Appendix A

Scientific Units: The Metric System

Much of the work chemists do involves measuring — things like the mass, volume, or length of a substance.

Because chemists must be able to communicate their measurements to other chemists all over the world, they need to speak the same measurement language. This language is the SI system of measurement (from the French *Systeme International*), commonly referred to as the *metric system.* There are actually minor differences between the SI and metric systems, but for the most part, they're interchangeable.

The SI system is a decimal system. There are base units for mass, length, volume, and so on, and there are prefixes that modify the base units. For example, *kilo-* means 1,000; a kilogram is 1,000 grams, and a kilometer is 1,000 meters.

This appendix lists the SI prefixes, base units for physical quantities in the SI system, and some useful SI-English conversions.

SI Prefixes

Use Table A-1 as a handy reference for the abbreviations and meanings of various SI prefixes.

Table A-1 SI (Metric) Prefixes

Prefix	*Abbreviation*	*Meaning*
Tera-	T	1,000,000,000,000 or 10^{12}
Giga-	G	1,000,000,000 or 10^{9}
Mega-	M	1,000,000 or 10^{6}

(continued)

Table A-1 ***(continued)***

Prefix	*Abbreviation*	*Meaning*
Kilo-	K	1,000 or 10^3
Hecto-	H	100 or 10^2
Deka-	Da	10 or 10^1
Deci-	D	0.1 or 10^{-1}
Centi-	C	0.01 or 10^{-2}
Milli-	M	0.001 or 10^{-3}
Micro-	μ	0.000001 or 10^{-6}
Nano-	N	0.000000001 or 10^{-9}
Pico-	P	0.000000000001 or 10^{-12}

Length

The base unit for length in the SI system is the *meter.* The exact definition of meter has changed over the years, but it's now defined as the distance that light travels in a vacuum in $\frac{1}{299,792,458}$ of a second. Here are some SI units of length:

1 millimeter (mm) = 1,000 micrometers (μm)

1 centimeter (cm) = 10 millimeters (mm)

1 meter (m) = 100 centimeters (cm)

1 kilometer (km) = 1,000 meters (m)

Some common English to SI system length conversions are

1 mile (mi) = 1.61 kilometers (km)

1 yard (yd) = 0.914 meters (m)

1 inch (in) = 2.54 centimeters (cm)

Mass

The base unit for mass in the SI system is the *kilogram.* It's the weight of the standard platinum-iridium bar found at the International Bureau of Weights and Measures. Here are some SI units of mass:

1 milligram (mg) = 1,000 micrograms (μg)

1 gram (g) = 1,000 milligrams (mg)

1 kilogram (kg) = 1,000 grams (g)

Some common English to SI system mass conversions are

1 pound (lb) = 454 grams (g)

1 ounce (oz) = 28.4 grams (g)

1 pound (lb) = 0.454 kilograms (kg)

1 grain (gr) = 0.0648 grams (g)

1 carat (car) = 200 milligrams (mg)

Volume

The base unit for volume in the SI system is the *cubic meter.* But chemists normally use the *liter.* A liter is 0.001 m^3. Here are some SI units of volume:

1 milliliter (mL) = 1 cubic centimeter (cm^3)

1 milliliter (mL) = 1,000 microliters (μL)

1 liter (L) = 1,000 milliliters (mL)

Some common English to SI system volume conversions are

1 quart (qt) = 0.946 liters (L)

1 pint (pt) = 0.473 liter (L)

1 fluid ounce (fl oz) = 29.6 milliliters (mL)

1 gallon (gal) = 3.78 liters (L)

Temperature

The base unit for temperature in the SI system is *Kelvin.* Here are the three major temperature conversion formulas:

Celsius to Fahrenheit: °F = (9/5)°C + 32

Fahrenheit to Celsius: °C = (5/9)(°F–32)

Celsius to Kelvin: °K = °C + 273

Pressure

The SI unit for pressure is the *pascal*, where 1 pascal equals 1 newton per square meter. But pressure can also be expressed in a number of different ways, so here are some common pressure conversions:

