Chemistry 120 Fall 2016

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Test Dates:

September 23, **2016 (Test 1): Chapter 1,2 &3**
October 13, **2016 (Test 2): Chapter 4 & 5**
October 31, **2016 (Test 3): Chapter 6, 7 & 8**
November 15, **2016 (Test 4): Chapter 9, 10 & 11**
November 17, **2016 (Make-up test) comprehensive: Chapters 1-11**

Chapter 2. Measurements in Chemistry

2-1 Measurement Systems

2-2 Metric System Units

- Metric Length Units
- Metric Mass Units
- Metric Volume Units

2-3 Exact and Inexact Numbers

2-4 Uncertainty in Measurement and Significant Figures

- Origin of Measurement Uncertainty
- Guidelines for Determining Significant Figures

2-5 Significant Figures and Mathematical Operations

- Rounding Off Numbers
- Operational Rules

2-6 Scientific Notation

- Converting from Decimal to Scientific Notation
- Significant Figures and Scientific Notation
- Multiplication and Division in Scientific Notation
- Calculators and Scientific Notation
- Uncertainty and Scientific Notation

Chapter 2. Measurements in Chemistry

2-7 Conversion Factors

Conversion Factors Within a System of Units

Conversion Factors between Systems of Units

2-8 Dimensional Analysis

2-9 Density

Density as a Conversion Factor

2-10 Temperature Scales

Conversions Between Temperature Scales

Temperature Readings and Significant Figures

What’s covered in this chapter?

- *Science and the scientific method*
- *Measurements* – what they are and what do the numbers really mean?
- *Units* – metric system and imperial system
- *Numbers* – exact and inexact
- *Significant figures and uncertainty*
- *Scientific notation*
- *Dimensional anaylsis* (conversion factors)

The scientific method

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D

- In order to be able to develop explanations for phenomena.
- After defining a problem
  - Experiments must be designed and conducted
  - Measurements must be made
  - Information must be collected
  - Guidelines are then formulated based on a pool of observations
- Hypotheses (predictions) are made, using this data, and then tested, repeatedly.
- Hypotheses eventually evolve to become laws and these are modified as new data become available
- An objective point of view is crucial in this process. Personal biases must not surface.

The scientific method

- At some level, everything is based on a model of behavior.
- Even scientific saws change because there are no absolutes.

Measurements

- An important part of most experiments involves the determination (often, the estimation) of quantity, volume, dimensions, capacity, or extent of something – these determinations are measurements

- In many cases, some sort of scale is used to determine a value such as this. In these cases, estimations rather than exact determinations need to be made.

SI Units

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s[a]
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

[a] The abbreviation sec is frequently used.

- *Système International d'Unités*

Prefix-Base Unit System

Prefixes convert the base units into units that are appropriate for the item being measured.

