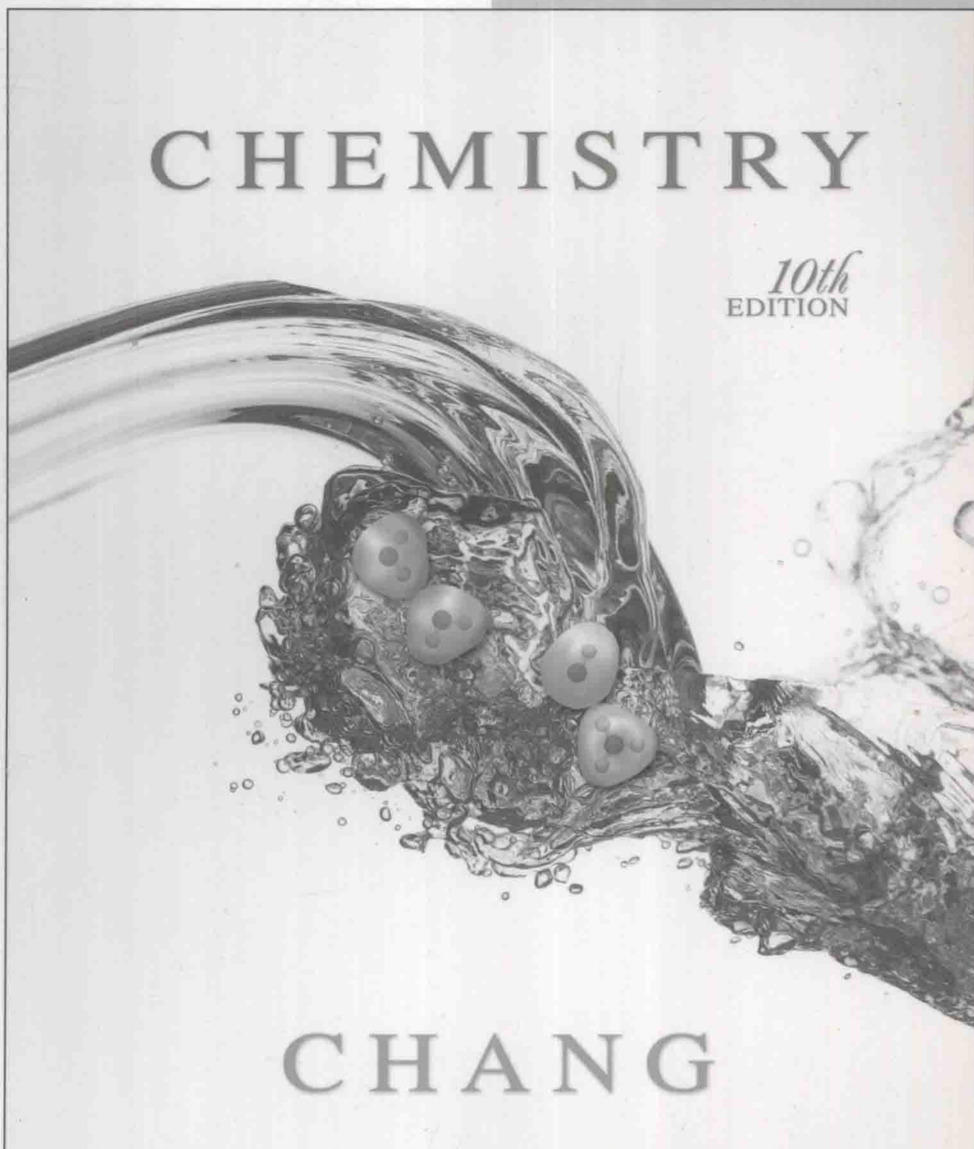
Student Study Guide

to accompany



Prepared by
Kim Woodrum

Student Study Guide

to accompany

Chemistry

Tenth Edition

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Student Study Guide to accompany
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RAYMOND CHANG

