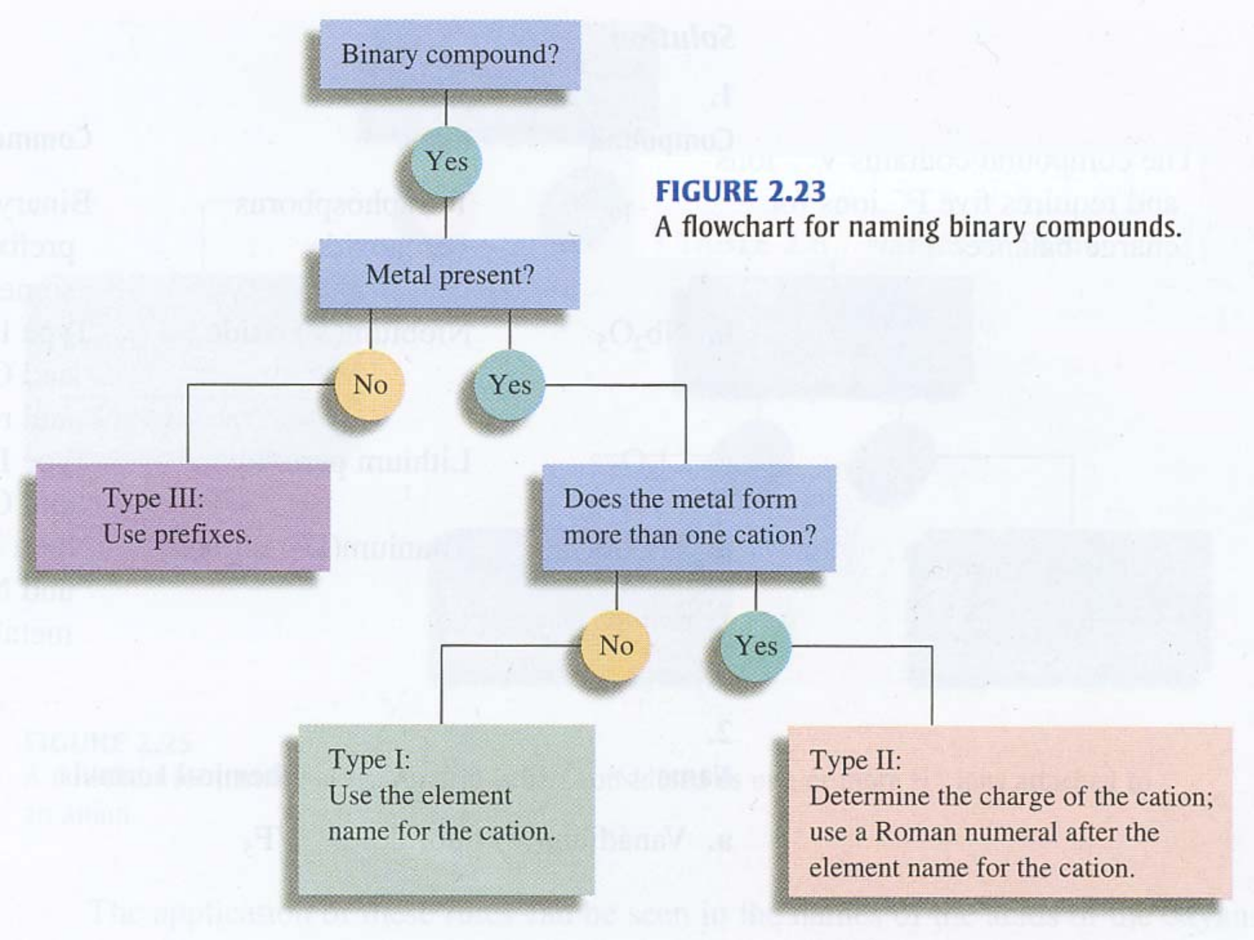


FIGURE 2.23
A flowchart for naming binary compounds.

