Faculty of Science

Unit 1: Measurement and Matter

CHEM 1503
Chemical Bonding and Organic Chemistry

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Overview

Unit 1 introduces you to three very important topics: the language of scientific measurement, the language of chemistry, and the mole concept. These topics are interrelated and, as you will see in later units, form the base on which you will build your knowledge of chemistry.

Your study of the language of scientific measurement will introduce you to units, conversion factors, significant figures, and exponential notation. Since chemistry is a quantitative science, chemical measurements must be expressed both accurately and precisely. The international system of units (SI) simplifies many calculations, reduces the number of units in use, and standardizes international usage.

The language of chemistry includes the classification and properties of matter, chemical symbols and formulas, the empirical laws of chemical combinations, basic chemical nomenclature, and the chemical equation. You can think of these topics as the vocabulary for the language of chemistry.

The last topic, the mole concept, provides the link between *mass* and *amount* of substances. The study of the relative quantities of elements combined in compounds, and the relative quantities of elements and compounds involved in chemical reactions is often called stoichiometry. This word is derived from the Greek word *stoicheion*, “element,” and thus means literally “the measurement of elements.” Stoichiometry is often more simply defined as chemical arithmetic.

Learning Outcomes

By the end of this unit you should be able to:

- Correctly use SI units, significant figures, and scientific notation.
- Explain the basic chemical terms introduced in the unit (e.g., element, compound, metal, nonmetal, metalloid, homogeneous, heterogeneous, atom, ion, molecule, isotope).
- Distinguish between examples of homogeneous and heterogeneous matter, mixtures, pure substances, compounds, and elements.
- Name and give formulas for simple inorganic compounds.
- Write balanced chemical equations.
- Carry out calculations involving moles, molecular weights, masses, and numbers of molecules.
- Calculate empirical formulas from mass composition data and vice versa, and relate these to molecular formulas.
- Perform stoichiometric calculations involving chemical reactions.
- Interrelate volume, molarity, and density of a solution.
- Describe types of energy and energy changes.

The Science of Chemistry

Read Sections 1.1, 1.2, and 1.3 of your textbook. The material covered in these introductory sections introduces you to the scientific method and provides you with some idea of what a chemist does. Study Sections 1.4 and 1.5. The definitions in these sections are important.

You probably already know most of the terms printed in *boldface* type in the textbook; however, you should be aware that terms used in chemistry sometimes have a more specific meaning than the same words used in everyday life. For example, from a consumer's point of view, "pure" orange juice contains juice, some orange pulp, and possibly even an orange seed or two. From a chemist's point of view, the same orange juice is not a pure substance. It does not fit the criterion of having the same definite chemical composition throughout, nor is its composition well defined. In this course you will find other examples of everyday words used with a more specific meaning in chemistry.

At this point you should know (i.e., memorize if you do not already know) the symbols for those elements given in Table 1.1 on the following page. The date of discovery and derivations of names are only for your interest. You should also learn the names and symbols of other elements as you encounter them in the course.

Table 1.1: Symbols of common elements and derivation of their names
(Gr. = Greek; Lat. = Latin)

Name	Symbol	Date of Discovery	Derivation
aluminum	Al	1825	*alumen* (Lat., alum)
argon	Ar	1894	*argos* (Gr., inactive)
arsenic	As	1250	*arsenikon* (Gr., yellow pigment)
barium	Ba	1808	*barys* (Gr., heavy)
beryllium	Be	1828	*beryl* (Lat., sweet)
boron	B	1808	*borak* (Persian, white)
bromine	Br	1826	*bromos* (Gr., stench)
cadmium	Cd	1817	*kadmia* (Gr., earth)
calcium	Ca	1808	*calx* (Lat., lime)
carbon	C	Ancient	*carbo* (Lat., charcoal)
chlorine	Cl	1774	*chloros* (Gr., green)
chromium	Cr	1797	*chroma* (Gr., colour)
cobalt	Co	1735	*kobold* (German, goblin)
copper	Cu	Ancient	*cuprum* (Lat., found in Cyprus)
fluorine	F	1886	*fluere* (Lat., flow)
gallium	Ga	1875	*Gallia* (Lat., France)
germanium	Ge	1886	*Germania* (Lat., Germany)
gold	Au	Ancient	*aurum* (Lat., shining dawn)
helium	He	1868	*helios* (Gr., sun)
hydrogen	H	1766	*hydro, genes* (Gr., water, forming)
iodine	I	1811	*iodes* (Gr., violet)
iron	Fe	Ancient	*ferrum* (Lat., iron)
krypton	Kr	1898	*kryptos* (Gr., hidden)
lead	Pb	Ancient	*plumbum* (Lat., heavy)
lithium	Li	1817	*lithos* (Gr., rock)
magnesium	Mg	1808	*Magnesia* (a district in Greece)
manganese	Mn	1774	*magnes* (Lat., magnet)
mercury	Hg	Ancient	*hydrargyrum* (Lat., water-silver)
neon	Ne	1898	*neos* (Gr., new)
nickel	Ni	1751	*kopparnickel* (Swedish, false copper)
nitrogen	N	1772	*nitron, genes* (Gr., native soda, forming)
oxygen	O	1772	*oxys, genes* (Gr., acid, forming)
phosphorus	P	1669	*phosphoros* (Gr., light bearing)
platinum	Pt	1735	*platina* (Spanish, little silver)
potassium	K	1807	*kalium* (Lat., potash)
scandium	Sc	1879	*Scandinavia*
selenium	Se	1817	*selene* (Gr., moon)
silicon	Si	1824	*silex* (Lat., flint)
silver	Ag	Ancient	*argentum* (Lat., silver)
sodium	Na	1807	*natrium* (Lat., soda)
sulphur	S	Ancient	*sulvere* (Sanskrit, sulphur)
tin	Sn	Ancient	*stannum* (Lat., tin)
titanium	Ti	1791	*Titans* (giants in Greek mythology)
uranium	U	1789	*Uranus* (Gr., planet discovered in 1781)
vanadium	V	1801	*Vanadis* (Norse goddess of beauty)
zinc	Zn	1746	*zink* (German, of obscure origin)