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Preface

General Chemistry is a course required by many different majors. Most students who are required to take this course are not chemistry majors. At most institutions, chemistry majors make up only a small percentage of the total enrollment. Chemistry is sometimes referred to as the "Central Science" because it is used in so many other areas. And while many students understand that the chemistry they learn will serve them in their chosen major, as I stand in front of my students, I know that many wonder why they are there. The civil engineer, for example, could be sitting there thinking, "This course is a waste of time. I'll never use this for my degree." To put energy into an endeavor, it is important to understand why. Why bother putting the 5-10 hours a week into a class? Even if you don't see the application of the material, right now, you need to know that the skills you will develop in this course will serve you in all your courses and beyond, even if the chemistry is never used again. The courses you are required to take were not put together lightly. Much thought and debate goes into deciding what courses a student must take. But what you develop in this course is much more than chemistry knowledge. It is a road (and for many students, their first road) on a journey to become a "problem solver". I find that some freshmen have entered college being pretty good memorizers and mimics. They can quote back definitions, they can follow a set of procedures, but if given a new challenge (a new problem to solve) which uses the same skills, they are lost.

Your time in General Chemistry as not only a time to learn theories and laws on the subject but also to develop problem solving skills. Let's compare being a good problem solver to being a good carpenter. A good carpenter learns how to use certain tools; a hammer, a saw, a sander, a planer. He knows what the tools do and when to use them. When the master carpenter builds a beautiful piece of furniture, he is not going to need someone to say, "Pick up a board. Now pick up the saw and cut 3-inches off one end…." The carpenter will use the tools and skills as he decides; based on the design he has planned. You, as a problem solver will need to do the same. In this course you must 1) learn how to use the tools and then 2) learn when these tools are useful so you can decide when to use them. You must get away from the notion that you must see an example just like an assigned problem in order to work the problem. You must also get away from thinking that a test should only consist of problems just like the ones you have already worked. Your instructor wants to test you on both the use of tools and whether you can pick them up and use them as necessary. They don't want to know if you merely memorized some facts and can mimic how someone else solves problems.

The Student Study Guide will summarize and emphasize points as discussed in Chang's Chemistry 9th Edition. In it, you will learn to use the "tools" of chemistry.

The icon to the left will point out the tools that must be mastered. As you master the use of the tools, you will then be able to develop problem solving skills. Often the tools needed are mathematical tools. There is no way around the fact that much of your success in General Chemistry is dependent on being able to use the pre-algebra and algebra skills you have already developed.

Each chapter contains:

- A summary of the material covered in the text, with tools indicated where appropriate. Some students may want to start here, reading through the summary to gain a different perspective of topics covered in the textbook. The summary content is built upon the work of Sharon Neal (8th Edition) and Ken Watkis (7th Edition.)

- A glossary of new terms used in the chapter, organized by association. This is the same list which is present in your textbook.
- A list of equations presented in the chapter.
- Worked Examples, with references to matched "Exercises & Problems."
- Exercises & Problems section. Some students may want to start here, and refer to the summary and Worked Examples, when needed.
- Practice test. To give you a good sense of how well you know the material, try the following:
    - Detach Practice Test pages which apply to the exam you will be taking.
    - Cut the questions apart along the dashed lines and randomly pick several from each chapter.
    - Find a quiet place, without notes, and work through the questions.
    - Grade it. The back of each question gives the chapter number. This will aide you in finding the answer in the appendix.
    - If you didn't do well, repeat after going over the ones you missed. *Seek out help, when needed.*

Thank you to Raymond Chang, and the McGraw-Hill team, for granting me this opportunity. Thanks to Kevin (my husband) and Logan, Jenni and Megan (my kids) for their patience with me as I spent too much time in front of the computer!

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SUMMARY

Introduction (Sections 1.2 – 1.3)

Chemistry & the Scientific Method. Chemistry is the study of matter and the changes matter undergoes. Matter is defined as anything that has mass and occupies space. The substances that make up matter have unique properties that chemists use to identify them. Chemists are concerned with developing tools to study matter and concepts to describe the properties of matter. They direct their efforts toward purposefully changing given forms of matter into new and different substances, and to the discovery of the properties and uses of these new materials. Chemists usually observe matter and the changes it undergoes using macroscopic (large enough to be seen using the naked eye) amounts of material. However, they interpret the properties and changes in matter in terms of atoms and molecules, microscopic entities that are too small to be seen.

