Chem 170

Stoichiometric Calculations

Module Two

Atoms, Molecules and Moles

Introduction to Module Two

In the early fifth century BC, the Greek philosopher Democritus proposed that matter consists of small indivisible particles that he named *atomos* (meaning uncuttable). Democritus' vision didn't gain much favor with his contemporaries and almost 2300 years passed before John Dalton reintroduced the idea of atoms in the early 1800s. Although Democritus' conception of *atomos* doesn't reflect our modern understanding of atoms, the idea that matter is composed of indivisible particles remains a simple but powerful idea.

At present, we have identified 113 elements[†]. Most of these elements occur naturally, although fewer than 20 make up more than 99.7% of the earth's crust or your body. Other elements, such as Einsteinium, have been synthesized only in the laboratory. Just as the 26 letters of our alphabet provide an abundance of words (with new words being coined everyday), the elements combine to make the many chemical compounds that surround you (and new chemical compounds are synthesized every day). The paper in your hand, the ink used to print these letters, and your retina, which helps you see the paper and ink, all consist of chemicals made from the elements.

How elements combine together to make chemical compounds is the focus of other courses in our curriculum. The main goal of this course is to learn how to handle the quantitative relationships between elements and chemical compounds, and between different chemical compounds. Several important concepts will help us achieve this goal, including atoms, molecules, and compounds, the mole, and molar mass; these are the subjects of this module.

Objectives for Module Two

In completing this module you will master the following objectives:

- to identify an element's atomic number and mass number
- to understand the similarities and differences between an element's isotopes
- to determine an element's average atomic mass and molar mass
- to understand the difference between a molecular formula and an empirical formula
- to calculate a molecule's or compound's molar mass
- to convert between grams, moles, and numbers of atoms or molecules

[†] This number is correct as of 5/24/2005, but changes as new elements are synthesized. Interestingly, as this module was first written in June 2001 the number of elements was 115; however, scientists at the Lawrence Berkeley National Labs had to retract their 1999 discovery of two elements.

Dalton's Atomic Theory

We trace the modern era of chemistry to John Dalton's development of atomic theory, which consists of three essential hypotheses:

1. Elements, which are the smallest division of matter with distinct chemical properties, are composed of atoms. All atoms of a given element are identical (*this is not strictly true, as we will see, but we won't hold that against Dalton)* and different from the atoms of other elements. The element carbon is made of carbon atoms, which are different from the atoms of oxygen that make up elemental oxygen.

2. Compounds are composed of atoms from two or more elements. Because atoms cannot be subdivided, the elements making up a compound are always present in ratios of whole numbers. A compound containing carbon and oxygen, for example, can have 1 carbon atom and 2 oxygen atoms, but cannot have 1.5 carbon atoms.

3. In a chemical reaction the elements making up compounds rearrange to make new compounds. The atoms making up these compounds, however, are not destroyed, nor are new atoms created.

Dalton's first hypothesis simply recognizes the atom as the basic building block of chemistry. Water, for example, is made from atoms of hydrogen and oxygen. The second hypothesis recognizes that for every compound there is a fixed combination of atoms. Regardless of its source (rain, tears, Big Walnut Creek, or a bottle of Evian) water always consists of two hydrogen atoms for every atom of oxygen. Dalton's third hypothesis is a statement that atoms are conserved in a reaction; this is more commonly known as the conservation of mass.

The Structure of the Atom

Although Dalton believed atoms to be indivisible, we know now that they are made from three smaller subatomic particles – the electron, the proton, and the neutron. The atom, however, remains the smallest division of matter with distinct chemical properties.

Electrons, Protons, and Neutrons. The characteristic properties of electrons, protons, and neutrons, are shown in Table 1.

Table 1. Mass and Charge of Subatomic Particles

		Charge	
Particle	Mass (g)	Coulomb	Charge Unit
electron	9.10939×10^{-28}	-1.6022×10^{-19}	-1
proton	1.67262×10^{-24}	$+1.6022\times 10^{-19}$	+1
neutron	1.67493×10^{-24}	0	0

The proton and neutron make up the atom's nucleus, which is located at the center of the atom and has a radius of approximately 5×10^{-3} pm. The remainder of the atom, which has radius of approximately 100 pm, is mostly empty space in which the electrons exist. Of the three subatomic particles, only the electron and proton carry a charge. Because elements have no net charge (that is, they are neutral), the number of electrons and protons in an element must be the same.

Atomic Numbers. Why is an atom of carbon different from an atom of hydrogen or helium? One possible explanation is that carbon and hydrogen have different numbers of electrons, protons, or neutrons. Table 2 provides the relevant numbers.

