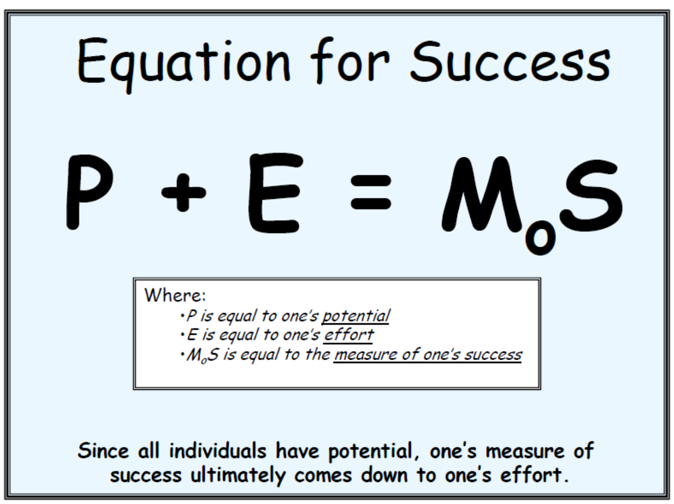
Unit 1: Nature of Science Unit Notes
Ms. Randall

Unit Objectives:

- *Differentiate between a pattern and a process*
- *Differentiate between a theory and a law*
- *Differentiate between a hypothesis and a theory*
- *Explain how models are used in authentic scientific inquiry*
- *Convert between units of measurements*
- *Differentiate between accuracy and precision*
- *Write numbers in scientific notation*
- *State rules to determine significant figures*
- *Count significant figures*
- *Understand the importance of significant figures*
- *Calculate the volume and density of an object*
- *Locate density of elements on the reference table*

Focus Questions for the Unit:

- How can matter be measured?
- Why do scientists need to make measurements?



Lesson 1: Metric Conversions

Objective:

- ***Define chemistry, matter and differentiate how matter can be described.***
- ***Convert between units of measurement***

Designed during the French Revolution of the 1790's, the metric system brought order out of the conflicting and confusing traditional systems of weights and measures then being used in Europe. Prior to the introduction of the metric system, it was common for units of length, land area, and weight to vary, not just from one country to another but from one region to another within the same country. As the modern nations were gradually assembled from smaller kingdoms and principalities, confusion simply multiplied. Merchants, scientists, and educated people throughout Europe realized that a uniform system was needed, but it was only in the climate of a complete political upheaval that such a radical change could actually be considered. In Chemistry we measure matter using SI units. This is an abbreviation for **System International.** The Metric System of measurement is **based on multiples of 10**. **Prefixes** are used to indicate what multiple of 10 the base unit is being multiplied or divided by.

SI BASE UNITS (AKA Base Units: **If you forget, use **Table D** in your Reference Tables!

Table D
Selected Units

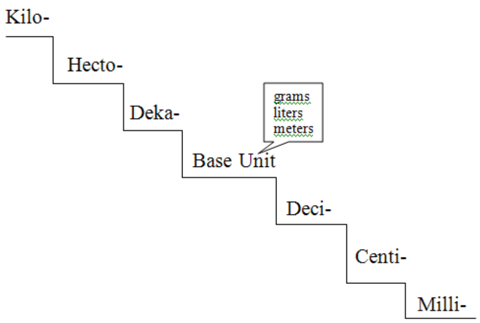
Symbol	Name	Quantity
m	meter	length
g	gram	mass
Pa	pascal	pressure
K	kelvin	temperature
mol	mole	amount of substance
J	joule	energy, work, quantity of heat
s	second	time
L	liter	volume
ppm	part per million	concentration
M	molarity	solution concentration

SI Metric Prefixes Conversion Factors – a mathematical expression that relates two units that measure the same type of quantity

Examples: 1 min = 60 sec 1000 g = 1 kg 1 L = 1000 mL

Methods for conversions:

<u>Staircase Method Instructions</u>-Look at the prefix you have and count how many steps you need to get to the prefix you want. Then move the decimal that many steps and in the same direction to convert the number to the new unit.