Measurement in Chemistry

Study Sections 1.6 and 1.7 in your textbook. Canada converted to the metric system many years ago, so you are probably familiar with metric units. You may not be as familiar with SI units, but don't worry if the SI system seems somewhat complicated at first. It is a system that was developed because it is simple to use. Once you have some experience in performing chemical calculations with the SI system, you'll find that you have quite painlessly become "SI-ed." For now, you need only know the base units for length, mass, time, and temperature. Other SI units will be introduced as you need them.

The SI system is a logical extension of our everyday metric system. The following example shows how any physical quantity can be expressed as derived units from SI base units.

Problem: Given that acceleration is defined as force per unit mass, convert the SI units for acceleration into derived units containing SI base units.

Solution: acceleration = force ÷ mass

$$\text{acceleration} = \frac{\text{force}}{\text{mass}} = \frac{\text{N}}{\text{kg}} \text{ or } \frac{\text{J}}{\text{m kg}} \left(\text{commonly used SI units}\right)$$

$$\text{and acceleration} = \frac{\text{N}}{\text{kg}} = \frac{\text{kg m s}^{-2}}{\text{kg}} = \text{m s}^{-2} \left(\text{expressed in basic SI units}\right)$$

Many places in your textbook give temperature in the Kelvin scale (K); this is the SI unit of temperature. Remember that, in order to convert temperature on the Kelvin scale (T) to temperature on the Celsius scale (centigrade scale, t), you must subtract 273.15 degrees, as in

$$t = T - 273.15$$

Conversely, $T = t + 273.15$

Note that we only use the letter K to represent a unit of temperature in the Kelvin scale, and not °K.

A temperature of 298 K may sound very hot, but on conversion,

$$t = 298 - 273 = 25°\text{C},$$

you will find that it is the temperature of a warm summer's day.

Dealing with Numbers

Significant Figures

Study Section 1.8 in your textbook. Make sure that you understand the “Guidelines for Using Significant Figures” explained in this section.

These rules will be followed in all the examples in this course. “Sig figs,” as they are called, are really only common sense. If you can weigh out 6.37 g of a substance (3 sig figs) and divide it into 2.4 parts (2 sig figs) a calculator will show the following results for this division.

$$\frac{6.37}{2.4} = 2.654166667 \quad \text{(10 sig figs)}$$

But common sense should tell you that your result cannot be more accurate than your least accurate measurement, which in this case has two significant figures. Your answer should be 2.7 (2 sig figs, rounded up). Remember that calculators give you arithmetic accuracy, not significant figures. In other words, calculators lack common sense.

Mathematical Procedures

CHEM 1503 and CHEM 1523 require a scientific calculator with exponential, logarithm, antilogarithm, reciprocal, and square root keys. Except for exponential notation, you will not use these mathematical functions in this course.

The Factor-Label Method

Study Section 1.9 in your textbook and work through Examples 1.6, 1.7, and 1.8. Do the practice exercises given in the text after each of these examples. The answers are at the very end of the chapter. The reason you can multiply by a *conversion factor* to convert from one type of unit of measurement to another is that a conversion factor always equals 1, and multiplying an amount by 1 does not change its value. This is a very useful procedure that you should know how to use.

Read the summary at the end of the chapter.

Basic Atomic Structure

Carefully read Sections 2.1 and 2.2 in your textbook. You do not need to study the historical material now, but you should know the meaning of the terms given in bold italics in these sections (e.g., electrons, protons, nucleus, radioactivity, etc.). You will return to the historical aspects at the beginning of Unit 2.