Table 2. Comparison of the Elements Hydrogen and Carbon

Element	Number of Protons	Number of Neutrons†	Number of Electrons
hydrogen	1	0, 1, or 2	1
helium	2	2	2
carbon	6	6, 7, or 8	6

Note that atoms of hydrogen and carbon each have three possibilities for the numbers of neutrons, and that it is even possible for a hydrogen atom to exist without a neutron. Clearly the number of neutrons is not crucial to determining if an atom is carbon, hydrogen, or helium. Although hydrogen, helium, and carbon have different numbers of electrons, the number is not critical to an element's identity. For example, it is possible to strip an electron away from helium forming a helium ion‡ with a charge of +1 that has the same number of electrons as hydrogen.

What makes an atom carbon is the presence of 6 protons, whereas every atom of hydrogen has 1 proton and every atom of helium has 2 protons. The number of protons in an atom is called its atomic number (Z).

Atomic Mass and Isotopes. Protons and neutrons are of similar mass and much heavier than electrons (see Table 1); thus, most of an atom's mass is in its nucleus. Because not all of an element's atoms necessarily have the same number of neutrons, it is possible for two atoms of an element to differ in mass. For this reason, the sum of an atom's protons and neutrons is known as its mass number (A). Carbon, for example, can have mass number of 12, 13, or 14 (six protons and six, seven, or eight neutrons), and hydrogen can have mass numbers of 1, 2, and 3 (one proton and zero, one, or two neutrons).

Atoms of the same element (same Z), but different numbers of neutrons (different A) are called isotopes. Hydrogen, for example has three isotopes (see Table 2). The isotope with 0 neutrons is the most abundant, accounting for 99.985% of all stable hydrogen atoms, and is known, somewhat self-referentially, as hydrogen. Deuterium, which accounts for 0.015% of all stable hydrogen atoms, has 1 neutron. The isotope of

† Only the most important examples are included in this table.
‡ An ion with a positive charge forms when electrons are removed and is called a cation. Anions, which have negative charges, form when electrons are added.

hydrogen with 2 neutrons is called tritium. Because tritium is radioactive it is unstable and disappears with time.

The usual way to represent isotopes is with the symbol

$${}^{A}_{Z}\mathrm{X}$$

where X is the atomic symbol for the element. The three isotopes of hydrogen, which has an elemental symbol of H, are

$${}^{1}_{1}\mathrm{H} \qquad {}^{2}_{1}\mathrm{H} \qquad {}^{3}_{1}\mathrm{H}$$

Because the elemental symbol (X) and the atomic number (*Z*) provide redundant information, we often omit the atomic number; thus, deuterium becomes

$${}^{2}\mathrm{H}$$

Unlike hydrogen, the isotopes of other elements do not have specific names. Instead they are named by taking the element's name and appending the atomic mass. For example, the isotopes of carbon are called carbon-12, carbon-13, and carbon-14.

Example 1. Uranium has two important isotopes. One of these, uranium-235, is fissionable and is used as a fuel in nuclear reactors; its symbol is

$${}^{235}_{92}\mathrm{U}$$

The other important isotope of uranium, which has 146 neutrons, does not undergo fission and cannot be used as a nuclear fuel. How many protons, neutrons, and electrons are there in uranium-235? What is the name and symbol for the non-fissionable isotope of uranium?

Solution. Because uranium's atomic number is 92, uranium-235 has 92 protons. The number of neutrons is A - Z,; thus, uranium-235 has 235 - 92 = 143 neutrons. Because elemental uranium is neutral, uranium-235 has 92 electrons.

The non-fissionable isotope of uranium has an mass number of A = 92 + 146 = 238. This isotope is known as uranium-238 and has a symbol of ${}^{238}_{92}\mathrm{U}$

Atomic Mass

Individual atoms weigh very little, typically about 10^{-24} g to 10^{-22} g. This amount is so small that there is no easy way to measure the mass of a single atom. To assign masses to atoms it is necessary to assign a mass to one atom and report the masses of other atoms relative to that absolute standard. By agreement, atomic mass is stated in terms of atomic mass units (amu), where 1 amu is defined as 1/12 of the mass of an atom of carbon-12. The atomic mass of carbon-12, therefore, is exactly 12 amu. The atomic mass of carbon-13 is 13.00335 amu because the mass of an atom of carbon-13 is 1.0836125 times greater than the mass of an atom of carbon-12.[†]

Average Atomic Mass. Because carbon exists in several isotopes, the atomic mass of an "average" carbon atom is neither $12.\overline{0}$ nor 13.00335. Instead it is 12.011, a value that is closer to $12.\overline{0}$ because 98.90% of all carbon atoms are carbon-12. As shown in the following example, if you know the percent abundance and atomic masses of an element's isotopes, then you can calculate it's average atomic mass.

Example 2. The element magnesium, Mg, has three stable isotopes with the following atomic masses and percent abundances:

^{24}Mg	23.9924 amu	78.70%
^{25}Mg	24.9938 amu	10.13%
^{26}Mg	25.9898 amu	11.17%

Calculate the average atomic mass for magnesium.