Base Unit 10^0

<u>Example:</u> Convert 52 mm to km

- starting at milli- it is 6 steps up the staircase to get to kilo
- move the decimal six places to the left
- so 52 becomes 0.000052km

<u>Using TABLE C in your reference table you can the same simple metric conversions as above without having to memorize anything. The table acts like the staircase.</u>

Table C
Selected Prefixes

Factor	Prefix	Symbol
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

Example: In the word kilometer, the root word (base unit) is "meter" and the prefix is "kilo." Kilo means multiply the root word by 1000. Therefore, one kilometer is 1000 meters (1 km = 1000 m).

1. Locate the prefix assigned to the measurement unit that you are starting with and then find the prefix that you want to convert to.

2. Count the number difference between the factors and then move your decimal that many places. Notice that a number of "stairs" are missing. Pretend they are there when you count. Don't forget that the base unit is understood to be present but is not listed!

Base Unit 10^0

Move decimal left ↑

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10^{-6}	micro-	μ
10^{-9}	nano-	n
10^{-12}	pico-	p

Move decimal right ↓

Example:

- The factor for centi is -2 and the factor for milli is -3.
- The difference between the two is “1 step”. Since you are moving down the chart you move the decimal one place to the right.

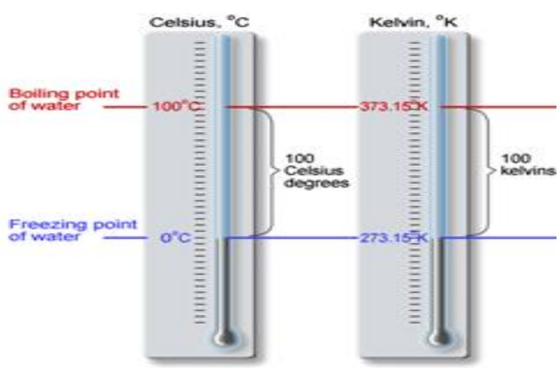
5.2 cm = 52. mm

*Rest Assured! For the Regents, the most you will have to convert will be between the milli-/kilo-/base unit (g, L, etc.). This is always a matter of 3 or 6 decimal places. You must also make sure you move the decimal the correct direction (right or left, which depends on whether you are converting from small to big or vice versa).

Temperature Conversions

Temperature is a measure of the **average kinetic energy** of the particles in an object. Adding energy (heat) to an object increases the kinetic energy of the particles which is observed as an increase in temperature.
When an object loses energy, kinetic energy of the particles decreases and is observed as a decrease in temperature. Several temperature scales are used in Chemistry and can be found on reference table A:

- **Kelvin Scale** based on the kelvin unit (K), the SI unit of temperature.
- **Celsius Scale** based on degrees Celsius (°C)



- The Celsius scale divides the range into 100 degrees—0°C to 100°C
- The Kelvin scale is divided into a range of 273K to 373K (the same scale as Celsius)

Temperature Conversions

The formula can be found on Table T of your Reference Tables

Temperature	K = °C + 273	K = kelvin °C = degree Celsius

Example: What is the temperature in Kelvin of an object that is 55 ° C ?

K= 273 + ° C
= 273 + 55 ° C =328K

Example: What is the temperature in Celsius of an object that is 150 K?

K= 273 + ° C therefore ° C =K – 273
150 K – 273= - 123 ° C

Lesson 2: Dimensional Analysis

Objective:

- ***To apply dimensional analysis to conversion problems***

What is Dimensional Analysis

Often you will be required to solve a problem with mixed units, or to convert from one set of units to another. Dimensional analysis is a simple method to accomplish this task.

Why Dimensional Analysis????