Chemists use the scientific method to guide their search for new knowledge and to deduce facts about the microscopic world of atoms and molecules from macroscopic observations of reactions and properties. The scientific method can be broken down into four stages: observation and experiment, hypothesis formation, law development, and theory adoption. The first stage involves defining the problem clearly and suggesting a systematic, logical approach towards a possible solution. The investigator must decide what s/he is trying to find out and what approach to take. During experiments, observations and measurements are made and information collected about the phenomenon being studied. After significant amounts of data related to the phenomenon have been collected, the relationships indicated by the data are collected into a simple verbal or mathematical statement, often called a law. At this point, a hypothesis may be formulated to explain the facts summarized by the law. A hypothesis is only temporary and is meant to serve as a working model that is adjusted as new information about the phenomenon is discovered. As a hypothesis develops and survives many experimental tests, it becomes a theory. A theory is a unifying principle that explains a large body of observations and laws. Theories are tested using carefully devised experiments that confirm or disprove their predictions. A theory is important when it leads to predictions across a wide range of observations. The theory of gravitational motion is a famous example of this process. Holy men, stargazers and sailors mapped the position of the stars and planets for centuries. In the late seventeenth century, Newton summarized these observations using a single inverse squared law. The validity of the law was established by the use of Newton's theories to locate unknown planets. Newton's theory stood uncontested for more than one hundred years until Einstein's theory of general relativity explained its limitations around very dense objects.

Examples 1.1, 1.2

Exercises 1, 2

Matter and Its Properties (Sections 1.4 – 1.6)

Elements and Compounds. A pure substance is a form of matter that has a fixed composition and distinct properties. Examples are water, table salt, and aluminum. Just as each individual person has a set of unique characteristics, such as fingerprints, voiceprint and retinal pattern, each pure substance has characteristic properties, such as density, color, and hardness. There are two types of pure substances: elements and compounds. An element is a pure substance that cannot be decomposed (broken down) into simpler substances by chemical reaction. (One element may be converted to another by nuclear reactions, more about this in Chapter 23.) Elements are the building blocks of which all compounds are composed. Oxygen (O), mercury (Hg) and iron (Fe) are examples of elements. Table 1.1 below, lists common elements and their symbols. Memorize these elements soon unless your instructor has other plans. Compounds are pure substances that are composed of two or more elements in fixed proportions. Water is a compound that contains twice as much hydrogen as oxygen (H_2O). Compounds can be broken down into simpler substances by chemical reactions: electrolysis (a reaction induced by passing current) of water produces hydrogen and oxygen gas; decomposition (a reaction induced by heating) of chalk produces calcium oxide and carbon dioxide.

Table 1.1 Common Elements & Symbols

Symbol	Element	Symbol	Element	Symbol	Element
Na	sodium	C	carbon	O	oxygen
Al	aluminum	N	nitrogen	Cl	chlorine
Fe	iron	S	sulfur	I	iodine
Cu	copper	Si	silicon	F	fluorine
Au	gold	H	hydrogen	He	helium
Ag	silver	K	potassium	Ne	neon

Mixtures. When pure substances are combined, they form mixtures. Unlike elements and compounds, mixtures do not always consist of specific proportions of substances. For example, we can buy milk that contains >3.5% (whole), 2% (reduced fat) or 0% (skim) fat. All three of these are mixtures of water, lactose (sugar), protein and fat. Mixtures can be homogeneous or heterogeneous depending on the extent of intermingling of the pure components. Sugar water is a homogeneous mixture; a mixture which is uniform throughout. The sugar crystals dissolve in the water and disperse evenly throughout the water. Homogeneous mixtures are called solutions, especially when the mixture is fluid. Granite is a heterogeneous mixture of rocks, namely quartz, feldspar and mica. The component rocks are not evenly dispersed; the characteristics of a heterogeneous mixture. The pieces of the components in the mixture are large enough for us to see them by eye. A mixture can be separated into its components by physical methods, such as filtration.