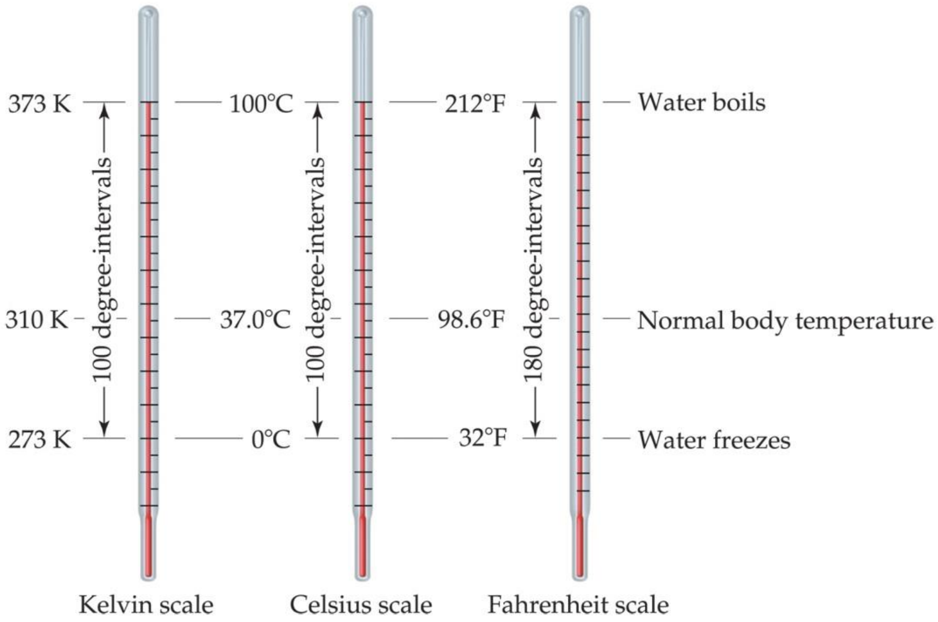
Know these prefixes and conversions

Prefix	Abbreviation	Meaning	Example
Giga	G	10^9	1 gigameter (Gm) = 1×10^9 m
Mega	M	10^6	1 megameter (Mm) = 1×10^6 m
Kilo	k	10^3	1 kilometer (km) = 1×10^3 m
Deci	d	10^{-1}	1 decimeter (dm) = 0.1 m
Centi	c	10^{-2}	1 centimeter (cm) = 0.01 m
Milli	m	10^{-3}	1 millimeter (mm) = 0.001 m
Micro	μ[a]	10^{-6}	1 micrometer (μm) = 1×10^{-6} m
Nano	n	10^{-9}	1 nanometer (nm) = 1×10^{-9} m
Pico	p	10^{-12}	1 picometer (pm) = 1×10^{-12} m
Femto	f	10^{-15}	1 femtometer (fm) = 1×10^{-15} m

[a] This is the Greek letter mu (pronounced "mew").

So, 3.5 Gm = 3.5 x 10^9 m = 3500000000 m
and 0.002 A = 2 mA

Temperature:



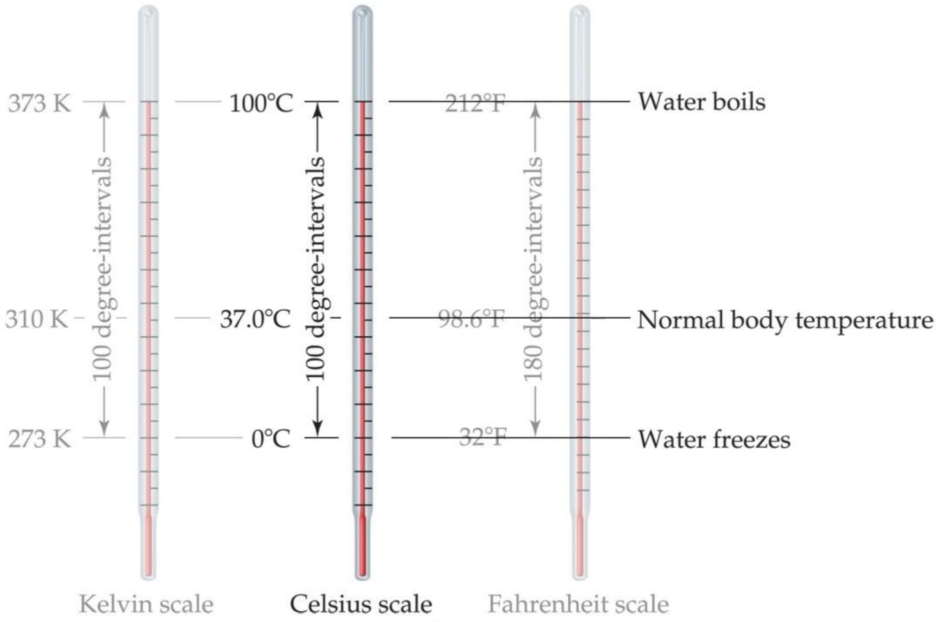
A measure of the average kinetic energy of the particles in a sample.

Kinetic energy is the energy an object possesses by virtue of its motion

As an object heats up, its molecules/atoms begin to vibrate in place. Thus the temperature of an object indicates how much kinetic energy it possesses.