On September 23, 1999, NASA's $125 million Mars Climate Orbiter approached the red planet under guidance from a team of flight controllers at the Jet Propulsion Laboratory. The probe was one of several planned for Mars exploration, and would stay in orbit around the planet as the first extraterrestrial weather satellite. It had been in flight for over nine months, covering more than 415 million miles of empty space on its way to Mars. As the Orbiter reached its final destination, the flight controllers began to realize that something was wrong. They had planned for the probe to reach an orbit approximately 180 km off the surface of Mars – well beyond the planet's thin atmosphere. But new calculations based on the current flight trajectory showed the Orbiter skimming within 60 km of the Martian surface. Now the probe would actually enter the planet's thin atmosphere, something for which it was never designed. The consequences were catastrophic: when the scientists and engineers commanding the probe lost communication, they could only assume that the spacecraft was incinerated by the friction from an atmospheric entry that it was never supposed to make.

What caused this disaster? The problem arose in part from a simple, seemingly innocent, mistake. Throughout the journey from Earth, solar winds pushed against the solar panels of the probe, throwing the spacecraft off course by a small amount. The designers had planned for this, and jet thrusters were turned on by the flight controllers to apply a force, making numerous small corrections to readjust its course. Unfortunately, the NASA engineers measured this force in pounds (a non-metric unit), while the JPL team worked in Newtons (a metric unit), and the software that calculated how long the thrusters should be fired did not make the proper conversion. Since 1 pound = 4.45 Newtons, 4.45 times too much thrust was applied each time the thrusters were used. While each individual adjustment mistake was very small, this mistake grew larger and larger over multiple adjustments, resulting in the craft's premature demise in the Martian atmosphere.

The Orbiter loss illustrates the need for consistent use of units. Most people, however, are most comfortable working in whatever units they grew up using. As a result, unit consistency may not be possible within or between teams around the world. Ideally, people should be comfortable with a variety of ways of converting units in order to allow for collaboration among individuals from a variety of backgrounds.

While most people are not controlling NASA space probes, unit conversion is something that happens every day, in all walks of life. Even such a simple problem as figuring out that two dozen eggs equals 24 eggs is, at its heart, a unit conversion problem. Whether you realize it or not, when you do this problem in your head, you're figuring it out like this:

$$2\ \cancel{\text{dozen eggs}} \times \frac{12\ \text{eggs}}{1\ \cancel{\text{dozen eggs}}} = 24\ \text{eggs}$$

Watch the following video:

[Dimensional Analysis/Factor Label Method - Chemistry Tutorial]

We will be using dimensional analysis for the rest of the school year as this is how chemists do their conversions.

Read through the chart that models the process below.

MODEL: Convert 6.0 cm to km

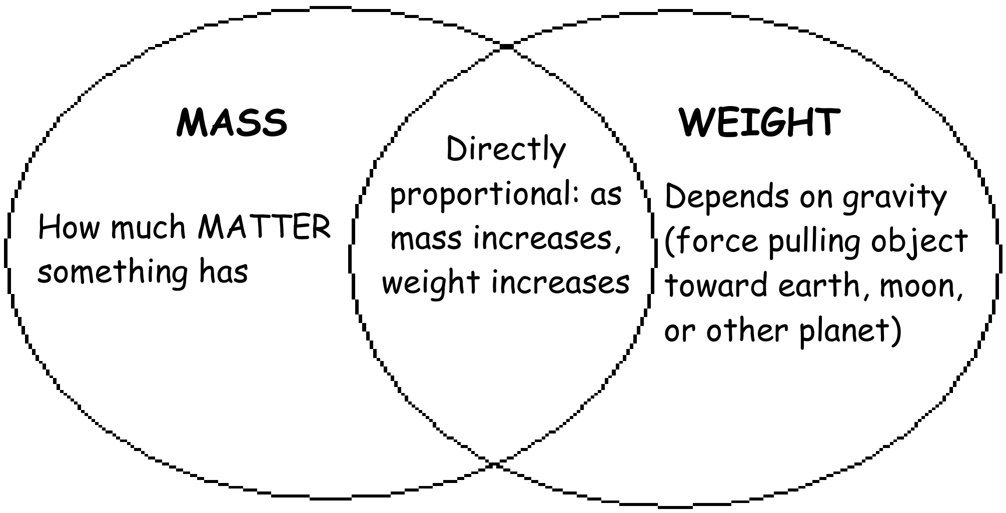
Steps	
1. Write the term to be converted (include both the number and the unit)	6.0cm
2. Write the conversion formula (see Ref Tables)	100000cm=1km
3. Make a fraction of the conversion formula such that the denominator units are the same as the units from step 1 and the numerator contains the units you want to convert to.	$\frac{1km}{100000cm}$
4. Multiply the term in step 1 by the fraction in step 3.	$6.0cm \text{ X } \frac{1km}{100000cm} =$
5. Cancel out "like" units	$6.0\cancel{cm} \text{ X } \frac{1km}{100000\cancel{cm}} =$
6. Solve (everything on top of fraction is multiplied and divided by everything on bottom)	.00006km

