CH 301 REVIEW: CHEMICAL FORMULAS AND COMPOSITION STOICHIOMETRY

Chapter Summary

Much of high school chemistry was devoted to introducing the concept of **Dalton's Atomic Theory**, which provides a foundation for a more quantitative description of matter. By combining the idea of the atom with the **Law of Multiple Proportions**, the concept of the molecule was developed. Chemical formulas and the associated names for small molecules and ions were introduced. From the idea that the individual atom representing each element has its own unique mass, the concept of atomic weight was developed. Then, in order to operate with experimentally-manageable amounts of matter, the mole was introduced, and the fun with calculations began.

The simple numerical relationship between atoms in molecules allows a collection of calculations called compositional stoichiometry problems to be performed. Most of this chapter is devoted to reviewing every kind of problem in which mass, numbers of atoms, moles, atomic and molecular formulas, elemental formulas, and percent composition are related to one another.

The good news is that there is a common approach to solving all stoichiometry problems: the use of a unit factor. Thus, in the same way that you have learned to convert inches into miles, by making the units in unit factors cancel, you can perform all the conversions in compositional stoichiometry. You will start with some known quantity like:

atoms of A and be asked to convert it into **grams of molecule AB**

Now, as you work problems like this, you will find that there are patterns that you will use over and over in putting together unit factors. The last half of the chapter is spent looking at almost every permutation of unit factors. We will look at a collection of examples, and then try to come up with a general procedure for working the problems.

One word of caution: You must learn to do compositional stoichiometry problems in this chapter effortlessly because the next review material considers reaction stoichiometry, which assumes you can do compositional stoichiometry blindfolded.

ATOMIC THEORY

Early scientists had figured out that a very complex world could be reduced to a collection of elemental materials. John Dalton came along in the early 1800s and proposed that these **elemental materials were made up of very small, indivisible particles he called atoms**. Dalton was to provide the framework for a theory, which although not perfect, launched the modern age of chemistry and physics.

Here are some ideas of **Dalton's Atomic Theory**:

1. An element is composed of small, indivisible particles called **atoms**.

2. All atoms of a given element have identical properties, which are unique from other elements.

3. Atoms cannot be created, destroyed, or converted into other elements.

4. **Molecules** are formed when atoms of different elements combine in simple, whole number ratios.

5. The relative number and type of atoms is constant for a given molecule.

Flaws in Dalton's Theory

Note some of the flaws in this theory. Obviously in the nuclear age, atoms can be created and destroyed. In addition, the theory does not recognize the existence of different **isotopes** of an atom. Nonetheless, the theory works extremely well as a platform for describing almost every chemical and physical process we will examine in this course.

Time out! We all know that atoms are made up of **electrons**, **protons**, and **neutrons**, but Dalton didn't in the early 1800's. And we don't need to know this either to appreciate the descriptive chemistry and stoichiometry calculations covered in this review material. So kick back and enjoy the simplicity of believing that atoms are small, indivisible particles: tiny billiard balls of different sizes.

CHEMICAL FORMULAS: ATOMS

Before we can start working all of our stoichiometry problems, we have to be able to develop a **nomenclature**—a shorthand language—for describing atoms and molecules. We will use, as a starting place, the shorthand notation for elements found in every periodic table. Try memorizing a few dozen elements and their abbreviations from early in the periodic table to speed up your test taking. Here are a dozen to get you started.

Atomic Number	Name	Symbol	Period	Group	Chemical Series	Mass (g/mol)
1	Hydrogen	H	1	1	Nonmetal	1.00794
2	Helium	He	1	18	Noble Gas	4.002602
3	Lithium	Li	2	1	Alkali Metal	6.941
4	Beryllium	Be	2	2	Alkaline Earth	9.012182
5	Boron	B	2	13	Metalloid	10.811
6	Carbon	C	2	14	Nonmetal	12.0107
7	Nitrogen	N	2	15	Nonmetal	14.0067
8	Oxygen	O	2	16	Nonmetal	15.9994
9	Fluorine	F	2	17	Halogen	18.9984032
10	Neon	Ne	2	18	Noble Gas	20.1797
11	Sodium	Na	3	1	Alkali Metal	22.98976928
12	Magnesium	Mg	3	2	Alkaline Earth	24.3050

By the way, do you have to memorize the elements? Of course not. There is always a periodic table handy. However, if you find yourself having to look up oxygen's atomic mass of 16 for the 300th time, you are probably not finishing the exams on time.

CHEMICAL FORMULAS: MOLECULES

Recall that Dalton’s theory suggests that molecules can be made from simple combinations of atoms. We can write a shorthand notation for this by using a subscript beside the atom abbreviation. In its simplest form with a single type of element, we can have

Monatomic molecules:	He	Helium
	Ne	Neon
Diatomic molecules:	O_2	Dioxygen
	H_2	Dihydrogen
Polyatomic molecules:	P_4	Tetraphosphorous
	S_8	Octasulfur

We can also mix elements together in various combinations to form the vast array of molecules that make up our universe:

One hydrogen and one chloride	HCl	Hydrochloric acid
Two hydrogens and one oxygen	H_2O	Water
One nitrogen and three hydrogens	NH_3	Ammonia
Three carbons and eight hydrogens	C_3H_8	Propane

The entire first half of the course is spent discussing bonding theories so that you can draw three dimensional structures of the compounds above, and thousands more, in your sleep.

IONS: COMPOUNDS WITH CHARGE

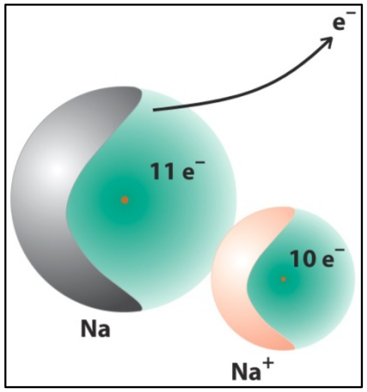
Life is never simple – neither is chemistry. For reasons we will learn about later, it is possible to find compounds that are not neutral but instead possess a charge.

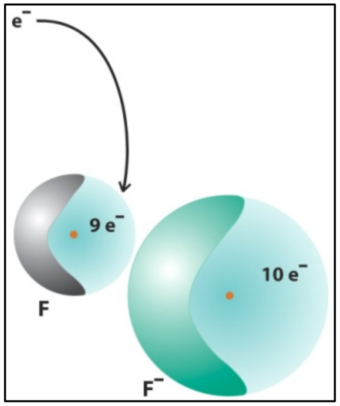
Cation: a positively-charged compound.

- **Examples:** Na^+, NH_4^+

Anion: a negatively-charged compound.

- **Examples:** F^-, SO_4^{-2}





Ionic compounds: What you get when you put together anions and cations to obtain a neutral species (no charge)

- **Examples:** NaCl, $(NH_4)_2SO_4$

IONS ARE NOT MOLECULES

Now Davis makes a big deal about **ions** not being able to form molecules. Thus, NaCl is not a molecule; it is just a simple way of describing the appropriate ratio of ions which come together to form regularly packed crystals. For example, shown in Figure 2 are ball and stick and space-filled models of table salt. So what kind of word do we use if we can't use the word molecule to describe NaCl?

Formula unit: the simplest whole number ratio of ions in a compound.

Thus, NaCl is the **formula unit** for sodium chloride.
But NaCl is **not** a **molecular unit** for sodium chloride

NAMING COMPOUNDS: IONS

As a start on the very important but very difficult process of naming chemical compounds, we will look at some rules of naming ionic compounds.

- The formula of an ionic compound is adjusted to make the **formula unit** neutral.
- The positive cation species are written first and followed by the negative anion species.
- The common ionic nomenclature for individual ions is used to assemble an ion's **formula unit**.

Examples:

$NaCl$	$Na^{+}\ Cl^{-}$	Sodium chloride
$(NH_4)_2S$	$2NH_4^{+}\ S^{=}$	Ammonium sulfide
$MgBr_2$	$Mg^{+2}\ 2Br^{-}$	Magnesium bromide

ATOMIC WEIGHT

You'll recall that early scientists were able to reduce mixtures to their elemental forms. In the process, they weighed these materials and not only came up with the law of multiple proportions, but also a relative ratio of weights of the different elements. For example, they found that by assigning hydrogen, the lightest element, an **atomic mass unit** of one, the following approximate relative ratios of other elements were:

Hydrogen	1	Magnesium	24
Carbon	12	and so on...	

With more precise measurements, it was found that these whole number values were not exact (we'll learn about isotopes and mass defects later). However, this concept did give way to the idea of **atomic weight**, which can be found for each element in the periodic table. This value represents an average weight for *naturally occurring* amounts of each element. Hence, the atomic weights are not even close to whole numbers.

For example, Cl has an atomic weight of 35.45 amu.

The rest of the atomic weights are listed in the periodic table.

THE MOLE

Recall that Dalton suggested atoms were very small. In contrast, we are very big. So, if we want to work with reasonable quantities of materials (like amounts we can hold in our hand and see), we have to deal with incredibly large numbers of atoms. It is out of this need to hold an amount we can see, but not want to have to deal with such big numbers, that the idea of a **mole** was born—in exactly the same way that chicken farmers dealt with dozens of eggs to reduce the magnitude of the number describing eggs, and Lincoln used score as a way to reduce the size of the number describing years ("four score and seven years ago" let him use the numbers four and seven rather than the number 87). In the same way, we use the mole to reduce the size of the numbers we use to describe atoms and molecules.

So: dozen = 12; score = 20; and **mole = 6.02×10^{23}**.

Where did 6.02×10^{23} (Avogadro's number) come from? Well, arbitrarily, it was decided to allow a mole to be equal to the atomic weight of an element in atomic mass units. This mass is called the molar mass and has units of grams/mole. So:

> A mole of hydrogen atoms weighs about 1 gram.
> A mole of oxygen atoms weighs about 16 grams.

And exactly how many molecules are in a mole of a sample? Like we said, Avogadro's number is: $6.02213167 \times 10^{23}$ particles.

This is an extraordinarily large number—a larger number than you will ever deal with in just about any other area of science. This magnitude makes talking about billions (10^9) and trillions (10^{12}) of things (like people or dollars) look minuscule.

WHY DO WE BOTHER TO USE MOLES?

It allows us to scale up using the law of multiple proportions in putting together simple ratios of atoms to make a molecule. This way, when we get a can of diet coke, we can say:

> "Hey, 360 mL = 360 grams = 20 moles of Diet Coke = 40 moles of H_2 combined with 20 moles of O_2, sure tastes good!"

Instead of saying:

> "Hey, 360 mL = 360 grams = 120×10^{23} molecules = 240×10^{23} atoms of H_2 combined with 120×10^{23} atoms of O_2, sure tastes good!"

That is all there is to it. We use moles so we don't have to run around saying "times ten to the twenty-third" all the time.

LET'S DO SOME PROBLEMS

Stoichiometry problem solving is all about setting up and canceling unit factors, until the unit you want results from the unit with which you begin.

The best news of all is that there are only three unit factors you will use commonly:

Density	Volume to mass
Avogadro's number	Moles to # of particles
Atomic weight	Moles to mass

The second best news is that the problems are **always** set up exactly the same way:

what you don't know = (*what you do know*)**(unit factors)(unit factors)(unit factors)** = *answer*

You find the answer to what you don't know by taking the value you are given in the problem and then using unit factors that turn the units of what you do know into the units of what you don't know.

It is that simple. Proof, let's look at some examples.