Properties of Matter. Chemistry is the study of matter and the changes it undergoes. Changes in matter are detected by observing changes in matter properties. Physical properties are those that can be observed and measured without changing the identity or chemical composition of the substance. Physical properties include color, hardness, density, melting point and boiling point. Heating ice (solid) above its melting point produces water (liquid). Additional heating above the boiling point produces steam (gas), but all three forms are the same substance: H_2O. Solid, liquid, and gas are three states of matter. In a solid, particles of matter are close and move very little. (The particles will be described in Chapter 2.) The distance between particles <u>and</u> extent of particle motion increases a little in liquids, but increases substantially in gases. Observation of chemical properties can only occur as a substance is converted to another species (element or compound) in a chemical reaction. Chemical properties include reactivity to

Examples 1.3 - 1.4 Exercises 3. 4

acids, bases, hydrogen, oxygen or water. Potassium, a metal, reacts violently (bubbling and fire!) with water to produce potassium hydroxide, KOH, but gold, also a metal, is stable in water indefinitely. All properties of matter are either extensive or intensive. Extensive properties vary with the amount of matter being considered. For example, the more material there is, the larger the mass or volume. Intensive properties are independent of the amount of material. For example, the color of a small piece of pure gold is the same as the color of a large piece of gold.

Recording Measurements (Sections 1.7 – 1.8)

The International System of Units. Scientists and engineers have adopted a uniform, international system of units for labeling observations. SI (abbreviation for Le Systeme International) is a metric system built on a foundation of base units for fundamental phenomena such as mass, length (distance) and time. A list of the base SI units used frequently in this text is given in Table 1.2. Units for other phenomena can be derived (calculated) from the base units.

Table 1.2 Base SI Units

Quantity	Name of Unit	Symbol
mass	kilogram	kg
length	meter	m
time	second	s
temperature	Kelvin	K
amount	mole	mol

Base and Derived Units. The mass of an object is a measure of its resistance to external forces. The greater the mass, the less external forces will affect an object's motion. A tractor-trailer requires brakes that exert more friction (force) on the wheel than a compact car does, because it has a much larger mass. The SI base unit for mass is the kilogram (kg). It would make more sense if the base unit was the gram (no prefix), but grams are small relative to the size of people and most of our possessions, so the kilogram wins out. The terms mass and weight are often used interchangeably, although they actually refer to distinct properties. The mass is a measure of the amount of matter in an object and is the same everywhere in the universe. The weight is the force exerted on the object by gravity. The weight of an object on the moon is $1/6^{th}$ its weight on the earth because the moon's gravity is six times smaller than the earth's. If you weigh 150 lbs on earth, you would weigh only 25 lbs on the moon.

The volume of an object is a measure of the space it occupies. The larger the volume, the more space the object requires. The SI unit for volume is a derived unit. The volume of a cube or rectangular solid is the product of the length, width and height, all measured in meters.

$$\begin{aligned} V &= \text{length} \times \text{width} \times \text{height} \\ &= \quad m \times m \times m \quad = m^3 \end{aligned}$$

Of course, the units most scientists use for volume are based on the liter (L). The liter is defined as the volume of a cube that has 10 cm (= 0.1 m) edges.

$$V = 0.1\ m \times 0.1\ m \times 0.1\ m = 0.001\ m^3 = 1\ L$$

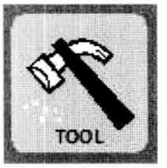
A convenient relationship to remember is 1 mL = 1 cm^3.

Example 1.5

Exercises 5, 6

The density of an object is the ratio of the mass to the volume, or the mass per unit volume, so it also has a derived unit. The ratio of the fundamental units for mass and volume is kg/m^3, but this unit is rarely used. Densities of liquids and solids are usually reported in g/mL or g/cm^3. Densities of gases are usually reported in g/L. Look back at the definition of intensive vs. extensive properties. Is density intensive or extensive? If you consider two pieces of gold, one twice as big as the other, will they have different densities? The mass of the two pieces doubles but so does the volume. Density = mass/volume so the value for density does not change. Density is intensive.

Conversion between	Tool
g ←→ mL (or cm^3)	density

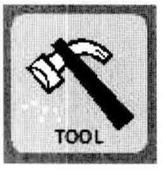


Unit Prefixes. Rather than report observations using very small or very large numbers, SI units are scaled (multiplied) by decimal prefixes. See Table 1.3. To make sense of this table, let us focus on the prefix kilo. Look at the table: What does kilo mean? Kilo means 1000. Therefore if something has as mass of 5 kg, it has a mass of 5000 g. If the prefix means something smaller than one, you can think of it in two different ways. For example, look at the table again: what does centi mean? Centi means 10^{-2} (one hundredth). If something is 1 cm long, it is 0.01 m, or it takes 100 cm to make 1 m. You need to memorize the prefix and the definition of each. The ones in boldface type are used most frequently in general chemistry.

