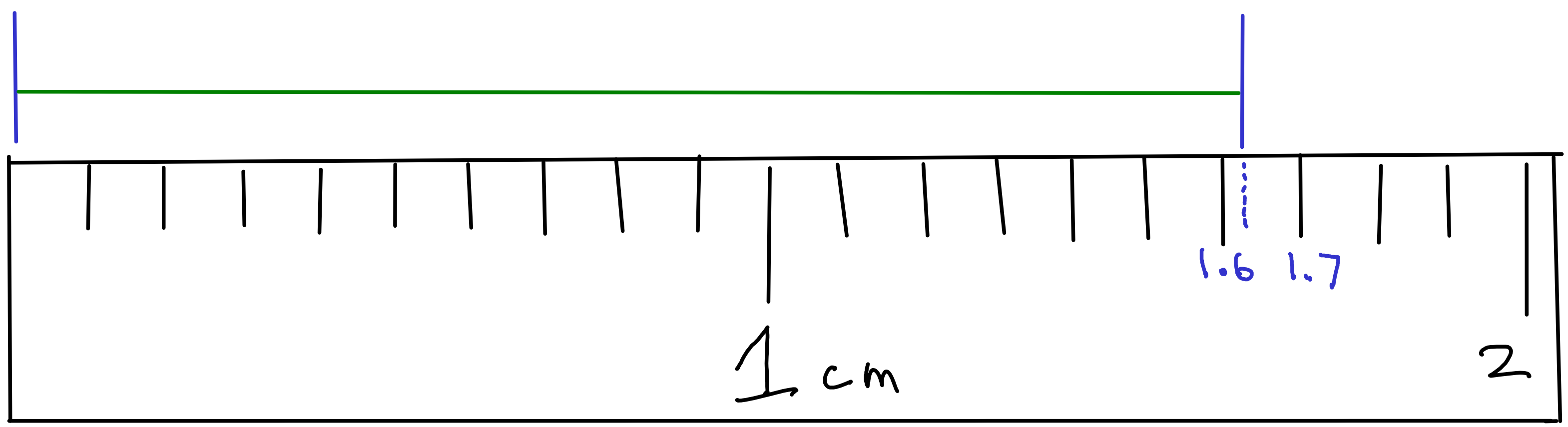
More on precison

Every measurement contains some amount of ERROR, or some amount of deviation from the true value of what is being measured.

RANDOM ERROR is the variability in a measurement that cannot be traced back to a single cause. Random errors cause measurements to fluctuate around the true value, but can be averaged out given enough measurements.

When reporting measurements, we want to indicate how much random error we think is present. How?

Form: X.XX cm



How long is the green line?

Write your answer on the card, then pass the card up to the front!

Our classroom experiment: Results

After throwing away obvious mistakes in reading the scale, we had:

Value	# students
1.60	1
1.61	1
1.62	9
1.63	6
1.64	1

18 measurements

Overall average

$\bar{x} = 1.62277778$ cm (unrounded)

$= \underbrace{1.6}\underbrace{2} \pm 0.01$ cm

CERTAIN DIGITS: Appear in nearly all repeats of the measurement

UNCERTAIN DIGITS: Vary.. Variation caused by estimation or other sources of random error.

When reading measurements from a scale, record all CERTAIN digits and one UNCERTAIN (or estimated) digit.

Significant figures

SIGNIFICANT FIGURES are a way to indicate the amount of uncertainty in a measurement.

The significant figures in a measurement are all of the CERTAIN DIGITS plus one and only one UNCERTAIN (or estimated) DIGIT

Example:

From our classroom experiment,

$\underbrace{1.6}\underbrace{2}$ cm

We estimated the last digit, It's uncertain

These digits were obtained in all measurements. They are certain

THIS MEASUREMENT HAS "THREE SIGNIFICANT FIGURES"!

Determining significant figures

When you read a measurement that someone has written using the significant figures convention, you can tell how precisely that measurement was made.

1.473 g ± 0.001

This was measured to the nearest +/- 0.001 g
The last digit is always UNCERTAIN (or estimated)

21 m ± 1

37.26 kg ± 0.01

Some other examples

3.2076 g ± 0.0001 g

uncertain digits

27.3 m ± 0.1 m

A small problem

The number ZERO has several uses. It may be a measured number, but it may also be a mere "placeholder" that wasn't measured at all!

So how do we tell a measured zero from a placeholder? There are a few ways:

1: BEGINNING ZEROS: Beginning zeros are NEVER considered significant.