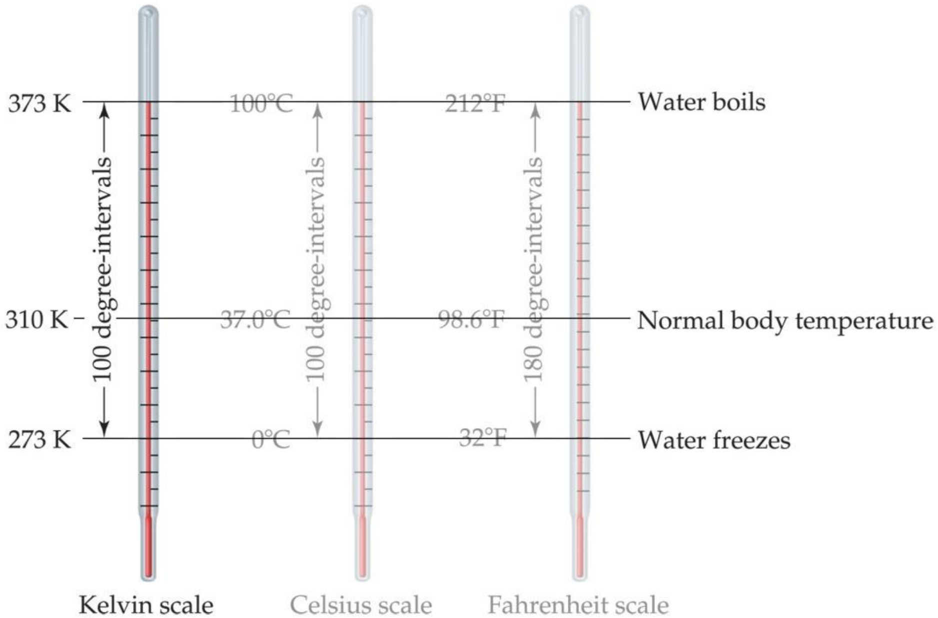
Farenheit: $^oF = (9/5)(^oC) + 32\ ^oF$

Temperature



- In scientific measurements, the Celsius and Kelvin scales are most often used.
- The Celsius scale is based on the properties of water.
  - 0°C is the freezing point of water.
  - 100°C is the boiling point of water.

Temperature



- The Kelvin is the SI unit of temperature.
- It is based on the properties of gases.
- There are no negative Kelvin temperatures.

$K = °C + 273$

0 (zero) K = *absolute zero* = -273 ^{o}C

Volume

- The most commonly used metric units for volume are the liter (L) and the milliliter (mL).

    A liter is a cube 1 dm long on each side.

    A milliliter is a cube 1 cm long on each side.