Lesson 3: Measuring Matter/Density

Objective:

1. ***Determine the volume of a substance***
2. ***Calculate Density/Mass/Volume***

Mass vs. Weight



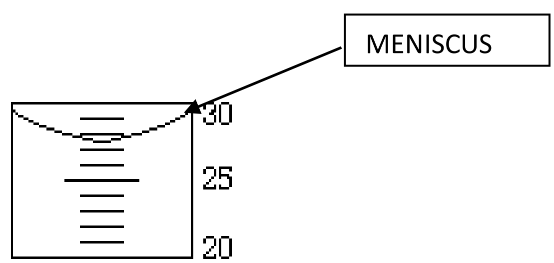
*We really only work with MASS in chemistry class!
** We have the same MASS whether we are on earth or on the moon. The different forces of gravity on each cause us to weigh more on earth than on the moon though (this is why we float on the moon!)

3. **Volume** – the amount of SPACE an object takes up

- Techniques:

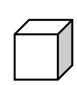
  a. Liquids → use graduated cylinder

  Reading a graduated cylinder:



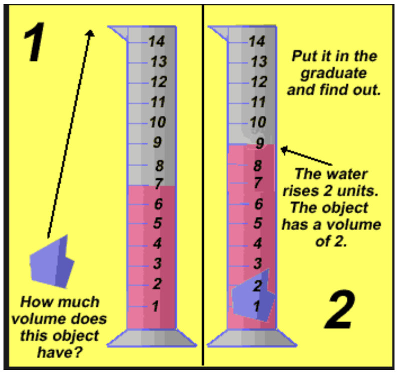
  ► Measurements are read from the bottom of the MENISCUS

  b. Regular Solids → measure dimensions and use l x w x h formula
  ***you need to MEMORIZE:

  1 cubic cm (1cm x 1cm x 1cm) = 1 cm^3 = 1 milliLiter (mL) 

  c. Irregular Solids → Water Displacement method

Intial volume-final volume= volume of object



4. **Density**: the amount of mass in a given volume (space); ratio of mass to volume. The formula can be found on Reference Table T

Density	$d = \frac{m}{V}$	d = density m = mass V = volume

Example: What is the density of an object with a mass of 60 g and a volume of 2 cm^3?

D=M/V

D=60g/2 cm^3

D=30g/cm^3

Densities values for all many of the elements can be found on Reference Table S in g/cm^3 at STP (standard pressure and temperature) when asked to solve for mass or volume.

Lesson 4: Precision (Significant Figures in Measurement)

Objective:
- ***Determine the accuracy of a measurement***
- ***Identify the precision of a measuring device.***
- ***Identify the amount of significant figures in a number***

Chemistry, and in fact all science, involves mathematical calculations. We need to be able to measure **quantitatively** (use numbers) as well as observe **qualitatively**.

Qualitative Data	Quantitative Data
• Deals with descriptions. • Data can be observed but not measured. • Colors, textures, smells, tastes, appearance, beauty, etc. • **Qualit**ative → **Qualit**y	• Deals with numbers. • Data which can be measured. • Length, height, area, volume, weight, speed, time, temperature, humidity, sound levels, cost, members, ages, etc. • **Quantit**ative → **Quantit**y

Quantitative Measurements in Chemistry:

1. Volume-the amount of space matter (an object) takes up

Metric base units of volume are: liters

2. Mass- the amount of matter an object contains (different than weight, which is mass + gravity… more on this in physics!)