- Another flow chart (Figure 2.24) from your text expands the naming process to encompass compounds containing polyatomic ions:

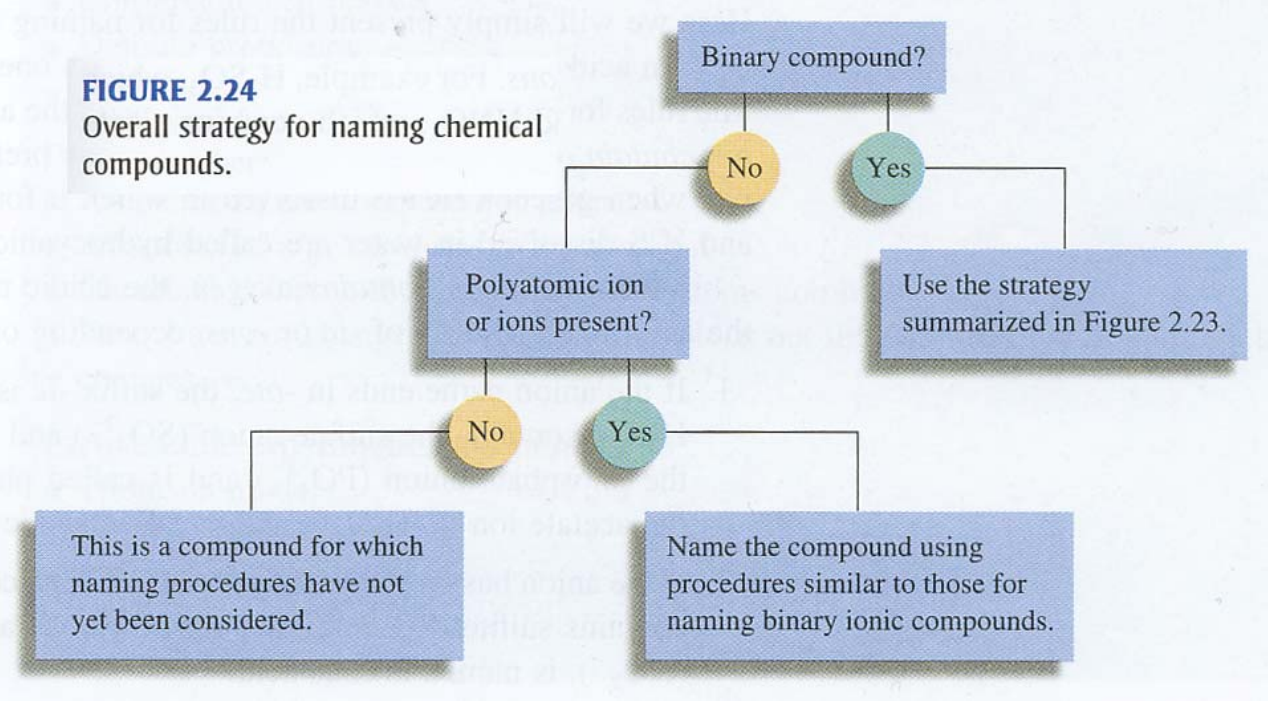


FIGURE 2.24
Overall strategy for naming chemical compounds.

- The following examples cover the naming of all the types of compounds we have discussed in this set of notes. You should be able to use the periodic table to deal with any ion whose name and formula you have not been asked to memorize.

Formula	**Name**
P_4O_{10}	
Nb_2O_5	
Li_2O_2	
$Ti(NO_3)_4$	
	Vanadium (V) fluoride
	Dioxygen difluoride
	Rubidium peroxide
	Gallium oxide

- **Names of Acids**
  - Acids are so important that three entire chapters of your text (Chapters 4, 14, and 15) are devoted to them. Here we will confine the discussion to a brief definition of acid and to some rules for naming acids.

- Acids are substances that, when dissolved in water, will produce hydrogen ions (H^+). An acid can be pictured as a molecule with one or more protons (H^+) attached to an anion.
- How to name acids:
  - If the anion does not contain oxygen, the acid's name includes the prefix, hydro-, and the suffix, -ic. The word, "acid," follows. For example, HCl contains a chloride anion, (Cl^-), and is named Hydrochloric acid. The most important acids of this class are listed in Table 2.7 from the text.