Example 1: Calculate the mass of a magnesium atom, Mg, in grams.

$$\text{? mass of Mg atom} = \left(24.3\,\frac{\text{g}}{\text{mole Mg}}\right)\left(\frac{1\text{ mole Mg}}{6.02\text{ X }10^{23}\text{ atoms}}\right)$$

$$= 4.04\text{ X }10^{-23}\text{ g / atom}$$

Example 2: Calculate the number of atoms in one-millionth of a gram of magnesium, Mg.

$$\text{? atoms} = \left(1\text{ X }10^{-6}\text{ g Mg}\right)\left(\frac{1\text{ mole Mg}}{24.3\text{ g}}\right)\left(\frac{6.02\text{ X }10^{23}\text{ atoms}}{1\text{ mole Mg}}\right)$$

$$= 2.47\text{ X }10^{16}\text{ atoms}$$

Example 3: How many atoms are in 1.67 moles of magnesium?

$$\text{? atoms} = \left(1.67\text{ moles Mg}\right)\left(\frac{6.02\text{ X }10^{23}\text{ atoms}}{1\text{ mole Mg}}\right)$$

$$= 1.00\text{ X }10^{24}\text{ atoms}$$

Example 4: How many moles of magnesium are in 73.4 grams of magnesium?

$$? \text{ moles of Mg atoms} = (73.4 \text{ g Mg})\left(\frac{1 \text{ mole Mg}}{24.3 \text{ g}}\right)$$

$$= 3.0 \text{ moles}$$

FORMULA WEIGHT (MOLECULAR WEIGHT)

To this point, we've dealt with moles of atoms. Now we deal with **moles** of **molecules** by introducing the concept of formula or molecular weight (these are the same if dealing with molecules. If dealing with ionic materials, it is more appropriate to describe a formula weight.)

Formula weight: The formula weight of a substance is found by summing the atomic weights of each atom in the molecular formula (or formula unit).

Example 5: Calculate the formula weight of propane, C_3H_8.

Atomic weight C = 12.01 g/mole
Atomic weight H = 1.0079 g/mole
Formula weight = 3 x 12.01 + 8 x 1.0079 = 44.09 g/mole

Now we can start to play games involving moles and molecules. One mole of a substance is 6 x 10^{23} particles of the substance and is equal to the formula weight in grams of the substance.

Example 6: Lets play fill in the blank:

The mole of the molecule, chlorine, Cl_2, weighs 2 x 35.45 = 70.9 grams.
It contains 2 moles of chlorine atoms, 6.02 x 10^{23} Cl_2 molecules and 2 x 6.02 x 10^{23} Cl atoms.

Example 7: Calculate the number of propane, C_3H_8 molecules, in 74.6 grams of propane.

$$? \text{ propane molecules} = 74.6 \text{ g}\left(\frac{1 \text{ mole}}{44.09 \text{ g}}\right)\left(\frac{6.02 \text{ X } 10^{23} \text{ molecules}}{1 \text{ mole}}\right)$$

$$= 1.01 \text{ X } 10^{24} \text{ molecules}$$

Example 8: What is the mass of 10.0 billion molecules of propane?

$$? \text{ g molecules} = \left(10.0 \text{ X } 10^{9} \text{ molecules}\right)\left(\frac{1 \text{ mole}}{6.02 \text{ X } 10^{23} \text{ molecules}}\right)\left(\frac{44.09 \text{ g}}{1 \text{ mole}}\right)$$

$$= 7.33 \text{ X } 10^{-13} \text{ g}$$

Example 9: How many moles, molecules, and oxygen atoms are contained in a 60 gram sample of ozone, O_3?

$$? \text{ moles} = \left(60 \text{ g } O_3\right)\left(\frac{1 \text{ mole}}{3 \text{ X } 15.9994 \text{ g}}\right) = 1.25 \text{ moles}$$

$$? \text{ molecules} = \left(1.25 \text{ moles}\right)\left(\frac{6.02 \text{ X } 10^{23}}{1 \text{ mole}}\right) = 7.53 \text{ X } 10^{23} \text{ molecules}$$

$$? \text{ atoms} = \left(7.53 \text{ X } 10^{23} \text{ molecules}\right)\left(\frac{3 \text{ O atoms}}{1 \text{ } O_3 \text{ molecules}}\right) = 2.26 \text{ X } 10^{24} \text{ atoms}$$

Now let's get a little fancier with our stoichiometry and some famous kinds of chemistry problems.

PERCENT COMPOSITION

The usual method for counting amounts of chemical compounds uses the mole as currency: a mole of propane, C_3H_8 contains 3 moles of carbon and eight moles of hydrogen.

Unfortunately, many of the devices we use for measuring chemicals actually measure the MASS of individual samples. Thus we need to be able to readily convert between MOLES of individual atoms in a molecule to percent by weight of each atom in the molecule.

Example 10: What is the percent composition of each element in propane?

$$1 \text{ mole } C_3H_8 = \left(3 \text{ moles C}\right)\left(\frac{12.01 \text{ g}}{1 \text{ mole}}\right) = 36.03 \text{ g C}$$

$$= \left(8 \text{ moles H}\right)\left(\frac{1.0079 \text{ g}}{1 \text{ mole}}\right) = 8.06 \text{ g H}$$

$$\text{total weight} = 36.03 \text{ g C} + 8.06 \text{ g H} = 44.09 \text{ g}$$

$$\% \text{ C} = \left(\frac{36.03 \text{ g C}}{44.09 \text{ g sample}}\right) \text{X } 100 = 81.7\%$$

$$\% \text{ H} = (100 - 81.7)\% = 18.3\%$$

EMPIRICAL FORMULA

We have learned by experimental means that it is possible to determine the % composition of an unknown compound. This percent composition can easily be turned into a molar ratio of elements, which the law of simple proportions tells us will be a simple whole number ratio. This simple whole number ratio is the smallest possible ratio of atoms in the sample: the **empirical formula** or **simplest formula**. However, the **molecular formula** may be some multiple of the **empirical formula**.

Example 11: A compound that is 24.74% K, 34.76% Mn, and 40.40% O, by mass has what empirical formula?

1. Convert grams to moles, assume 100 g sample

$$24.74 \text{ g K}\left(\frac{1 \text{ mole}}{39.09 \text{ g}}\right) = 0.633 \text{ mole}$$

$$34.76 \text{ g Mn}\left(\frac{1 \text{ mole}}{54.93 \text{ g}}\right) = 0.633 \text{ mole}$$

$$40.50 \text{ g O}\left(\frac{1 \text{ mole}}{15.9994 \text{ g}}\right) = 2.53 \text{ mole}$$

2. Divide by smallest number of moles

$$\frac{0.633}{0.633} = 1$$

$$\frac{2.53}{0.633} = 4$$

$\therefore$ the empirical formula is $KMnO_4$

DETERMINATION OF MOLECULAR FORMULA

We can take the elemental composition problem one step further by adding knowledge about a compound's actual molecular weight. There are a variety of analytical techniques for determining the molecular weight of a compound, **mass spectrometry** being the most famous of the bunch.) Thus if we know:

1. The empirical (simplest) formula
2. The molecular weight

We can obtain the molecular formula of the compound.

For example, if the empirical formula of a hydrocarbon is C_2H_5 (empirical weight = 12+12+5 = 29) and the molecular weight is 116, then the molecular formula is simply (116/29) = a four-fold multiple of C_2H_5 = C_8H_{20}.

Example 12: A compound is found to contain 85.63% C and 14.37% H by mass. In another experiment its molecular weight is found to be 56.1 grams/mole. What is the molecular formula of the compound?

First, find the empirical formula. Convert g → mole

$$85.63 \text{ g C}\left(\frac{1 \text{ mole}}{12 \text{ g}}\right) = 7.12 \text{ moles}$$

$$14.37 \text{ g H}\left(\frac{1 \text{ mole}}{1.0079 \text{ g}}\right) = 14.25 \text{ moles}$$

$$\text{so we have } \frac{14.25 \text{ moles H}}{7.12 \text{ moles C}} \Rightarrow CH_2 \quad \text{as the empirical formula}$$

Second, the empirical weight of $CH_2 \approx 14$

$$\text{so the } \frac{\text{molecular weight}}{\text{empirical weight}} \approx \frac{56}{14} = 4$$

$\therefore$ the molecular formula is C_4H_8

<u>Example 13:</u> What mass of phosphorous is contained in 45.3 grams of $(NH_4)_3PO_4$?

$$\left(\mathrm{NH_4}\right)_3\mathrm{PO_4} = \text{M.W. of } 149.08\ \mathrm{g/mole}$$

$$?\ \mathrm{g\ P} = 45.3\ \mathrm{g}\ \left(\mathrm{NH_4}\right)_3\mathrm{PO_4}\left(\frac{1\ \text{mole}}{149.08\ \mathrm{g}}\right) = 0.304\ \text{moles}$$

$$= 0.304\ \text{moles}\ \left(\mathrm{NH_4}\right)_3\mathrm{PO_4}\left(\frac{1\ \text{mole P atom}}{1\ \text{mole}\ \left(\mathrm{NH_4}\right)_3\mathrm{PO_4}}\right)\left(\frac{30.97\ \mathrm{g\ P}}{1\ \text{mole P}}\right)$$

$$= 9.41\ \mathrm{g\ P}$$

<u>Example 14:</u> What mass of ammonium phosphate would contain 15.0 g of nitrogen?

$$?\ \mathrm{g}\ \left(\mathrm{NH_4}\right)_3 PO_4 = 15\ \mathrm{g}\ N\left(\frac{1\ \text{mole N}}{14.00\ \mathrm{g}}\right)\left(\frac{1\ \text{mole}\ \left(\mathrm{NH_4}\right)_3 PO_4}{3\ \text{mole N atoms}}\right)\left(\frac{149.08\ \mathrm{g}}{1\ \text{mole}\ \left(\mathrm{NH_4}\right)_3 PO_4}\right)$$

$$= 53.2\ \mathrm{g}\ \left(\mathrm{NH_4}\right)_3\mathrm{PO_4}$$

<u>Example 15:</u> What mass of propane, C_3H_8, contains the same mass of carbon as is contained in 1.35 grams of barium carbonate, $BaCO_3$?

$$?\ \text{mole C in } \mathrm{BaCO_3} = (1.35\ \mathrm{g\ BaCO_3})\left(\frac{1\ \text{mole}}{197.3\ \mathrm{g}}\right)\left(\frac{1\ \text{mole C}}{1\ \text{mole}\ \mathrm{BaCO_3}}\right)$$

$$= 6.8\ \mathrm{X}\ 10^{-3}\ \text{moles of C}$$

want the same number of mole of C in C_3H_8 as $BaCO_3$

$$\mathrm{g\ C_3H_8} = (6.8\ \mathrm{X}\ 10^{-3}\ \text{mole C})\left(\frac{1\ \text{mole}\ \mathrm{C_3H_8}}{3\ \text{moles C atoms}}\right)\left(\frac{44\ \mathrm{g\ C_3H_8}}{1\ \text{mole}}\right)$$

$$= 0.1\ \text{g of}\ \mathrm{C_3H_8}$$

PURITY OF SAMPLES

The percent purity of a compound can be used in a collection of unit factors, which make problems more difficult. In these problems, it is necessary to have an appropriate scaling factor or multiplier added to the problem to get the right answer. For example, if you have a bottle of ethanol that contains 0.5% benzene impurity, the ethanol is 99.5 % pure. We can set up a collection of unit factors:

$$\frac{0.5\text{ g benzene}}{100\text{ g sample}} \qquad \frac{99.5\text{ ethanol}}{100\text{ g sample}} \qquad \frac{0.5\text{ g benzene}}{99.5\text{ g ethanol}}$$

Example 16: A bottle of sodium phosphate, Na_3PO_4, is 98.3% pure Na_3PO_4. What are the masses of sodium phosphate and impurities in 250 grams of sample?

$$250\text{ g sample}\left(\frac{98.3\text{ g }Na_3PO_4}{100\text{ g sample}}\right) = 245.8\text{ g }Na_3PO_4$$

$$250\text{ g sample}\left(\frac{1.7\text{ g impurity}}{100\text{ g sample}}\right) = 4.25\text{ g impurity}$$

And now on to reaction stoichiometry, in which we apply these ideas of **stoichiometry** to chemical reactions.