Solution. To find the average atomic mass we multiply each isotopes' atomic mass by its fractional abundance (the decimal equivalent of its percent abundance) and add together the results; thus

$$\text{avg. amu} = (0.7870)(23.994\text{ amu}) + (0.1013)(24.9938\text{ amu}) + (0.1117)(25.9898\text{ amu})$$

$$\text{avg. amu} = 24.32\text{ amu}$$

As the next example shows, we also can work such problems in reverse, using an element's average atomic mass and the atomic masses of its isotopes to find each isotope's percent abundance.

[†] By the way, if you calculate the masses of carbon-12 and carbon-13 by adding up the masses of each isotope's electrons, neutrons, and protons you will obtain a mass ratio of 1.08336, not 1.0836125. The reason for this is that the masses given in Table 2 are for "free" electrons, protons, and neutrons; that is, for electrons, protons, and neutrons that are not in an atom. When an atom forms, some of the mass is lost. "Where does it go?," you ask. Remember Einstein and $E=mc^2$? Mass can be converted to energy and the lost mass is the nuclear binding energy that holds the nucleus together.

Example 3. The element gallium, Ga, has two naturally occurring isotopes. The isotope ^{69}Ga has an atomic mass of 68.926 amu, whereas that for ^{71}Ga is 70.926 amu. The average atomic mass for gallium is 69.723. Find the percent abundances for gallium's two isotopes.

Solution. If we let x be the fractional abundance of ^{69}Ga, then the fractional abundance of ^{71}Ga is $1 - x$ (that is, the total amounts of ^{69}Ga and ^{71}Ga must add up to one). Using the same general approach as Example 2, we know that

$$\text{avg. amu} = 69.723 \text{ amu} = (x)(68.926 \text{ amu}) + (1 - x)(70.926 \text{ amu})$$

$$69.723 \text{ amu} = 68.926x \text{ amu} + 70.926 \text{ amu} - 70.926x \text{ amu}$$

$$2.000x \text{ amu} = 1.203 \text{ amu}$$

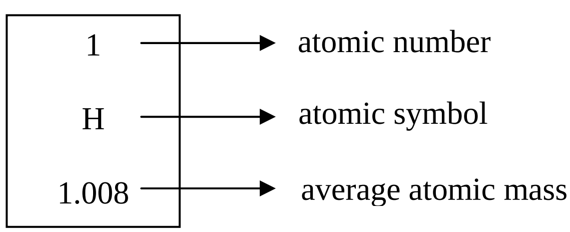
$$x = 0.6015$$

$$1 - x = 1 - 0.6015 = 0.3985$$

Thus, 60.15% of naturally occurring gallium is ^{69}Ga and 39.85% is ^{71}Ga.

The Periodic Table

With 113 elements to work with, it helps to have a convenient way to compile atomic numbers and average atomic masses. For over 100 years, chemists have found the periodic table to be the best way to organize this information. The periodic table is arranged into horizontal rows called periods and vertical columns called groups or families. You'll learn much more about the importance of this organization in other courses. For our purposes, it is important for you to know only how to use the periodic table to find an element's atomic symbol, its atomic number, and its average atomic mass.



A copy of the periodic table of the elements is shown here with average atomic masses given to a maximum of four significant figures. Average atomic masses shown in parentheses are approximate. At the end of this module you will find a list providing average atomic masses to each element's currently accepted number of significant figures.

Periodic Table of the Elements

1 H 1.008																	2 He 4.003
3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.90	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.70	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.0	89 Ac 227.0	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 * (269)	111 * (272)	112 * (277)		114 * (285)				

58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

* These elements were recently synthesized; however, they have not been named.

The Mole and the Molar Mass of Elements

Although the atomic mass unit provides a scale for comparing the relative masses of atoms, it is not a useful unit when working in the laboratory because it is too small (approximately 10^{-24} g). Additionally, the atomic mass unit applies to a single atom, whereas we work with gazillions of atoms at a time. To get around this problem we introduce another unit that is better suited for samples containing enormous numbers of atoms. Basically, the idea is to define a unit that represents a particular number of objects, just as we use a dozen to represent a collection of 12 eggs and a baker's dozen to represent 13 cookies.

In the SI system of units the mole is defined as the amount of a substance containing the same number of objects as there are atoms in exactly 12 g of carbon-12. This number has been determined experimentally to be 6.0221367×10^{23} and is known as Avogadro's number (N_A). For our purposes, we usually round Avogadro's number to 6.022×10^{23} particles/mol. A mole of zinc contains 6.022×10^{23} atoms of zinc and a mole of jellybeans contains 6.022×10^{23} jellybeans.

The advantage of defining a mole in this way is that an element's average atomic mass is identical to its molar mass.† Why is the true? A single atom of carbon-12 has an atomic mass of exactly 12 amu and a mole of carbon-12 atoms has a molar mass of exactly 12 g. A single atom of carbon-13 has an atomic mass of 13.00335 because its mass is 1.0836125 times greater than the mass of an atom of carbon-12. A mole of carbon-13, therefore, will have a mass of 13.00335 g.