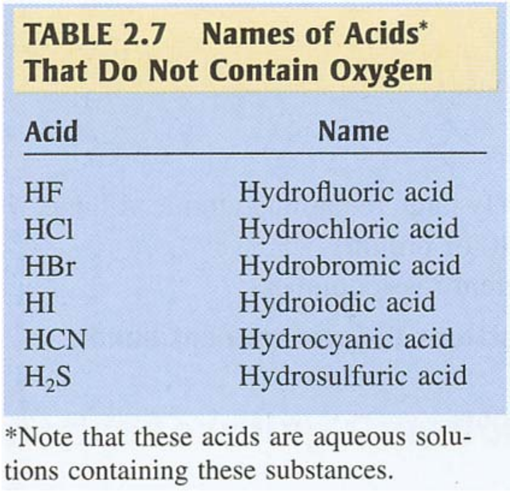


TABLE 2.7 Names of Acids* That Do Not Contain Oxygen

Acid	Name
HF	Hydrofluoric acid
HCl	Hydrochloric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
HCN	Hydrocyanic acid
H_2S	Hydrosulfuric acid

*Note that these acids are aqueous solutions containing these substances.

  - If the anion does contain oxygen, and the name of the anion ends in "–ate," the suffix, "-ic," replaces the "-ate," and the word, "acid," follows. For example, H_2SO_4 contains a sulfate anion (SO_4^{2-}), and is named sulfuric acid.
  - If the anion contains oxygen and the name of the anion ends in "-ite," the suffix "-ous," replaces the "-ite," and the word, "acid," follows. For example, H_2SO_3 contains a sulfite anion (SO_3^{2-}), and is named sulfurous acid. Some examples of oxygen-containing acids are listed in Table 2.8.

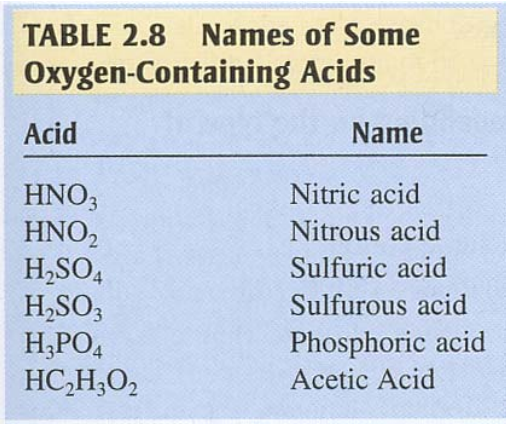
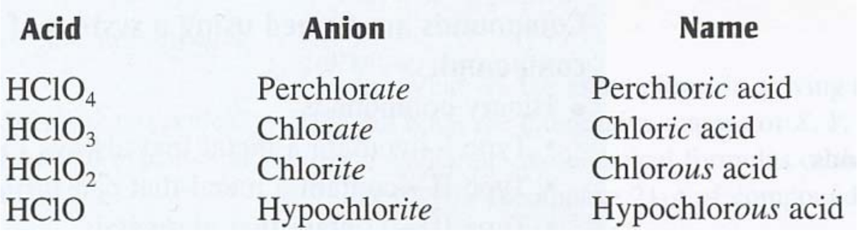


TABLE 2.8 Names of Some Oxygen-Containing Acids

Acid	Name
HNO_3	Nitric acid
HNO_2	Nitrous acid
H_2SO_4	Sulfuric acid
H_2SO_3	Sulfurous acid
H_3PO_4	Phosphoric acid
$HC_2H_3O_2$	Acetic Acid

- The oxyacids of chlorine illustrate these rules:



Acid	Anion	Name
$HClO_4$	Perchlor*ate*	Perchlor*ic* acid
$HClO_3$	Chlor*ate*	Chlor*ic* acid
$HClO_2$	Chlor*ite*	Chlor*ous* acid
$HClO$	Hypochlor*ite*	Hypochlor*ous* acid

- A flowchart for the naming of acids is given in the text (Figure 2.25).