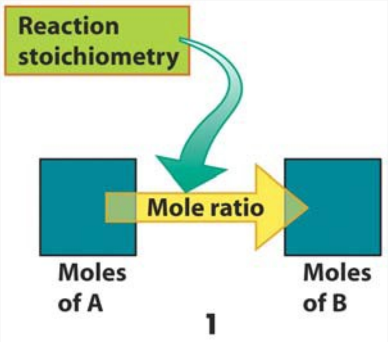
REACTION STOICHIOMETRY

Chapter Summary

We just learned that simple quantitative relationships based upon the idea of the law of simple proportions could be combined with other concepts from Dalton's Atomic Theory to create a host of problems based upon the quantitative relationships between atoms in molecules. We learned to use unit factors to convert between:

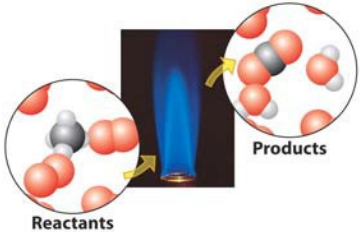
grams and moles
moles and number of atoms or molecules
atoms_A and molecules_A

Now we take things just a single step further by applying these concepts of stoichiometry to chemical reactions. After all, since we know there is conservation of mass in the world, then when a chemical change occurs and new compounds are formed by chemical reaction, there must be a collection of simple unit factors that describe these changes. The additional unit factor is simply asking for a stoichiometric relationship between numbers of atoms or molecules of one chemical species and another chemical species involved in a reaction.



Example: Consider the combustion of methane, CH_4:

methane + oxygen → water and carbon dioxide



Writing it without concern for stoichiometry we have:

$CH_4 + O_2 \rightarrow CO_2 + H_2O$ (which is unbalanced and violates the conservation of matter)

But by inspection, we can find a way for conserve matter in a balanced reaction by placing a coefficient of 2 in front of the diatomic oxygen and in front of the water:

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ (balanced, matter is conserved, all is right in the world)

From this balanced equation we can form all kinds of unit factors relating molecules or atoms of reactants to molecules or atoms of products.

For example:

- There is one mole of CH_4 for every one mole of H_2O
- There are two molecules of O_2 for every one molecule of CO_2
- There are even four atoms of H in CH_4 for every two atoms of O in two H_2O

…and many more unit factors. This means that we can do all stoichiometry problems from compositional stoichiometry with the added twist that we can add unit factors for reactions.

BALANCING CHEMICAL EQUATIONS

One annoying feature of using chemical reactions is that sometimes the equations you are given are not balanced (i.e., they don't satisfy the law of conservation of mass). For example, the first methane combustion equation above was not balanced. You can't use unit factors if the equations are not balanced, so we need some practice with balancing.

So let's try practicing with another reaction, this time propane combustion.

Coefficients are 1, 1, 1, 1 = 4 total molecules in unbalanced reaction.

$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$ Unbalanced

Coefficients are 1, 1, 3, 4 = 9 total molecules in unbalanced reaction.

$C_3H_8 + O_2 \rightarrow 3CO_2 + 4H_2O$ Unbalanced

Coefficients are 1, 5, 3, 4 = 13 total molecules in a balanced reaction.

$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$ Balanced

How did we do that? A few hints for balancing equations (these don't work for complicated redox problems, but they are a way to start with simpler equations):

1. First balance the compounds which contain elements that appear only once on each side of the equation. For example, in the propane combustion reaction above, note that C only appears once on the left (in C_3H_8) and once on the right (in CO_2) above. **Balance it first**. Note that oxygen atoms appear in two compounds on the right (CO_2 and H_2O), do it later.
2. Save the elements or atoms with a single element to balance last (He, H_2, N_2, O_2, P_4, S_8,....). For example, solve for O_2 last in the propane problem above.

If you follow these two rules for balancing equations, solving every equation will be a snap—unless the rules don't apply. Then you are on your own and will wish you'd practiced more.

Look, everyone can balance equations. It is just a matter of how much time it takes: practice a lot and solve your exams equation in 30 seconds or practice while taking the exam and do the problem in five minutes. It is up to you.

Here is a little practice for right now. Try to follow the rules.

Example 1. Balance $NaCl + H_2O \rightarrow NaOH + Cl_2 + H_2$

Start with Na that appears only once on either side and end with Cl_2 or H_2, which are elements:

Step 1. Na are already balanced: $NaCl + H_2O \rightarrow NaOH + Cl_2 + H_2$

Step 2. Two Cl on right means need to add two on left: $2NaCl + H_2O \rightarrow NaOH + Cl_2 + H_2$

Step 3. Now Na are not balanced so need to add two to right side

$2NaCl + H_2O \rightarrow 2NaOH + Cl_2 + H_2$

Step 4. Now balance the O: $2NaCl + 2H_2O \rightarrow 2NaOH + Cl_2 + H_2$

Step 5. Finally, balance the H--oh, it already is: $2NaCl + 2H_2O \rightarrow 2NaOH + Cl_2 + H_2$

Example 2. Balance: $H_2SO_4 + NH_3 \rightarrow (NH_4)_2SO_4$

By inspection, add a 2 in front of the ammonia $H_2SO_4 + 2NH_3 \rightarrow (NH_4)_2SO_4$

Example 3. Balance by yourself. $Ba(OH)_2 + P_4O_{10} \rightarrow Ba_3(PO_4)_2 + H_2O$

That wasn't so easy, was it? Here is the answer.

$$6Ba(OH)_2 + P_4O_{10} \rightarrow 2Ba_3(PO_{4)2} + 6H_2O$$

Okay, so the rules don't always work. Actually balancing chemical reactions can get very difficult when oxidation numbers change during a reaction. We'll explore more systematic ways to balance chemical reactions toward the end of CH302 when we start learning about electrochemistry. For now, assume that simple "inspection" will do the trick most of the time.

REACTION STOICHIOMETRY CALCULATIONS

Assuming that you have a balanced equation, you can do a whole host of problems that involve converting between grams of this and atoms of that which are conceptually identical to compositional stoichiometry problems. For example, consider how many units factors there are for:

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

There is a bunch, like:

- 1 molecule of propane for every 3 molecules of CO_2
- 3 atoms of C in propane for 4 molecules of water
- 9 atoms of hydrogen in propane for 8 atoms of hydrogen in water

…and on and on. How many all together? Would you believe hundreds?

Now on to some problems. For the following problems, use the reaction:

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

Example 1: How many CO molecules are required to react with 25 formula units of Fe_2O_3?

$$\text{? CO molecules} = (25\ Fe_2O_3\ \text{form. unit})\left(\frac{3\ \text{CO molecules}}{1\ Fe_2O_3\ \text{form. unit}}\right)$$
$$= 75\ \text{CO molecules}$$

Example 2: How many Fe atoms are produced by reaction of 2.5 x 10_5 formula units of **Fe_2O_3** with excess **CO**?

$$\text{? Fe atoms} = (2.5\ \text{X}\ 10^5\ Fe_2CO_3)\left(\frac{2\ \text{Fe atoms}}{1\ Fe_2CO_3}\right)$$
$$= 5\ \text{X}\ 10^5\ \text{Fe atoms}$$

Example 3: What mass of CO is required to react with 146 grams of **Fe_2O_3**?

$$?\text{g CO} = (146 \text{ g } Fe_2O_3)\left(\frac{1 \text{ mole } Fe_2CO_3}{159.69 \text{ g}}\right)\left(\frac{3 \text{ mole CO}}{1 \text{ mole } Fe_2CO_3}\right)\left(\frac{28 \text{ g CO}}{1 \text{ mole CO}}\right)$$

$$= 76.7 \text{ g CO}$$

Example 4: What mass of **CO_2** can be produced by the reaction of 0.540 moles of **Fe_2O_3** with excess **CO**?

$$? \text{ g } CO_2 = (0.54 \text{ mole } Fe_2O_3)\left(\frac{3 \text{ mole } CO_2}{1 \text{ mole } Fe_2O_3}\right)\left(\frac{44 \text{ g } CO_2}{1 \text{ mole } CO_2}\right)$$

$$= 71.3 \text{ g } CO_2$$

Example 5: What mass of **Fe_2O_3** reacts with excess **CO** if the reaction produces 8.65 grams of **CO_2**?

$$? \text{ g } Fe_2O_3 = (8.65 \text{ g } CO_2)\left(\frac{1 \text{ mole } CO_2}{44 \text{ g } CO_2}\right)\left(\frac{1 \text{ mole } Fe_2O_3}{3 \text{ mole } CO_2}\right)\left(\frac{159.69 \text{ g } Fe_2O_3}{1 \text{ mole } Fe_2O_3}\right)$$

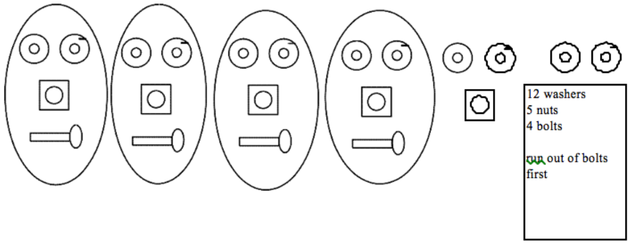
$$= 10.6 \text{ g } Fe_2O_3$$

And now we'll look at some specialty problems involving reaction stoichiometry.

LIMITING REAGENTS

You'll notice that, in several problems, I referred to "excess of this reagent" or "left over that reagent". In every reaction, we will either have a situation in which a stoichiometric amount of reagents react (for example, exactly three times as many **CO** as **Fe_2O_3**) or we will have one of the reagents left over. Practically speaking, we can never hope to use such accurate measures, so we will always have one reagent we run out first in a chemical reaction**.** This is known as the **limiting reagent.**

Example: I want to assemble a gadget that requires one nut, one bolt and two washers for every hole. I have in my garage a bucket filled with 12 washers, 4 bolts and five nuts. What is the limiting small metal object?



Of course, rather than small metal objects, I can use chemical reagents.