1 millimeter of mercury (mm Hg) = 1 torr

1 atmosphere (atm) = 760 millimeters of mercury (mm Hg) = 760 torr

1 atmosphere (atm) = 29.9 inches of mercury (in Hg)

1 atmosphere (atm) = 14.7 pounds per square inch (psi)

1 atmosphere (atm) = 101 kilopascals (kPa)

Energy

The SI unit for energy (heat being one form) is the *joule*, but most folks still use the metric unit of heat, the *calorie*. Here are some common energy conversions:

1 calorie (cal) = 4.184 joules (J)

1 food Calorie (Cal) = 1 kilocalorie (kcal) = 4,184 joules (J)

1 British thermal unit (BTU) = 252 calories (cal) = 1,053 joules (J)

Appendix B

How to Handle Really Big or Really Small Numbers

Those who work in chemistry become quite comfortable working with very large and very small numbers. For example, when chemists talk about the number of sucrose molecules in a gram of table sugar, they're talking about a very large number. But when they talk about how much a single sucrose molecule weighs in grams, they're talking about a very small number. Chemists can use regular longhand expressions, but they become very bulky. It's far easier and quicker to use exponential or scientific notation.

Exponential Notation

In *exponential notation,* a number is represented as a value raised to a power of ten. The decimal point can be located anywhere within the number as long as the power of ten is correct. In *scientific notation,* the decimal point is always located between the first and second digit — and the first digit must be a number other than zero.

Suppose, for example, that you have an object that's 0.00125 meters in length. You can express that number in a variety of exponential forms:

0.00125 m = 0.0125×10^{-1} m, or 0.125×10^{-2} m, or 1.25×10^{-3} m, or 12.5×10^{-4} m, and so on.

All these forms are mathematically correct as numbers expressed in exponential notation. In scientific notation, the decimal point is placed so that there's one digit other than zero to the left of the decimal point. In the preceding example, the number expressed in scientific notation is 1.25×10^{-3} m. Most scientists automatically express numbers in scientific notation.

Here are some positive and negative powers of ten and the numbers they represent:

$1 \times 10^0 = 1$

$1 \times 10^1 = 10$

$1 \times 10^2 = 1 \times 10 \times 10 = 100$

$1 \times 10^3 = 1 \times 10 \times 10 \times 10 = 1{,}000$

$1 \times 10^4 = 1 \times 10 \times 10 \times 10 \times 10 = 10{,}000$

$1 \times 10^5 = 1 \times 10 \times 10 \times 10 \times 10 \times 10 = 100{,}000$

$1 \times 10^{10} = 1 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10 = 10{,}000{,}000{,}000$

$1 \times 10^{-1} = 1/10 = 0.1$

$1 \times 10^{-2} = 1/100 = 0.01$

$1 \times 10^{-3} = 1/1000 = 0.001$

$1 \times 10^{-10} = 1/10{,}000{,}000{,}000 = 0.0000000001$

Addition and Subtraction

To add or subtract numbers in exponential or scientific notation, both numbers must have the same power of ten. If they don't, you must convert them to the same power. Here's an addition example:

$(1.5 \times 10^3\,g) + (2.3 \times 10^2\,g) = (15 \times 10^2\,g) + (2.3 \times 10^2\,g) =$
$17.3 \times 10^2\,g$ (exponential notation) $= 1.73 \times 10^3\,g$ (scientific notation)

Subtraction is done exactly the same way.