Section 2.1 in your textbook discusses the **law of definite proportions** (sometimes known as the **law of constant composition**) and the **law of multiple proportions**. These two laws might seem obvious since the idea of the existence of atoms will not be novel to you. But in the eighteenth century, these laws had a profound effect on the development of chemical theory. One way to remember the law of constant composition is “water is water is water.” In other words, water will always have the formula H_2O, regardless of the chemical reaction which formed it. If it does not have the formula H_2O, then it is not water. Another compound also composed of only hydrogen and oxygen is hydrogen peroxide, H_2O_2. It contains hydrogen and oxygen in whole number ratios, but in a different ratio to water. Hydrogen peroxide thus demonstrates the law of multiple proportions. You may already be familiar with the **law of conservation of mass** from physics courses.

You should memorize the charges of the subatomic particles and their relative masses, as shown in the table below.

Table 1.2: Comparative charges and masses of the subatomic particles

Name	Symbol	Charge	Relative mass
proton	p	+1	1
neutron	n	0	1
electron	e	−1	$\frac{1}{1840} \approx 0$

As the mass of an electron is very small compared to the mass of a proton or a neutron, an electron is often considered to have a relative mass of zero.

Atomic Number, Mass Number, and Isotopes

Study Section 2.3 in your textbook. This is a very important section, so make sure you follow the material being discussed. One way to ensure that you understand the principles is to try the worked examples as they appear in the textbook. For example, after you have studied atomic number, mass number, and isotopes, try to do Example 2.1 and the practice exercise. Don't look at the answers until you have attempted them for yourself. If you need more practice with this material, check the exercises at the end of the chapter where there is a section on atomic number, mass number, and isotopes. Answers to some of these problems are given at the back of the textbook.

The Periodic Table

Study Section 2.4 in your textbook. You should know the terms given in bold italics, but you do not need to memorize the periodic table (shown in Figure 2.10 in your textbook). You will be supplied with a periodic table for the final exam for this course and you will find that the more you refer to the table, the sooner it will become familiar.

Molecules and Ions

Study Sections 2.5 and 2.6 in your textbook. Although the term **molecule** is defined here as being two or more atoms held together by chemical bonding, you may find the term **monatomic molecules** used for the noble gases (helium, neon, argon, krypton, etc.). This may seem to be a contradiction of the definition but as these gases exist naturally as single atoms, they are often referred to in this way. You should know the difference between ***molecular*** and ***empirical*** **formulas**. A molecular formula will be the empirical formula times 1 or times 2 or times 3, etc. Some further examples of molecular and empirical formulas are given in Table 1.3 in this unit. Work through Examples 2.2, 2.3, and 2.4 in the textbook and their practice exercises, then check your answers at the very end of the chapter.

Table 1.3: Examples of empirical and molecular formulas

Empirical formula	Molecular formula	Name of compound
NH_2	N_2H_4	hydrazine
CH_2O	$C_2H_4O_2$	acetic acid
$HgCl$	Hg_2Cl_2	mercury(I) chloride
NO_2	NO_2	nitrogen dioxide
NO_2	N_2O_4	dinitrogen tetroxide

Chemical Formulas and Names

Sections 2.6 and 2.7 in your textbook are very important parts of this course. Be sure to study, understand, and where necessary memorize the materials in these sections.

Ionic Compounds

This part of Section 2.7 explains how to write formulas for inorganic compounds and how to name them using the IUPAC (pronounced I-u-pack; from International Union of Pure and Applied Chemistry) rules. To make it easier for you to learn this material, in Table 1.4 below is a summary of the more common anions, grouped by charge. For now, memorize the names, formulas, and charges of these anions. You'll eventually learn most of the other anions listed in Table 2.3 of your textbook.

Table 1.4: Names, formulas, and charges of common anions

−1 anions		**−2 anions**		**−3 anions**	
fluoride	F^-	oxide	O^{2-}	nitride	N^{3-}
chloride	Cl^-	sulphide	S^{2-}	phosphide	P^{3-}
bromide	Br^-	carbonate	CO_3^{2-}	phosphate	PO_4^{3-}
iodide	I^-	sulphate	SO_4^{2-}		
hydroxide	OH^-	sulphite	SO_3^{2-}		
hydrogen carbonate*	HCO_3^-	dichromate	$Cr_2O_7^{2-}$		
perchlorate	ClO_4^-				
chlorate	ClO_3^-				
chlorite	ClO_2^-				
hypochlorite	ClO^-				
permanganate	MnO_4^-				
nitrate	NO_3^-				
nitrite	NO_2^-				
acetate	$C_2H_3O_2^-$				
hydrogen sulphate*	HSO_4^-				

* The traditional names for hydrogen carbonate, HCO_3^-, and hydrogen sulphate, HSO_4^-, are bicarbonate and bisulphate. You will probably find these traditional names in chemistry as often as the IUPAC names.