$\underbrace{0}.\boxed{15}$ g

.15 g

This zero merely indicates that there is a decimal point coming up!

$\underbrace{0.0}\boxed{15}$ m (1.5 cm)

These zeros are placeholders. They'll disappear if you change the UNITS of this number!

$\underbrace{0.000}\boxed{63}$ mm

None of these zeros are considered significant

2: END ZEROS are sometimes considered significant. They are significant if

- there is a WRITTEN decimal point in the number
- there is another written indicator that the zero is significant. Usually this is a line drawn over or under the last zero that is significant!

$\boxed{1.50}$ km ± 0.01 km

This zero IS considered significant. There's a written decimal.

$\boxed{15}00$ m ± 100 m

These zeros ARE NOT considered significant (no written decimal, and no other indication that the zeros are significant)

$\boxed{143\bar{0}}00$ g ± 100 g

These zeros are not significant.

This zero IS significant. It's marked.

How many significant figures are there in each of these measurements?

76.070 g
5

85000. mm (± 1 mm)
5
decimal point

0.001030 kg
4

156.0002 g
7

0.10 s
2

17000000 mg (± 1,000,000 mg)
2

120000 km (± 100 km)
4

1350 ms (± 10 ms, not ± 1 ms)
3

Green: Approximate uncertainty. Matches up with the last significant figure (the uncertain digit)

Red: Significant digits.

Calculations with measurements

When you calculate something using measured numbers, you should try to make sure the ANSWER reflects the quality of the data used to make the calculation.

An ANSWER is only as good as the POOREST measurement that went into finding that answer!

$$\begin{array}{r} 14.20\underline{6} \\ 154.7\underline{2} \\ 1.\underline{6} \\ +\ 0.22\underline{2} \\ \hline 170.\underline{748} \end{array}$$

Round so that there's only one uncertain digit in the answer!

How should we report this answer? How much uncertainty is in this answer?

$\boxed{170.7}$

* If you add an uncertain number to either a certain or an uncertain number, then the result is uncertain!
* If you add certain numbers together, the result is certain!

For addition and subtraction, round FINAL ANSWERS to the same number of decimal places as the measurement with the fewest decimal places. This will give an answer that indicates the proper amount of uncertainty.

For multiplication and division, round FINAL ANSWERS to the same number of SIGNIFICANT FIGURES as the measurement with the fewest SIGNIFICANT FIGURES!

$\underbrace{15.62}_{4} \times \underbrace{0.0667}_{3} \times \underbrace{35.0}_{3} = 36.46489$

How should we report this answer?

$\boxed{36.5}$

$\underbrace{25.4}_{3} \times \underbrace{0.00023}_{2} \times \underbrace{15.201}_{5} = 0.088804242$

How should we report this answer?

$\boxed{0.089}$

A few more math with significant figures examples:

$$\overset{5}{15047} \times \overset{2}{11} \times \overset{4}{0.9876} = 163464.5892 \quad \boxed{160000}$$

~~16~~

1.6×10^5

Placeholder zeros, even though they aren't SIGNIFICANT, still need to be included, so we know how big the number is!

Addition:

```
   147.3
  2432
     0.97
+  111.6
  2691.67
```

$\boxed{2692}$

DENSITY CALCULATION

$$\frac{\overset{6}{14.7068}\ g}{\underset{2}{2.7}\ mL} = 5.44962963\ g/mL$$

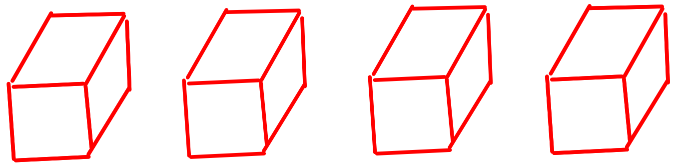
$\boxed{5.4\ g/mL}$

To improve (make more precise) the density measurement illustrated here, we must improve the precision of the VOLUME measurement. Improving the MASS measurement will not improve the precision of the calculated density.

Exact Numbers

- Some numbers do not have any uncertainty. In other words, they weren't measured!

1) Numbers that were determined by COUNTING!



How many blocks are to the left?

exactly 4

2) Numbers that arise from DEFINITIONS, often involving relationships between units

$12\ in = 1\ ft$

$Km = 10^3\ m$

$\boxed{\text{* All metric prefixes are exact!}}$

- Treat exact numbers as if they have INFINITE significant figures or decimal places!

Example

You'll need to round the answer to the right number of significant figures!
Convert 4.45 m to in, assuming that 2.54 cm = 1 in

EXACT!

