

CHM 101 GENERAL CHEMISTRY

FALL QUARTER 2008

Section 2

Lecture Notes – 9/29/2008

(last revised: 9/30/08)

1.5 Significant Figures

- Nonzero Integers: always significant (0.000**251**)
- Leading Zeroes: never significant (0.000**251**)
- Captive (Embedded) Zeroes: always significant (**204.502**)
- Trailing Zeroes in a number without a decimal point: never significant (**527000**)
- Trailing Zeroes in a number with a decimal point: always significant (**260.0**)
- Exact Numbers have an infinite number of significant figures.

Counts: (**3** apples, **17** peanuts, **25** students)

Numbers in formulas: ($c = 2\pi r$)

Some Definitions: (**1** inch = 2.54 cm)

- Scientific Notation:

The digits in the number multiplied by the power of 10 are all significant: (**7.213** x 10⁶)
- When you are multiplying and dividing, the simple rule of thumb is that the result contains the same number of significant digits as the least significant factor in the calculation.
- When you are adding and subtracting, you need to consider two cases:
 - 1) If all the numbers involved in the addition or subtraction contain decimal points, you must determine

which of the numbers has the fewest significant figures to the right of its decimal point. Your result will have this same number of significant figures to the right of its decimal point. (Example: $17.5 + 23.29 = 40.79 = 40.8$)

2) If one or more of the numbers in the addition or subtraction lacks a decimal point, you must determine which of these numbers has the greatest number of (non-significant) trailing zeroes. Your result will have this same number of trailing zeroes. (Examples: $1750 + 2329 = 4079 = 4080$; $22 - 17.8 = 4.2 = 4$)

- When performing a computation, carry all the digits your calculator gives you, then round the final result. If the first non-significant digit is 0-4, simply truncate all the non-significant digits. If it is 5-9, increase the rightmost significant digit by 1 after you truncate the non-significant digits. (e. g., 7.316793 becomes 7.3, 2.5469 becomes 2.55.)
- See Sample Exercises 1.4 a-c.

1.9 Classification of Matter

- Matter: Anything that occupies space and has mass.
- States of Matter:

Solid: rigid, with a fixed volume and shape.

Liquid: has a fixed volume, but no specific shape – will assume the shape of its container.

Gas: no fixed volume or shape – will take on the shape and volume of its container.

- Liquids and solids are said to be condensed and will compress only slightly when put under high pressure. Neighboring molecules in a liquid or solid can be regarded as “touching” each other.
- Neighboring molecules in a gas are relatively far apart, and gases are highly compressible.
- Pure substances have fixed compositions: Examples: Nitrogen gas in a cylinder, distilled water, a silicon wafer.

- Mixtures have compositions that can and do vary. They can be either:

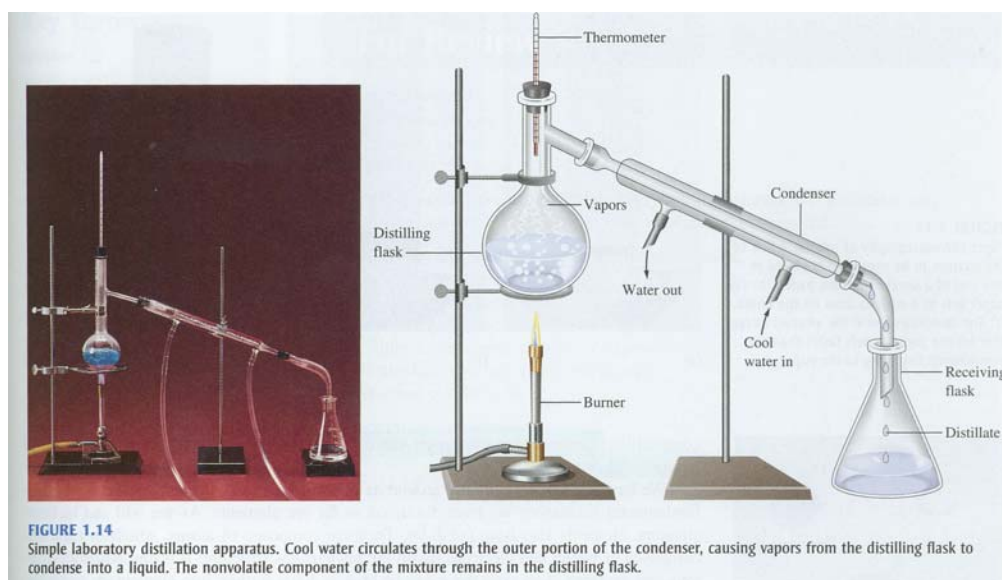
Heterogeneous mixtures have visibly distinguishable parts. (Examples: Wet sand is a mixture of sand and water. A granite boulder is a mixture of small crystals of mica, feldspar, quartz, and other minerals.)

