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.

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

.15g
↑

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

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

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

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

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!

$\boxed{1.5\underbrace{0}}$ km ± 0.01 km

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

$\boxed{15}\underbrace{00}$ m ± 100 m

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

$\boxed{143\underbrace{\bar{0}}}\underbrace{00}$ 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?

$\underline{76.070}$ g

5

$\underline{85000}$. mm

5

↖decimal point

$0.00\underline{1030}$ kg

4

$\underline{156.0002}$ g

7

$0.\underline{10}$ s

2

$\underline{17}000000$ mg

2

$\underline{120\bar{0}}00$ km

4

$\underline{135}0$ ms

3

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!

$$\begin{array}{rl} 14.20\underline{6} & \pm 0.001 \\ 154.7\underline{2} & \pm 0.01 \\ 1.\underline{6} & \pm 0.1 \\ +\ 0.22\underline{2} & \pm 0.001 \\ \hline 170.748 & \end{array}$$

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!

$$\overset{4}{\underline{15.62}} \times 0.0\overset{3}{\underline{667}} \times \overset{3}{\underline{35.0}} \approx 36.4|6489$$

How should we report this answer?

$$\boxed{36.5}$$

$$\overset{3}{\underline{25.4}} \times 0.000\overset{2}{\underline{23}} \times \overset{5}{\underline{15.201}} \approx 0.088|804242$$

How should we report this answer?

$$\boxed{0.089}$$

Note: The first significant figure can't be a beginning zero (they are never significant). The first significant figure is the leftmost "8"!

A few more math with significant figures examples:

$$\overset{5}{\underline{15047}} \times \overset{2}{\underline{11}} \times \overset{4}{\underline{0.9876}} = 163464.5892 \quad \boxed{160000} \quad (1.6 \times 10^5)$$

$$\cancel{16}$$

Placeholder zeroes (or scientific notation) required here since we need to know where the decimal goes!

Addition:

$$\begin{array}{rl} 147.3 & \pm 0.1 \\ 2432 & \pm 1 \\ 0.97 & \pm 0.01 \\ +\ 111.6 & \pm 0.1 \\ \hline 2691.67 & \end{array}$$

$$\boxed{2692}$$

DENSITY CALCULATION

$$\frac{\overset{6}{14.7068\text{ g}}}{\underset{2}{2.7\text{ mL}}} = 5.446962963\text{ g/mL}$$

$$\boxed{5.4\text{ g/mL}}$$

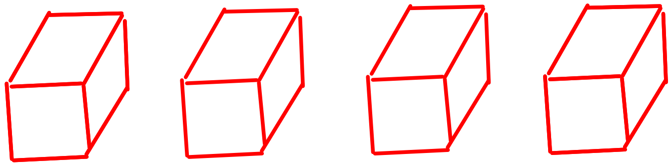
The only way to improve the precision of this density measurement is to improve the precision of the VOLUME measurement, since it limits the precision of the answer.

(We can actually use a LESS precise balance than the one we're currently using and still have the same quality density measurement!)

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 $\underline{\underline{4}}$

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

12 in = 1 ft

Km = 10^3 m

* 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 $\underbrace{2.54\ \text{cm} = 1\ \text{in}}_{\text{EXACT!}}$

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

$$\underset{\uparrow\ 3}{4.45\ \cancel{\text{m}}} \times \frac{\cancel{\text{cm}}}{\underset{\uparrow\ \infty}{10^{-2}\ \cancel{\text{m}}}} \times \frac{\text{in}}{\underset{\uparrow\ \infty}{2.54\ \cancel{\text{cm}}}} = 17\overset{\downarrow}{5}.1968504\ \text{in}$$

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

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