1 m = 10 dm

$(1\ m)^3 = (10\ dm)^3$

$1\ m^3 = 1000\ dm^3$

or

$0.001\ m^3 = 1\ dm^3$

These are conversion factors

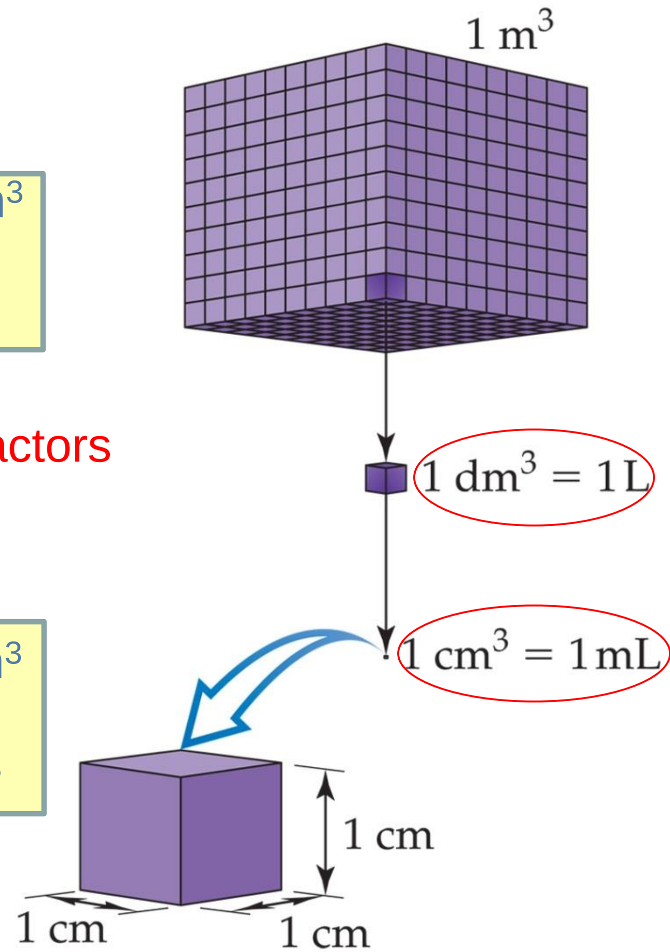
1 dm = 10 cm

$(1\ dm)^3 = (10\ cm)^3$

$1\ dm^3 = 1000\ cm^3$

or

$0.001\ dm^3 = 1\ cm^3$



1 m = 10 dm = 100 cm

Incidentally, $1\ m^3 = 1\text{x}10^6\ cm^3$

Density:

Another physical property of a substance – the amount of mass per unit volume

Density does not have an assigned SI unit – it's a combination of mass and length SI components.

$$d = \frac{\overset{\text{mass}}{m}}{\underset{\text{volume}}{V}}$$

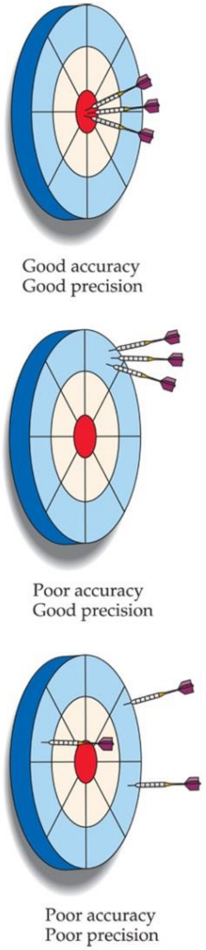
e.g. The density of water at room temperature ($25^{o}C$) is ~1.00 g/mL; at $100^{o}C$ = 0.96 g/mL

Density:

- Density is temperature-sensitive, because the volume that a sample occupies can change with temperature.
- Densities are often given with the temperature at which they were measured. If not, assume a temperature of about 25^{o}C.

Accuracy versus Precision

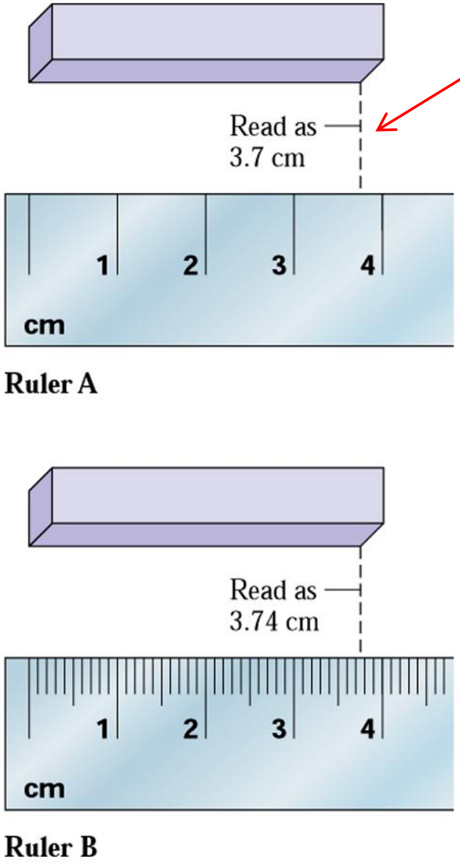
- **Accuracy** refers to the proximity of a measurement to the true value of a quantity.
- **Precision** refers to the proximity of several measurements to each other (Precision relates to the uncertainty of a measurement).