Table 1.3 Commonly Used SI Prefixes

Prefix	Symbol	Definitions	
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	**k**	**1,000**	$\mathbf{10^{3}}$
deci	d	0.1	10^{-1}
centi	**c**	**0.01**	$\mathbf{10^{-2}}$
milli	**m**	**0.001**	$\mathbf{10^{-3}}$
micro	μ	0.000001	10^{-6}
nano	**n**	**0.000000001**	$\mathbf{10^{-9}}$
pico	p	0.000000000001	10^{-12}

Conversion between	Tool
g ←→ kg (for example)	Prefix Meanings (Table 1.3 above)



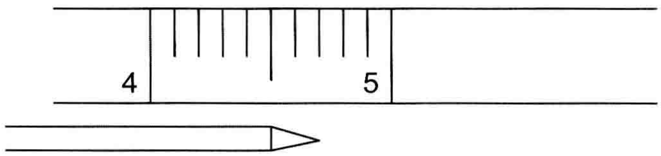
Significant Figures. It is not enough to label all measurements and calculations with appropriate units, the numerical values must be reported in a way that does not mislead the reader about the precision of the measurement. If a scientist reports that a sample weighs 0.12546 g, the number implies that the balance used could distinguish 0.1254 g from 0.1255 g. If that balance is only reliable to within 1 milligram, the value is being reported using too many digits, called significant figures. Scientists have adopted the convention of using all the certain digits and the first uncertain digit to report a measurement. Therefore, the scientist should report the weight as 0.1255 g, even if the

Examples
1.6 – 1.10

Exercises
8 – 13

display on the balance provides more digits. This notation informs the reader that the measurement is really 0.1255 $\pm$ 0.0001 g and the last digit is uncertain.

Consider the diagram below. The ruler is marked off in mm (that is, 1/10 of a cm). To measure the length of the pencil, we see that it is on or close to the mark "4.7 cm." We can then estimate one more place, by imagining 10 subdivisions between the 4.7 and 4.8. I'll estimate the pencil as just past the 4.7 mark and call it 4.71 cm. You might say, "It looks right on the line." If so, you would call it 4.70 cm.



We know two digits with certainty, but the last digit is uncertain but significant. It lets the reader of the measurement 4.71 cm know that the pencil is somewhere between 4.70 and 4.72 cm, that is, 4.71 ± 0.01 cm.

Algorithm for Identifying Significant Figures. To determine the number of significant digits in a written number use the following rules:

1. All nonzero digits are significant (this includes the 1 uncertain digit).
2. Zeros between nonzero digits are significant.
3. Leading zeros, to the left of the first nonzero digit, are **not** significant. These zeros only locate the decimal point and reveal nothing about certainty. Thus, 0.6 g and 0.0006 g both have 1 significant figure. Watch out. This is probably the most frequently forgotten rule.
4. Trailing zeros, to the right of the last nonzero digit, are definitely significant if the number contains a decimal point. Therefore, 1.0 mg has two significant figures, 0.500 mg has three significant figures and 600.00 g has five significant figures.
5. Trailing zeros are ambiguous if they appear in numbers without a decimal point. The number 450 g has 2 significant figures if it represents 450 $\pm$ 10 g and three if it represents 450 $\pm$ 1 g. To avoid this ambiguity, use scientific notation (see Math review in Appendix I). The number 450 g should be written as 4.5×10^2 g if it has two significant figures, and as 4.50×10^2, if it has three.
6. Numbers obtained from definitions are exact. For example, acceleration due to gravity (here on earth) is exactly 9.8 m/s^2. In a calculation, this value would be treated as if it had an infinite number of significant figures rather than two.

Algorithms for Calculating with Significant Figures. In many cases, the results we seek are calculated from the measurements we make. The precision (number of significant figures) is limited by the least precise measurement. For example, when we want to know the density of an object, we measure its mass and volume and then calculate their ratio. The number of significant figures depends on the accuracy of the balance and glassware we used to measure the mass and volume, not the number of digits our calculator displays when we calculate the ratio. Say an investigator needs to determine if an object is metal quickly and cheaply, its density is a good indicator. The mass of the object is 2.1572 g (5 significant figures) and the volume is measured as 1.66 mL (3 significant figures). The density of the object is 1.30 g/mL (3 significant figures) even though the calculator reads 1.29939759.