Homogeneous mixtures appear uniform; their parts are visibly indistinguishable. Another name for a homogeneous mixture is a solution. (Examples: Brass is a solid solution of copper and zinc. Air can be regarded as a gaseous solution of nitrogen, oxygen, argon, and trace gases. Sweetened tea can be regarded as a solution of water, sugar, and substances extracted from tea leaves by hot water.)

Compositions can vary among similar mixtures. Consider a cup of tea sweetened with one spoonful of sugar versus another cup sweetened with two spoons of sugar.

- Mixtures can be separated by physical methods. Examples:

Distillation: When a solution is boiled, the most volatile component will vaporize first and can be collected in a separate vessel.



Filtration: An easy way to separate a mixture of a solid and a liquid.

Chromatography: Pass a mobile phase (a liquid or gaseous solution) through a stationary phase (a porous solid). The various components of the solution will pass through the stationary phase at different rates and thus can be separated.

- Pure substances are either pure elements or chemical compounds.

Pure elements, such as a gold ingot, or nitrogen gas in a cylinder, cannot be separated (except, perhaps, by nuclear processes).

Chemical compounds can be separated by chemical processes into their constituent elements. Examples: Water can be broken down by electrolysis into hydrogen gas and oxygen gas. Common salt (sodium chloride) can be separated into sodium metal and chlorine gas.

1.6 Dimensional Analysis

- Dimensional Analysis (also known as the Unit Factor Method) is useful for many kinds of scientific computations. We will use it here for conversions of quantitative observations (a number and a unit) from one system of units to another.

- Here is a selection of conversion factors:

TABLE 1.4 English–Metric Equivalents

Length	1 m = 1.094 yd 2.54 cm = 1 in
Mass	1 kg = 2.205 lb 453.6 g = 1 lb
Volume	1 L = 1.06 qt 1 ft ³ = 28.32 L

- Many others are listed in Appendix A6.
- Significant Figures: Read the "1's" in Table 1.4 as having the same number of significant figures as the equivalent quantities on the opposite sides of the equations: e. g., 1.000 m = 1.094 yd; 2.54 cm = 1.00 in.
- Example from the text: How many inches long is a pin measuring 2.85 cm?

Step 1: Pick the conversion factor that relates inches to centimeters (Refer to Table 1.4 or Appendix A6):

$$2.54 \text{ cm} = 1 \text{ in} \quad (1)$$

Step 2: Note that if you multiply centimeters by (in/cm), you get inches. So divide both sides of (1) by 2.54 cm.

$$\frac{2.54 \text{ cm}}{2.54 \text{ cm}} = 1 = \frac{1 \text{ in}}{2.54 \text{ cm}} \quad (2)$$

The right hand side of **Error! Reference source not found.** is called a *unit factor*, because it is equal to 1. Note that it has the desired dimensions, (in/cm).

Step 3: Now you can multiply 2.85 cm by this unit factor. The result is the same length, only expressed in inches:

$$2.85 \text{ cm} = 2.85 \text{ cm} \times \frac{1 \text{ in}}{2.54 \text{ cm}} = \frac{2.85}{2.54} \text{ in} = 1.12 \text{ in} \quad (3)$$

- Another example (Sample Exercise 1.5 from your text): How many centimeters long is a pencil measuring 7.00 in?

Step 1: The conversion factor is the same one **Error! Reference source not found.** we used on the previous problem.

Step 2: This time we want to multiply inches by (cm/in) to get centimeters, so we divide both sides of (1) by 1 in.

$$\frac{1 \text{ in}}{1 \text{ in}} = 1 = \frac{2.54 \text{ cm}}{1 \text{ in}} \quad (4)$$

This time our unit factor has the dimensions, (cm/in).