$2.54\ cm = in \qquad cm = 10^{-2}\,m$

$$4.45\,\cancel{m} \times \frac{c\cancel{m}}{10^{-2}\,\cancel{m}} \times \frac{in}{2.54\,c\cancel{m}} = 175.1968504\ in$$

$$= \boxed{175\ in}$$

3 $\qquad$ ∞ $\qquad$ ∞

Usually, in unit conversions the answer will have the same number of significant figures as the original measurement did.

EXCEPTION: Temperature conversions, since these often involve ADDTION (different rule!)

A note on rounding: If possible, try to round only at the END of a multiple-step calculations. Avoid rounding intermediate numbers if possible, since extra rounding introduces ERROR into your calculations.

DALTON'S ATOMIC THEORY

- 1808: Publication of Dalton's "A New System of Chemical Philosophy", which contained the atomic theory

- Dalton's theory attempted to explain two things:

1 CONSERVATION OF MASS

- The total amount of mass remains constant in any process, chemical or physical!

2 LAW OF DEFINITE PROPORTIONS (also called the LAW OF CONSTANT COMPOSITION): All pure samples of a given compound contain the same proportion of elements by mass

The parts of Dalton's theory

1 Matter is composed of small, chemically indivisible ATOMS

2 ELEMENTS are kinds of matter that contain only a single kind of atom. All the atoms of an element have identical chemical properties.

3 COMPOUNDS are kinds of matter that are composed of atoms of two or more ELEMENTS which are combined in simple, whole number ratios.

Most importantly,

4 CHEMICAL REACTIONS are REARRANGEMENTS of atoms to form new compounds.

- Atoms are not gained or lost during a chemical reaction.

- Atoms do not change their identity during a chemical reaction.

- All the atoms that go into a chemical reaction must go out again!

Another look at chemical reactions

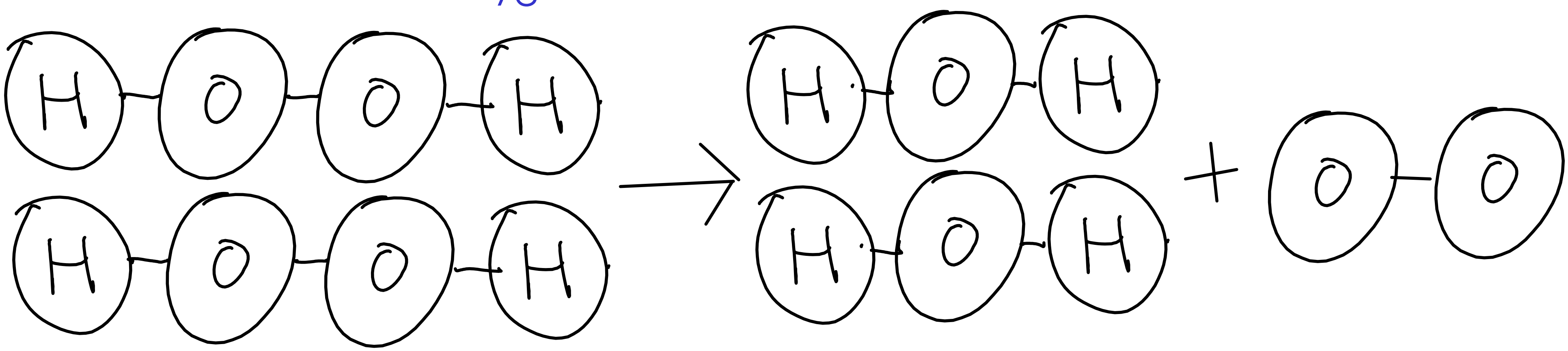
The decomposition of hydrogen peroxide over time (or when poured over a cut) works like this:

HYDROGEN PEROXIDE $\longrightarrow$ WATER + OXYGEN GAS

implying:... $H_2O_2 \longrightarrow H_2O + O_2$

... but wouldn't this mean that somehow an extra oxygen atom would form?

Not according to Dalton's theory. Dalton's theory would predict a different RATIO of water and oxygen would form:



$$\boxed{2\ H_2O_2 \longrightarrow 2\ H_2O + O_2}$$

- Dalton's theory sets LIMITS on what can be done with chemistry. For example:

1 Chemistry can't convert lead (an element) into gold (another element). Sorry, alchemists!

2 You can't have a compound form in a chemical reaction that contains an element that was not in your starting materials.

3 You can only make a certain amount of desired product from a fixed amount of starting material.