To determine the number of significant figures in a calculation use the following rules:

1. The number of significant figures in a product (x) or quotient (÷) is the number of significant figures in the least accurate factor
2. The number of significant figures in a sum (+) or difference (-) depends on the number of digits to the right of the decimal point (decimal digits) rather than the number of significant figures in the components. In general chemistry calculations we multiply and divide more frequently so students remember rule one and forget about this rule.

Addition	Subtraction
102.226 g	
2.51 g	102.25 g
+ 96 g	- 99.3 g
Sum = 200.736 g	Difference = 2.95 g
Correctly reported result = 201 g	Correctly reported result = 3.0 g

If adding two numbers which are in scientific notation but have different exponents, the numbers must be converted to a common exponent. For example,

$(4.5 \times 10^{-5}) + (1.5 \times 10^{-6})$	Convert one of the numbers to match the others exponent. We will convert the -6 to -5
$(4.5 \times 10^{-5}) + (0.15 \times 10^{-5})$	Now, add the coefficients (i.e., the numbers in front)
$(4.5 + 0.15) \times 10^{-5} = 4.65 \times 10^{-5}$	Use your rounding rules (least number of decimal places)
4.7×10^{-5}	This is YOUR ANSWER to the correct significant figures.

The significant figure convention was adopted to insure that measurement values reflect the precision of the measurement; the uncertainty associated with precise measurements is small. The precision says nothing about the accuracy of a measurement. Accurate measurements are close to the true value of the quantity under observation. Precise measurements can be inaccurate, as when a scale consistently reads 5 lbs too low. Accurate measurements can be imprecise, but this is less probable. The careful work that leads to accurate measurements, generally leads to small measurement uncertainty too.

Calculations with Observed Quantities (Section 1.9)

Dimensional Analysis. The key to dimensional analysis (also called the factor-label method) is to consider the units as an essential part of any quantity. We should not even think of a quantity without its units. The mass of a sample is never 5.00; it is 5.00 g. In everyday speech, it is acceptable to say that a man is "six-three" and leave the units implied. In scientific speech, he is always 6 feet 3 inches tall. The great benefit of this is that when quantities labeled with the correct units are subjected to calculations, the units are correctly computed as well.

Dimensional analysis uses two math facts.

1. A number divided by itself is one.
2. A number multiplied by 1 gives that number.

Let's use donuts as an example. A dozen donuts are equivalent to 12 donuts so;

$$\frac{1\text{ dozen donuts}}{12\text{ donuts}} = 1 \qquad \text{or} \qquad \frac{12\text{ donuts}}{1\text{ dozen donuts}} = 1 \qquad \text{(Fact 1)}$$

If you had 4 ½ dozen donuts and wanted to use dimensional analysis to determine the actual number of donuts, the conversion would look like this:

$$4.5\text{ dozen donuts} \times \frac{12\text{ donuts}}{1\text{ dozen donuts}} = 54\text{ donuts} \qquad \text{(Fact 2)}$$

Note that the units "dozen donuts" will now cancel out leaving you only the units "donuts."

> In general, conversion factors are simple ratios. Any time units of a given quantity are a ratio (as in m/s, g/mL, etc.) a corresponding conversion factor can be written. For example if the density of a substance is given as 1.78 g/mL, we can write the density as the following conversion factor:
>
> $$\frac{1.78\text{ g}}{1\text{ mL}}$$

Conversion factors can be used according to the following general form

$$\text{desired value} = \text{given value} \times \frac{\#\text{desired unit}}{\#\text{given unit}}.$$

Examples 1.11–1.15

Exercises 14 – 18

Some conversions, for example temperature conversions, are complicated by differences in the reference value. The conversion factor for donuts to dozen is a simple ratio because 0 donuts = 0 dozen. Freezing water is 32° on the Fahrenheit scale but 0° on the Celsius scale, so the conversion from °F to °C consists of a subtraction to adjust for the difference in the reference temperatures in addition to the ratio between units. The generalized formula becomes

$$\text{desired value} = (\text{given value} - \text{given reference}) \times \frac{\#\text{desired unit}}{\#\text{given unit}}$$

or

$$\text{desired value} = \text{given value} \times \frac{\#\text{desired unit}}{\#\text{given unit}} + \text{desired reference}$$