Molecules, Network Solids, and Ionic Compounds

Only a few elements (the "noble gases" He, Ne, Ar, Kr, Xe, and Rn) are present in nature as single atoms. These gases are called monatomic, meaning one atom. All other matter consists of molecules, network solids, or ionic compounds. You will learn much more about these forms of matter in other courses. For our purposes, the following simple descriptions will suffice.

Molecules. A molecule is a collection of two or more atoms that are held in a specific arrangement by chemical bonds. The simplest molecules are diatomic molecules containing two atoms of the same element. The oxygen necessary for life is a diatomic molecule containing two oxygen atoms. Other diatomic molecules contain an atom of two different elements; thus, the poisonous gas carbon monoxide contains one atom of carbon and one atom of oxygen. Most molecules, however, consist of three or more atoms. For example, a molecule of glucose, a simple sugar, contains six atoms of carbon, six atoms of oxygen, and 12 atoms of hydrogen.

Network Solids. A molecule is a discrete particle containing an exact and invariant number of atoms. All molecules of carbon monoxide, for example, contain exactly one carbon atom and one oxygen atom. Other compounds, however, contain an exact ratio of atoms, but the absolute number of atoms may vary from sample-to-sample. Sand, or silicon dioxide, for example, has two oxygen atoms for every silicon atom. A large grain of sand, however, has more total oxygen atoms than a smaller grain. Such compounds are called network solids. Diamonds, which are a network of carbon atoms, and copper wire are other examples.

Ionic Compounds. As noted earlier in this module, a cation has a positive charge and an anion has a negative charge. An ion can be as simple as an element that has lost or gained an electron (Na^+ or Cl^-) or a collection of atoms that together have a positive or negative charge (OH^-). An ionic compound consists of one or more cations and one or more anions, combined such that the compound does not have a net charge. Ordinary table salt, for example, is an ionic compound containing an equal number of Na^+ and Cl^- ions. Ionic compounds have a fixed ratio of anions and cations, but individual samples may have different absolute numbers of anions and cations. Thus, for example, a large cube of table salt has more sodium cations and chloride anions than a smaller cube.

† Another name for an atom's molar mass is its atomic weight. Strictly speaking, weight is the force that gravity exerts on an object, which depends on the local effect of gravity. An atom's weight on earth is different than its weight on the moon; its mass, however, is the same in both locations.

Molecular Formulas and Empirical Formulas

It is inefficient to describe a molecule of glucose as containing six carbon atoms, six oxygen atoms, and 12 hydrogen atoms, or to say that ordinary table salt contains an equal number of Na^+ and Cl^- ions. Instead, we use two types of chemical formulas to indicate the composition of molecules, network solids and ionic compounds.

Molecular Formulas. A molecular formula shows the exact number of atoms of each element in a molecule. To write a molecular formula we list the elements using their atomic symbols, appending a subscript to each to indicate how many atoms of that element are in the molecule. Here are some examples:

	contains	molecular formula
oxygen	2 oxygen atoms	O_2
water	1 oxygen atom, 2 hydrogen atoms	H_2O
glucose	6 carbon atoms, 6 oxygen atoms, 12 hydrogen atoms	$C_6H_{12}O_6$

Sometimes molecular formulas are written to show some of a molecule's structural information. For example, acetic acid is often written as CH_3COOH, instead of $C_2H_4O_2$, because the –COOH part of the molecule, called a carboxylic acid group, is typical of many organic acids. Another example is dimethylamine, which is usually written as $(CH_3)_2NH$ instead of its condensed form, C_2H_7N, to emphasize that the molecule is a derivative of ammonia, NH_3, in which two methyl groups, CH_3–, replace two hydrogen atoms. Don't worry about these different ways of writing a molecule's molecular formula. No matter how we write a molecular formula, the number of atoms of each element in the molecule is the same. Both $(CH_3)_2NH$ and C_2H_7N show that dimethylamine contains 2 carbon atoms, 7 hydrogen atoms, and 1 nitrogen atom.

Example 4. Malonic acid is an organic acid with two carboxylic acid groups. Its formula is often represented as CH2(COOH)2. Rewrite this formula in its condensed form.

Solution. Counting up the elements we find that there are 3 carbon atoms, 4 hydrogen atoms, and 4 oxygen atoms; thus the condensed formula is $C_3H_4O_4$.

Empirical Formula. The molecular formula for glucose tells us exactly how many carbon atoms, oxygen atoms, and hydrogen atoms are in one molecule of glucose. An empirical formula doesn't provide this level of detail. Instead, an empirical formula tells us only what elements are present and the simplest whole-number ratio of their atoms. Looking at the molecular formula for glucose, we see that there are 2 hydrogen atoms for each carbon atom and each oxygen atom, and that there is one carbon atom for each oxygen atom. The empirical formula for glucose, therefore, is CH_2O.