You do not have to memorize the charges of most cations. You can always work them out. However, you must memorize the names and formulas. Non-variable valence cations, that is, those having only one possible charge, are given in Table 1.5 below.

Table 1.5: Summary of non-variable valence cations

Charge	**Cations**
+ 1 cations	all group 1A metal cations, Ag^+, NH_4^+
+ 2 cations	all group 2A cations, Zn^{2+}, Cd^{2+}
+ 3 cations	Al^{3+}

The remainder of the metals that you are expected to know are in the variable charge category. You do not have to memorize that iron can have Fe^{2+} or Fe^{3+} cations, although you can if you wish. You can always calculate what the charge is on a cation from its formula, or from its name.

For example, $FeCl_3$ is an iron chloride, but which iron chloride? The chloride ion is always Cl^{1-}, and there are three of them in this compound. Therefore, their total charge is –3. This means that the iron must be Fe^{3+} to form a neutral (uncharged) compound; the name will be iron(III) chloride. Iron(II) chloride would be $FeCl_2$. Note the absence of a space between "iron" and the Roman numeral in parentheses. Although written with Roman numerals, these names are spoken as “iron three chloride” and “iron two chloride.” The numbers refer to the charge on the cation. Note that sodium chloride is not written as sodium(I) chloride nor said as “sodium one chloride.” To say sodium(I) chloride is incorrect. It implies that sodium has a +1 charge here, but could have a different charge in another compound. The use of Roman numerals in the name is reserved for variable charge metals.

Go over Examples 2.5 and 2.6 in your textbook and work through their practice exercises.

Do not memorize the older “ous-ic” naming system. You should read about it because you may already have heard of it or may encounter it later. This “ous-ic” system can be ambiguous, so we will stick to the Roman numeral system approved by the International Union of Pure and Applied Chemistry (IUPAC) in this course.

Molecular Compounds

These are compounds of nonmetal elements and their naming rules are straightforward. Remember that you never use this “mono-, di-, tri-…” system with ionic compounds. Work through Examples 2.7 and 2.8 in your textbook and their practice exercises.

Acids and Bases

Memorize the names, formulas, and corresponding anions of the acids listed in Tables 2.5 and 2.6 of your textbook as well as the following acids.

H_2SO_4	sulphuric acid
HNO_3	nitric acid
H_2CO_3	carbonic acid
$HC_2H_3O_2$	acetic acid
H_3PO_4	phosphoric acid
H_2SO_3	sulphurous acid
HNO_2	nitrous acid

You should know the following bases.

NaOH	sodium hydroxide
KOH	potassium hydroxide
$Ba(OH)_2$	barium hydroxide
NH_3	ammonia

Hydrates

Go over the material on hydrates in your textbook. You do not need to memorize any formulas here. All you need to know is how to name a hydrate using the Greek prefixes and how to write its formula. For example, $ZnSO_4 \cdot 7H_2O$ is zinc sulphate heptahydrate, and the formula for sodium sulphate decahydrate is $Na_2SO_4 \cdot 10H_2O$.

You do not need to memorize any of the common names listed in Table 2.7 in your textbook, but you may find that you have to consult this table occasionally.

Introduction to Organic Compounds

Read Section 2.8 in your textbook and become familiar with the names of the first ten straight-chain alkanes listed in Table 2.8, as well as the hydroxyl, amino, and carboxyl functional groups. In the second half of this course, we will investigate organic chemistry in much more depth.

At this point you may feel overwhelmed, especially if your background in chemistry is not particularly strong, or if it has been quite a while since you studied chemistry. This is a natural reaction and you should not feel discouraged. When you are advised in this course to memorize, or know, or understand some material, you are not expected to achieve instant enlightenment. If you keep working steadily on the course, you will find that things start to fall into place. You are not expected to be immediately perfect in nomenclature and formulas. Like everything else, familiarity will help. From now on, every time you come across a formula or a name, make sure you understand why it has that formula or that name. Eventually, you will automatically be able to give the names of many chemical compounds, and you will instinctively know if a formula is correct.

Read the summary at the end of Chapter 2 of your textbook. You may find it useful to attempt some of the chapter questions now, as recommended at the end of this unit.

Atomic and Molecular Masses

Study Sections 3.1 and 3.2 in your textbook. Note that you do not need to memorize average atomic masses (e.g., 63.55 amu for copper) or molar masses of elements (e.g., 63.55 g for copper). However, make sure that you can explain the difference between the mass of an atom and the molar mass of an element. Work through Examples 3.1 to 3.4 and their associated practice exercises. Answers are given at the very end of Chapter 3.

Study Section 3.3 in your textbook and work through Examples 3.5, 3.6 and 3.7 and their practice exercises.