Example 1.16

Exercises 19, 20

The ratio of "desired units" to "given units" is called a conversion factor. The temperature of boiling water is 212 °F. The conversion factor is 5 °C for every 9 °F, or 5 °C per 9 °F. On the Celsius scale 212 °F is

$$(212\,^\circ\text{F} - 32\,^\circ\text{F}) \times \frac{5\,^\circ\text{C}}{9\,^\circ\text{F}} = 100\,^\circ\text{C}.$$

The formula for conversions from Celsius to Fahrenheit can be demonstrated as follows: to convert 100 °C to °F:

$$100\ ^\circ\mathrm{C} \times \frac{9\ ^\circ\mathrm{F}}{5\ ^\circ\mathrm{C}} + 32\ ^\circ\mathrm{F} = 212\ ^\circ\mathrm{F}$$

The formula for conversions from Celsius to Kelvin is

$$(100\ ^\circ\mathrm{C} + 273.15\ ^\circ\mathrm{C}) \times \frac{1\ \mathrm{K}}{1\ ^\circ\mathrm{C}} = 373.15\ \mathrm{K}$$

GLOSSARY LIST

chemistry	substance	extensive properties	density
scientific method	element	intensive properties	Kelvin
qualitative	compound	macroscopic properties	significant figures
quantitative	mixture	microscopic properties	accuracy
hypothesis	homogeneous mixture	international system of units	precision
law	heterogeneous mixture	weight	
theory	physical properties	mass	
matter	chemical properties	volume	

EQUATIONS

Algebraic Equation	English Translation
$d = \frac{m}{V}$	The density is the mass per unit volume. Usually in g/mL or g/cm^3.
$y\ ^\circ\mathrm{C} = (x\ ^\circ\mathrm{F} - 32\ ^\circ\mathrm{F}) \times \frac{5\ ^\circ\mathrm{C}}{9\ ^\circ\mathrm{F}}$	Used to convert from °F to °C. Formula simplified: °C=(°F-32)5/9
$y\ ^\circ\mathrm{F} = x\ ^\circ\mathrm{C} \times \frac{9\ ^\circ\mathrm{F}}{5\ ^\circ\mathrm{C}} + 32\ ^\circ\mathrm{F}$	Used to convert from °F to °C. Formula simplified: °F=(9/5)°C + 32
$y\ \mathrm{K} = (x\ ^\circ\mathrm{C} + 273.15\ ^\circ\mathrm{C}) \times \frac{1\ \mathrm{K}}{1\ ^\circ\mathrm{C}}$	Used to convert from °C to K. Formula simplified: K=°C + 273.15
$y\ ^\circ\mathrm{C} = (x\ \mathrm{K} - 273.15\ \mathrm{K}) \times \frac{1\ ^\circ\mathrm{C}}{1\ \mathrm{K}}$	Used to convert from K to °C. Formula simplified: °C=K – 273.15

WORKED EXAMPLES

EXAMPLE 1.1 Element, Compound or Mixture

Classify each of the following as an element, compound or mixture.

a. Aluminum foil
b. Table salt
c. Kool-aid

• Solution

At the early stages of your chemical education, you may not have the body of knowledge to address these topics properly. If it appears on the periodic table, it will be an element. To decide between a compound and a mixture is sometimes a bit more difficult.

a. Aluminum foil is simply made of aluminum flattened into sheets. This is an element (atomic number 13).
b. Table salt is a compound it consists of sodium and chlorine chemically combined with the formula NaCl.
c. If you think of how you make Kool-aid, you will decide it is a mixture. Sugar, water and flavoring are mixed together. Each substance retains its identity. No reaction occurs.

Work EXERCISES & PROBLEMS: 1

EXAMPLE 1.2 Homogeneous or Heterogeneous Mixture

Classify each of the following as a homogeneous or a heterogeneous mixture.

a. The beverage tea
b. Oil and water
c. Cow's milk
d. Wine

• Solution

Recall that homogeneous mixtures are uniform throughout, while heterogeneous mixtures have components that can be observed to be physically separate.

a. The mixture called tea is uniform throughout and no particles of tea can be observed (if there are no tea leaves). It is a homogeneous mixture.
b. Since oil floats on water, the oil and water components can be observed to be separate. Samples from the top part of the mixture have different properties from samples taken from the bottom. It is a heterogeneous mixture.
c. Cow's milk contains fats and solids suspended in water. On standing, the cream (fat) will rise to the top and the solids will settle. It is a heterogeneous mixture.
d. Wine contains ethyl alcohol, water, flavor components, color components, and other substances. The composition is uniform throughout. No solid particles are visible in the liquid; it is clear. It is a homogeneous mixture.

