Significant figures in a number include all digits that are known for certain, plus one estimated digit.

Rules for Determining the Number of Significant Figures

Significant figures in a number can be determined using Atlantic – Pacific rule.

If a Decimal is Absent in a number (whole number)

- Start counting with the first nonzero from the **A**tlantic (right) side of the number
- Count toward the left
- Stop counting with the last non zero digit.

How many you counted is the number of significant figure in that number.

If a Decimal is Present in a number (decimal fraction)

- Start counting with the first non-zero from the **Pacific** (left) side of the number
- Count toward the right and count all numbers (including zeros) once you have started counting.

How many you counted is the number of significant figure in that number.

NOTE:

All zeros between two non-zero numbers are always counted as being significant.

All zeros to start a number are never counted as being significant.

Examples:

405	has **3**	significant figures
4050	has **3**	significant figures
200	has **1**	significant figure
02	has **1**	significant figure
0.00**36**	has **2**	significant figures
0.0**936**	has **3**	significant figures
0.0**9360**	has **4**	significant figures
200.	has 3	significant figures
1.04	has **3**	significant figures
01.0	has **2**	significant figures

When multiplying or dividing

Limit or round the calculated result so that it has the same number of significant figures as the factor with the *least number of significant figures.*

Example 1

How much heat is absorbed by a 17 gram sample of ice to melt? *Leave answer in the correct number of significant figures.*

Heat = mass x H_f

Heat = 17 x 334

Heat = 5678 J

Heat = $\boxed{\textbf{5700 J}}$

17	factor with least number of sig fig.: 2
5678	(calculator result) has 4 sig fig. It must be rounded to 2 sig. fig
5700	(answer) is rounded to 2 sig fig.

Example 2

What is the density of an unknown substance if a 42.6 cm^3 sample has a mass of 22.43 g?

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

$$\text{Density} = \frac{22.43\text{ g}}{42.6\text{ cm}^3} = \boxed{\mathbf{0.527}\ \frac{\mathbf{g}}{\mathbf{cm^3}}}$$

42.6	factor with the least number of sig fig: **3**
0.527	(answer) also has 3 significant figures

When adding or subtracting

Limit or round the calculated result so it has the same number of decimal places as the factor with the *least number of decimal places (numbers after the decimal point)*

Example 3

What is the sum of 0.31 , 1.310 and 1.3205 to the correct number of significant figures?

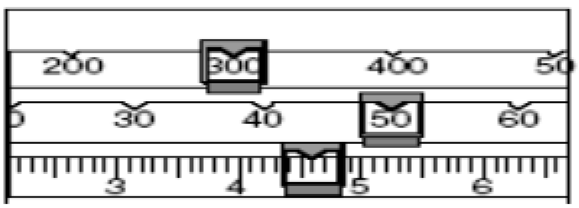
0.31 + 1.310 + 1.3205 = 2.9405 = $\boxed{\textbf{2.94}}$

0.31	Factor with the least number of decimal places: 2
2.9405	(calculator result) has 4 decimal places. It must be rounded and limited to 2 decimal places.
2.94	(answer) has 2 decimal places.

Sig Figs When Recording a Measurement

- All laboratory measurements should include a value and unit.
- Measurements should be given to the right significant figures.
- A measurement has the correct number of significant figures when it includes all digits known with certainty, and one estimated digit determined between two of the smallest unit markings on the measuring equipment.

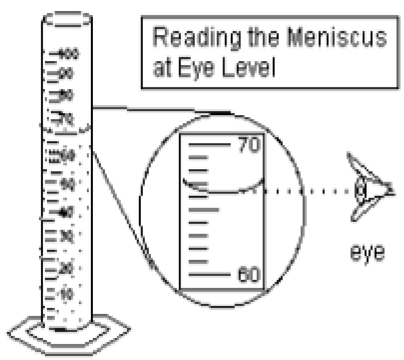
Examples:



Known with certainty

$\underbrace{354.5}8$ g

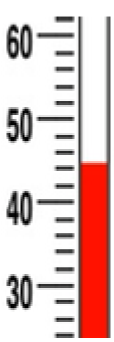
Estimated pointer position.



Known with certainty

$\underbrace{66.}1$°C

Estimated based on the bottom of the meniscus level.



Known with certainty

$\underbrace{44.}7$°C

Estimated based on the level being between the two smallest markings.

