

Measurement and Significant Figures (*cont'd*)

- To count the number of significant figures in a measurement, observe the following rules:
 - All nonzero digits are significant.
 - Zeros between significant figures are significant.
 - Zeros preceding the first nonzero digit are not significant.
 - Zeros to the right of the decimal after a nonzero digit are significant.
 - Zeros at the end of a nondecimal number may or may not be significant. (Use scientific notation.)

Measurement and Significant Figures (*cont'd*)

- Number of significant figures refers to the number of digits reported for the value of a measured or calculated quantity, indicating the precision of the value.
 - When multiplying and dividing measured quantities, give as many significant figures as the least found in the measurements used.
 - When adding or subtracting measured quantities, give the same number of decimals as the least found in the measurements used.

Example:

Calculate concentration of solution in g/mL

when 14.0 g of NaCl is dissolved in 102.4 mL of solution:

$$14.0 \text{ g} / 102.4 \text{ mL} = 0.137 \text{ g/mL}$$

only significant figures

Measurement and Significant Figures (*cont'd*)

- An exact number is a number that arises when you count items or when you define a unit.
 - For example, when you say you have nine coins in a bottle, you mean exactly nine.
 - When you say there are twelve inches in a foot, you mean exactly twelve.
 - Note that exact numbers have no effect on significant figures in a calculation.

Table 1.2 SI Base Units:

In 1960, the General Conference of Weights and Measures adopted the International System of units (or SI), which is a particular choice of metric units. This system has seven SI base units, the SI units from which all others can be derived

Quantity	Unit	Symbol
Length	Meter	m
Mass	Kilogram	Kg
Time	Second	S
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

Table 1.3 SI Prefixes:

The advantage of the metric system is that it is a decimal system. A larger or smaller unit is indicated by a SI prefix -- that is, *a prefix used in the International System to indicate a power of 10.*

Multiple	Prefix	Symbol
10^6	mega	M
10^3	kilo	k
10^{-1}	deci	D
10^{-2}	centi	C
10^{-3}	milli	m
10^{-6}	micro	μ
10^{-9}	nano	n
10^{-12}	pico	p

These have to be memorized!!!

Temperature

- The Celsius scale (formerly the Centigrade scale) is the temperature scale in general scientific use.
 - However, the SI base unit of temperature is the kelvin (K), a unit based on the absolute temperature scale.
 - The conversion from Celsius to Kelvin is simple since the two scales are simply offset by 273.15o.

$$K = ^\circ C + 273.15$$

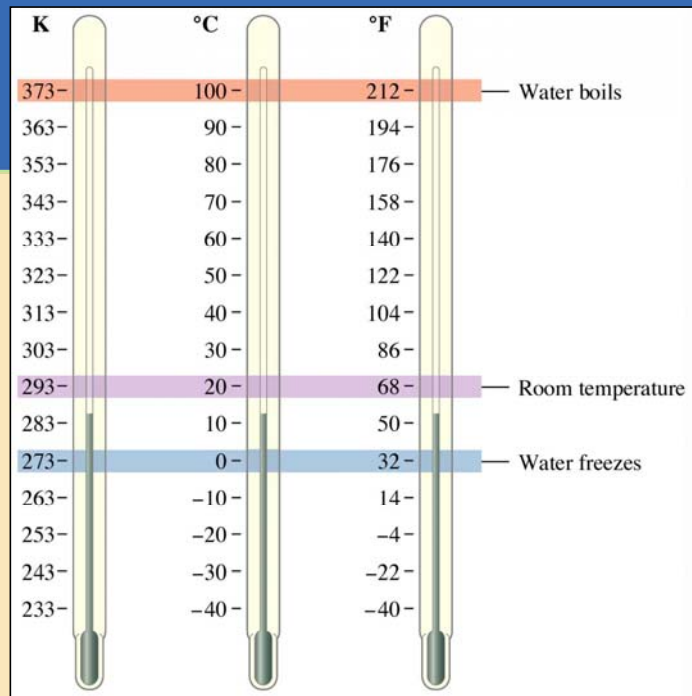
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Presentation of Lecture Outlines, 1-7

Temperature Scales:

- The Fahrenheit scale is at present the common temperature scale in the United States.
 - The conversion of Fahrenheit to Celsius, and vice versa, can be accomplished with the following formulas:

$$^\circ C = \frac{^\circ F - 32}{1.8}$$



$$^\circ F = 1.8 (^\circ C) + 32$$

Example: converting 113 F to Celsius:

Derived Units

- The SI unit for speed is meters per second, or m/s.
 - This is an example of an SI derived unit, created by combining SI base units.
 - Volume is defined as length cubed and has an SI unit of cubic meters (m³).
 - Traditionally, chemists have used the liter (L), which is a unit of volume equal to one cubic decimeter.

$$1 \text{ L} = 1 \text{ dm}^3 \quad \text{and} \quad 1 \text{ mL} = 1 \text{ cm}^3$$

Derived Units

- The density of an object is its mass per unit volume,

$$d = \frac{m}{V}$$

where **d** is the density, **m** is the mass, and **V** is the volume.

- Generally the unit of mass is the **gram**.
- The unit of volume is the **mL** for liquids; **cm³** for solids; and **L** for gases.

A Density Example

- A sample of the mineral galena (lead sulfide) weighs 12.4 g and has a volume of 1.64 cm³. What is the density of galena?