Work EXERCISES & PROBLEMS: 2

EXAMPLE 1.3 Chemical and Physical Changes

Classify the following changes as physical or chemical.

a. Solid shortening turns to liquid when heated
b. Gasoline burning in air

• Solution

a. Solid fats melt at a fairly low temperature. Melting is a physical change. Melting can be reversed by lowering the temperature. Removing the source of heat in this case will cause the fat to solidify.
b. Gasoline undergoes combustion with oxygen (burning) to produce carbon dioxide, water and heat. Removing the heat does not regenerate gasoline, so burning is a chemical change.

Work EXERCISES & PROBLEMS: 3

EXAMPLE 1.4 Chemical and Physical Properties

The following are properties of the element silicon; classify them as physical or chemical properties.
a. Melting point, 1410 °C
b. Reacts with fluorine to form silicon tetrafluoride
c. Gray color
d. Not affected by most acids

• Solution

Physical properties can be observed without a change in composition, while chemical properties describe reactions with other substances.
a. Melting involves a change in physical state but no chemical change. The melting point is a physical property.
b. This statement describes the change of silicon into another substance on reaction with fluorine. The reaction is a chemical property.
c. The color of a substance is a physical property. No change in composition occurs while observing the color.
d. The lack of reactivity with another substance or class of substances such as acids is a chemical property.

Work EXERCISES & PROBLEMS: 4

EXAMPLE 1.5 Density

A flask filled to the 25.0 mL mark contained 29.97 g of a concentrated salt–water solution. What is the density of the solution?

• Solution

The density (d) of an object or a solution is defined as the ratio of its mass (m) to its volume (V). Substituting the given quantities:

$$d = \frac{m}{V} = \frac{29.97\ \text{g}}{25.0\ \text{mL}} = 1.20\ \text{g/mL (rounded to 3 significant figures)}$$

Work EXERCISES & PROBLEMS: 5, 6

EXAMPLE 1.6 Scientific Notation

Write the following numbers using scientific (exponential) notation.
a. 7,620,000 b. 0.000495

• Solution

a. 7,620,000—Count the number of places that the decimal point must be moved to the *left* to give a number between 1 and 10, in this case, 7.62. Since this requires six places, the exponent of 10 must be 6. The value is 7.62×10^6.

b. 0.000495—Fractional numbers like this have negative exponents in scientific notation. Count the number of places that the decimal point must be moved to the *right* to give 4.95. Since this requires a move of four places, the exponent must be –4. The value is 4.95×10^{-4}.

Work EXERCISES & PROBLEMS: 7, 8

EXAMPLE 1.7 Math Operations with Scientific Notation

Multiply 4.0×10^4 by 3.5×10^{-6}.

• Solution

$$(4.0 \times 10^4)(3.5 \times 10^{-6})$$

Regroup so that the coefficients are separated from the exponentials, and add the exponents.

$$(4.0 \times 3.5) \times 10^{4+(-6)} = 14 \times 10^{-2} = 1.4 \times 10^{-1}$$

• Calculator Calculation

Most people use an electronic calculator for these types of calculations. On most, but not all, calculators, you would carry out the above multiplication problem in the following way:

$$(4.0 \times 10^4)(3.5 \times 10^{-6})$$

When entering the exponential terms, the exponents (4) and (–6) are entered as [EXP] then 4, and [EXP] then [+/–] 6. Therefore, 4.0×10^4 is entered as 4.0 [EXP] 4, and 3.5×10^{-6} is entered as 3.5 [EXP] [+/–] 6. Note that the × sign within the exponential must not be entered. Some calculators have an [EE] key for exponents.

The calculation can be carried out with the following sequence of key strokes:

4.0 [EXP] 4 [×] 3.5 [EXP] [+/–] 6 [=]

The calculator will display the answer as 0.14, or in exponential form as 1.4^{-01}, or as 1.4 –01 depending on the brand of calculator. Write your answer in exponential notation.