Multiplication and Division

To multiply numbers expressed in exponential notation, multiply the coefficients (the numbers) and add the exponents (powers of ten):

$(9.25 \times 10^{-2}\,m) \times (1.37 \times 10^{-5}\,m) = (9.25 \times 1.37) \times$
$10^{(-2+-5)} = 12.7 \times 10^{-7} = 1.27 \times 10^{-6}$

To divide numbers expressed in exponential notation, divide the coefficients and subtract the exponent of the denominator from the exponent of the numerator:

$$(8.27 \times 10^{5}\,\text{g}) \div (3.25 \times 10^{3}\,\text{mL}) = (8.27 \div 3.25) \times 10^{5-3}\,\text{g/mL} = 2.54 \times 10^{2}\,\text{g/mL}$$

Raising a Number to a Power

To raise a number in exponential notation to a certain power, raise the coefficient to the power and then multiply the exponent by the power:

$$(4.33 \times 10^{-5}\text{cm})^{3} = (4.33)^{3} \times 10^{-5\times3}\,\text{cm}^{3} = 81.2 \times 10^{-15}\,\text{cm}^{3} = 8.12 \times 10^{-14}\,\text{cm}^{3}$$

Using a Calculator

Scientific calculators take a lot of drudgery out of doing calculations. They enable you to spend more time thinking about the problem itself.

You can use a calculator to add and subtract numbers in exponential notation without first converting them to the same power of ten. The only thing you need to be careful about is entering the exponential number correctly. I'm going to show you how to do that right now:

I assume that your calculator has a key labeled *EXP.* The EXP stands for $\times$ *10.* After you press the EXP key, you enter the power. For example, to enter the number 6.25×10^{3}, you type 6.25, press the EXP key, and then type 3.

What about a negative exponent? If you want to enter the number 6.05×10^{-12}, you type 6.05, press the EXP key, type 12, and then press the +/- key.

REMEMBER

When using a scientific calculator, *don't* enter the $\times$ *10* part of your exponential number. Press the EXP key to enter this part of the number.

Appendix C

Unit Conversion Method

You'll find that it's often unclear how to actually set up chemistry problems to solve them. A scientific calculator will handle the math, but it won't tell you what you need to multiply or what you need to divide.

That's why you need to know about the *unit conversion method,* which is sometimes called the *factor label method.* It will help you set up chemistry problems and calculate them correctly. Two basic rules are associated with the unit conversion method:

- **Rule 1:** Always write the unit and the number associated with the unit. Rarely in chemistry will you have a number without a unit. Pi is the major exception that comes to mind.
- **Rule 2:** Carry out mathematical operations *with* the units, canceling them until you end up with the unit you want in the final answer. In every step, you must have a correct mathematical statement.

How about an example so you can see those rules in action? Suppose that you have an object traveling at 75 miles per hour, and you want to calculate its speed in kilometers per second. The first thing you do is write down what you start with:

$$\frac{75\,\text{mi}}{1\,\text{hr}}$$

Note that per Rule #1, the equation shows the unit and the number associated with it.

Now convert miles to feet, canceling the unit of miles per Rule #2:

$$\frac{75\,\cancel{\text{mi}}}{1\,\text{hr}} \times \frac{5{,}280\,\text{ft}}{1\,\cancel{\text{mi}}}$$

Next, convert feet to inches:

$$\frac{75\,\cancel{\text{mi}}}{1\,\text{hr}} \times \frac{5{,}280\,\cancel{\text{ft}}}{1\,\cancel{\text{mi}}} \times \frac{12\,\text{in}}{1\,\cancel{\text{ft}}}$$

Convert inches to centimeters:

$$\frac{75\,\cancel{mi}}{1\,hr}\times\frac{5{,}280\,\cancel{ft}}{1\,\cancel{mi}}\times\frac{12\,\cancel{in}}{1\,\cancel{ft}}\times\frac{2.54\,cm}{1\,\cancel{in}}$$

Convert centimeters to meters:

$$\frac{75\,\cancel{mi}}{1\,hr}\times\frac{5{,}280\,\cancel{ft}}{1\,\cancel{mi}}\times\frac{12\,\cancel{in}}{1\,\cancel{ft}}\times\frac{2.54\,\cancel{cm}}{1\,\cancel{in}}\times\frac{1\,m}{100\,\cancel{cm}}$$