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{12.4 \text{ g}}{1.64 \text{ cm}^3} =$$

Units: Dimensional Analysis

- In performing numerical calculations, it is good practice to associate units with each quantity.
 - The advantage of this approach is that the units for the answer will come out of the calculation.
 - And, if you make an error in arranging factors in the calculation, it will be apparent because the final units will be nonsense.

Units: Dimensional Analysis

- Dimensional analysis (or the factor-label method) is the method of calculation in which one carries along the units for quantities.
 - Suppose you simply wish to convert 20 yards to feet.

$$20 \text{ yards} \times \frac{3 \text{ feet}}{1 \text{ yard}} = 60 \text{ feet}$$

- Note that the units have cancelled properly to give the final unit of feet.

Units: Dimensional Analysis

- The ratio (3 feet/1 yard) is called a conversion factor.
 - The conversion-factor method may be used to convert any unit to another, provided a conversion equation exists.
 - Relationships between certain U.S. units and metric units are given in Table 1.5.

Unit Conversion

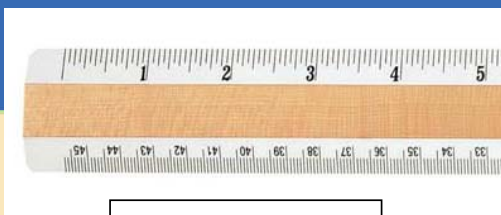
- Sodium hydrogen carbonate (baking soda) reacts with acidic materials such as vinegar to release carbon dioxide gas. Given an experiment calling for 0.348 kg of sodium hydrogen carbonate, express this mass in milligrams.

$$0.348 \text{ kg} \times \frac{10^3 \text{ g}}{1 \text{ kg}} \times \frac{10^3 \text{ mg}}{1 \text{ g}} = 3.48 \times 10^5 \text{ mg}$$

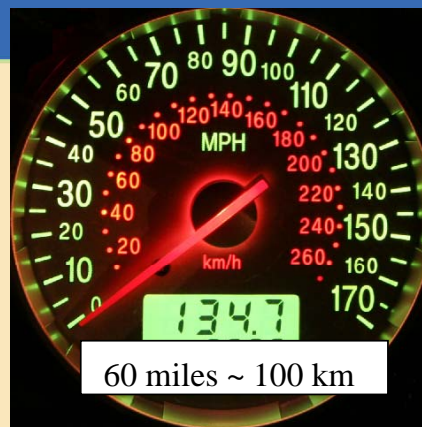
Table 1.5 Relationships of Some U.S. and Metric Units

Length	Mass	Volume
1 in = 2.54 cm	1 lb = 0.4536 kg	1 qt = 0.9464 L
1 yd = 0.9144 m	1 lb = 16 oz	4 qt = 1 gal
1 mi = 1.609 km	1 oz = 28.35 g	
1 mi = 5280 ft		

USEFUL ballpark (not exact!) conversion estimates to keep in mind



1 inch ~ 2.5 cm



60 miles ~ 100 km



1 gallon ~ 4 liters



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Unit Conversion

- Suppose you wish to convert 0.547 lb to grams.
 - From Table 1.5, note that 1 lb = 453.6 g, so the conversion factor from pounds to grams is 453.6 g/1 lb. Therefore,



EXERCISE Conversions Involving Density

What is the mass in grams of 1.00 gal of water? The density of water is 1.00 g/mL.

Solution Before we begin solving this exercise, we note the following:

1. We are given volume and density and are asked to calculate mass. Density = Mass/Volume, then: Mass = Volume x Density
2. The conditions involve SI and non-SI units, and the answer is asked in SI – units. This means we will need to use conversion factors between SI/non-SI units. Those will be given to you, - no need to memorize.

1L = 1.057 qt; 1 gal = 4qt ← note all numbers in conversion factors are exact, that is they do not affect the sig figs count in your results

We will calculate the volume in SI units first:

The first of these conversion factors must be used as written (with grams in the numerator) to give the desired result, whereas the last conversion factor must be rearranged in order to cancel gallons:

$$\text{VOLUME} = (1.00 \text{ gal}) \left(\frac{4 \text{ qt}}{1 \text{ gal}} \right) \left(\frac{1 \text{ L}}{1.057 \text{ qt}} \right) = 3.788 \text{ L} \quad (\text{we left some extra sig,figs})$$

Now we can calculate mass according to equation Mass = Volume x Density

But note that our volume is in L (SI base unit) while our density uses mL.

We will need conversion factor based on power of ten. See table on next page (which you WILL need to memorize):

1mL = 10⁻³ L (little "m" stands for milli and means multiply by 10⁻³)

At this point you can choose to convert either density from g/mL to g/L or volume from L to milliliters.

The point is they have to match. I chose to convert L to mL, and then multiplied by density in g/mL as was given.

$$\begin{aligned} \text{Mass in grams} &= 3.788 \text{ L} \times \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) \left(\frac{1.00 \text{ g}}{1 \text{ mL}} \right) \\ &= 3.78 \times 10^3 \text{ g water} \end{aligned}$$

The units of our final answer are appropriate, and we've also taken care of our significant figures. The answer has 3 sig. fig, since we did multiplication/ division, and the given information (volume and density) had 3 sig. figs. (Rule 1)

Operational Skills

- Using the law of conservation of mass.
- Using significant figures in calculations.
- Converting from one temperature scale to another.
- Calculating the density of a substance.
- Converting units.