If an empirical formula contains less information than a molecular formula, then why do we use them? There are two reasons. First, molecular formulas are written only for compounds that exist in identical units; that is, for molecules. For example, every molecule of glucose has exactly 6 carbons, 6 oxygens, and 12 hydrogens. Ionic compounds and network solids, on the other hand, don't exist in identical units. If you pour some table salt on a table and examine it with a hand lens, you will see that the crystals differ in size. Although the absolute number of sodium ions and chloride ions is different in each crystal, each has exactly one Na^+ for every Cl^-, which we show by writing its empirical formula as NaCl. The same is true for silicon dioxide, whose empirical formula is SiO_2.

The second reason for discussing empirical formulas is that when we analyze molecules we can determine only the relative amounts of the elements making up the compound. For example, if we analyze a sample of glucose to determine its chemical composition, we will find that there are 2 hydrogen atoms and 1 oxygen atom for every carbon atom. From this we can deduce that glucose's empirical formula is CH_2O, but we cannot distinguish between the many possible molecular formulas, three of which are $C_2H_4O_2$, $C_6H_{12}O_6$, and $C_{100}H_{200}O_{100}$, without some additional information. We will return to this idea in Module 3.

Molar Mass of Molecules, Network Solids, and Ionic Compounds

On average, an atom of oxygen has an atomic mass of 16.00 amu and a molar mass of 16.00 g/mol. A molecule of O_2, which contains two oxygen atoms, has a molecular mass equivalent to two oxygen atoms, or 32.00 amu, and a molar mass of 32.00 g/mol.[†] This approach is general. As shown by the following four examples, the molar mass of any molecule, network solid or ionic compound is simply the sum of the molar masses for each atom of each element in the compound's molecular or empirical formula.

Example 5. Glucose has a molecular formula of $C_6H_{12}O_6$. Determine its molar mass.

Solution. To find glucose's molar mass we add together the molar mass for each atom in the molecule being sure to account for each element's stoichiometry; thus

$$\text{molar mass of } C_6H_{12}O_6 = 6\times \text{molar mass for C} + 12\times \text{molar mass for H} + 6\times \text{molar mass for O}$$

$$\text{molar mass of } C_6H_{12}O_6 = 6\times 12.01 \text{ g/mol} + 12\times 1.008 \text{ g/mol} + 6\times 16.00 \text{ g/mol}$$

$$\text{molar mass of } C_6H_{12}O_6 = 180.16 \text{ g/mol}$$

[†] Just as an atom's molar mass is often called its atomic weight, a molecule's molar mass is also known as its molecular weight or formula weight.

Example 6. Trimethylamine has a molecular formula of $(CH_3)_3N$; determine its molar mass.

Solution. Let's begin by rewriting the molecular formula as C_3H_9N to clarify the number of atoms present for each element. Proceeding as in Example 5 gives

$$\text{molar mass of } (CH_3)_3N = 3 \times 12.01 \text{ g/mol} + 9 \times 1.008 \text{ g/mol} + 1 \times 14.01 \text{ g/mol}$$

$$\text{molar mass of } (CH_3)_3N = 59.11 \text{ g/mol}$$

Example 7. Baking soda, which is also known as sodium bicarbonate, is an ionic compound with an empirical formula of $NaHCO_3$. Determine the molar mass for $NaHCO_3$.

Solution. To find sodium bicarbonate's molar mass we add together the molar mass for each atom in the compound's empirical formula being sure to account for each element's stoichiometry; thus

$$\text{molar mass of } NaHCO_3 = 22.99 \text{ g/mol} + 1.008 \text{ g/mol} + 12.01 \text{ g/mol} + 3 \times 16.00 \text{ g/mol}$$

$$\text{molar mass of } NaHCO_3 = 84.01 \text{ g/mol}$$

Example 8. Prussian blue is a blue pigment discovered in the early 1700s that is used as a dye and as a pigment for paints. Its empirical formula is $Fe_4[Fe(CN)_6]_3$. The use of the square brackets separates the cation, Fe^{3+}, from the anion, $Fe(CN)_6^{4-}$. Determine the molar mass for $Fe_4[Fe(CN)_6]_3$.

Solution. Let's begin by rewriting the molecular formula as $Fe_7C_{18}N_{18}$ to clarify the number of atoms present for each element. Proceeding as in earlier examples gives

$$\text{molar mass of } Fe_4[Fe(CN)_6]_3 = 7 \times 55.85 \text{ g/mol} + 18 \times 12.01 \text{ g/mol} + 18 \times 14.01 \text{ g/mol}$$

$$\text{molar mass of } Fe_4[Fe(CN)_6]_3 = 859.31 \text{ g/mol}$$

Working With Avogadro's Number and Molar Mass

Avogadro's number and an element's or compound's molar mass provide a means to convert between mass and moles, and between moles and number of particles. Using these two numbers we can, for example, calculate the moles of glucose in a sugar packet, or the number of sodium ions in a box of baking soda. If you need to, review the discussion of dimensional analysis in Module 1 before studying the following four examples.