Percent Error

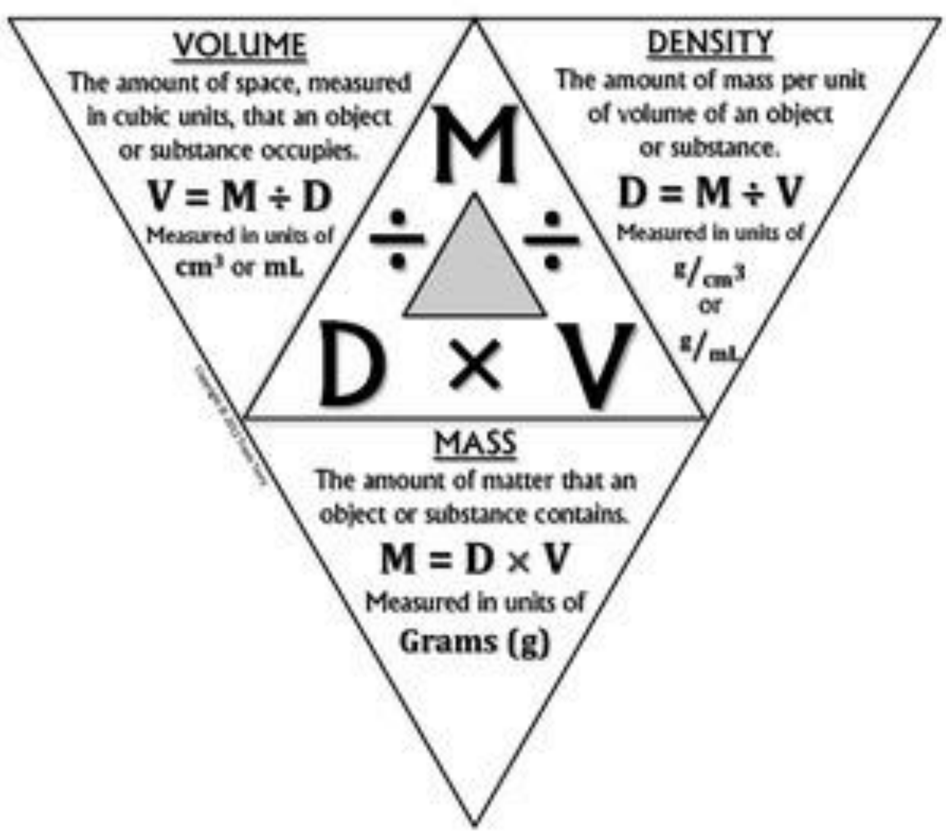
Data from a laboratory experiment is often different from what it should it be.
Percent error is an expression of the difference between the measured values of an experiment and the actual (accepted) values for that same experiment.

- Percent error should always be positive
- The smaller the percent error, the more accurate the measured value.
- Human errors and imprecision of measuring equipment are the common causes of high percent errors in lab experiments.

Percent error = $\dfrac{\text{measured value} - \text{accepted value}}{\text{accepted value}} \times 100$	The lower the percent error, the more accurate the lab result. **Reference Table T equation**

Density

Density is an intensive property that depends only on the composition of a substance, not on the size of the sample. The density of a substance generally decreases as its temperature increases, and increases as its temperature decreases. So the relationship between an object's phases and it's relative density is as follows: S>l>g. The only exception to this is $H_2O(l)$ and $H_2O_{(s)}$. $H_2O_{(l)}$ is more dense than $H_2O_{(s)}$.



Measuring in chemistry is multifaceted. Measurements that use tools, such as thermometers, electronic balances, rulers, etc. will require you to make quantitative measurements. They have numbers with units. Quantitative measurements will be made in lab nearly all the time. Without units all you have are numbers. Both numbers **and** proper units are necessary. Examples include 197 pounds, 23.45 grams, 10.0 mL, 13.56 g/cm^3 and 6.02 x 10^{23} molecules.

A qualitative measurement is one that uses descriptions only, no numbers or units are required. Examples include: the solutions is blue or cold. "How blue or hot" are not described by such qualitative measurements.

Precise and Accurate Measurements

When we measure in chemistry we hope to make perfect measurements. That means we do our best, using our instruments correctly, to get measurements that are close to true, so we can prove to ourselves that the chemistry works the same in the lab as we'd expect from figuring out chemical reactions on paper.

The better our measuring, the closer our experiments will match our expectations. The chemistry always works properly, if we can be careful enough in lab we'll be able to show that to ourselves.

When you make a measurement that is in fact very close or perfectly correct, that measurement is said to be accurate. An accurate measurement is right, it is the same as the ACTUAL VALUE. This is what we strive for.

If we can repeat our measurements and always get the same (or very close to the same) results, these measurements are said to be precise. Precise measurements are close together. They might be accurate also (as in the 3rd circle) or might not be accurate (as in the 2nd circle). That second circle indicates you are measuring properly, but your tool is not working correctly. In chemistry class the plan is to be both accurate & precise every time we measure.

Precision and Accuracy

- Precision (repeatability) = the degree to which repeated measurements show the same result
- Accuracy = the degree of closeness of measurements of a quantity to the actual (or accepted) value

High Accuracy Low Precision | **High Precision Low Accuracy** | **High Accuracy High Precision**

Metric Prefixes and Conversions

Commonly Used Metric Prefixes			
Prefix	Symbol	Meaning	Factor
mega	M	1 million times larger than the unit it precedes	10^6
kilo	k	1000 times larger than the unit it precedes	10^3
deci	d	10 times smaller than the unit it precedes	10^{-1}
centi	c	100 times smaller than the unit it precedes	10^{-2}
milli	m	1000 times smaller than the unit it precedes	10^{-3}
micro	μ	1 million times smaller than the unit it precedes	10^{-6}
nano	n	1 billion times smaller than the unit it precedes	10^{-9}
pico	p	1 trillion times smaller than the unit it precedes	10^{-12}

Ex #1: Convert 558nm to m.

$$\frac{1\ m}{10^9 nm} = \frac{x\ m}{558\ nm}$$

X= 5.58 x 10^{-7} m

Ex #2: Convert 500mL to L.

$$\frac{1\ L}{1000\ mL} = \frac{x\ L}{500\ mL}$$

X = .5 L