And convert meters to kilometers:

$$\frac{75\,\cancel{mi}}{1\,hr}\times\frac{5{,}280\,\cancel{ft}}{1\,\cancel{mi}}\times\frac{12\,\cancel{in}}{1\,\cancel{ft}}\times\frac{2.54\,\cancel{cm}}{1\,\cancel{in}}\times\frac{1\,\cancel{m}}{100\,\cancel{cm}}\times\frac{1\,km}{1000\,\cancel{m}}$$

Stop and stretch. Now you can start working on the denominator of the original fraction by converting hours to minutes:

$$\frac{75\,\cancel{mi}}{1\,\cancel{hr}}\times\frac{5{,}280\,\cancel{ft}}{1\,\cancel{mi}}\times\frac{12\,\cancel{in}}{1\,\cancel{ft}}\times\frac{2.54\,\cancel{cm}}{1\,\cancel{in}}\times\frac{1\,\cancel{m}}{100\,\cancel{cm}}\times\frac{1\,km}{1000\,\cancel{m}}\times\frac{1\,\cancel{hr}}{60\,min}$$

Next, convert minutes to seconds:

$$\frac{75\,\cancel{mi}}{1\,\cancel{hr}}\times\frac{5{,}280\,\cancel{ft}}{1\,\cancel{mi}}\times\frac{12\,\cancel{in}}{1\,\cancel{ft}}\times\frac{2.54\,\cancel{cm}}{1\,\cancel{in}}\times\frac{1\,\cancel{m}}{100\,\cancel{cm}}\times\frac{1\,km}{1000\,\cancel{m}}\times\frac{1\,\cancel{hr}}{60\,\cancel{min}}\times\frac{1\,\cancel{min}}{60\,s}$$

Now that you have the units of kilometers per second (km/s), you can do the math to get the answer:

0.033528 km/s

Note that you can round off your answer to the correct number of significant figures. Appendix D gives you details on how to do so, if you're interested. The rounded-off answer to this problem is

0.034 km/s or 3.4×10^{-2} km/s

Note that although the setup of the preceding example is correct, it's certainly not the only correct setup. Depending on what conversion factors you know and use, there may be many correct ways to set up a problem and get the correct answer.

Now I want to show you one more example to illustrate an additional point. Suppose that you have an object with an area of 35 inches squared, and you want to figure out the area in meters squared. Again, the first step is to write down what you start with:

$$\frac{35.0\,\text{in}^2}{1}$$

Now convert from inches to centimeters, but remember that you have to cancel inches *squared*. You must square the inches in the new fraction, and if you square the unit, you have to square the number also. And if you square the denominator, you have to square the numerator, too:

$$\frac{35.0\,\cancel{\text{in}^2}}{1}\,\frac{(2.54\,\text{cm})^2}{(1\,\cancel{\text{in}})^2}$$

Now convert from centimeters squared to meters squared in the same way:

$$\frac{35.0\,\cancel{\text{in}^2}}{1}\,\frac{(2.54\,\cancel{\text{cm}})^2}{(1\,\cancel{\text{in}})^2}\times\frac{(1\,\text{m})^2}{(100\,\cancel{\text{cm}})^2}$$

Now that you have the units of meters squared (m^2), you can do the math to get your answer:

$$0.0225806\ \text{m}^2$$

And if you want to round off your answer to the correct number of significant figures (see Appendix D for details), you get

$$0.023\ \text{m}^2 \text{ or } 2.3\times10^{-2}\ \text{m}^2$$

With a little practice, you'll really like and appreciate the unit conversion method. It got me through my introductory physics course!

Appendix D

Significant Figures and Rounding Off

Significant figures (no, I'm not talking about some supermodel) are the number of digits that you report in the final answer of the mathematical problem you are calculating. If I told you that one student determined the density of an object to be 2.3 g/mL and another student figured the density of the same object to be 2.272589 g/mL, I bet that you would naturally believe that the second figure was the result of a more accurate experiment. You might be right, but then again, you might be wrong. You have no way of knowing whether the second student's experiment was more accurate unless both students obeyed the significant figure convention. The number of digits that a person reports in his or her final answer is going to give a reader some information about how accurately the measurements were made. The number of the significant figures is limited by the accuracy of the measurement. This appendix shows you how to determine the number of significant figures in a number, how to determine how many significant figures you need to report in your final answer, and how to round your answer off to the correct number of significant figures.