Example 6: What is the maximum mass of SO_2 that can be produced by reacting 95.6 grams of CS_2 with 110 grams of O_2?

$$CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2$$

First, find limiting reagent

$$? \text{ moles } CS_2 = 95.6 \text{ g } CS_2\left(\frac{1 \text{ mole } CS_2}{76 \text{ g } CS_2}\right) = 1.26 \text{ moles } CS_2$$

$$? \text{ moles } O_2 = 110 \text{ g } O_2\left(\frac{1 \text{ mole } O_2}{32 \text{ g } O_2}\right) = 3.44 \text{ moles } O_2$$

But we need 3 moles O_2 for every mole CS_2. So if we have 1.26 mole CS_2, then we need 3 X 1.26 = 3.77 moles O_2. We have not quite enough O_2, so it is the limiting reagent.

$$? \text{ g } SO_2 = (3.44 \text{ moles } O_2)\left(\frac{2 \text{ mole } SO_2}{3 \text{ moles } O_2}\right)\left(\frac{64 \text{ g } SO_2}{1 \text{ mole } SO_2}\right) = 146.8 \text{ g } SO_2$$

PERCENT YIELD IN A REACTION

Another concept of interest is the **percent yield**. Ideally, if you react every single molecule in one mole of CS_2 in the above reaction, you can make 2 moles of SO_2. This would correspond to a 100% yield of product. Sadly, life doesn't work that way. If instead we ended up with only one mole of SO_2, we would say that we had achieved only a 50% yield. When you get to organic chemistry lab, you will learn that your chances for medical school might hang in the balance on how high a percent yield you can get from your starting material.

Example 7: A 10.0 gram sample of ethanol, C_2H_5OH, was boiled with acetic acid, CH_3COOH, to produce a 14.8 gram yield of ethyl acetate, $CH_3COOC_2H_5$. What is the percent yield?

$$CH_3COOH + C_2H_5OH \rightarrow CH_3COOC_2H_5 + H_2O$$

$$\text{? moles } C_2H_5OH = 10.0\text{ g}\left(\frac{1\text{ mole}}{46\text{ g } C_2H_5OH}\right) = 0.22\text{ moles of reagent}$$

$$\text{? moles } CH_3COOC_2H_5 = 14.8\text{ g}\left(\frac{1\text{ mole}}{88\text{ g } CH_3COOC_2H_5}\right) = 0.17\text{ moles of product}$$

formed in 1:1 mole ratio so

$$\%\text{ yield} = \frac{0.17\text{ mole product}}{0.22\text{ mole reagent}} = 78\%$$

SOLUTION STOICHIOMETRY: BECAUSE THE INTERESTING STUFF HAPPENS IN WATER

What follows is a discussion of now to determine the concentration of a solution and how to determine the concentration after the reaction of a solution. How important is this material? Well most chemical reactions of use to human beings occur in solution (specifically water) where the reactant molecules of interest can scurry around in the solvent looking for other reagents. This makes for a heck of a lot more exciting reaction chemistry than if you placed to rocks (solids) on top of each other and waited for the reaction to occur. So learning about compositional stoichiometry and reaction stoichiometry in solutions is a very important thing.

A couple of definitions before we start:

- **Solute** is the compound of interest that you put into a solution.
- **Solvent** is the non-reactive excess liquid to which you add the solute.

So if you dump salt in water, you are dumping a **solute** into the **solvent**.

SOLUTION CALCULATIONS

Once you add salt to the water, you might want to know how salty the water was. Is it salty or really salty? Clearly we need some quantitative measure of solute concentration. There are many ways to do this. I give you the two most common.

Percent by mass: This measure of concentration is used because it follows simply from doing a measurement on a balance.

Example:

Add 10 grams of salt to 90 grams of water, then:
The solution is 10% by weight, salt, and 90 % by weight water.
(Remember we found out the Coca-Cola is a 10% by weight sugar in water.)

Example 8: What mass of NaOH is required to prepare a 250 gram solution that is 8% NaOH?

$$8\%\ \text{NaOH} = \frac{8\ \text{g NaOH}}{92\ \text{g H}_2\text{O}} \quad \text{or} \quad \frac{8\ \text{g NaOH}}{100\ \text{g solution}}$$

$$?\ \text{mass NaOH} = 250\ \text{g soln}\left(\frac{8\ \text{g NaOH}}{100\ \text{g soln}}\right)$$

$$= 20\ \text{g NaOH}$$

Of course, most often we weigh the solute and we measure the solvent (a liquid) by volume, for example, using a graduated cylinder. Thus, we have to use **density** to convert from volume to weight.

Example 9: Calculate the mass of NaOH in 300 ml of solution that is 8% NaOH if the density of solution is 1.09 g/ml.

$$?\ \text{g NaOH} = 300\ \text{ml soln}\left(\frac{1.09\ \text{g}}{1\ \text{mL}}\right)\left(\frac{8\ \text{g NaOH}}{100\ \text{g soln}}\right) = 26\ \text{g NaOH}$$

Molarity

By far the most common method for expressing concentration is not on a mass basis but on a MOLE basis. Thus the MOLARITY is defined as

$$\frac{\text{moles}}{\text{liter}} = \frac{\text{grams / molecular weight}}{\text{liter}}$$

If you stay in science, you will make about 200,000 different molarity solutions of things, so you might as well learn how to do it now.

Problem 10. Calculate the molarity of a solution that contains 12.5 grams of H_2SO_4 in 1.75 liters of solution.

$$\text{Molarity } H_2SO_4\left(\frac{\text{moles}}{\text{Liter}}\right) = \frac{\dfrac{12.5 \text{ g}}{98 \text{ g / mole}}}{1.75 \text{ L}} = 0.07 \text{ M}$$

Problem 11. Calculate the mass of $Ca(NO_3)_2$ required to prepare 3.5 liters of a 0.8 M solution.

$$? \text{ g } Ca(NO_3)_2 = \left(\frac{0.8 \text{ mole}}{1 \text{ L}}\right)(3.5 \text{ L})\left(\frac{128 \text{ g}}{\text{mole}}\right) = 358 \text{ g}$$

TIME OUT For SOLUTION REACTION EQUATIONS Here is something important to know. For the rest of the chapter as we learn about neutralizations, or dilutions, or titrations, we will use the following startling facts:

$$\text{Molarity (M)} = \frac{\text{moles}}{\text{Volume (V)}}$$

So on rearranging the equation we have

$$\text{moles} = \text{(Molarity)(Volume)} = \text{MV}$$

But you know that stoichiometry tells us that

moles **a** are related in simple proportions to moles **b**

So we can make all kinds of useful equations:

moles **a** = M_bV_b or M_aV_a = moles **b**

or even

$$M_aV_a = M_bV_b$$

But there is more. Since

moles = grams/molecular weight (MW)

then we have a bunch of other possible equations:

moles **a** = grams **a**/MW_a = M_bV_b = moles **b**

or any other similar combination.

These are the simple equations you use to do DILUTION problems and TITRATION problems and SOLUTION REACTIONS.

A DILUTION CALCULATION

Example 12: What volume of 18.0 M H_2SO_4 is required to prepare 2.5 liters of 2.4 M H_2SO_4.

$$\begin{aligned} \text{moles } H_2SO_4 \text{ before} &= \text{moles } H_2SO_4 \text{ after} \\ M_bV_b &= M_aV_a \\ (18\text{ M})(?\ V_b) &= (2.4\text{ M})(2.5\text{ L}) \\ V_b &= 0.33\text{ L} \end{aligned}$$

A CHEMICAL REACTION IN SOLUTION

Example 13: What volume of 0.5 M $BaCl_2$ solution is required to complete a reaction with 4.32 grams of Na_2SO_4?

$$Na_2SO_4 \;+\; BaCl_2 \;\text{---->}\; BaSO_4 \;+\; 2\,NaCl$$

$$\begin{aligned} \text{moles } BaCl_2 &= \text{moles } Na_2SO_4 \\ M_a V_a &= \frac{g_b}{MW_b} \\ (0.5\text{ M})(?\text{ L } BaCl_2) &= \frac{4.32\text{ g}}{142\text{ g / mole}} \\ V_{BaCl_2} &= 0.61\text{ mL} = 61\text{ mL} \end{aligned}$$

TITRATIONS

If you have had a chemistry class, you probably have done a titration. A titration is nothing more than using:

- The known concentration of something.
- The stoichiometry of a reaction.
- One of the titration equations like $M_1V_1 = M_2\,V_2$.

Let's see how it works.

Example 14: What is the molarity of a KOH solution if 38.7 mL of KOH is titrated with 43.2 mL of a 0.223 M HCl solution?

$$\begin{aligned} \text{moles}_b &= \text{moles}_a \\ M_b V_b &= M_a V_a \\ (?\text{ M}_b)(38.7\text{ mL}) &= (0.223\text{ M HCl})(43.2\text{ mL HCl}) \\ M_{KOH} &= 0.25\text{ M} \end{aligned}$$

AN INTRO INTO DESCRIPTIVE CHEMISTRY

To this point in our review, we've worked on methods for establishing a **quantitative** foundation in chemistry. We can now manipulate all manner of unit factors to solve problems involving amounts of chemical materials, and while there was some modest requirement that you understand the nature of the substances involved in the problems, you might just as well have applied the techniques you learned to solve problems associated with preparing a good white sauce—make the units cancel and you've learned the right answer—whatever those **numbers** meant. There was nothing particularly relevant to chemistry in the problem solving. Well, now there is a radical change in material. For one thing, there aren't any problems to work. Reading this material is like taking a botany class. Suddenly, our primary focus is on **describing** chemistry. So when you get down to it, in this chapter there is a whole lot of **memorizing** going on. However, I will also introduce you to the basic theoretical foundation that **significantly reduces memorization** by learning the following basic concepts.

Concept 1: Periodic Trends Like Metallic Character

- You learn basic trends in the elements that prompted the creation of a periodic table of those elements. The first is which elements are metals and which are not?

Concept 2: Acids Base Chemistry

- You will learn what things do and don't behave as acid and bases.

Concept 3: Which Ions dissolve in Water?

- You will learn elementary solubility rules for ions in water.

Concept 4: Oxidation/Reduction

- You will learn to assign oxidation numbers to elements. The oxidation number is you first indication of where electron density is located in chemical compounds.

Concept 5: Active Metals

- You will learn about displacement reactions, which means you will learn to predict what happens when you dump metals into water.

Concept 6: Chemical Nomenclature

- You will learn the common way to name binary and ternary ionic compounds. A hint: **Don't** memorize compound names, memorize naming rules.

THE PERIODIC TABLE

Long ago in about 1870, Mendeleev was inspired to put together a tabulation of the known elements in such a way as to describe, in a periodic way, their physical and chemical properties. The modern version of Mendeleev's Periodic Table is an attractive addition to most science lecture halls—it is worth studying during drier moments of lectures if for no other reason than to prepare for chemistry tests.

Families of Elements

Note that the table of elements consists of columns (groups or families) of elements in which similar properties are observed.

Group IA	Alkali Metal	Li, Na, K, Rb, Cs
Group IIA	Alkaline Earth Metals	Be, Mg, Ca, Sr, Ba
Group VIIA	Halogens	F, Cl, Br, I
Group 0	Noble or Rare Gases	He, Ne, Ar, Kr, Xe

Periods: Chemical Trends Along Rows of Elements

You can also look at the table as a collection of horizontal rows called periods:

Row 2	Li, Be, B, C, N, O, F, Ne

As we will learn, important trends in chemical and physical properties are observed as we move across a row. For example...

TREND 1: METALS AND NON-METALS

We will look at numerous trends within the periodic table during the semester. Our first trend will be with respect to metallic character in elements. Note in Table 1 the qualities of metals.

Table 1: Some Physical Properties of Metals and Nonmetals

Metals	Nonmetals
High electrical conductivity	Poor electrical conductivity
High thermal conductivity	Good heat insulators
Metallic gray or silver luster	No metallic luster
Almost all are solids	Solids, liquids, or gases
Malleable	Brittle in solid state
Ductile (can be drawn into wires)	Non-ductile

Your first periodic table trend

Metallic: Character increases from right to left and top to bottom. Thus, Cs is the most active metal. Note that most elements are metals.