For a measured quantity, we can generally improve its accuracy by making more measurements

Measured Quantities and Uncertainty

The measured quantity, 3.7, is an estimation; however, we have different degrees of confidence in the 3 and the 7 (we are sure of the 3, but not so sure of the 7).



Whenever possible, you should estimate a measured quantity to one decimal place smaller than the smallest graduation on a scale.

Uncertainty in Measured Quantities

- When measuring, for example, how much an apple weighs, the mass can be measured on a balance. The balance might be able to report quantities in grams, milligrams, etc.
- Let's say the apple has a true mass of 55.51 g. The balance we are using reports mass to the nearest gram and has an uncertainty of +/- 0.5 g.
- The balance indicates a mass of 56 g
- The measured quantity (56 g) is true to some extent and misleading to some extent.
- The quantity indicated (56 g) means that the apple has a true mass which should lie within the range 56 +/- 0.5 g (or between 55.5 g and 56.5 g).

Significant Figures

- The term significant figures refers to the meaningful digits of a measurement.
- <u>The significant digit farthest to the right in the measured quantity is the uncertain one</u> (e.g. for the 56 g apple)
- When rounding calculated numbers, we pay attention to significant figures so we do not overstate the accuracy of our answers.

In any *measured* quantity, there will be some uncertainty associated with the measured value. This uncertainty is related to limitations of the technique used to make the measurement.

Exact quantities

- In certain cases, some situations will utilize relationships that are exact, defined quantities.
  - For example, a dozen is defined as exactly 12 objects (eggs, cars, donuts, whatever…)
  - 1 km is defined as exactly 1000 m.
  - 1 minute is defined as exactly 60 seconds.
- Each of these relationships involves an infinite number of significant figures following the decimal place when being used in a calculation.

Relationships between metric units are exact (e.g. 1 m = 1000 mm, exactly)
Relationships between imperial units are exact (e.g. 1 yd = 3 ft, exactly)
Relationships between metric and imperial units are not exact (e.g. 1.00 in = 2.54 cm)

Significant Figures

When a measurement is presented to you in a problem, you need to know how many of the digits in the measurement are actually significant.

1. All nonzero digits are significant. (1.644 has <u>four</u> significant figures)
2. Zeroes between two non-zero figures are themselves significant. (1.6044 has five sig figs)
3. Zeroes at the beginning (far left) of a number are never significant. (0.0054 has two sig figs)
4. Zeroes at the end of a number (far right) are significant *if a decimal point is written in the number*. (1500. has four sig figs, 1500.0 has five sig figs)

   (For the number 1500, assume there are two significant figures, since this number could be written as 1.5×10^3.)

Rounding

- Reporting the correct number of significant figures for some calculation you carry out often requires that you round the answer to the correct number of significant figures.
- Rules: round the following numbers to 3 sig figs
  - 5.483 (this would round to 5.48, since 5.483 is closer to 5.48 than it is to 5.49)
  - 5.486 (this would round to 5.49)

If calculating an answer through more than one step, only round at the final step of the calculation.

Significant Figures

- When addition or subtraction is performed, answers are rounded to the least significant decimal place.

Example: 20.4 + 1.332 + 83 = 104.732 = 105

“rounded”

- When multiplication or division is performed, answers are rounded to the number of digits that corresponds to the *least* number of significant figures in any of the numbers used in the calculation.

Example: 6.2/5.90 = 1.0508… = 1.1

Significant Figures

- <u>If both addition/subtraction and multiplication/division are used in a problem</u>, you need to follow the order of operations, keeping track of sig figs at each step, before reporting the final answer.

$$\frac{[\mathbf{104.6} \times (\mathbf{68.2} + \mathbf{14})]}{\mathbf{22.58}} = ?$$

1) Calculate (68.2 + 14). Do not round the answer, but keep in mind how many sig figs the answer possesses.
2) Calculate [104.6 x (answer from 1st step)]. Again, do not round the answer yet, but keep in mind how many sig figs are involved in the calculation at this point.
3) $\frac{(answer\ from\ 2^{nd}\ step)}{22.58}$, and then round the answer to the correct sig figs.