Numbers: Exact and Counted Versus Measured

If I ask you to count the number of automobiles that you and your family own, you can do it without any guesswork involved. Your answer might be 0, 1, 2, or 10, but you would know exactly how many autos you have. Those are what are called *counted numbers*. If I ask you how many inches there are in a foot, your answer will be 12. That is an *exact number*. Another exact number is the number of centimeters per inch — 2.54. This number is exact by definition. In both exact and counted numbers, there is no doubt what the answer is. When you work with these types of numbers, you don't have to worry about significant figures.

Now suppose that I ask you and four friends to individually measure the length of an object as accurately as you possibly can with a meter stick. You then report the results of your measurements: 2.67 meters, 2.65 meters, 2.68 meters, 2.61 meters, and 2.63 meters. Which of you is right? You are all within experimental error. These measurements are measured numbers, and measured values always have some error associated with them. You determine the number of significant figures in your answer by your least reliable *measured* number.

Determining the Number of Significant Figures in a Measured Number

Here are the rules you need to determine the number of significant figures, or *sig. figs.*, in a measured number.

- **Rule 1:** All nonzero digits are significant. All numbers, one through nine, are significant, so 676 contains three sig. figs., 5.3×10^5 contains two, and 0.2456 contains four. The zeroes are the only numbers that you have to worry about.
- **Rule 2:** All of the zeroes between nonzero digits are significant. For example, 303 contains 3 sig. figs., 425003704 contains nine, and 2.037×10^6 contains four.
- **Rule 3:** All zeros to the left of the first nonzero digit are *not* significant. For example, 0.0023 contains two sig. figs. and 0.0000050023 contains five (expressed in scientific notation it would be 5.0023×10^{-6}).
- **Rule 4:** Zeroes to the right of the last nonzero digit are significant if there is a decimal point present. For example, 3030.0 contains five sig. figs., 0.000230340 contains six, and 6.30300×10^7 also contains six sig. figs.
- **Rule 5:** Zeroes to the right of the last nonzero digit are *not* significant if there is not a decimal point present. (Actually, a more correct statement is that I really don't know about those zeroes if there is not a decimal point. I would have to know something about how the value was measured. But most scientists use the convention that if there is no decimal point present, the zeroes to the right of the last nonzero digit are not significant.) For example, 72000 would contain two sig. figs and 50500 would contain three.

Reporting the Correct Number of Significant Figures

In general, the number of significant figures that you will report in your calculation will be determined by the *least* precise measured value. What values qualify as the least precise measurement will vary depending on the mathematical operations involved.

Addition and subtraction

In addition and subtraction, your answer should be reported to the number of decimal places used in the number that has the fewest decimal places. For example, suppose you're adding the following amounts:

2.675 g + 3.25 g + 8.872 g + 4.5675 g

Your calculator will show 19.3645, but you are going to round off to the hundredths place based on the 3.25, because it has the fewest number of decimal places. You then round the figure off to 19.36.

Multiplication and division

In multiplication and division, you can report the answer to the same number of significant figures as the number that has the *least* significant figures. Remember that counted and exact numbers don't count in the consideration of significant numbers. For example, suppose that you are calculating the density in grams per liter of an object that weighs 25.3573 (6 sig. figs.) grams and has a volume of 10.50 milliliters (4 sig. figs.). The setup looks like this:

(25.3573 grams/10.50 mL) × 1000 mL/L

Your calculator will read 2414.981000. You have six significant figures in the first number and four in the second number (the 1000 mL/L does not count because it is a exact conversion). You should have four significant figures in your final answer, so round the answer off to 2415 g/L. Only round off your final answer. Do not round off any intermediate values.