Non-metallic: Character increases from left to right and bottom to top. The most active non-metal is F.

Metalloids: The switch hitters of the world, along the interface between metals and nonmetals. These include elements like B, Si, As, Ge. You may notice that the world of semiconductors revolves around the metalloids.

AQUEOUS CHEMISTRY

If there is one molecule worth giving too much attention, it is water. Water makes up 75 % of the earth's surface, 70% of the human body and 100% of a Diet Coke. It is not surprising then that so much effort is devoted to understanding the chemistry of water—almost every reaction we will examine between now and the end of the course will occur in water.

ELECTROLYTES: IONS MOVING THROUGH WATER

The chemistry of water is primarily the chemistry of dissolved ions. Water is a polar solvent with the ability to promote the dissociation of ions. One property of these ions is that they conduct electricity in water.

STRONG ELECTROLYTES

Compounds that dissociate completely in water to form ions are called strong electrolytes. Examples of these include alkali metal salts and strong acids and bases:

$NaCl \rightarrow Na^+ + Cl^-$

$KNO_3 \rightarrow K^+ + NO_3^-$

$HCl \rightarrow H^+ + Cl^-$

$NaOH \rightarrow Na^+ + OH^-$

We assume that essentially 100% of the salts or acids of strong electrolytes dissociate to form ions.

WEAK ELECTROLYTES

Some compounds dissociate only slightly, preferring to remain primarily in a non-dissociated form. These molecules are referred to as weak electrolytes and include a wide range of sparingly soluble salts as well as weak acids and bases.

$AgCl \rightarrow Ag^+ + Cl^-$
$CaCO_3 \rightarrow Ca^{++} + CO_3^=$
$CH_3COOH \rightarrow H^+ + CH_3COO^-$

NON-ELECTROLYTES

Of course, there are molecules that are too tough to have their bonds ripped apart in water. These molecules possess covalent rather than ionic bonds, such as sugar. Water may be capable of solubilizing these molecules (dissolving them), but no ions are formed and electricity is not conducted in solution.

$SUGAR_{(s)} \rightarrow SUGAR_{(aq)}$

Note that ions are not formed, but the sugar is dissolved in the water (aq means aqueous.)

ACID-BASE CHEMISTRY

What is the big deal with acids and bases? Again, it is because we are working with water, which has its own chemistry that produces ions like H^+ and OH^-, and which, as everyone knows, are what Arrhenius called acids and bases. Consequently, we will spend a lot of time looking at what happens with these two ions, the proton or hydronium ion, H^+, and the hydroxide ion, OH^-. In CH 302, we will develop many equations that attempt to determine their concentration in solution (i.e., their pH). But again, understand the context: we are working in water, and we are examining electrolytes. In other environments, H^+ and OH^- would have no significance. For example, how often do you hear people talk about the pH of gasoline?

Bronsted/Lowry Definition Acids and Bases:

Acid: capable of donating a proton

Base: capable of accepting a proton

The general equation for the dissociation of an acid is simply:

$$HA \rightarrow H^+ + A^-$$

For example:

$$HCl \rightarrow H^+ + Cl^-$$

$$CH_3COOH \rightarrow H^+ + CH_3COO^-$$

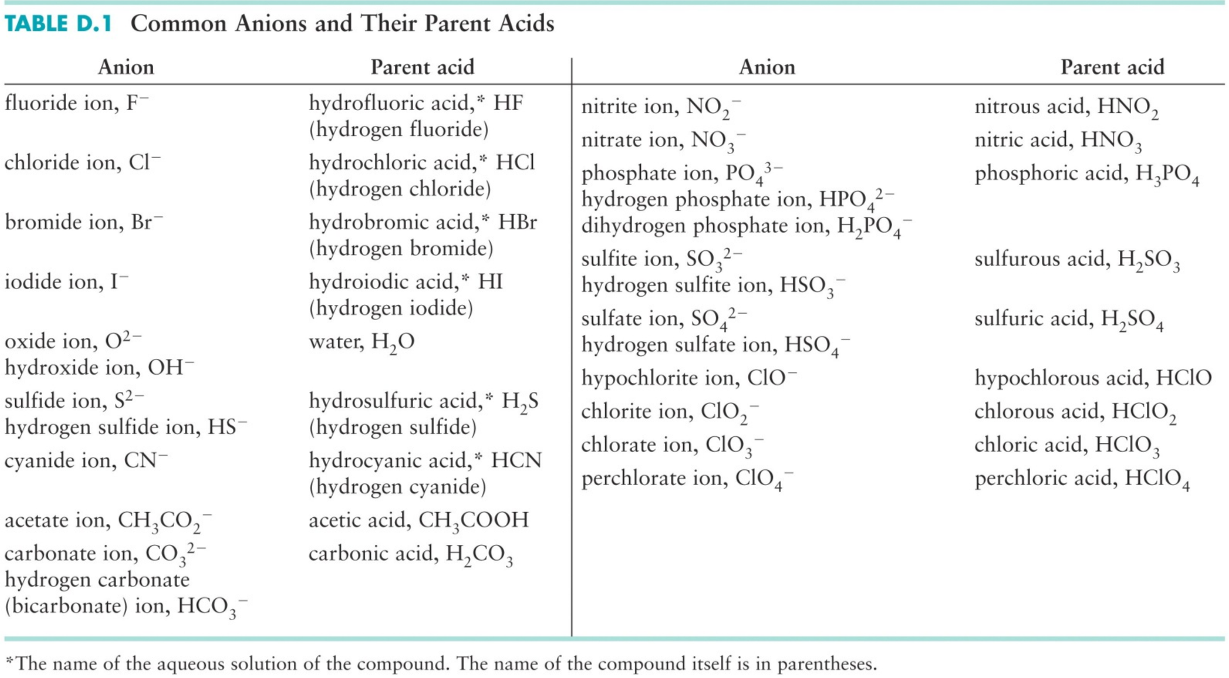


TABLE D.1 Common Anions and Their Parent Acids

Anion	Parent acid
fluoride ion, F^-	hydrofluoric acid,* HF (hydrogen fluoride)
chloride ion, Cl^-	hydrochloric acid,* HCl (hydrogen chloride)
bromide ion, Br^-	hydrobromic acid,* HBr (hydrogen bromide)
iodide ion, I^-	hydroiodic acid,* HI (hydrogen iodide)
oxide ion, O^{2-} hydroxide ion, OH^-	water, H_2O
sulfide ion, S^{2-} hydrogen sulfide ion, HS^-	hydrosulfuric acid,* H_2S (hydrogen sulfide)
cyanide ion, CN^-	hydrocyanic acid,* HCN (hydrogen cyanide)
acetate ion, $CH_3CO_2^-$	acetic acid, CH_3COOH
carbonate ion, CO_3^{2-} hydrogen carbonate (bicarbonate) ion, HCO_3^-	carbonic acid, H_2CO_3

Anion	Parent acid
nitrite ion, NO_2^-	nitrous acid, HNO_2
nitrate ion, NO_3^-	nitric acid, HNO_3
phosphate ion, PO_4^{3-} hydrogen phosphate ion, HPO_4^{2-} dihydrogen phosphate ion, $H_2PO_4^-$	phosphoric acid, H_3PO_4
sulfite ion, SO_3^{2-} hydrogen sulfite ion, HSO_3^-	sulfurous acid, H_2SO_3
sulfate ion, SO_4^{2-} hydrogen sulfate ion, HSO_4^-	sulfuric acid, H_2SO_4
hypochlorite ion, ClO^-	hypochlorous acid, HClO
chlorite ion, ClO_2^-	chlorous acid, $HClO_2$
chlorate ion, ClO_3^-	chloric acid, $HClO_3$
perchlorate ion, ClO_4^-	perchloric acid, $HClO_4$

*The name of the aqueous solution of the compound. The name of the compound itself is in parentheses.

DISTINGUISHING AND DEFINING STRONG AND WEAK ACIDS AND BASES

Now, just as there are strong electrolytes that dissociate completely and weak electrolytes that only partially dissociate, there are the equivalent strong and weak acids and bases.

- Strong acids are strong electrolytes
- Weak acids are weak electrolytes

When we start to do a lot of problems involving acid-base chemistry, it is important to be able to distinguish which compounds are strong acids and bases and which are weak acids and bases. Remember that something is defined as an acid or base in water by an ability to donate a proton or hydroxide, respectively. Whether a compound is a strong or weak acid or base has to do with the extent to which it dissociates.

ACIDS

<u>Strong acids</u>: compounds that are assumed to dissociate completely upon addition to water.

	$HA \rightarrow$	H^+	+	A^-
Initial	100 %	0%		0%
Equilibrium	0%	100 %		100%

Common Strong Acids			**Anions of These Strong Acids**	
Formula	**Name**		**Formula**	**Name**
HCl	hydrochloric acid		Cl^-	chloride ion
HBr	hydrobromic acid		Br^-	bromide ion
HI	hydroiodic acid		I^-	iodide ion
HNO_3	nitric acid		NO_3^-	nitrate ion
$HClO_4$	perchloric acid		ClO_4^-	perchlorate ion
$HClO_3$	chloric acid		ClO_3^-	chlorate ion
H_2SO_4	sulfuric acid		HSO_4^-	hydrogen sulfate ion

How do we know which compounds are strong acids or bases? We memorize the table.

Of course if an acid or base doesn't dissociate completely, it behaves as a **weak electrolyte**. Most of the acid stays in the molecular form. Only a small amount dissociates to form a proton and the anion.

	CH_3COOH	→	H^+	+	CH_3COO^-
Initial	100 %		0%		0%
Equilibrium	99%		1 %		1%

How do we know which acids are weak? Simple. It is any acid that isn't one of the seven strong acids. Several examples are shown in the table above.

BASES

We can describe an analogous collection of strong and weak bases that produce OH^- in aqueous solution. The strong bases include the Group IA and IIA metal hydroxides. For example:

	$NaOH$	→	Na^+	+	OH^-
Initial	100 %		0%		0%
Equilibrium	0%		100 %		100%

Table of Strong Bases

$LiOH$	Lithium hydroxide	$CsOH$	Cesium hydroxide
$NaOH$	Sodium hydroxide	$Ca(OH)_2$	Calcium hydroxide
KOH	Potassium hydroxide	$Sr(OH)_2$	Strontium hydroxide
$RbOH$	Rubidium hydroxide	$Ba(OH)_2$	Barium hydroxide

Weak bases: an electrolyte that does not dissociate significantly in solution. The most famous weak base is ammonia, NH_3:

	NH_3	→	NH_4^+	+	OH^-
Initial	100 %		0%		0%
Equilibrium	99%		1 %		1%

SOLUBILITY

Throw a compound into water. Does it dissolve or does it sink to the bottom of the flask? Whether this happens or not determines the **solubility** of a compound. A soluble compound **dissolves**. An **insoluble** compound usually sinks to the bottom of the flask as a **precipitate**, although it may form a **suspension** in solution. For the most part, we are interested in what ionic compounds do in water. How do we know what these ionic materials do in solution? We learn rules about solubility—see the next page.

SOLUBILITY OF COMMON IONIC COMPOUNDS IN WATER

There are two ways to learn the solubility rules.

- Memorize the individual solubility rules for each kind of cation and anion. The two tables below list the ions that are generally soluble, and the ions that are generally insoluble.

- If the thought of memorizing a lengthy table makes you sad, see if you can spot some generalizations that make it easy to slide by with some simple rules that account for most of the ion solubilities.