Example 9. Urea, $(NH_2)_2CO$, is a fertilizer used to add nitrogen to soil. How many grams of nitrogen are added if 16.8 kg of urea is applied to a farmer's field?

Solution. Here we need to convert the mass of urea into the moles of urea, the moles of urea into the moles of nitrogen, and the moles of nitrogen into the mass of nitrogen. We can do each step separately or string the conversions together as shown here

$$16.8 \text{ kg urea} \times \frac{1000 \text{ g urea}}{\text{kg urea}} \times \frac{1 \text{ mol urea}}{60.06 \text{ g urea}} \times \frac{2 \text{ mol N}}{\text{mol urea}} \times \frac{14.01 \text{ g N}}{\text{mol N}} = 7.84 \times 10^3 \text{ g N}$$

Example 10. Female insects emit pheromones to attract males for mating. One such pheromone has a molecular formula of $C_{19}H_{38}O$. Typically, the female insect releases approximately 2.2×10^9 molecules at a time. How many pg of the pheromone is this?

Solution. Beginning with the molecules of the pheromone we proceed as shown here

$$2.2 \times 10^9 \text{ molecules} \times \frac{1 \text{ mol } C_{19}H_{38}O}{6.022 \times 10^{23} \text{ molecules}} \times \frac{282.5 \text{ g } C_{19}H_{38}O}{\text{mol } C_{19}H_{38}O}$$

$$\times \frac{1 \times 10^{12} \text{ pg } C_{19}H_{38}O}{1 \text{ g } C_{19}H_{38}O} = 1.0 \text{ pg } C_{19}H_{38}O$$

Example 11. The molecular formula for aspartame is $C_{14}H_{18}N_2O_5$. Suppose you have a 589-mg sample of aspartame (Asp). Calculate the moles of aspartame, the number of molecules of aspartame, and the number of atoms of carbon in the sample.

Solution. To find the moles of Asp we use its molar mass

$$589 \text{ mg Asp} \times \frac{1 \text{ g Asp}}{1000 \text{ mg Asp}} \times \frac{1 \text{ mol Asp}}{294.3 \text{ g Asp}} = 2.00 \times 10^{-3} \text{ mol Asp}$$

To convert this result to molecules of Asp, we use Avogadro's number.

$$\left(2.00 \times 10^{-3} \text{ mol}\right)\left(\frac{6.022 \times 10^{23} \text{molecules}}{\text{mol}}\right) = 1.20 \times 10^{21} \text{ molecules}$$

Finally, to convert this result to atoms of carbon we note that each molecule of Asp contains 14 carbon atoms.

$$\left(1.20 \times 10^{21} \text{ molecules}\right)\left(\frac{14 \text{ atoms C}}{\text{molecule}}\right) = 1.68 \times 10^{22} \text{ atoms of C}$$

Example 12. Aluminum oxide, Al_2O_3, occurs in nature as the mineral corundum. The density of corundum is 3.97 g/cm^3. How many atoms of Al are there in a cube of corundum whose volume is 15.6 cm^3?

Solution. The mass of corundum in the cube is

$$15.6\ \text{cm}^3 \times \frac{3.97\ \text{g Al}_2\text{O}_3}{\text{cm}^3} = 61.93\ \text{g Al}_2\text{O}_3$$

Note that we've retained an extra significant figure because we will use this number in the next calculation.

$$61.93\ \text{g Al}_2\text{O}_3 \times \frac{1\ \text{mol Al}_2\text{O}_3}{101.96\ \text{g}} \times \frac{2\ \text{mol Al}}{\text{mol Al}_2\text{O}_3}$$

$$\times \frac{6.022 \times 10^{23}\ \text{atoms Al}}{\text{mol Al}} = 7.32 \times 10^{23}\ \text{atoms Al}$$

$$2.2 \times 10^9\ \text{molecules} \times \frac{1\ \text{mol C}_{19}\text{H}_{38}\text{O}}{6.022 \times 10^{23}\ \text{molecules}} \times \frac{282.5\ \text{g C}_{19}\text{H}_{38}\text{O}}{\text{mol C}_{19}\text{H}_{38}\text{O}}$$

$$\times \frac{1 \times 10^{12}\ \text{pg C}_{19}\text{H}_{38}\text{O}}{1\ \text{g C}_{19}\text{H}_{38}\text{O}} = 1.0\ \text{pg C}_{19}\text{H}_{38}\text{O}$$

Practice Problems

The following problems provide practice in meeting this module's objectives. Answers are provided on the last page. Be sure to seek assistance if you experience difficulty with any of these problems. When you are ready, schedule an appointment for the module's exam.