Step 3: Now we multiply 7.00 in by this new unit factor to get the length of the pencil in cm:

$$7.00 \text{ in} = 7.00 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 7.00 \times 2.54 \text{ cm} = 17.8 \text{ cm} \quad (5)$$

- Please be aware of what we did not do in the above two examples. We did not have to remember whether to multiply by 2.54 or divide by 2.54 in either of the two above examples. Instead, by paying attention to the dimensions and making sure

that the unwanted ones canceled out, we were able to figure that we should divide in the first example and multiply in the second.

- Another example, like Sample Exercise 1.8 in your text, but in reverse. You order a car for European delivery, but equipped for use in the USA, so the odometer reads miles, and the speedometer reads miles per hour, but speed limits and distances are expressed on the road signs in km/hr and km.
- The posted speed limit is 90 km/hr. How fast is that in miles per hour?

Step 1: We know the following equivalence statements:

$$1 \text{ km} = 1000 \text{ m} \quad (\text{a})$$

$$1 \text{ m} = 1.094 \text{ yd} \quad (\text{b})$$

$$1760 \text{ yd} = 1 \text{ mi} \quad (\text{c})$$

Step 2: We need to convert kilometers into meters, then meters into yards, and finally yards into miles. (We do not need to convert hours.) Thus we derive the following unit factors:

$$1 = \frac{1000 \text{ m}}{1 \text{ km}} \quad (\text{a}')$$

$$1 = \frac{1.094 \text{ yd}}{1 \text{ m}} \quad (\text{b}')$$

$$1 = \frac{1 \text{ mi}}{1760 \text{ yd}} \quad (\text{c}')$$

Step 3: Now we multiply 90 km/hr by the appropriate unit factors, making sure all the unwanted units cancel out:

$$90 \frac{\text{km}}{\text{hr}} = 90 \frac{\cancel{\text{km}}}{\text{hr}} \times \frac{1000 \cancel{\text{m}}}{1 \cancel{\text{km}}} \times \frac{1.094 \cancel{\text{yd}}}{1 \cancel{\text{m}}} \times \frac{1 \text{ mi}}{1760 \cancel{\text{yd}}}$$

Step 4: Perform the final computation:

$$90 \frac{\text{km}}{\text{hr}} = \left(\frac{90 \times 1000 \times 1.094}{1760} \right) \frac{\text{mi}}{\text{hr}} = 57 \frac{\text{mi}}{\text{hr}}$$

- Now we are driving down the road with our speedometer hovering at 57 mi/hr so we can obey the 90 km/hr speed limit. Our owner's manual directs us to get the car serviced after we have driven 500 miles. How many kilometers can we go before having this service?

Step 1: Using the same equivalence statements as before, we derive the following unit factors. Notice that we inverted these vs. the previous ones:

$$1 = \frac{1760 \text{ yd}}{1 \text{ mi}} \quad (\text{c}'')$$

$$1 = \frac{1 \text{ m}}{1.094 \text{ yd}} \quad (\text{b}'')$$

$$1 = \frac{1 \text{ km}}{1000 \text{ m}} \quad (\text{a}'')$$

Step 2: Now we multiply these by 500 mi:

$$500 \text{ mi} = 500 \cancel{\text{mi}} \times \frac{1760 \cancel{\text{yd}}}{1 \cancel{\text{mi}}} \times \frac{1 \cancel{\text{m}}}{1.094 \cancel{\text{yd}}} \times \frac{1 \text{ km}}{1 \cancel{\text{m}}}$$

Step 3: And do the computation:

$$500 \text{ mi} = \frac{500 \times 1760}{1.094} \text{ km} = 804 \text{ km}$$

1.7 Temperature Conversion

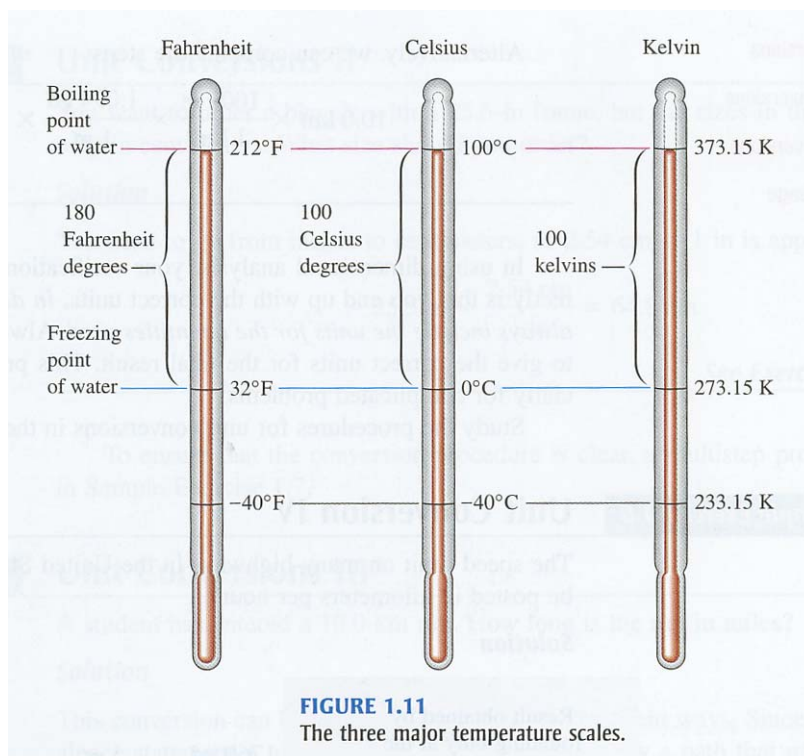
- The three major temperature scales are

Fahrenheit: Part of the system of English Units and commonly used only in the USA.

Kelvin: The fundamental temperature scale in SI (Systeme International).

Celsius: The common temperature scale for most of the world.

- The essential relationships among these three scales are shown in the following figure (Figure 1.11 in your text)



- On the Fahrenheit Scale, water freezes at 32 °F (degrees Fahrenheit) and boils at 212 °F (at sea level atmospheric pressure).
- The corresponding temperatures on the Celsius scale are 0 °C (freezing) and 100 °C (boiling).
- There is a 180 F° (Fahrenheit degree) difference between the freezing and boiling points of water.
- The same difference on the Celsius Scale is 100 C° (Celsius degrees).
- The equivalence statement between Fahrenheit and Celsius degrees is

$$9 \text{ F}^\circ = 5 \text{ C}^\circ$$

- When doing conversions between Fahrenheit and Celsius temperatures, we need to shift the Fahrenheit temperature by 32 F° to compensate for the difference between the zero points of the two scales.

When converting from Fahrenheit to Celsius, we do the 32 F° shift first, then multiply by the unit factor (5 C°/9 F°).

$$T_C = (T_F - 32\text{F}^\circ) \frac{5\text{C}^\circ}{9\text{F}^\circ}$$

When converting from Celsius to Fahrenheit, we first multiply by the unit factor (9 F°/5 C°), then we apply the 32 F° shift.

$$T_F = T_C \frac{9\text{F}^\circ}{5\text{C}^\circ} + 32\text{F}^\circ$$

My advice is to remember the equivalence statement and the need to shift the Fahrenheit temperature by 32 degrees, rather than trying to memorize these two equations. (What if you remember them wrong?)

- Kelvin degrees (i. e., temperature differences on the Kelvin scale) are the same as Celsius degrees).
- Water freezes at 273.15 K (i. e., on the Kelvin scale), so to convert from Celsius to Kelvin, you add 273.15.
- To convert from Kelvin to Celsius, you subtract 273.15.

1.8 Density

- Density is a very important property for chemists. It is defined as the amount of mass of a substance in a unit volume.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \quad \text{or} \quad \rho = \frac{m}{v}$$

- Sample Exercise 1.13: A chemist is trying to identify an unknown liquid. She weighs a clean beaker, then adds 25.00 cm³ of the unknown liquid to it with a pipette. The difference in mass between the beaker with added sample and the empty beaker is

19.625 g (at 20 °C). The unknown might be any of the five compounds whose densities are listed to the right. What is the most likely identity for the unknown?

Step 1: She computes the density:

$$\rho \frac{\text{g}}{\text{cm}^3} = \frac{m \text{ g}}{v \text{ cm}^3} = \frac{19.625 \text{ g}}{25.00 \text{ cm}^3} = 0.7850 \frac{\text{g}}{\text{cm}^3}$$

Step 2: She compares the result against the density data in the table. The closest match is isopropyl alcohol, but ethanol is also fairly close, so she might want to repeat the measurement several times to confirm that 0.7850 g/cm³ is the true density.