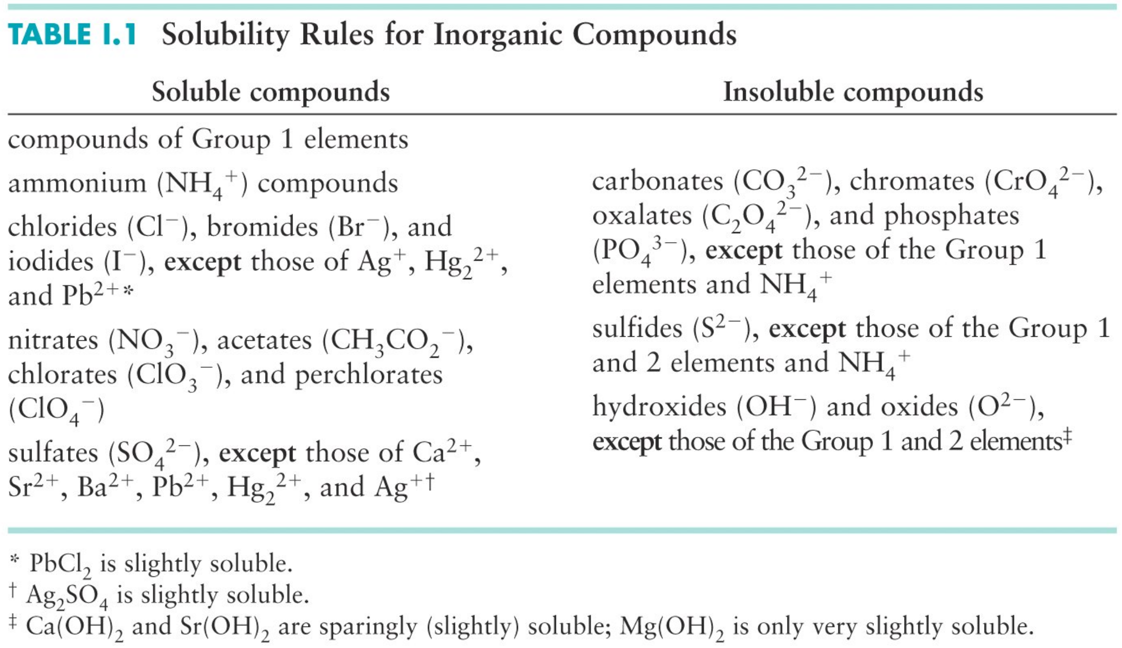


TABLE I.1 Solubility Rules for Inorganic Compounds

Soluble compounds	Insoluble compounds
compounds of Group 1 elements	
ammonium (NH_4^+) compounds	carbonates (CO_3^{2-}), chromates (CrO_4^{2-}), oxalates ($C_2O_4^{2-}$), and phosphates (PO_4^{3-}), **except** those of the Group 1 elements and NH_4^+
chlorides (Cl^-), bromides (Br^-), and iodides (I^-), **except** those of Ag^+, Hg_2^{2+}, and Pb^{2+}*	sulfides (S^{2-}), **except** those of the Group 1 and 2 elements and NH_4^+
nitrates (NO_3^-), acetates ($CH_3CO_2^-$), chlorates (ClO_3^-), and perchlorates (ClO_4^-)	hydroxides (OH^-) and oxides (O^{2-}), **except** those of the Group 1 and 2 elements‡
sulfates (SO_4^{2-}), **except** those of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Hg_2^{2+}, and Ag^+†	

* $PbCl_2$ is slightly soluble.

† Ag_2SO_4 is slightly soluble.

‡ $Ca(OH)_2$ and $Sr(OH)_2$ are sparingly (slightly) soluble; $Mg(OH)_2$ is only very slightly soluble.

There has to be an easier way!

Some cation solubility rules:

- Rule 1: All of the alkali metals (K^+, Na^+. etc.) ions plus NH_4^+ are always soluble.
- Rule 2: Heavy cations (down the periodic table) like Ba^{++} and Ag^+ and Pb^{++} are insoluble.

Some anion solubility rules:

- Rule 3: Anions that are always insoluble except for Rule 1 are mostly the multiply charged species ($S^=$, $CO_3^=$, PO_4^{-3}, $O^=$, etc.)
- Rule 4: The anions that are primarily soluble are the conjugate bases of strong acids like Cl^- from HCl, Br^- from HBr, NO_3^- from HNO_3, ClO_4^- from $HClO_4$ etc.

These rules aren't perfect but they are good enough to get an A in CH 301 without doing a lot of memorizing.

THE BEST SOLUBILITY RULE OF ALL

But the biggest trend of all, the rule of thumb if you only want to remember one thing, is this:

The singly charged ions are usually soluble.
and
The multiply charged ions are usually insoluble.

The reason for this is something we will learn in CH 302.

Solubility example: We can use this information to decide the types of reactions that occur when we place compounds in solution. Consider the reaction of $CaNO_3$ and KCO_3.

Note from above that nitrates and alkali metals are soluble. So when these compounds are added to solution, four ions are formed:

Ca^{++} NO_3^- K^+ $CO_3^=$

Note from above that carbonates are generally insoluble, so it is expected that the following precipitation reaction occurs:

$$Ca^{++}(aq) + CO_3^=(aq) \rightarrow CaCO_3(s)$$

OXIDATION NUMBERS

Reactions in which substances undergo changes in oxidation number are called oxidation-reduction reactions or redox reactions. Redox reactions involve the transfer of electrons. Some definitions:

- **Oxidation:** an algebraic increase in oxidation number in which electrons are lost from a compound.
- **Reduction:** an algebraic decrease in oxidation number in which electrons are gained from a compound.
- **Oxidizing agent:** substance that gains electrons and oxidizes other substances by being reduced.
- **Reducing agent:** substance that loses electrons and reduces other substances by being oxidized.

OXIDATION RULES

Assignment of the oxidation number for a compound follows certain rules:

1. The oxidation of a free element is zero.
   Example: Na is 0
2. The oxidation number of an element in a monatomic ion is the charge on the ion.
   Example: Na^+ is +1
3. The charge of a monatomic ion is usually determined by its group (column on periodic table). See table below.
4. In the formula for any compound, the sum of the oxidation numbers of all elements in the compound is zero.
   Example: NaCl is 0
5. In a polyatomic ion, the sum of the oxidation numbers of the constituent elements is equal to the charge in the ion.
   Example: NO_3^- is -1

Some Typical Oxidation Numbers

Element	**Oxidation Number**
H	+1 (but sometimes –1)
O	-2 (but sometimes –1 or –1/2)
Group IA	+1 always
Group IIA	+2
Group IIIA	+3
Group VIIA in reaction with metals	-1
Group VIA in reaction with metals	-2
Group VA in reaction with metals	-3

DISPLACEMENT REACTIONS

The world is filled with two kinds of metals: those that dissolve in your pant pockets, and those that don't.

Coinage metals—like gold, silver, and copper—don't dissolve in your pants or when you bury them in the ground. This is why they are so valuable.

But valuable also means boring. If you want real fun, you hang out with active metals—like potassium, magnesium, and lithium—that explode at the slightest provocation (like placing them in water).

So what does happen when you throw a metal in water?

$$M + H_2O \rightarrow M^+ + H_2 + OH^-$$

- The metal is oxidized (becomes a cation).
- The water is reduced making hydrogen (this is why Na explodes in water).
- The water turns basic (you can see this with an indicator turning color).

Wow, what a famous reaction!

There is a special kind of reaction in which one element displaces another element from a compound. For example:

$$Cu(II)SO_4 + Zn \rightarrow ZnSO_4 + Cu$$

In general, active metals displace less active metals and hydrogen. There are three categories to consider:

- Reactions that displace hydrogen from non-oxidizing acids.
- Reactions involving metals that displace hydrogen from steam.
- Reaction of metals that displace hydrogen from cold water.

These results are tabulated on the next page.

Type I Elements		**Common Reduced Form**	**Common Oxidized Forms**
Li		Li	Li^{+}
K	Displace hydrogen from cold water	K	K^{+}
Ca		Ca	Ca^{2+}
Na		Na	Na^{+}
Type II Elements			
Mg		Mg	Mg^{2+}
Al		Al	Al^{3+}
Mn	Displace hydrogen from steam	Mn	Mn^{2+}
Zn		Zn	Zn^{2+}
Cr		Cr	Cr^{3+}, Cr^{6+}
Fe		Fe	Fe^{2+}, Fe^{3+}
Type III Elements			
Cd		Cd	Cd^{2+}
Co	Displace hydrogen from non-oxidizing acids	Co	Co^{2+}
Ni		Ni	Ni^{2+}
Sn		Sn	Sn^{2+}, Sn^{4+}
Pb		Pb	Pb^{2+}, Pb^{4+}
Type IV Elements			
Cu		Cu	Cu^{+}, Cu^{2+}
Hg	Don’t react	Hg	Hg_2^{2+}, Hg^{2+}
Ag		Ag	Ag^{+}
Pt		Pt	Pt^{2+}, Pt^{4+}

In reading the **metal activity table**, the metals at the top of the chart displace salts of the less active metals below them. An easy way to think of this: a reaction occurs if the more stable metal (the one lower on the chart) is a reaction product.

Again, the general rule to remember is: the farther to the right, the more reactive the metal.

- **Type I**: The alkali metals (column 1)explode in cold water.
- **Type II**: The alkali earths (column 2) dissolve in hot water.
- **Type III**: The transition metals dissolve in acid.
- **Type IV**: The coinage metals don’t do anything.

FOREIGN LANGUAGE CLASS: CHEMISTRY 506

All of you must have taken a foreign language at some time. This means you've had practice at memorizing large collections of random letters that make up the names we assign to different types of chemical compounds. We are specifically interested in procedures for naming **binary** compounds—formed by combining two elements—and **ternary** compounds—formed by putting together three elements. To assist us in our endeavor, it will be of great assistance to memorize the now expanded list of common cations and anions shown in the Table below. Try putting these on flash cards as well.

NAMING BINARY COMPOUNDS

Let's look at rules for **binary ionic compounds**. In putting together two elements as ions, the following rules apply:

Rule 1: The element that is more metallic (more electronegative) is listed first.
Example: Sodium is always listed first because it is always +1.

Rule 2: If the element has multiple common oxidation states, that oxidation number is included in parenthesis after the element.
Example: Copper is either Cu(I) if it is +1 oxidation state; or, Cu(II) if it is +2 oxidation state.

Rule 3: The second element listed has less metallic character, and it's named by adding the suffix "-ide" to the element's stem.

Naming the Nonmetals

Element	Symbol	Stem	Name
Boron	B	bor	boride
Carbon	C	carb	carbide
Silicon	Si	silic	silicide
Nitrogen	N	nitr	nitride
Phosphorous	P	phosph	phosphide
Arsenic	As	arsen	arsenide
Antimony	Sb	antimon	antimonide
Oxygen	O	ox	oxide
Sulfur	S	sul	sulfide
Selenium	Se	selen	selenide
Hydrogen	H	hydr	hydride
Fluorine	F	fluor	fluoride
Chlorine	Cl	chlor	chloride
Bromine	Br	brom	bromide
Iodine	I	iod	iodide

Examples of Binary Compound Names

Compound	Name
LiBr	lithium bromide
$MgCl_2$	magnesium chloride
Li_2S	lithium sulfide
Al_2O_3	aluminum oxide
Na_3P	sodium phosphide
Mg_3N_2	magnesium nitride

MULTIPLE OXIDATION NUMBERS

If only life were this simple. Recall that not all elements have a single common oxidation number in ionic form. The idea of using a Roman numeral after the name was proposed to name these compounds. However, there is an older way of naming these ambiguous compounds using -ous and -ic to terminate the metal. The -ic stood for the higher oxidation state, and -ous was attached to the metal with the lower oxidation number.

