Chapter 1 – Keys to the Study of Chemistry

The first four chapters of this book provide a large number of definitions necessary to get through this course. You need to know these definitions, as well as the ones in the subsequent chapters. Like all of the sciences, chemistry has a unique vocabulary. Some technical words used in this discipline will have meanings similar to the same words used in everyday speech; others will have very different meanings. For example, to someone working in an appliance store, the word “element” might refer to the heating coil in a toaster. Needless to say, it means something entirely different to a chemist. Just as one cannot speak a foreign language without knowing its vocabulary, one cannot understand chemistry without learning its terminology. In this chapter we will cover some of basic concepts necessary to begin discussing chemistry. We start with the definition of chemistry your textbook uses: chemistry is the study of matter and its properties, then changes that matter undergoes, and the energy associated with those changes. Know the definitions of underlined words.

1.1 Some Fundamental Definitions

Matter is anything that has mass and occupies space (has volume). Mass is the quantity of a substance that is present. (Note that mass and weight are different. Weight is the force that gravity exerts on an object. Thus, the same object weighs more on a more massive planet because there is more gravity.) Matter can be made up of many different types of substances, which have defined, fixed compositions. The other major entity in chemistry is energy, which is the capacity to transfer heat or do work.

Two ways to describe matter are through their physical and chemical properties. Physical properties are ones that do not result in a change in the identity of a substance. Mass is a physical
property because weighing a substance doesn’t change what it is. In contrast, chemical properties (almost always called chemical reactions) can change the identity of a substance. Thus, this is typically the interaction of a substance with a different substance. For example, adding ammonia (a base) to vinegar (an acid) results in the formation of ammonium acetate, which is a new substance. (It is also a chemical property when the substances don’t react.)

Besides matter, energy is a major interest to chemists. Just like there are many ways to describe matter, many ways of describing energy also exist. Broadly speaking, energy may be divided into kinetic and potential energies. **Potential energy** arises from the position of an object relative to another object. For example, consider two magnets with north and south poles facing each other and relatively close together. You can feel them pulling together. That is potential energy. **Kinetic energy** is the energy associated with motion. When the magnets are allowed to come together, the potential energy is released as kinetic energy.

Your book goes through four examples of this, and I’ll discuss the fourth one briefly. Consider burning gasoline, which is the reaction of oxygen gas with an organic chemical. Both are at room temperature, but the product molecules (water and carbon dioxide (CO₂)) are more stable than the reacting molecules. The attractions between the atoms in the product are stronger than those in the reactants and burning releases that energy, which we experience as heat.

Matter may take one of three forms (phases or states) of interest to chemists: solids, liquids, and gases.

- **solid** - the shape and volume of the material are definite (the material is rigid)
- **liquid** - the volume is definite, but the shape may vary
- **gas** - both the shape and volume can vary (the material is compressible)

Thinking about solids, liquids, and gases is at the sub-microscopic level may help. Imagine
what a solid, liquid, and gas would look like if you could see elementary particles. The solid would have them packed very tightly together, probably in some sort of array. The bricks in a house wall provide an analogy. Liquids would be particles packed together nearly as tightly but moving past each other. A way of thinking about this is pouring a box of marbles onto the floor. A gas has a lot of separation between the molecules and would look like one of the lotto machines if the gas was represented by the ping pong balls. Generally speaking a solid will convert to a liquid and a liquid to a gas as the temperature of a sample increases (i.e. heat energy is added).

1.2 The Scientific Approach: Developing a Model

Read this on your own.

1.3 Chemical Problem Solving

Read this section on your own. The basic methodology here is very sound and what most chemists really do when solving problems, even if they don’t explicitly break it down into parts.

How many nickels are in a dollar and ten cents? The method for doing this is called dimensional analysis. It is simply the best way of solving problems in chemistry and every practicing chemist I know uses it. *You must use this method on tests.* Dimensional analysis starts with the premise that if you know the values you begin with and the units you will end with you can solve the problem. You start with the given information and try to divide out units to wind up with the desired units. The book provides several examples, I’ll provide two, starting with the question at the beginning of this paragraph.
Examples:

How many nickels are in $1.10?

Nickels = \((\$1.10) \left( \frac{100 \text{ cents}}{\$1} \right) \left( \frac{1 \text{ nickel}}{5 \text{ cents}} \right) = 22 \text{ nickels} \)

You probably do a process very similar to this in your head. Maybe breaking it into the number of nickels in $1.00 and in $0.10 separately, then adding them, but it’s still basically this method.