Read Section 3.4, where the modern and highly accurate method of atomic and molecular mass determination using the mass spectrometer is discussed.

Mass Composition of Compounds

Study Section 3.5 in your textbook. A chemical formula tells you the mole ratio of elements in the compound (or the ratio of the number of each sort of atom). The mass composition, generally expressed as percent composition or percent by mass, gives you the weight ratio. Given a table of atomic masses, you should be able to calculate the percent composition by mass from a chemical formula as shown in Example 3.8 and its practice exercise. Also work through Example 3.10.

Determining Empirical Formulas

Example 3.9 and its practice exercise show how to determine the empirical formula from mass percent composition data. As this is a very basic type of chemistry problem, provided below is another example showing the problem laid out in a table format.

Problem: A compound contains 26.58% by weight potassium, 35.35% by weight chromium, and 38.07% by weight oxygen. Find its empirical formula.

Solution: First assume you have 100 g of the compound. You may of course assume any mass you like, but in this case choosing 100 g keeps the arithmetic simple. So, 100 g of the compound contains 26.58 g of potassium, 35.35 g of chromium, and 38.07 g of oxygen. You cannot say that the formula is

$$K_{26.58}\,Cr_{35.35}\,O_{38.07}$$

because formulas should represent mole ratios rather than mass ratios.

The second step is to convert these masses to moles of atoms by dividing by the appropriate atomic weight.

$$K\ \frac{26.58\ \text{g}}{39.10\ \text{g mol}^{-1}} \quad Cr\ \frac{35.35\ \text{g}}{52.00\ \text{g mol}^{-1}} \quad O\ \frac{38.07\ \text{g}}{16.00\ \text{g mol}^{-1}}$$

which will give you mole ratios of

$$K_{0.6798}\,Cr_{0.6798}\,O_{2.379}$$

However, in an empirical formula, mole ratios must be expressed in small integers.

So in the third step, dividing through by the smallest number

$$K \frac{0.6798}{0.6798} \quad Cr \frac{0.6798}{0.6798} \quad O \frac{2.379}{0.6798}$$

gives $K_1Cr_1O_{3.5}$.

In many cases, this third step will give you the smallest possible integer ratio. At other times you will get a ratio which will obviously need to be multiplied by 2 or 3, or even 4 or 5 in order to have all integer subscripts.

$K_1Cr_1O_{3.5}$ should be multiplied by 2 to give $K_2Cr_2O_7$, which is the correct empirical formula.

The solution to this type of problem can be given in a table format, using the same three steps.

Element	**K**	**Cr**	**O**
Express % composition as part of 100 g.	26.58 g	35.35 g	38.07 g
Get mole ratio by dividing by atomic weight.	$\frac{26.58\ g}{39.10\ g\ mol^{-1}}$ = 0.6798 mol	$\frac{35.35\ g}{52.00\ g\ mol^{-1}}$ = 0.6798 mol	$\frac{38.07\ g}{16.00\ g\ mol^{-1}}$ = 2.379 mol
Divide by the smallest mole value.	$\frac{0.6798\ mol}{0.6798\ mol} = 1$	$\frac{0.6798\ mol}{0.6798\ mol} = 1$	$\frac{2.379\ mol}{0.6798\ mol} = 3.5$

Formula with Mole ratios: $K_1Cr_1O_{3.5}$

or multiplying by 2 to get the empirical formula: $K_2Cr_2O_7$

The advantage of making a table like this is that you organize the steps in the calculation. You can ensure that you do not inadvertently leave out one step of an element. If you don't get a reasonable answer for the empirical formula, check your arithmetic since this is a common mistake.

When you start empirical formula calculation problems, you should make a quick check that the percent by weight values do indeed add to 100% (within experimental limits). If you find a calculation, such as 60.0% carbon and 4.48% hydrogen which adds up to 64.5%, you can usually assume that the remaining 35.5% is oxygen. This happens because, unlike carbon and hydrogen, oxygen cannot be analyzed for directly. Hydrogen and carbon are converted into H_2O and CO_2, respectively. Similarly nitrogen, phosphorus, and sulphur can be converted into their oxides. Oxygen remains unanalyzed since it cannot be converted to its oxide.

In some problems, you may not be given weights of elements in the compound, or the mass per cent of the elements. Instead you'll be given analysis data featuring amounts of H_2O and/or CO_2. Such problems require some preliminary analysis before you can apply the same three steps tabulated above.

Study Section 3.6 in your textbook and work through Example 3.11, its practice exercise, and the following problems.

Problem: A 2.50 g sample of the compound nicotine was burned in oxygen to produce 6.78 g of CO_2, 1.94 g of H_2O, and 0.420 g of N_2. What is the empirical formula of nicotine?