Significant Figures

- If both addition/subtraction and multiplication/division are used in a problem, you need to follow the order of operations, keeping track of sig figs at each step, before reporting the final answer.

$$\frac{[104.6 \times (68.2 + 14)]}{22.58} =$$

Despite what our calculator tells us, we know that this number only has 2 sig figs.

$$\frac{[104.6 \times 82.2]}{22.58} =$$

Despite what our calculator tells us, we know that this number only has 2 sig figs.

$$\frac{8598.12}{22.58} = 380.7847652790079716563330380868\ldots$$

$$= 380 \; or \; 3.8 \times 10^{2}$$

Our final answer should be reported with 2 sig figs.

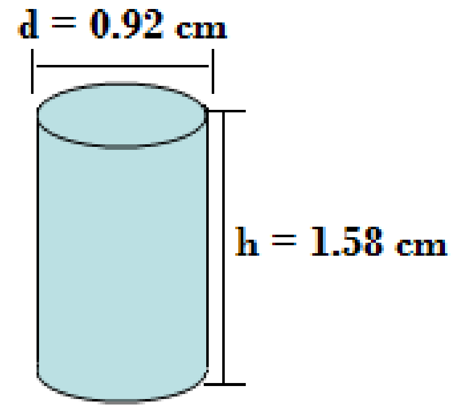
An example using sig figs

- In the first lab, you are required to measure the height and diameter of a metal cylinder, in order to get its volume
- Sample data:

$V = \pi r^2 h$

height (h) = 1.58 cm

diameter = 0.92 cm; radius (r) = 0.46 cm



2 sig figs 3 sig figs

$$\text{Volume} = \pi r^2 h = \pi(0.46\ \text{cm})^2(1.58\ \text{cm})$$

$$= 1.050322389\ \text{cm}^3$$

Only operation here is multiplication

$\boxed{\text{Answer} = 1.1\ \text{cm}^3}$

If you are asked to report the volume, you should round your answer to 2 sig figs

Calculation of Density

- If your goal is to report the density of the cylinder (knowing that its mass is 1.7 g), you would carry out this calculation as follows:

Then round the answer to the proper number of sig figs

$$d = \frac{m}{V} = \frac{1.7g}{1.050322389cm^3} = 1.61855066...\frac{g}{cm^3} = 1.6\frac{g}{cm^3}$$

Please keep in mind that although the “non-rounded” volume figure is used in this calculation, it is still understood that for the purposes of rounding in this problem, it contains only two significant figures (as determined on the last slide)

Use the non-rounded volume figure for the calculation of the density. If a rounded volume of 1.1 cm^3 were used, your answer would come to 1.5 g/cm^3

Dimensional Analysis
(conversion factors)

- The term, “dimensional analysis,” refers to a procedure that yields the conversion of units, and follows the general formula:

$$Given\_Units\underbrace{\left(\frac{Desired\_Units}{Given\_Units}\right)}_{\text{conversion factor}} = Desired\_Units$$

Some useful conversions

This chart shows all metric – imperial (and imperial – metric) system conversions. They each involve a certain number of sig figs.

Metric - to – metric and imperial – to – imperial conversions are exact quantities.

Examples:

16 ounces = 1 pound

1 kg = 1000 g

exact relationships

	Metric to English	English to Metric
Length		
1.00 inch = 2.54 centimeters	$\frac{1.00\text{ in.}}{2.54\text{ cm}}$	$\frac{2.54\text{ cm}}{1.00\text{ in.}}$
1.00 meter = 39.4 inches	$\frac{39.4\text{ in.}}{1.00\text{ m}}$	$\frac{1.00\text{ m}}{39.4\text{ in.}}$
1.00 kilometer = 0.621 mile	$\frac{0.621\text{ mi}}{1.00\text{ km}}$	$\frac{1.00\text{ km}}{0.621\text{ mi}}$
Mass		
1.00 pound = 454 grams	$\frac{1.00\text{ lb}}{454\text{ g}}$	$\frac{454\text{ g}}{1.00\text{ lb}}$
1.00 kilogram = 2.20 pounds	$\frac{2.20\text{ lb}}{1.00\text{ kg}}$	$\frac{1.00\text{ kg}}{2.20\text{ lb}}$
1.00 ounce = 28.3 grams	$\frac{1.00\text{ oz}}{28.3\text{ g}}$	$\frac{28.3\text{ g}}{1.00\text{ oz}}$
Volume		
1.00 quart = 0.946 liter	$\frac{1.00\text{ qt}}{0.946\text{ L}}$	$\frac{0.946\text{ L}}{1.00\text{ qt}}$
1.00 liter = 0.265 gallon	$\frac{0.265\text{ gal}}{1.00\text{ L}}$	$\frac{1.00\text{ L}}{0.265\text{ gal}}$
1.00 milliliter = 0.034 fluid ounce	$\frac{0.034\text{ fl oz}}{1.00\text{ mL}}$	$\frac{1.00\text{ mL}}{0.034\text{ fl oz}}$

Sample Problem

- A calculator weighs 180.5 g. What is its mass, in kilograms?

"given units" are grams, g

$$\boxed{Given\_Units\left(\frac{Desired\_Units}{Given\_Units}\right)=Desired\_Units}$$

"desired units" are kilograms. Make a ratio that involves both units.

Since $\boxed{1\ kg = 1000g}$ conversion factor is made using this relationship

$$180.5g\left(\frac{Desired\_Units}{Given\_Units}\right)=180.5g\left(\frac{1kg}{1000g}\right)=0.1805kg$$

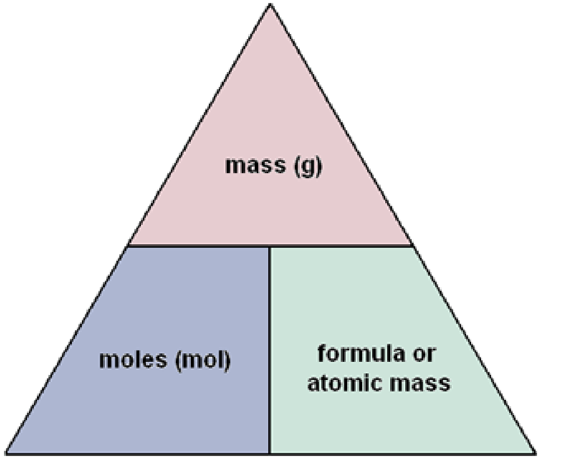
The mass of the calculator has four sig figs.
(the other numbers have many more sig figs)
The answer should be reported with four sig figs

Both 1 kg and 1000 g are exact numbers here (1 kg is defined as exactly 1000 g); assume an infinite number of decimal places for these

Dimensional Analysis

- Advantages of learning/using dimensional analysis for problem solving:
  - Reinforces the use of units of measurement
  - You don’t need to have a formula for solving most problems

How many moles of H_2O are present in 27.03g H_2O?



Sample Probelm

a) A prescription for nifedipine calls for a dose of 0.2 mg/kg of body weight. The drug is packaged in capsules containing 5.0 mg per capsule. How many capsules should be given to a patient who weighs 75 kg?

Sample Problem

- A car travels at a speed of 50.0 miles per hour (mi/h). What is its speed in units of meters per second (m/s)?
- Two steps involved here:
  - Convert miles to meters
  - Convert hours to seconds

a measured quantity

0.621 mi = 1.00 km
1 km = 1000 m
1 h = 60 min
1 min = 60 s

$$Given\_Units\left(\frac{Desired\_Units}{Given\_Units}\right) = Desired\_Units$$

$$50.0\frac{\cancel{mi}}{\cancel{h}}\left(\frac{1\cancel{km}}{0.621\cancel{mi}}\right)\left(\frac{1000m}{1\cancel{km}}\right)\left(\frac{1\cancel{h}}{60\cancel{\min}}\right)\left(\frac{1\cancel{\min}}{60s}\right) = 22.3653605296..\frac{m}{s} = 22.4\frac{m}{s}$$

should be 3 sig figs