What is the length in feet of a 2.00 m long stick?

\[
\text{length (ft)} = (2.00 \text{ m}) \left( \frac{100 \text{ cm}}{1 \text{ m}} \right) \left( \frac{1 \text{ in}}{2.54 \text{ cm}} \right) \left( \frac{1 \text{ ft}}{12 \text{ in}} \right) = 6.56 \text{ ft} \quad \text{(remember SF)}
\]

Now, is this reasonable? If you remember that a meter is a little longer than a yard and a yard is 3 feet, you can see this answer is reasonable.

1.4 Measurement in Scientific Study

You should be familiar with the common units used in chemistry. With rare exceptions, chemists use metric based units such as the centimeter, liter, and gram. In particular, you should be familiar with the common prefixes in Table 1.2 (p. 15) and what they mean numerically. In particular you will see kilo-, centi-, milli-, micro-, and nano- often.

If you travel abroad, you will find almost all other nations use metric units for mass, distance, and volume. This is not simply a matter of memorizing conversion factors for working problems; it can help you avoid mistakes in calculations. For example, if someone calculated the weight in ounces of the textbook and told you it was about 0.10 ounces, you would know immediately the person had made a mistake and that, in all likelihood, s/he had divided the weight (in pounds) by 16 instead of multiplying by 16. You need to develop the same intuitive
feel for metric values.

For this reason, you should know one English-to-metric conversion factor for each type of unit. For example, there are 454 grams in a pound (mass), 1.05 quarts in a liter (volume), and 2.54 centimeters in an inch (length). At the beginning, most people who have trouble with metric units do so because they are unfamiliar. These conversion factors will allow you to quickly check a number for reasonableness.

1.4 Measurement in Scientific Study

A physical property that should be discussed separately is density. Density is the ratio of the mass of an object to its volume and is measured in g/mL for liquids and solids and g/L for gases. Confusing density and weight is almost as common as confusing mass and weight. A quart of water has more mass (weighs more) than a lead-fishing anchor, but has a lower density. That is, if you put the fishing weight on a scale it would weigh less than the water, but the anchor would sink to the bottom of the water because the lead anchor would weigh more per equivalent volume. Another way to think about this is: Which has more mass: a kilogram of lead or a kilogram of feathers? Answer: they have the same mass: 1 kilogram. The feathers would be in a much larger bundle, so would have a much lower density.

Example: A substance has a density of 7.33 g/cm³. You are given a cube of material that has a mass of 56.22 g. What are the dimensions of the cube in cm?

\[
\text{volume (cm}^3) = (56.22 \text{ g}) \left(\frac{1 \text{ cm}^3}{7.33 \text{ g}}\right) = 7.67 \text{ cm}^3
\]

\[
\text{length (cm)} = \sqrt[3]{7.67 \text{ cm}^3} = 1.97 \text{ cm}
\]

There are two temperature scales commonly used in chemistry. You have already encountered the Celsius scale either in a previous science class or in temperature displays on digital
outdoor signs. The Celsius thermometer uses the freezing and boiling points of pure water as its reference points. They are 0 °C and 100 °C, respectively. On the Fahrenheit scale (which is not used in chemistry), the melting point of pure water is 32 °F and water boils at 212 °F. Ever wonder why? Daniel Fahrenheit was a physicist and chose the normal body temperature as 96 °F and a saturated salt ice bath as 0 °F because the first thermometers used alcohol as the fluid and it boiled at a lower temperature than water. The odd numbers come about because he wanted to use a 12 point scale with 8 gradations between the major marks (96 points). This gives water a value of 32 °F. Improvements in the thermometer showed the body temperature actually to be 98 °F. Fahrenheit later invented the mercury thermometer. To convert between the scales the following equations are used:

\[ ^\circ C = \left(\frac{5}{9}\right) (^\circ F - 32) \quad \text{or} \quad ^\circ F = \frac{9}{5} (^\circ C) + 32 \]

Please note that the 32 °F used here is an exact number. Thus, it has an infinite number of significant digits (see below, p. 8 of notes).

The other temperature scale you’ll use in chemistry is the Kelvin scale. A Kelvin is the same size as a degree Celsius. The difference is that the Kelvin scale assigns the lowest possible temperature as 0 K. To convert from Celsius to Kelvins use the following formula:

\[ \text{Kelvin temperature} = ^\circ C + 273.15 \]

The Kelvin scale is an absolute scale so no degree sign is used. It is used in nearly every equation that requires a temperature in the calculation. Forgetting to use Kelvin temperatures is one of the most common errors made when working problems.
**Example**: Normal body temperature is 98.6 °F. What is normal body temperature in on the Celsius and Kelvin scales?

a) \[ \text{temp} = \left( \frac{5 \, ^\circ C}{9 \, ^\circ F} \right) (98.6 \, ^\circ F - 32 \, ^\circ F) = \left( \frac{5 \, ^\circ C}{9 \, ^\circ F} \right) (66.6 \, ^\circ F) = 37.0 \, ^\circ C \]

b) \[ \text{temp} = 37.0 \, ^\circ C + 273.15 \, K = 310.2 \, K \]

Many physical properties are temperature dependent. This means that the property changes as the temperature of the substance changes. For example, metal hardness decreases as temperature increases (i.e. a hot piece of metal deforms more easily than a cold one) and substances become superconducting only below a certain temperature.