Metric base units of mass are: grams

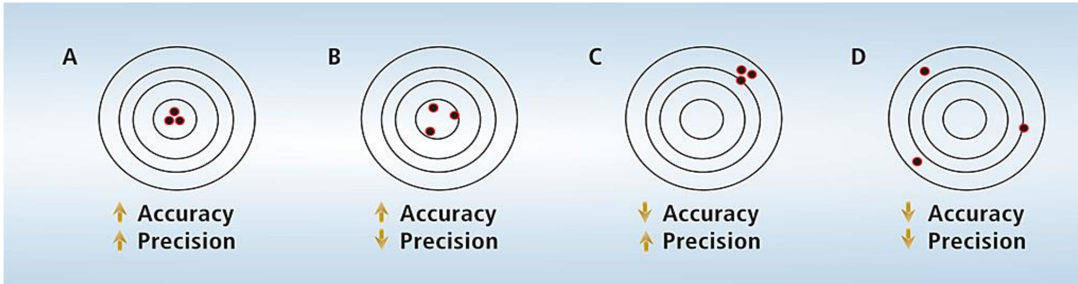
3. For **quantitative** measurements, we need to know the difference between **accurate** measurements and **precise** measurements.

Video Watch this!!! Accuracy & Precision

Accuracy: How close is a measurement to the true value?

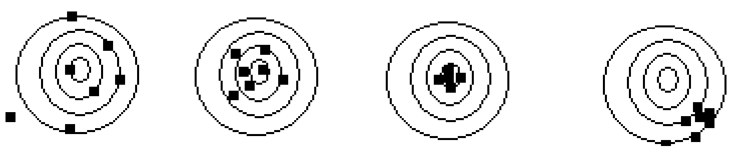
Precision: How reproducible is a measurement?

Dart board example:



Example: Cheryl, Cynthia, Carmen, and Casey shot the targets above at camp. Match each target with the proper description (assume bulls eye is the desired result)

(a) Accurate and precise________
(b) Accurate but not precise___________
(c) Precise but not accurate____________
(d) Neither precise nor accurate____________



(Answer: Carmen, Cynthia, Casie, Cheryl)

We can compensate for lack of **<u>precision</u>** by making many measurements or by using better equipment. You can improve **<u>accuracy</u>** by fixing the equipment you are using or correcting errors in use. When considering **<u>precision</u>** of a measurement, it is important to consider **<u>significant figures.</u>** **<u>Significant Figures</u>** in a measurement include all the digits that can be known precisely plus a **<u>last digit that can be estimated.</u>**

<u>Significant Figures Rules and Precision</u>

There are three rules on determining how many significant figures are in a number:

1. Non-zero digits are always significant.
2. Any zeros between two significant digits are significant.
3. A final zero or trailing zeros in the decimal portion <u>ONLY</u> are significant.

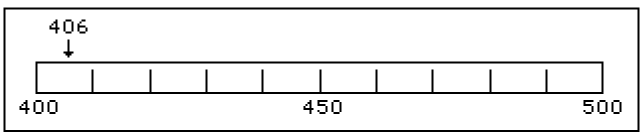
Rule 1: Non-zero digits are always significant.

If you measure something and the device you use (ruler, thermometer, triple-beam balance, etc.) returns a number to you, then you have made a measurement decision and that ACT of measuring gives significance to that particular numeral (or digit) in the overall value you obtain.

Hence a number like 26.38 would have four significant figures and 7.94 would have three.

Rule 2: Any zeros between two significant digits are significant.

Suppose you had a number like 406. By the first rule, the 4 and the 6 are significant. However, to make a measurement decision on the 4 (in the hundred's place) and the 6 (in the unit's place), you HAD to have decided on the ten's place. The measurement scale for this number would have hundreds and tens marked with an estimation made in the unit's place. Like this:



These are sometimes called "captured zeros."

Rule 3: A final zero or trailing zeros in the decimal portion <u>ONLY</u> are significant.