Example: Copper (I) is also called cuprous.
Example: Copper (II) is also called cupric.
Examples:

Compound	Older Name	Modern Name
$FeBr_2$	ferrous bromide	iron (II) bromide
$FeBr_3$	ferric bromide	iron (III) bromide
SnO	stannic oxide	tin (II) oxide
SnO_2	stannous oxide	tin (I) oxide
$TiCl_2$		titanium (II) chloride
$TiCl_3$		titanium (III) chloride
$TiCl_4$		titanium (IV) chloride

PSEUDO BINARY COMPOUNDS NAMES (POLYATOMIC IONS)

But what about all the compounds containing polyatomic anions and cations? Ions like hydroxide (OH^-), ammonium (NH_4^+), and sulfate (SO_4^{2-}). Well, these ions behave as simple monatomic anions and cations, so we include them in our naming of binary ionic species using the same rules as above, except that we have to memorize all of the polyatomic ion exceptions. There are only 50 or 100 listed on the next page for your memorization pleasure.

Common Cations

Li^+	1+	lithium ion
Na^+	1+	sodium ion
K^+	1+	potassium ion
NH_4^+	1+	ammonium ion
Ag^+	1+	silver ion
Co^{2+}	2+	cobalt (II) ion
Mg^{2+}	2+	magnesium ion
Ca^{2+}	2+	calcium ion
Ba^{2+}	2+	barium ion
Cd^{2+}	2+	cadmium ion
Zn^{2+}	2+	zinc ion
Cu^{2+}	2+	copper (II) ion
Hg_2^{2+}	2+	mercury (I) ion
Hg^{2+}	2+	mercury (II) ion
Mn^{2+}	2+	manganese (II) ion
Ni^{2+}	2+	nickel (II) ion
Pb^{2+}	2+	lead(II) ion
Sn^{2+}	2+	tin (II) ion
Fe^{2+}	2+	iron (II) ion
Fe^{3+}	3+	iron (III) ion
Al^{3+}	3+	aluminum ion
Cr^{3+}	3+	chromium (III) ion

Common Anions

F^-	1-	fluoride ion
Cl^-	1-	chloride ion
Br^-	1-	bromide ion
I^-	1-	iodide ion
OH^-	1-	hydroxide ion
CN^-	1-	cyanide ion
ClO^-	1-	hypochlorite
ClO_2^-	1-	chlorite ion
ClO_3^-	1-	chlorate ion
ClO_4^-	1-	perchlorate
CH_3COO^-	1-	acetate ion
MnO_4^-	1-	permanganate
NO_2^-	1-	nitrite ion
NO_3^-	1-	nitrate ion
SCN^-	1-	thiocyanate
O^{2-}	2-	oxide ion
S^{2-}	2-	sulfide ion
HSO_3^-	1-	bisulfite ion
SO_3^{2-}	2-	sulfite ion
HSO_4^-	1-	bisulfate ion
SO_4^{2-}	2-	sulfate ion
PO_4^{3-}	3-	phosphate ion
HCO_3^-	1-	bicarbonate
CO_3^{2-}	2-	carbonate ion
CrO_4^{2-}	2-	chromate ion
$Cr_2O_7^{2-}$	2-	dichromate ion

Examples of Compounds Made with Pseudobinary Ion

Compound	Name
Na_2SO_4	sodium sulfate
$Cu(NO_3)_2$	copper (II) nitrate
LiCN	lithium cyanide
NH_4Cl	ammonium chloride

So what if it isn't an ion?…

NAMING BINARY MOLECULES (COVALENT BONDS)

Often two non-metals are put together. In these cases neither has a great desire to rip away or give up electrons, so ionic species aren't formed. Moreover, these types of compounds may exist in a variety of oxidation states. Rather than deal with oxidation state, the number of each atom is listed with a prefix noting the number of atoms of each element.

# Atoms	2	3	4	5	6	7	8
Prefix	di-	tri-	tetra-	penta-	hexa-	hepta-	octa-

Examples of Molecules

Compound	Name
CO_2	carbon dioxide
SO_3	sulfur trioxide
OF_2	oxygen difluoride
P_4O_6	tetraphosphorus hexaoxide
Cl_2O_7	dichorine heptaoxide

TERNARY COMPOUNDS (AFTER TWO COMES THREE)

Ternary compounds contain combinations of atoms from three different elements. Things can get really complicated with the additional element, so we will stick to just a few simple classes.

First, in working with ionic compounds there were a variety of polyatomic ions that you were asked to memorize. These followed the basic rules of binary ions and are easy to name IF you have them all memorized.

TERNARY ACIDS AND THEIR SALTS

There are a collection of well know ternary acids which consist of hydrogen, oxygen and a nonmetal. The most famous example is H_2SO_4, sulfuric acid.

RULES FOR NAMING TERNARY ACIDS

There are some basic rules for naming such compounds which depends upon the oxidation state of the central atom, XXX :

	Ternary Acid	**Anion Salt**	
Decreasing Oxidation Number of X	perXXXic acid	perXXXate	Decreasing Oxygens on X
	XXXic	XXXate	
	XXXous acid	XXXite	
	hypoXXXous acid	hypoXXXite	

Examples of Ternary Acids

Acid	Name	Sodium Salt	Name
HNO_2	nitrous acid	$NaNO_2$	sodium nitrite
HNO_3	nitric acid	$NaNO_3$	sodium nitrate
H_2SO_3	sulfurous acid	Na_2SO_3	sodium sulfite
H_2SO_4	sulfuric acid	Na_2SO_4	sodium sulfate
$HClO$	hypochloric acid	$NaClO$	sodium hypochlorite
$HClO_2$	chlorous acid	$NaClO_2$	sodium chlorite
$HClO_3$	chloric acid	$NaClO_3$	sodium chlorate
$HClO_4$	perchloric acid	$NaClO_4$	sodium perchlorate

I have never done anything so boring in my life as type this particular set of notes.

The end.

ORGANIC CHEMISTRY: AN INTRODUCTION

I am not the best person to be giving a lecture on organic chemistry—the world is filled with two kinds of people, the ones who despise organic, and the ones who love it. During the year I took organic in college, I was also plastering the words from sad songs all over my dorm room walls—looking back I realize there was a cause and effect relationship between organic and typing up those songs. Honestly, the lunch each Tuesday afternoon before organic lab made me feel pretty much like a person going before a firing squad, only it happened over and over and over for an entire year.

With that off my chest, a nice lecture that surveys the important foundational concepts for what to expect in organic chemistry. Mostly you will get a feel for the fact that organic is a lot of nomenclature and a lot of systematic rules. For those who like languages and lists, and for those who can close their eyes and imagine a three-dimensional world of molecular structures, organic is quite easy. For those, like me, who would rather reduce the subject to a couple of simple mathematical formulas, it is quite difficult. I mean, what kind of science course doesn't need a calculator?

But enough rant.

NOMENCLATURE

Why does it take a year to do organic—the study of a single atom, carbon? The reason is that pesky carbon atom is remarkably adept at forming different kinds of molecules. For example, if I were to draw all the legal ways to put together 12 C atoms and 26 H atoms (to satisfy the Noble gas configuration), there are 355 correct structures, each with its own unique properties. And then, if we let the C react with all the compounds that form covalent bonds—like N, O, Cl, P, S, etc.—the possibilities explode exponentially. And organic chemists want you to be able to draw and name every one of those compounds and then explain why one of them evaporates just a little faster, or reacts just a little slower, than another.

So I guess they can have their year. But it is still just stuff you learned in my class.

Alkanes: The basic organic compound is an alkane—it has only C and H bonds, and every one of those bonds is a sigma bond. Alkanes are not very exciting because they don't do much reacting. But they are a great place to start learning nomenclature.

To begin, every organic molecule has three parts defined by a **very** systematic set of rules: a prefix, a parent, and a suffix.

- The parent is the length of the longest carbon chain—like **methane** for one carbon, **ethane** for two carbons, and **heptane** for seven carbons.
- The prefix contains any substituents on the molecule besides the parent C chain—like an extra methyl or hydroxyl group
- The suffix is the family in which the compound falls based upon the functional group attached to the parent chain—for example, the suffix **–ol** indicates that ethan**ol** is an alcoh**ol**.

For the first molecules we are studying, the alk**ane** family has a suffix **-ane** to indicate that nothing more exciting than a bunch of hydrocarbon sigma bonds are found in the molecule.

So whether you have pentane, decane, 2-methylpropane, or 2,4 dimethylheptane, the **-ane** suffix indicates you have an alkane.

So what are all the possible parent names? Not surprisingly, they kind of make sense if you know your Greek, and are so consistent that you pick them up quickly even if you don't speak the language.

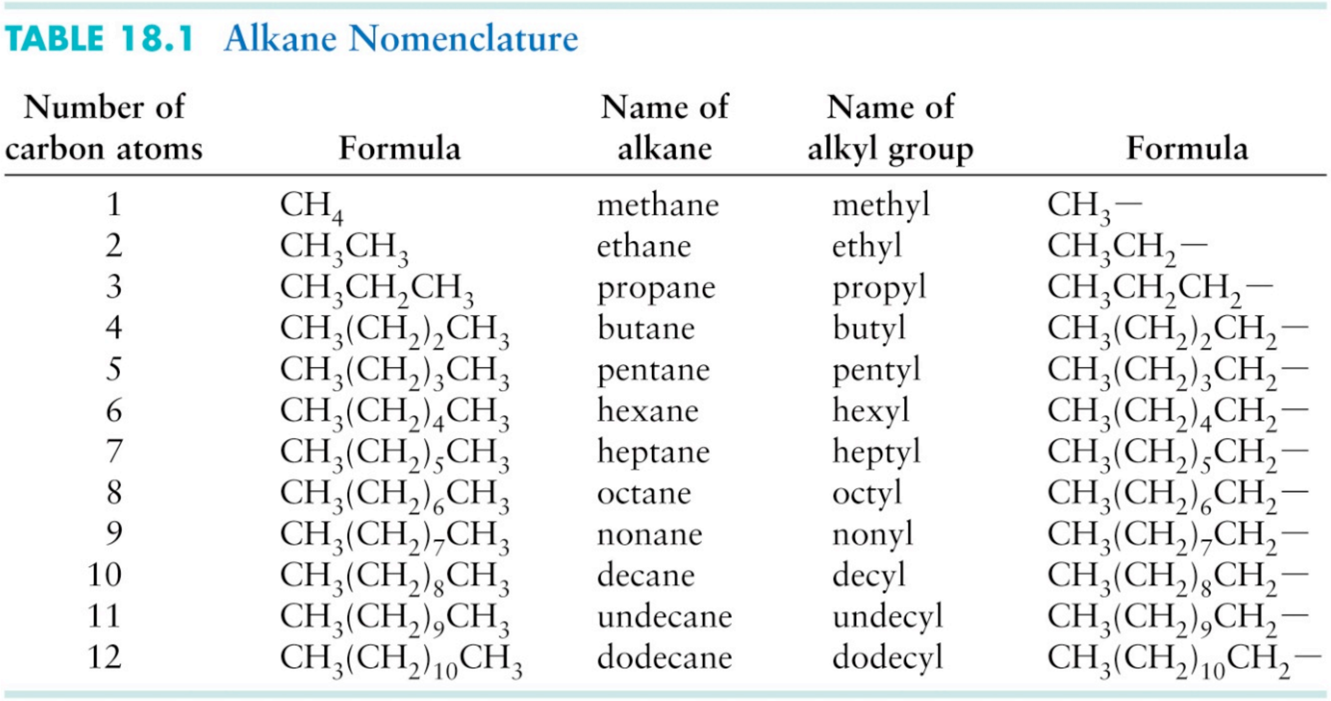
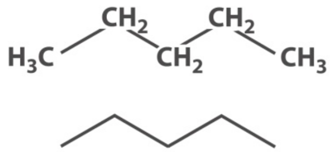