Another method of classifying properties is as extensive or intensive. **Extensive properties** depend on the amount of the material present, while **intensive** ones don’t. Thus, mass and energy are extensive properties, while color and melting temperature are not. The labels here are not that important, but the concept behind them is. That is, sometimes you have to worry about how much of a substance is present (like when determining how much energy is needed to boil a beaker of water) and sometimes you don’t (when you find out the temperature at which boiling occurs).

1.5 **Uncertainty in Measurement: Significant Figures (SF)**

Aspects of significant figures are subtle, so you’ll probably need to work on this a little harder than some of the other material. I will try to always use the correct number of significant digits in all problems in the notes, homework, and in class. You should do the same thing to get in the habit.

**Significant digits** are those digits in a number that have physical meaning. If I told you I am 70 inches tall, you probably would just assume I meant near 70 inches, not exactly 70. Significant figures are a measure of how close the number is to the “real” value. I’m really 68 inches tall, so
the “7” in 70 would be significant, but the “0” would not be because it isn’t giving you any information. In effect, the statement “I am 70 inches tall.” means I am between 60 and 80 inches tall, while saying I am 68 inches tall means I am between 67 and 69 inches tall. In general, measured numbers have a certain number of significant digits that depend on the measuring device. Unless you are told otherwise, assume the number is accurate to ±1 in the last significant digit.

All digits are significant except i) zeros to the left of the first non-zero digit (e.g. 0.002 has 1 SF) and ii) numbers with zeros to the right of a non-zero digit but without a decimal point. (100 has only 1 SF) Note that zeros both to the right of a non-zero digit and the decimal point are always significant. (3.00 has 3 SF)

A way around situation (ii) is to use scientific notation. This expresses a number as a power of ten. Thus, if 100 has 1 significant digit, it is $1 \times 10^2$; if it has 2 significant digits it is $1.0 \times 10^2$; while if it has 3 significant digits it is $1.00 \times 10^2$.

Exact numbers have an infinite number of significant digits. Exact numbers include numbers obtained by counting. For example, a dozen eggs has exactly 12 eggs. It cannot have 13 eggs, 12.12 eggs, etc. Thus, there is no error in this number.

When multiplying or dividing numbers, the answer will have as many significant figures as the number with the fewest significant figures used in the calculation. Thus:

$$(2.032)(1.4) = 2.8 \text{ not 2.8448}$$

Why? We know the second value is between 1.3 and 1.5, right? Let’s calculate using those numbers.

$$(2.032)(1.3) = 2.6416$$

$$(2.032)(1.5) = 3.0480$$

Putting in the extra 3 digits implies you know the answer more accurately than you actually do.
The book goes over addition and subtraction and I’ll leave it to you to read this on your own. You’ll encounter it less than multiplication and division. **You must be able to do this.** Returning to my height: If you measured my height as 68” tall using a stick with only inch marks, my height in feet would be 5.7 ft, not 5.6667 ft. Why? Because if I were actually 5’8.5” tall, the ruler couldn’t measure the extra half inch, so my actual height would be larger than 5.6667 ft. The extra digits would suggest to the reader that my ruler was accurate enough to measure them.

**Accuracy** is a measure of how close a measured value is to the “real” or “true” value. **Precision** is a measure of how reproducible a measurement is (how close the measurements are to each other). For example, if I asked 3 of you to guess my height and you came up with 5’2”, 5’8”, and 6’4”, it would yield an accurate average (5’8”). The average would be imprecise, though, because the values were so far apart. On the other hand if 3 different people guessed, 5’5”, 5’6”, and 5’7” the average would be fairly precise, but at 5’6” it would be inaccurate. Percent error provides a guide to the accuracy of a particular measurement (or set of measurements). The formula for percent error is:

\[
\% \text{ error} = \frac{(\text{actual value}) - (\text{measured value})}{\text{actual value}} \times 100\%
\]

**Example:** What is the percent error in my height if three estimates were 5’5”, 5’6”, and 5’7” and the true value is 5’8”?

First you must convert height to inches (or fractional feet if you’re daring).

Height = \((1/3)(65” + 66” + 67”)) = 66”

\[
\% \text{ error} = \frac{68” - 66”}{68”} \times 100\% = 3\%
\]

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