1. What is the mass number of an atom of cobalt that has 32 neutrons?

2. One of the most useful radioactive isotopes in medicine is ^{99}Tc, which is used to obtain images of organs such as the heart and lungs. How many protons, neutrons, and electrons are in this isotope?

3. Which of the isotopes shown below has the greatest number of neutrons and how many neutrons does it have?

   ^{52}Fe ^{57}Co ^{56}Ni ^{56}Mn

4. Another medically useful radioactive isotope is iodine-131, which is used in evaluating the thyroid. How many protons, neutrons, and electron are in this isotope?

5. Naturally occurring argon, Ar, consists of three isotopes with the following percent abundances and atomic masses

isotope	atomic mass (amu)	percent abundance
^{36}Ar	35.96755	0.34%
^{38}Ar	37.96273	0.07%
^{40}Ar	39.96238	99.59%

   What is the average atomic mass for Ar?

6. Naturally occurring chlorine, Cl, consists of the isotopes ^{35}Cl and ^{37}Cl with atomic masses of 34.968 amu and 36.956 amu, respectively. The average atomic mass of chlorine is 35.453. What are the percent abundances for ^{35}Cl and ^{37}Cl?

7. Glycine is one of the amino acids present in proteins. The following facts are known about a molecule of glycine:

   - it contains exactly one nitrogen atom
   - there are twice as many carbon atoms as there are nitrogen atoms
   - the number of oxygen atoms is equal to the number of carbon atoms
   - the ratio of hydrogen atoms to carbon atoms is 2.5:1

   What is the molecular formula for glycine?

8. One of the first treatments for bipolar depression was the ionic compound lithium carbonate. Lithium is present as the cation Li^+ and carbonate as the anion CO_3^{2-}. What is the empirical formula for lithium carbonate?

9. Calculate the molar mass to four significant figures for the following compounds:

    $C_6H_8O_6$, which is ascorbic acid

    $C_3H_5N_3O_9$, which is nitroglycerine

    $Ca_3(PO_4)_2$, a major component of bones

    $PtCl_2(NH_3)_2$, a chemotherapeutic reagent

    $Cu_3(OH)_2(CO_3)_2$, the mineral azurite

10. If an Avogadro's number of pennies was equally distributed to the approximately 290 million people living in the United States, how many dollars would each person receive?

11. The characteristic smell of garlic is due to the molecule allicin, $C_6H_{10}OS_2$. How many moles are in 15.3 mg of allicin?

12. One of the compounds that makes chili peppers hot is capsaicin, $C_{18}H_{27}NO_3$. How many atoms of carbon are in 0.662 g of capsaicin?

13. Chloral hydrate, $C_2H_3Cl_3O_2$, is a drug once used as a sedative and as a hypnotic. It is the compound popularized in detective stories as a "Mickey Finn." Suppose you have 1.00 g of chloral hydrate. How many atoms of chlorine do you have? How many grams of oxygen do you have? How many moles of carbon do you have?

14. A single molecule of the antibiotic penicillin G has a mass of 5.56×10^{-22} g. What is the molar mass of penicillin G?

15. Erythorcin-B, $C_{20}H_6O_5I_4Na_2$, is a red dye that was once used to color maraschino cherries. As few as 2×10^{12} molecules are sufficient to impart a noticeable color to 1 mL of water. How many grams of erythrocin-B is this?

16. Humolone, $C_{21}H_{30}O_5$, is one of the components of hops that imparts a bitter flavor to beer. What is the mass in grams of 1.00×10^9 molecules of humolone?

17. A drop of water, H_2O, has a volume of approximately 0.05 mL. Assuming that the density of water is 1.00 g/mL, how many molecules of water are in 3 drops of water?

18. Gold, Au, is present in sea water to the extent of 0.15 mg Au/ton seawater. How many Au atoms are present in 3.00×10^{2} g seawater, which is equivalent to approximately a glassful of seawater? Note: 1 ton is exactly 2000 pounds.

19. Several compounds can be used to add nitrogen to soil. One is ammonia, NH_3, and another is ammonium nitrate, NH_4NO_3. How many grams of ammonium nitrate are needed to provide the same amount of nitrogen as 1.00 gram of ammonia?

20. If a dime weighs 2.35 g and is 83% Ni by mass, how many Ni atoms does it contain? Suppose that each Ni atom has a diameter of 3.0×10^{-10} m. If you place these Ni atoms in a straight line so that they just touch each other, how long will the line be in meters?