TABLE 18.1 Alkane Nomenclature

Number of carbon atoms	Formula	Name of alkane	Name of alkyl group	Formula
1	CH_4	methane	methyl	CH_3-
2	CH_3CH_3	ethane	ethyl	CH_3CH_2-
3	$CH_3CH_2CH_3$	propane	propyl	$CH_3CH_2CH_2-$
4	$CH_3(CH_2)_2CH_3$	butane	butyl	$CH_3(CH_2)_2CH_2-$
5	$CH_3(CH_2)_3CH_3$	pentane	pentyl	$CH_3(CH_2)_3CH_2-$
6	$CH_3(CH_2)_4CH_3$	hexane	hexyl	$CH_3(CH_2)_4CH_2-$
7	$CH_3(CH_2)_5CH_3$	heptane	heptyl	$CH_3(CH_2)_5CH_2-$
8	$CH_3(CH_2)_6CH_3$	octane	octyl	$CH_3(CH_2)_6CH_2-$
9	$CH_3(CH_2)_7CH_3$	nonane	nonyl	$CH_3(CH_2)_7CH_2-$
10	$CH_3(CH_2)_8CH_3$	decane	decyl	$CH_3(CH_2)_8CH_2-$
11	$CH_3(CH_2)_9CH_3$	undecane	undecyl	$CH_3(CH_2)_9CH_2-$
12	$CH_3(CH_2)_{10}CH_3$	dodecane	dodecyl	$CH_3(CH_2)_{10}CH_2-$

DRAWING AN ORGANIC MOLECULE IN TWO-DIMENSIONS

So a simple five carbon alkane is called pentane and has the following structure:



2 **Pentane, C_5H_{12}**

Note that in trying to tell people about the molecule, you can give the compound a formal name, like pentane, or you can draw a simple structure in which the bonds between the carbons are shown adjacent to the CH_2 and CH_3 groups. Or, if you are a real organic chemist, you don't waste your time ever drawing C and H, since that is 99% of what you have in your molecules anyway, and instead you just draw the bonds (so zig-zag sticks above), and let everyone else make a mental note that at the vertices are carbons and that every carbon is fitted out with enough hydrogens to satisfy the octet rule.

By the way, if you wanted, you could turn pentane into a nice three dimensional picture just like the ethanol above, that is, if you wanted to, and after awhile, you just automatically imagine that 3-dimensional picture in your head and don't bother with any more than squiggly lines.

STRUCTURAL ISOMERS

Recall that I told you that there were 355 ways you could write forms of $C_{12}H_{26}$? Well now you see how this can be done. A **structural isomer** is a compound that has the same molecular formula, but a different structure. While smaller chain hydrocarbons don't have many structural isomers (for example, butane with four carbons has just two structural isomers), by the time you reach the C_{32} alkanes, there are about 3×10^7 structural isomers (trust me on this.)

As we near the end of the section on alkanes, a question, can you state a general formula for hydrocarbons of length n? This is as close as we get to math:

Answer: C_nH_{2n+2} – take a look at the table and you will see it is so.

Now some chemistry of alkanes. Basically, alkanes burn, and cleanly, I might add, following the combustion equation below:

$$2C_nH_{2n+2} + (3n)O_2 \rightarrow (2(n+2))H_2O + 2nCO_2$$

In the table below is a who's who of fuels which are basically hydrocarbons, mostly alkanes, all of them fractional cuts of petroleum, separated by boiling point range as the number of carbons increases. And as you can see, even though alkanes are the most boring of organic molecules, they are the only ones over which people fight wars.

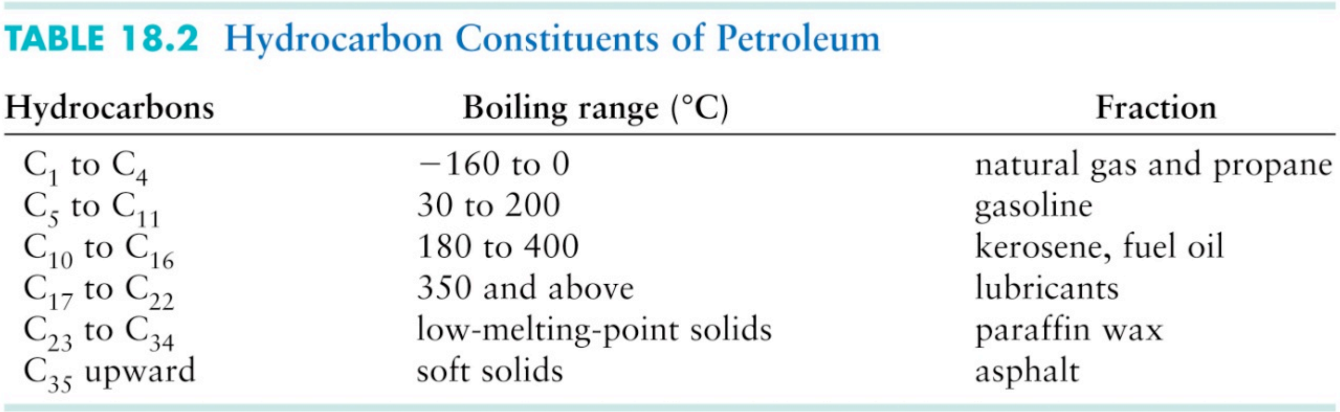


TABLE 18.2 Hydrocarbon Constituents of Petroleum

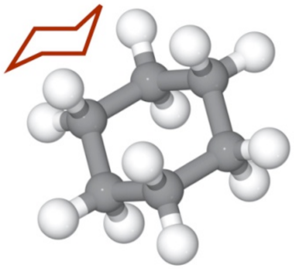
Hydrocarbons	Boiling range (°C)	Fraction
C_1 to C_4	−160 to 0	natural gas and propane
C_5 to C_{11}	30 to 200	gasoline
C_{10} to C_{16}	180 to 400	kerosene, fuel oil
C_{17} to C_{22}	350 and above	lubricants
C_{23} to C_{34}	low-melting-point solids	paraffin wax
C_{35} upward	soft solids	asphalt

LET’S GET CYCLIC

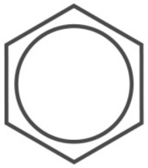
A bit more nomenclature involving hydrocarbons. It is possible for a molecule to form a loop on itself. These are two kinds of cyclic compounds.

Cyclic hydrocarbons are known as cyclic compounds, and we designate them with the prefix, **cyclo—.**

Benzene and its aromatic friends. The really important cyclic species are aromatic compounds—those based on the ring structure shown below:

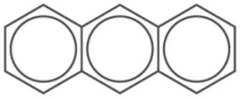


16 Cyclohexane, C_6H_{12}

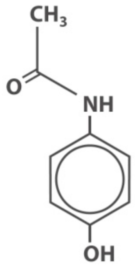


31 Benzene, C_6H_6

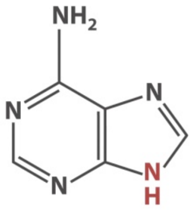
As we learned while discussing molecular orbital theory in CH301, aromatic compounds are hexagonal ring structures formed from the overlap of C_{sp2}-C_{sp2} that form σ bonds. What make aromatic compounds are the leftover 6 electrons that form a sea of delocalized π electrons positioned above and below the ring. It is this p-bonding structure that creates a unique stability and reactivity for aromatic species. While we won’t spend any more time on these structures, it is enough to know that aromatic species are ubiquitous in nature, whether in petrochemicals like anthracene or drugs like Tylenol or biological-active molecules like adenine.



33 Anthracene, $C_{14}H_{10}$



14 Tylenol

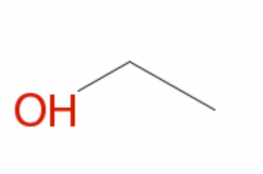


22 Adenine

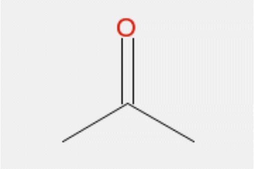
OTHER FUNCTIONAL GROUPS IN ORGANIC MOLECULES

If all of our compounds had only hydrogens and carbons, organic chemistry would be easy and life would be boring. So there are hundreds of functional groups that can be added. I want you to know just four of them because they pop up all the time:

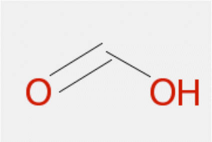
- **Alcohols**: Ethane plus **–ol** (an alcohol) make ethanol, a 2 carbon alcohol



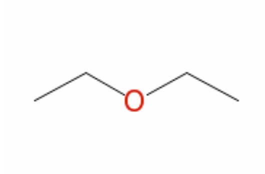
- **Ketones**: Ethane plus **–one** (a ketone) make acetone, a 2 carbon ketone



- **Carboxylic acids**: Methane plus **–oic acid** (a carboxylic acid) makes metanoic acid or formic acid, a 1 carbon acid that is the stuff in ant bites



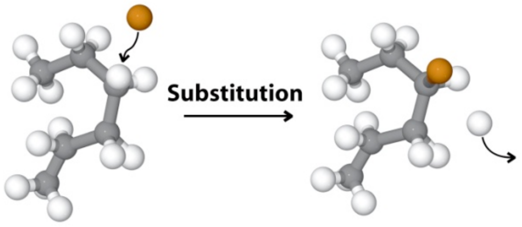
- **Ethers**: An ethane plus an ethane with an O in the middle is an ether, specifically diethyl ether, the stuff bad guys put in handkerchiefs to kidnap you.

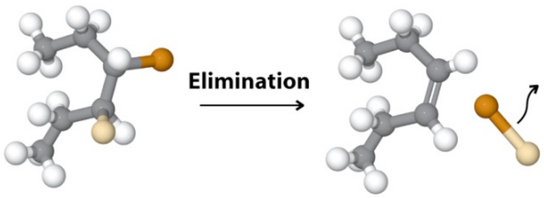


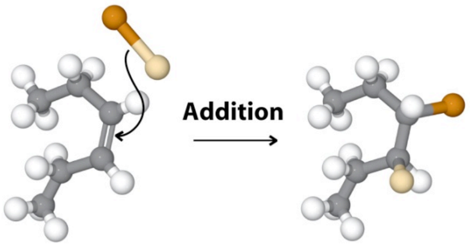
ORGANIC CHEMISTRY REACTIVITY

What makes organic such a challenge isn't simply being able to assign names to compounds, but it is being able to predict the reactions that this incredibly broad array of compounds undergo. To say an understanding of organic reactivity is beyond the scope of this course, is about as understated a comment as one can make. To understand organic reactivity is to understand the chemistry of not just these relatively small molecules, but also the fields of polymer chemistry and biochemistry—you know, pretty much everything.

But we can create three classes of general reaction mechanisms. They are listed below in some nice pictures from the text that suggest why the names are given to these reactions.







And that is more than enough to know about organic.