Rounding Off Numbers

When rounding off numbers, use the following rules:

- **Rule 1:** Look at the first number to be dropped; if it is 5 or greater, drop it and all the numbers that follow it, and increase the last retained number by 1. For example, suppose that you want to round off 237.768 to four significant figures. You drop the 6 and the 8. The 6, the first dropped number, is greater than 5, so you increase the retained 7 to 8. Your final answer is 237.8.
- **Rule 2:** If the first number to be dropped is less than 5, drop it and all the numbers that follow it, and leave the last retained number unchanged. If you're rounding 2.35427 to three significant figures, you drop the 4, the 2, and the 7. The first number to be dropped is 4, which is less than 5. The 5, the last retained number, stays the same. So you report your answer as 2.35.

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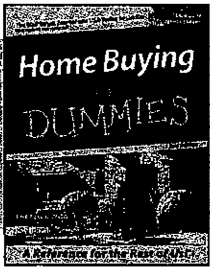

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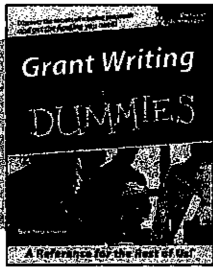

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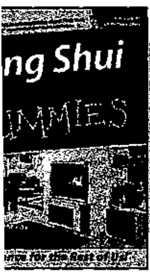

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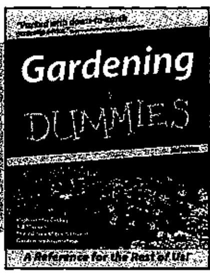

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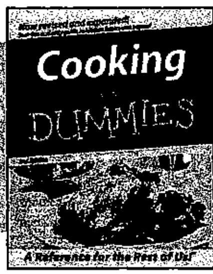

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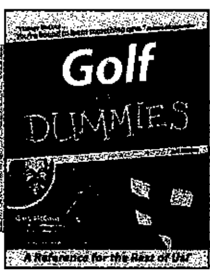

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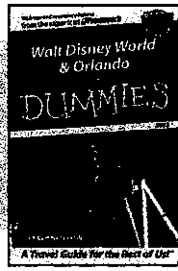

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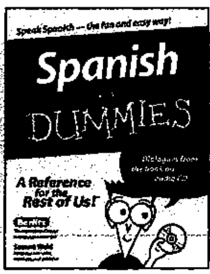

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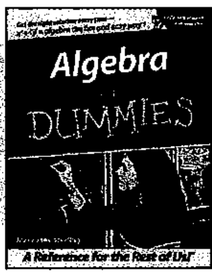

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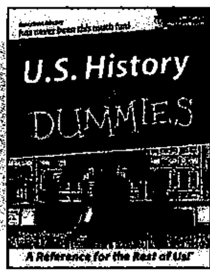

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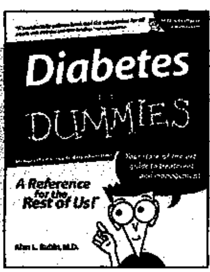

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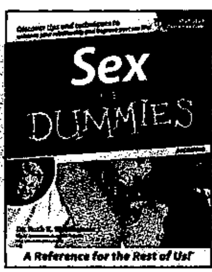

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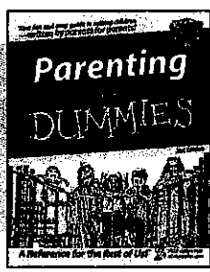

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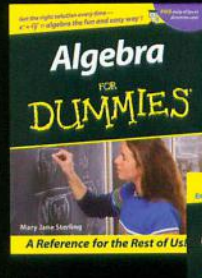

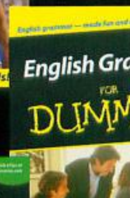

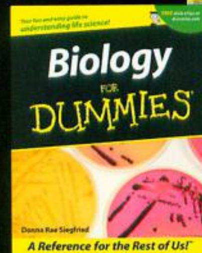

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