This rule causes the most difficulty with students. Here are two examples of this rule with the zeros this rule affects in boldface:

4. 0.005**00**
5. 0.0304**0**

Here are two more examples where the significant zeros are in boldface:

6. 2.3**0** x 10^{-5}
7. 4.5**00** x 10^{12}

<u>What is experimental error?</u> It's the process of screwing stuff up in the lab. It's all about the **<u>accuracy.</u>** In my experience, there are three main ways that experiments are screwed up in the lab:

1. **Human error**: *You* screwed up! Directions were not followed correctly (ie – used the wrong amount of chemicals, mixed chemicals together in the wrong order) or values were miscalculated.

2. **Procedural error**: Before you are offended that your teacher thinks you would screw up in the lab, please understand that it's impossible not to make mistakes in the laboratory. These mistakes can range from splashing chemicals while pouring them together, leaving residue on the inside of glassware, and any number of procedural errors. It's impossible not to make any errors, making it inevitable that percent error values are almost never 0%. ***Most of the reported sources of error in your written lab abstracts should be procedural error.***

3. **Instrumentation error**: *The black box* screwed up! Modern technology has given us a wide variety of machines that give us data in magical and mysterious ways. Take an electronic balance, for example: Place an object on the pan and the mass magically shows up on the screen. Unfortunately, it's hard to tell when these machines are mistaken, making them a possible source of error. You may, however, be unhappy to find that most instruments work reliably, making this sort of error unlikely, but not impossible.

4. **Unknown error**: *Something* screwed up! Was atmospheric humidity the culprit? Was there a mini-earthquake that threw off the data? Probably not – most unknown errors are just human or procedural errors that we can't even imagine we made in the first place. No matter what the source, you should be made aware that error is inevitable, no matter how carefully an experiment is performed. This may make you feel better if you do something wrong in an experiment.

We can calculate experimental error using the Percent error formula. This is a measure of accuracy!

Percent error is defined as the Measurement of the % that the measured value is "off" from accepted value

Measured value = value you "get"
Accepted value = value you "should get"

The formula for Percent error as found in Table T of your Reference Tables

Percent Error	$\% \text{ error} = \dfrac{\text{measured value} - \text{accepted value}}{\text{accepted value}} \times 100$

If value is negative, your measured value is LESS THAN the accepted value
If value is positive, your measured value is GREATER THAN the accepted value
*It is very important that you put the given values into the proper place in the formula!

Lesson 5: Rounding with Sig Figs

Objective:

- ***Round answers to proper sig figs in calculations***

What do Significant Figures Mean?

Precision is shown by significant figures.

For example, measuring 5.14 mL is more precise than measuring 5.1 mL, which is more precise than measuring 5 mL.

In the lab you could measure the same quantity of liquid, using different graduated cylinders.

5.14 has 3 sig figs ←most precise
5.1 has 2 sig figs
5 has 1 sig fig ←least precise

When massing matter, you could use a balance that will give you a mass of 19.6 grams, or a more expensive balance that will give you 19.58g. A very expensive balance could give you 19.582g.

NOTE: *The last number in any measurement is always estimated. On an analytical balance, which measures to 3 decimal places, the last number will often fluctuate because of the balance's sensitivity to air flow. In any case, the last digit is an estimate.*

A calculation using measured values gives an answer with the number of sig figs equal to the **lowest number of measured sig figs.**

NOTE: Definitions do not have significant figure!
For example, 100cm = 1m (this is ***not*** *1 sig fig)*

So: 1.02 cm x 2.3 cm = 2.3 cm^2, even though your calculator says 2.3460. That's because the *precision* of your answer is determined by the *least precise measurement*.

To calculate density, you divide mass by volume. If you measure the mass very precisely, but measure the volume less precisely, your answer will depend on the significant figures of the volume measurement.

So: A sample has a mass of 19.587 g and a volume of 3.2 cm$^{3.}$ Its density is 6.1g/cm^3, even though your calculator says 6.1209.