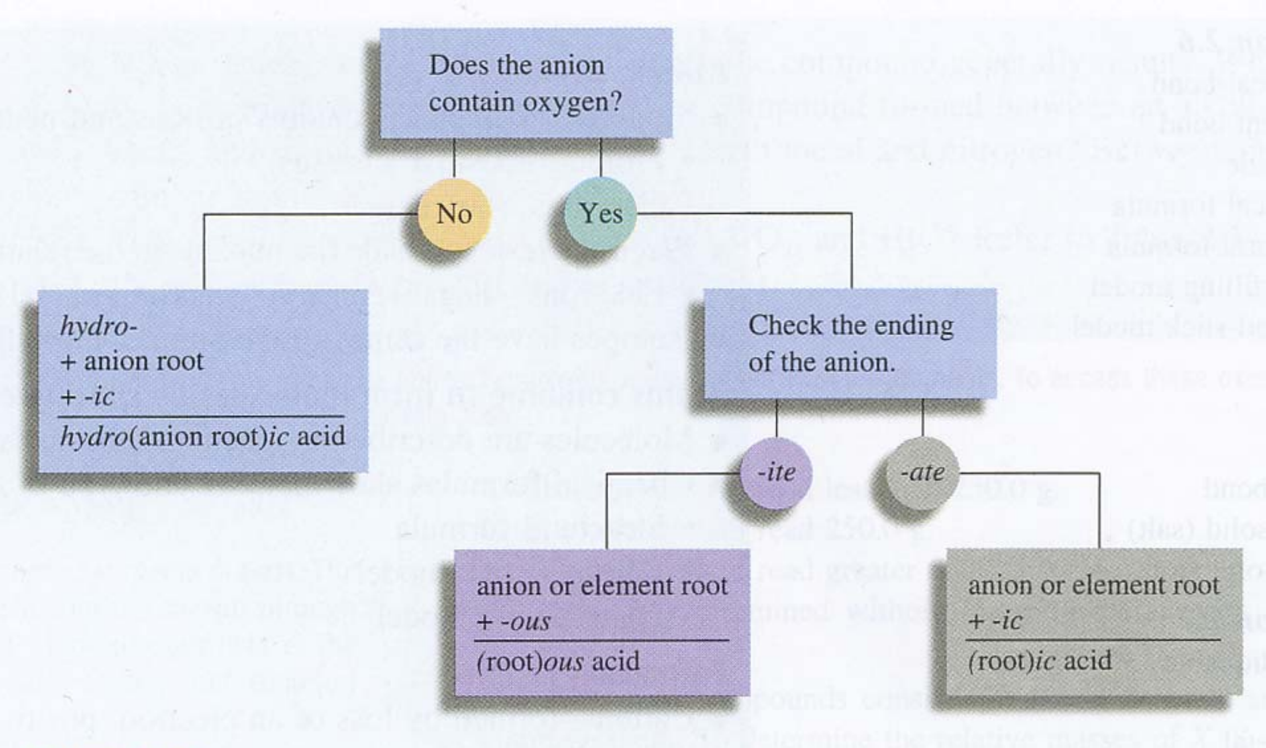


FIGURE 2.25
A flowchart for naming acids. An acid is best considered as one or more H^+ ions attached to an anion.

Chapter 3 – Stoichiometry

Chemistry is a quantitative science, and the counting and figuring that underlies it is called stoichiometry. You can also call it chemical arithmetic.

3.1 Counting by Weighing

- It is very tedious and time-consuming to count large numbers of small objects. Suppose you have a jar full of pennies, and you would like to know how many there are. Suppose also that you know the mass of a single penny. Then, the most efficient way to count the pennies in the jar is to weigh its contents to get the total mass and divide that by the mass of a single penny. This gives you the count, if you assume that each penny has the same mass as any other penny.

- Suppose instead that you have objects like jelly beans that don't all have exactly the same masses, but whose masses are all fairly close to each other. The sample of 10 beans in your text (p. 77) has a distribution of masses (The totals for each column appear in the bottom row):

Number of Beans	**Mass (g)**	**% of Total**
1	4.8	10.
2	4.9	20.
4	5.0	40.
2	5.1	20.
1	5.2	10.
10	50.0	100.