Answers to Practice Problems

1. 59

2. 43 protons, 56 neutrons, 43 electrons

3. ^{56}Mn, 31 neutrons

4. 53 protons, 78 neutrons, 53 electrons

5. 39.95 amu

6. 75.60%, 24.40%

7. $C_2H_5NO_2$

8. Li_2CO_3

9. 176.1 g/mol

   227.1 g/mol

   310.2 g/mol

   300.1 g/mol

   344.7 g/mol

10. $\$2.07 \times 10^{13}$

11. 9.43×10^{-5} mol

12. 2.35×10^{22} C atoms

13. 1.09×10^{22} atoms Cl, 0.193 g O, 1.21×10^{-2} mol C

14. 335 g/mol

15. 3×10^{-9} g

16. 6.02×10^{-13} g

17. 5×10^{21} molecules

18. 1.5×10^{14} atoms Au

19. 2.35 g

20. 2.0×10^{22} atoms Ni, 6.0×10^{12} m

List of the Elements with Their Symbols, Atomic Numbers, and Atomic Masses*

Element	Symbol	Atomic Number	Atomic Mass	Element	Symbol	Atomic Number	Atomic Mass
Actinium	Ac	89	(227)	Mendelevium	Md	101	(256)
Aluminum	Al	13	26.981538	Mercury	Hg	80	200.59
Americium	Am	95	(243)	Molybdenum	Mo	42	95.94
Antimony	Sb	51	121.760	Neodymium	Nd	60	144.24
Argon	Ar	18	39.948	Neon	Ne	10	20.1797
Arsenic	As	33	74.92160	Neptunium	Np	93	(237)
Astatine	At	85	(210)	Nickel	Ni	28	58.6934
Barium	Ba	56	137.327	Niobium	Nb	41	92.90638
Berkelium	Bk	97	(247)	Nitrogen	N	7	14.00674
Beryllium	Be	4	9.012182	Nobelium	No	102	(253)
Bismuth	Bi	83	208.98038	Osmium	Os	76	190.23
Bohrium	Bh	107	(262)	Oxygen	O	8	15.9994
Boron	B	5	10.811	Palladium	Pd	46	106.42
Bromine	Br	35	79.904	Phosphorous	P	15	30.973762
Cadmium	Cd	48	112.411	Platinum	Pt	78	195.078
Calcium	Ca	20	40.078	Plutonium	Pu	94	(242)
Californium	Cf	98	(249)	Polonium	Po	84	(210)
Carbon	C	6	12.0107	Potassium	K	19	39.0983
Cerium	Ce	58	140.116	Praseodymium	Pr	59	140.90765
Cesium	Cs	55	132.90545	Promethium	Pm	61	(147)
Chlorine	Cl	17	35.4527	Protactinium	Pa	91	231.03588
Chromium	Cr	24	51.9961	Radium	Ra	88	(226)
Cobalt	Co	27	58.933200	Radon	Rn	86	(222)
Copper	Cu	29	63.546	Rhenium	Re	75	186.207
Curium	Cm	96	(247)	Rhodium	Rh	45	102.90550
Dubnium	Db	105	(260)	Rubidium	Rb	37	85.4678
Dysprosium	Dy	66	162.50	Ruthenium	Ru	44	101.07
Einsteinium	Es	99	(254)	Rutherfordium	Rf	104	(257)
Erbium	Er	68	167.26	Samarium	Sm	62	150.36
Europium	Eu	63	151.964	Scandium	Sc	21	44.95591
Fermium	Fm	100	(253)	Seaborgium	Sg	106	(263)
Fluorine	F	9	18.9984032	Selenium	Se	34	78.96
Francium	Fr	87	(223)	Silicon	Si	14	28.0855
Gadolinium	Gd	64	157.25	Silver	Ag	47	107.8682
Gallium	Ga	31	69.723	Sodium	Na	11	22.98977
Germanium	Ge	32	72.61	Strontium	Sr	38	87.62
Gold	Au	79	196.96655	Sulfur	S	16	32.066
Hafnium	Hf	72	178.49	Tantalum	Ta	73	180.9479
Hassiuim	Hs	108	(265)	Technetium	Tc	43	(99)
Helium	He	2	4.002602	Tellurium	Te	52	127.60
Holmium	Ho	67	164.93032	Terbium	Tb	65	158.92534
Hydrogen	H	1	1.00794	Thallium	Tl	81	204.3833
Indium	In	49	114.818	Thorium	Th	90	232.0381
Iodine	I	53	126.90447	Thulium	Tm	69	168.93421
Iridium	Ir	77	192.217	Tin	Sn	50	118.710
Iron	Fe	26	55.845	Titanium	Ti	22	47.867
Krypton	Kr	36	83.80	Tungsten	W	74	183.84
Lanthanum	La	57	138.9055	Uranium	U	92	238.0289
Lawrencium	Lr	103	(257)	Vanadium	V	23	50.9415
Lead	Pb	82	207.2	Xenon	Xe	54	131.29
Lithium	Li	3	6.941	Ytterbium	Yb	70	173.04
Lutetium	Lu	71	174.967	Yttrium	Y	39	88.90585
Magnesium	Mg	12	24.3050	Zinc	Zn	30	65.39
Manganese	Mn	25	54.938049	Zirconium	Zr	40	91.224
Meitnerium	Mt	109	(266)				

* Approximate atomic masses for radioactive elements are given in parentheses.