Solution: The first step involves determining the mass of C, H, and N atoms in the nicotine sample.

$$\text{mass of C from } CO_2 = \quad 6.78\text{g } CO_2 \times \frac{12.01\text{g C}}{44.01\text{g } CO_2} = 1.85\text{g C}$$

$$\text{mass of H from } H_2O = \quad 1.94\text{g } H_2O \times \frac{2.016\text{g H}}{18.01\text{g } H_2O} = 0.217\text{g H}$$

Since all the carbon in the nicotine sample is converted to CO_2 and all the hydrogen to H_2O, the masses of the elements present in the 2.50 g sample are: C 1.85 g; H 0.217 g; and, by difference, N 0.433 g. (Note that 0.433 g of N_2 molecules is the same mass as 0.433g of N atoms.)

You can use this data to find the empirical formula of nicotine by turning mass ratios to mole ratios, and then to the simplest integer ratio.

Element	**C**	**H**	**N**
Mass ratio	1.85 g	0.217 g	0.433 g
Divide by atomic weight to get mole ratio.	$\frac{1.85\text{g}}{12.01\text{g mol}^{-1}}$ = 0.154 mol	$\frac{0.217\text{g}}{1.008\text{g mol}^{-1}}$ = 0.215 mol	$\frac{0.433\text{g}}{14.01\text{g mol}^{-1}}$ = 0.0309 mol
Divide by the smallest mole value.	$\frac{0.154}{0.0309} = 4.98$	$\frac{0.215}{0.0309} = 6.96$	$\frac{0.0309}{0.0309} = 1.0$

The empirical formula for nicotine is C_5H_7N.

Determining Molecular Formulas

If you are given the molar mass of an unknown compound or information that allows you to determine molar mass, you can convert an empirical formula to a molecular formula. This is because the molecular formula is an integer multiple of the empirical formula. Thus, the molecular molar mass is this same integer multiple of the empirical molar mass. Example 3.11 and its practice exercise demonstrate this procedure.

Chemical Reactions and Equations

Study Section 3.7 in your textbook. Note that the physical state of the compound is often included in the equation.

Balancing Chemical Equations

Sometimes equations can be balanced simply by looking at the molecules on either side. For others, several steps may be required. These processes are explained in Section 3.7. Go over Example 3.12 in your textbook. Practice is the best way to learn how to balance equations, so try appropriate examples at the end of Chapter 3. A word of warning: you will not be able to balance equations correctly unless the formulas involved are correct.

Mass Relationships in Chemical Reactions

Study Sections 3.8, 3.9, and 3.10 in your textbook. Make sure you work through all the examples in these sections as you come to them.

Stoichiometric Calculations

This is a very important part of Unit 1. You will be using your knowledge of chemical reactions to calculate the amount of material formed (**the product**) from a given amount of starting material (**the reactant**).

In stoichiometric calculations, the steps that you follow are:

- Write a balanced equation $A \longrightarrow 2B$
- Convert to $moles_A$ using molar mass $mass_A \longrightarrow moles_A$
- Convert to $moles_B$ using the mole ratio $moles_A \longrightarrow moles_B$
- Convert to $mass_B$ using molar mass $moles_B \longrightarrow mass_B$

You should realize by now how important it is to write balanced equations that give the mole ratios in the reactions. It is equally important to convert masses into moles and moles into masses in order to do stoichiometric calculations. The key "bridging" step always goes through moles, as shown in Figure 3.8 in your textbook. The best method of learning to do this type of problem is by practice. See the practice exercises and suggested supplementary exercises at the end of this unit.

Work through Examples 3.13 and 3.14 and their practice exercises in your textbook.

Limiting Reagents

If you are given two or more masses, you will first have to calculate which will be used up first (the limiting reagent), and which reactant or reactants are present in excess. Then the calculation to determine the amounts of products proceeds as before, based on the amounts of limiting reagent reactant.

Sometimes you are asked to calculate both the mass of product and the mass of the reagent in excess that remains after the reaction has gone to completion. Example 3.15 and its practice exercise show this. Remember to identify the limiting reagent before you start the main calculation. You may, as in Example 3.15, be asked to determine the limiting reagent as part of the calculation. In other cases, although the problem does not specifically ask this, it is necessary to determine the limiting reagent first (as shown in the following problem).

Problem: Find the number of grams of N_2 gas formed when 18.1 g of ammonia react with 90.4 grams of copper(II) oxide in the reaction

$$2NH_3(g) + 3CuO(s) \longrightarrow N_2(g) + 3Cu(s) + 3H_2O(g)$$

Solution: First determine the number of moles of each reactant:

$$\frac{18.1 \text{ g } NH_3}{17.0 \text{ g} \cdot \text{mol}^{-1}} = 1.06 \text{ mol } NH_3 \quad \frac{90.4 \text{ g } CuO}{79.5 \text{ g} \cdot \text{mol}^{-1}} = 1.14 \text{ mol CuO}$$

The mole ratio (the stoichiometric ratio in the equation) shows that you need three moles of CuO for every two moles of NH_3. Thus, to use up all the NH_3, you have to have

$$1.06 \times \frac{3}{2} = 1.59 \text{ moles of CuO}$$

There is *not* enough CuO available (only 1.14 moles) so CuO is the limiting reagent. The calculation to determine the mass of N_2 gas produced is thus based on the 1.14 moles of CuO present. (Complete the calculation and confirm the answer is 10.6 g of $N_2(g)$.)