The above masses are all close to 5.0 g each, so it makes sense to compute the average mass of a bean:

$$m_{av} \tfrac{g}{\text{bean}} = \frac{1\times 4.8 + 2\times 4.9 + 4\times 5.0 + 2\times 5.1 + 1\times 5.2}{10 \text{ beans}} = \frac{50.0g}{10 \text{ beans}} = 5.0 \tfrac{g}{\text{bean}}$$

Now if we want to measure out 1,000 beans, we don't need to count them. All we need to do is calculate the mass of the 1,000 beans by multiplying 1,000 beans by the average mass of a single bean:

$$1000 \text{ beans} \times 5.0 \tfrac{g}{\text{bean}} = 5000g$$

We count pennies and beans by weighing because of convenience, not because there is no way to count them one-by-one. It is different with atoms because they are so small and because there are so many of them in the amounts we would normally handle. Under nearly all circumstances, we must weigh our atoms in order to count them.

3.2 Atomic Masses

- Each element has its own characteristic atomic mass, and you can look these up in the periodic table (inside front cover of your text) or in the table of atomic masses (facing the inside front cover). How were these masses determined? We saw in Chapter 2 that pioneering chemists like Lavoisier, Dalton, and Berzelius performed careful analytical work to determine the relative combining weights of all the then-known elements, and that Gay-Lussac determined the relative combining volumes of elements and compounds involved in gaseous reactions. Then Dalton's atomic theory and Avogadro's principle allowed this information to be interpreted in terms of relative combining numbers of atoms. Thus water (for example) was known to contain 2 grams of hydrogen for each 16 grams of oxygen, and a water molecule was known to contain 2 hydrogen atoms and 1 oxygen atom. Thus if we define the atomic mass of hydrogen to be exactly 1 atomic mass unit, we can say that the atomic mass of oxygen is 16 atomic mass units.
- **Defining the atomic mass unit:** The earliest tables of atomic masses were based on a scale where the atomic mass of **hydrogen** is defined as **1**. This was unsatisfactory because the atomic masses of the heaviest elements deviated substantially from values clustered around .9, .0, and .1. The standard I learned when I took General Chemistry 50 years ago was that the atomic mass of **oxygen** was defined as exactly **16**. Even this was unsatisfactory, because it depended on the isotopic composition of oxygen, i. e., how much ${}^{17}_{8}\mathrm{O}$ and ${}^{18}_{8}\mathrm{O}$ were mixed with the predominant isotope, ${}^{16}_{8}\mathrm{O}$. So in 1961, the IUPAC adopted the now-current standard where the mass of ${}^{12}_{6}\mathrm{C}$ is defined as exactly 12 atomic mass units (amu).
- **Dealing with Isotopic Composition:** Most elements are like oxygen, in that they have two or more isotopes. For example, carbon consists of 98.89% ${}^{12}_{6}\mathrm{C}$, 1.11% ${}^{13}_{6}\mathrm{C}$, and a negligible amount of the (radioactive) isotope, ${}^{14}_{6}\mathrm{C}$. You might ask, Isn't it like trying to add apples and oranges to use a single value for the atomic mass of carbon? The answer is that we are dealing with such large numbers of atoms at once, that we can depend on the proportions of ${}^{12}_{6}\mathrm{C}$ and ${}^{13}_{6}\mathrm{C}$ remaining constant from sample to sample. Therefore we can compute an average atomic mass for carbon, given that the atomic mass of ${}^{12}_{6}\mathrm{C}$ is 12 amu

(exactly) and the atomic mass of ${}^{13}_{6}\mathrm{C}$ is 13.003355 amu. The computation technique is called calculating a weighted average.

$$(0.9889)(12)\ \text{amu} + (0.0111)(13.003355)\ \text{amu} = 12.01\ \text{amu}$$

Just as we used 5.0 g as the average mass of a jelly bean when we counted jelly beans by weighing them, we can use 12.01 amu as the average mass of a carbon atom when we count carbons by weighing them.

- From here on, we will consider (average) atomic masses to be known properties of their elements that can be looked up in the periodic table or in the table of atomic masses.
- **Atomic Masses by Mass Spectrometry:** This is an interesting subject that you are encouraged to read about in Section 3.2, but none of your homework problems require you to know how a mass spectrometer works, and we will not discuss the topic in class.