Chapter 1 – Essential Ideas

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 chemistry 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 the basic concepts necessary to begin discussing chemistry. **Know the definitions of underlined words**.

1.1 Chemistry in Context

We begin with the definition of chemistry your textbook uses: chemistry is the study of composition, properties, and interactions of matter.

Scientific laws are statements that summarize a vast number of scientific observations and describe or predict some aspect of the natural world. They are facts, not explanations, usually made as very simple declarative sentences. For example, the Law of Conservation of Mass states that matter cannot be created or destroyed in chemical processes. There is no explanation, it's just a statement of fact. In contrast, theories are well-substantiated, comprehensive, testable explanations of a particular aspect of nature. These can change, sometimes quite significantly, over time. In contrast, laws either don't change or require only minor adjustments. For example, the addition of "in chemical processes" came about because of the transmutation of elements in nuclear processes, but in a chemical reaction, you don't ever observe mass increasing or decreasing. Hypotheses are tentative explanations of observations that act as guides for gathering and checking information. They are generally refined and either discarded or become theories after enough testing.

Read the rest of this section on your own and know the difference between the 3 domains.

1.2 Phases and Classification of Matter

Matter is anything that has mass and occupies space (has volume). 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 takes the shape of its container

gas - both the shape and volume follow that of the container (the material is compressible)

Thinking about solids, liquids, and gases 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).

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.) There is an important law in chemistry, the law of conservation of matter, which states that the sum of the masses of the reacting species in a chemical reaction equals the sum of the masses of the products. Your book phrases it differently, but this version may be easier to follow.

Matter can be made up of many different types of substances, which have defined, fixed compositions. Substances can be divided into two broad categories: pure and mixtures. A pure substance is composed of a single component and cannot be divided into parts by physical processes. 24-Carat gold and distilled water are pure substances. They differ in that the gold is an element, while water is a compound. Elements are pure substances that consist of only one type of atom and cannot be broken into simpler substances. (Atoms are the smallest particles of an element.) Thus, they are the building blocks of all complex materials. There are over 110 known elements, although only about 90 occur naturally on Earth. **You should know the names and symbols of the first 30 elements because you will use them throughout the course.** You will find a few others (e.g. Br = bromine) commonly used as well. Compounds are pure substances made up of atoms of two or more different elements chemically bonded together. Each unit of a compound has exactly the same number of each type of atom attached in the same way. Compounds can be broken up into their constituent elements by chemical processes.

In contrast, a mixture is a substance that is composed of two or more pure substances that can be separated by physical processes. 18-Carat gold, soda, and milk are mixtures. Mixtures can further be broken down into two categories: homogeneous and heterogeneous. Those that are distributed uniformly throughout and visibly appear uniform are called homogeneous mixtures or solutions. 18-Carat gold, soda and a cup of coffee are examples of solutions. An important feature of mixtures is that they can be separated into their constituent components by physical means (e.g. water can be separated from coffee by evaporation). Further, it is important to remember that solutions may involve all three phases of matter. For example, air is a gaseous solution of oxygen, nitrogen, and other gases, while brass is a solid solution of zinc and copper. Heterogeneous mixtures would then be those whose compositions vary throughout the mixture. Trail mix would be an example of a heterogeneous mixture. If you withdrew two scoops of trail mix and counted the number of each ingredient, you would very likely get two different amounts of each ingredient.

1.3 Physical and Chemical Properties

Two ways to describe matter are through their physical and chemical properties. Physical properties are ones that do not change the identity of a substance. Mass is a physical property because weighing a substance doesn't change what it is. In contrast, chemical properties (usually chemical reactions or classes of 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.)

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 pot of water) and sometimes you don't (when you find out the temperature at which boiling occurs).

1.4 Measurements

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 1.3 (pp. 30-31) and what they mean numerically. 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. Using the metric system helps 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.

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 ºC and 100 ºC, respectively. On the Fahrenheit scale (which is not used in chemistry), the melting point of pure water is 32 ºF and water boils at 212 ºF. Ever wonder why? Daniel Fahrenheit was a physicist and chose the normal body temperature as 96 ºF and a saturated salt ice bath as 0 ºF 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 freezing point value of 32 ºF. Improvements in the thermometer showed the body temperature actually to be 98 ºF. Fahrenheit later invented the mercury thermometer. To convert between the scales, the following equations are used:

$$^{\circ}\text{C} = \left(\frac{5\,^{\circ}\text{C}}{9\,^{\circ}\text{F}}\right)(^{\circ}\text{F} - 32\,^{\circ}\text{F}) \qquad \text{or} \qquad ^{\circ}\text{F} = {}^{\circ}\text{C}\left(\frac{9\,^{\circ}\text{F}}{5\,^{\circ}\text{C}}\right) + 32\,^{\circ}\text{F}$$

Please note that the 32 ºF 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:

Kelvin temperature = ºC + 273.15 K

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.

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/cm^3. You are given a cube of material that has a mass of 56.22 g. What are the dimensions of the cube in cm?

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

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

1.5 Measurement Uncertainty, Accuracy and Precision

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 1×10^2; if it has 2 significant digits it is 1.0×10^2; while if it has 3 significant digits it is 1.00×10^2.

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 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 them 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 am 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:

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

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”

$$\%\ \text{error} = \frac{68'' - 66''}{68''} \times 100\% = 3\%$$

1.6 <u>Mathematical Treatment of Measurement Results</u>

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 we will use for determining this is called dimensional analysis. It is simply the best way of solving problems in chemistry and every practicing chemist I know uses it. ***<u>You must use this method on tests</u>***. 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 end with the desired units. The book provides several examples, I’ll provide three, 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 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?

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 ft (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.

Normal body temperature is 98.6 ºF. What is it on the Celsius and Kelvin scales?

a) temp = $\left(\frac{5\,^{\circ}\text{C}}{9\,^{\circ}\text{F}}\right)$(98.6 ºF – 32 ºF) = $\left(\frac{5\,^{\circ}\text{C}}{9\,^{\circ}\text{F}}\right)$(66.6 ºF) = 37.0 ºC

b) temp = 37.0 ºC + 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.