Reaction Yields

You need to know the distinction between *theoretical yield*, *actual yield*, and *percent yield* as explained in Section 3.10. Example 3.17 is a calculation that uses these concepts. You may also be asked to determine how much reactant is necessary to obtain a certain mass of product *if the reaction is less than 100 percent efficient*. The following problem illustrates this.

Problem: How many grams of CuO must you start with to produce 65.5 g of nitrogen gas if the reaction

$$2NH_3(g) + 3CuO(s) \longrightarrow N_2(g) + 3Cu(s) + 3H_2O(g)$$

is only 86.0% efficient?

Solution: First calculate the required weight of CuO assuming the reaction is 100% efficient.

$$\frac{65.6\ \text{g N}_2}{28.0\ \text{g} \cdot \text{mol}^{-1}\ \text{N}_2} \times \frac{3\ \text{mol CuO}}{1\ \text{mol N}_2} \times 79.5\ \text{g} \cdot \text{mol}^{-1}\ \text{CuO}$$

$$= 5.59 \times 10^2\ \text{g CuO}$$

But this is if the reaction is 100% efficient. If the reaction is less than 100% efficient, you will need to start with *more* CuO to obtain the 65.6 g of nitrogen gas, that is,

$$5.59 \times 10^2\ \text{g} \times \frac{100}{86.0} = 6.50 \times 10^2\ \text{g CuO}.$$

Read the summary at the end of Chapter 3. Then try some of the stoichiometry problems from the textbook, as noted at the end of this unit.

Properties of Aqueous Solutions

Study Section 4.1 in your textbook. From this section you must know what is meant by the terms: solution, solute, solvent, electrolyte, nonelectrolyte, reversible, and chemical equilibrium. These terms will appear often throughout the course units. Do not memorize Table 4.1 in your textbook, but understand the distinction between *weak* and *strong electrolytes* and *nonelectrolytes*.

Types of Chemical Reactions

Study Sections 4.2, 4.3, and 4.4 in your textbook so that you can classify a reaction as one of the following:

- Precipitation
- Acid-base
- Oxidation-reduction

These sections go into a good deal of detail in describing these types of chemical reactions. You should read about these reactions and look at the excellent illustrative pictures. At this stage, however, you are not expected to know all of these chemical reactions. For example, if you saw the reaction

$$BiCl_3 + NH_3 + H_2O \longrightarrow$$

you would not be expected to know that the products are $Bi(OH)_3$ and NH_4Cl, that is,

$$BiCl_3 + NH_3 + H_2O \longrightarrow Bi(OH)_3 + NH_4Cl.$$

But you should be able to balance it once you know what the products are, that is,

$$BiCl_3 + 3NH_3 + 3H_2O \longrightarrow Bi(OH)_3 + 3NH_4Cl.$$

In more obvious cases, you may be asked to complete and balance a reaction, as in the following example.

$$Al(s) + O_2(g) \longrightarrow \textbf{?}$$

Here, a reasonable assumption is that this is an oxidation-reduction (“redox”) reaction and that the product is aluminum oxide.

$$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$$

Reactions involving ions in solution can be written as ionic equations. If the species that do not take part in the reaction, the **spectator ions**, are eliminated from both sides, the result is a net ionic equation that involves only the “participating ions.” Work through Examples 4.2, 4.3, and 4.6 in your textbook. Although you do not have to memorize oxidation number rules yet, you may find them useful in, for example, naming ionic compounds.

Stoichiometry of Solutions

Study Section 4.5 in your textbook.

Molarity

Be sure to memorize the definition of molarity, *M*, since molarity is a very common way of expressing the concentration of a solution.

Examples 4.7 and 4.8, and their practice exercises, demonstrate typical molarity problems.

Dilution of Solutions

The process of dilution will be covered in detail in the next course, CHEM 1523. Since you may find this useful in your laboratory course, you should study the material and make sure you can do calculations like those in Example 4.9, based on the equation

$$M_{\text{initial}}V_{\text{initial}} = M_{\text{final}}V_{\text{final}}$$

Density

To convert from mass to volume, you will need density values given in units of g/cm^3. These values will usually be provided for you.