AN ANALOGY: *Consider the weakest link in a chain. The chain's strength is dependent on that weakest link. The chain can only be made stronger by strengthening* *that link. Likewise, the lowest precision measurement is the "weakest link" in an experiment.*

Rules for Using Sig Figs in Calculations

Operation	Rule	Examples
Multiplication/Division	Perform operation as normal & express answer in least # sig figs that were given to you	12.257 x 1.16<u>2</u> = 14.2<u>4</u>2634 **14.24**
Addition/Subtraction	Line decimal points up; round final answer to lowest decimal place (least accurate) value given	3.95 2.879 + 213.6 220.<u>4</u>29 **220.4**

Additional Notes:

- **Sometimes a whole number in a calculation can be considered to have an unlimited number of significant digits ... which means essentially that you ignore it.**
  For example, if an object's mass is 2.5 kg (two significant digits), and you have seven of them, the calculation for total mass would look like this:
  7 x 2.5 = 17.5 kg = 18 kg (rounded to two significant digits to match the 2.5).
  The number 7's significant digits are ignored.

- **When doing multi-step calculations, keep at least one more significant digit in intermediate results than you will need in your final answer.**
  For example, if a final answer requires two significant digits, then carry at least *three* significant digits in intermediate calculations. If you round-off all your intermediate answers to only two digits, you are discarding the information contained in the third digit, and as a result the second digit in your final answer might be incorrect. (This is known as 'round-off error')

Lesson 6: Scientific Notation

Objective:

- ***Convert numbers into scientific notation and standard notation***
- ***Calculate mathematical operations using scientific notation***

Scientific Notation is used for expressing very large or small numbers easily. It is also necessary to convert some numbers to scientific notation in order to have the correct number of significant figures. All numbers, regardless of magnitude, can be expressed in the form:

$N \times 10^n$ where

- N is a number, either an integer or decimal, between 1 and 10.
- n is a positive or negative integer.

When written in this form there must be one digit, and only one digit to the left of the decimal point in the number N.

Example: Standard notation 1,230,000: Scientific notation 1.23×10^6

Positive exponents

Example: 36,600

1. a number greater than 1:
   exponent of 10 is a positive whole number
2. value of the exponent:
   number of places the decimal point must be moved so that the notation is in standard form
3. $36{,}600 \times 10^0$
   For each place the decimal point is moved to the left, add 1 to the original exponent 3.66×10^4

Negative exponents

Example: 0.00563

1. a number less than 1:
   exponent of 10 is a positive whole number
2. value of the exponent:
   number of places the decimal point must be moved so that the notation is in standard form
3. 0.00563×10^0
   For each place the decimal point is moved to the right, subtract 1 from the original exponent 5.63×10^{-3}

Scientific notation: Multiplication

When multiplying numbers written in exponential notation:

1. Multiply digit terms in the normal fashion.
2. Obtain the exponent in the product by adding the exponents of the factors multiplied.
3. If necessary, adjust the exponent to leave just one digit to the left of the decimal point.

$(1.25 x 10^{5}) \text{ x } (4.0 x 10^{-2}) = (1.25 x 4.0) \text{ x } 10^{5+(-2)} = 5.0 x 10^{3}$

Scientific notation: Division

When dividing numbers written in exponential notation:

1. Divide the digit terms in the normal fashion.
2. Obtain the exponent in the quotient by subtracting the exponent of the divisor from the exponent of the dividend.
3. If necessary, adjust the exponent to leave just one digit to the left of the decimal point.

$(7.5 x 10^{6}) / (3.0 x 10^{-2}) = (7.5/3.0) \text{ x } 10^{6-(-2)} = 2.5 x 10^{8}$
dividend divisor

<u>**NOTE**</u>: If you need to plug these values into your **calculator** at any time, follow these steps:

Make sure that the number in scientific notation is put into your calculator correctly.
Read the directions for your particular calculator. For inexpensive scientific calculators:

1. Punch the number (the digit number) into your calculator.
2. Push the EE or EXP button. Do **NOT** use the “x” (times) button!!
3. Enter the exponent number. Use the +/- button to change its sign.
4. Voila! Treat this number normally in all subsequent calculations.

To check yourself, multiply $6.0 \text{ x } 10^{5}$ times $4.0 \text{ x } 10^{3}$ on your calculator.

Your answer should be **$2.4 \text{ x } 10^{9}$.**