Gravimetric, Acid-Base, and Redox Stoichiometry

Study Sections 4.6, 4.7, and 4.8 in your textbook. You do not need to learn the different
techniques described here, such as precipitating, filtering, drying, and titrating. You will encounter these procedures in your laboratory course. You should, however, be able to do calculations on the chemistry involved in gravimetric analysis and titrations.

Examples 4.10, 4.11, 4.12, and 4.13 will give you some practice. Also, do the practice exercises. You are not expected to balance the redox equations, except for simple ones that can be balanced by inspection. Read the summary at the end of Chapter 4.

Energy Changes in Chemical Processes

Study Sections 6.1 and 6.2 of your textbook. The difference between heat (heat energy or thermal energy) and temperature is important. Heat energy can be directly transferred from one object to another. When the heat energy is absorbed by the other object, then its total energy increases by an increase in its kinetic energy, or its potential energy, or both. The temperature of an object is a measurement of the average random or disorganized kinetic energy of its component particles. The rest of Chapter 6 of the textbook will be covered in CHEM 1523.

Self-Assessment

Having reached the end of Unit 1, you should be able to:

- Correctly use SI units, significant figures, and scientific notation.
- Explain the basic chemical terms introduced in the unit (e.g., element, compound, metal, nonmetal, metalloid, homogeneous, heterogeneous, atom, ion, molecule, isotope).
- Distinguish between examples of homogeneous and heterogeneous matter, mixtures, pure substances, compounds, and elements.
- Name and give formulas for simple inorganic compounds.
- Write balanced chemical equations.
- Carry out calculations involving moles, molecular weights, masses, and numbers of molecules.
- Calculate empirical formulas from mass composition data and vice versa, and relate these to molecular formulas.
- Perform stoichiometric calculations involving chemical reactions.
- Interrelate volume, molarity, and density of a solution.
- Describe types of energy and energy changes.

Practice Exercise 1

Finish working through Unit 1 before starting this practice exercise which is found listed under the Practice Exercises section of course. That is because all the material in Unit 1 is integrated throughout this exercise. The solutions to these problems will be provided once you have completed this practice exercise.

Before you try this practice exercise, you may wish to try some of the supplementary exercises suggested under the next heading. However, if you feel that you have a reasonable understanding of the material, you can begin the practice exercise immediately. If you have no difficulties with this practice exercise, begin Assignment 1. If you feel you need more practice, go to the following Suggested Supplementary Exercises before starting Assignment 1.

Before consulting the Solutions to Practice Exercise 1, you should make a serious attempt to solve each problem in Practice Exercise 1 by yourself, using as a guide similar examples given in the textbook and in this unit. If your attempt is serious, you will learn far more from the solutions provided.

Suggested Supplementary Exercises

At the end of each chapter in your textbook are further exercises on the material covered. We suggest that you do as many of these exercises as necessary. Doing such problems helps you understand and apply the principles involved in the concepts discussed. Fully worked solutions for all even-numbered problems are given in the *Student Solutions Manual*, by B. J. Cruickshank and Raymond Chang. Some of the even-numbered exercises at the end of the chapter have their answers in the back of the textbook. If you decide to try the odd-numbered problems and need help, consult your Open Learning Faculty Member. The odd-numbered exercises can usually be solved in a similar manner to adjacent even-numbered problems.

Important: You should make a serious attempt to solve a problem by yourself before looking at a worked solution.

From Chapter 1:

1.8, 1.12, 1.16, 1.18, 1.22, 1.26, 1.30, 1.32, 1.34, 1.36, 1.40, 1.46, 1.50, 152, 154, 1.58, 1.64, 1.72, 1.76, 1.80, 1.84, 196 (b, c)

From Chapter 2:

2.8, 2.14, 2.16, 2.18, 2.24, 2.26, 2.32, 2.34, 2.36, 2.38, 2.42, 2.44, 2.46, 2.56, 2.58, 2.60, 2.64, 2.68, 2.70, 2.72, 2.74, 2.78, 2.82, 2.90, 2.100, 2.102

From Chapter 3:

3.6, 3.8, 3.14, 3.16, 3.20, 3.22, 3.24, 3.28, 3.30, 3.40, 3.44, 3.46, 3.52, 3.54, 3.60, 3.64, 3.66, 3.70, 3.72, 3.80, 3.86, 3.88, 3.90, 3.92, 3.108, 3.110, 3.134

From Chapter 4:

4.8, 4.10, 4.16, 4.18, 4.22, 4.30, 4.50, 4.56, 4.60, 4.62, 4.68, 4.78, 4.80, 4.84, 4.88

Assignment 1

Now refer to your *Assignments* and complete Assignment 1. Consult your Course Guide for the week this assignment is due.

You may send the assignment to your Open Learning Faculty Member using the assignment tool in Blackboard or by mail with a Marked Assignment Form.

Be sure to keep a copy of the assignment—it will be useful if you wish to discuss your work with your Open Learning Faculty Member.