# The BASICS

This is your required reading.  It would be helpful to have one clean sheet of paper to write down questions for your teacher that come up during your reading.  Ask them and they will be answered.  You are expected to know this material, know the material that is reviewed in class and in the labs. There is overlap, but some material is just from class, while some material may be just from the BASICS.

It is written and provided for you as a supplement to your class time, your lab time, and the lab reports you will write.  Don't rely ONLY on these Basics, or your teacher.  Science is about seeking knowledge.  This is a good source for you, but not the only source for you to learn from.  Look for any discrepancies between what is done in class, and what is written here.  Sometimes you may be confused, and sometimes you might catch your teacher mistyping.  Always think.  Read this early, not the night before your Celebration of Knowledge.

## Measurement BASICS

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 of quantitative measurements include 197 pounds, 23.45 grams,  and 10.0 mL.

A qualitative measurement is one that uses descriptions only, no numbers or units are used.  Examples include: the solutions is blue or cold.

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Precise Measurement  vs. Accurate Measurements

When we measure in chemistry we hope to make perfect measurements. That means we use our instruments correctly, to get measurements that are close to the actual or true values, so we can prove to ourselves that the chemistry works as well in the lab as we'd can prove on paper.

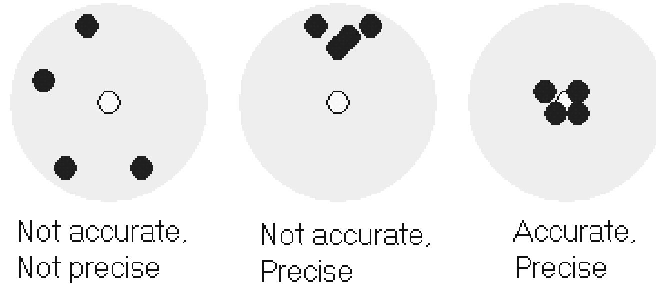
The better our measuring, the closer our experiments will match our expectations.  The chemistry always works perfectly, if we can be careful enough in lab we'll be able to prove that to ourselves.

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.

This drawing represents darts shot at a target.
At left, the darts are random. They are not accurate (near the center) nor near each other (not precise).

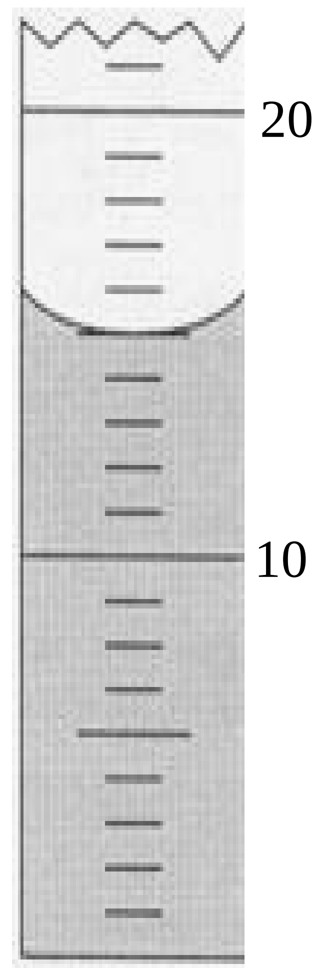



The center diagram shows the darts are all together but not near the center (these are precise but inaccurate).

At the far right the darts are all very close to the center, which makes them (precise + accurate).

## Significant Figures

When we make measurements in chemistry, or use formulas to convert measurements as needed, our calculators will "do the math" for us. It is very important to understand the significance of significant figures, they "keep you honest" in your measuring.




When you measure something, say how many milliliters of water are in a graduated cylinder, you can see the lines and make a measurement. This cylinder close up at right shows us that the tube contains 15 mL of water. Each line represents 1 mL.

It turns out that the rules of measuring require a bit more thinking on your part. You can see 15 mL, but you are required to estimate one more place to the best of your ability. This tube really shows 15.0 mL of water. You are going to have to estimate one place more than your tool shows you to get the most accurate measurement.

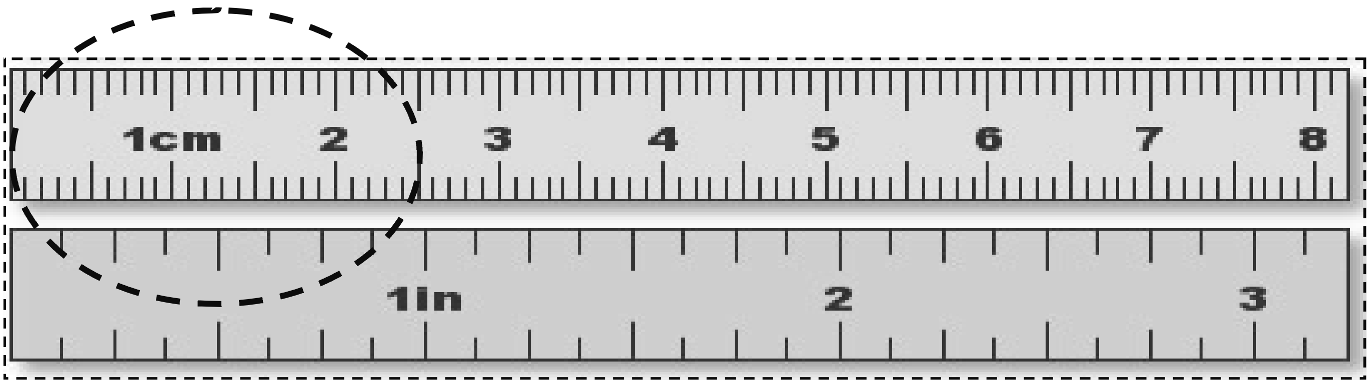
If you thought it was 15.00 mL that would be wrong too. Your eyes can never "see" that accurately. 15.000 mL would be an even worse measure, because you just can't really see, or estimate properly, to the thousandth of a mL unit.

The measurements you make require you to measure as accurately as your tools and eyes let you, not more and no less. The number of units you measure to are called the significant figures.

There are rules to using significant figures (a handout is coming called the Significance of Significant Figures). Rules exist for measuring, and rules for rounding in math problems as well.

Once you measure properly, you must keep your measurements as accurate as possible in any calculations you do with these measurements, not falsely gaining accuracy, or giving it away.

In this close up of two rulers, top metric in millimeters and centimeters, the bottom in English, with eighths of inches and inches. If we try to measure the length of the oval, we could say it's about 1 inch long. Or we could measure it to be 2.5 cm, or 25.0 mm. The smaller the increment, the more accurate the measurement. If we look at the significant figures here, 1 inch, 2.5 cm, or 25.0 mm, the number of significant figures is one, two, and three. The more significant figures, the closer the measurement it to the actual value.




Quickly, the rules of significant figures are as follows:
All digits 1 to 9 are always significant.
All zeroes *between* significant figures are also significant.
Zeroes on the right end of a number, *after* a decimal point are also significant (25.0 mm for example). Zeroes at the right end of a number, *before* a decimal point are also significant (100. meter dash).
Zeroes at the right end of a number *without a decimal point* are NOT significant (100 yard swim = 1 SF).

When using math, the one rule to follow is this: the answer must have the same number of significant figures as the *least number* of significant figures in the problem.
How do we calculate this area problem of: 3.67mm x 2.0mm = ?

3.67 mm x 2.0 mm = 7.34mm$^2$ **but that is INCORRECT** (your answer can have just 2 SF)
3.67 mm x 2.0 mm = 7.3 mm$^2$ with the correct number of SF in your answer.

Your answer cannot become "more accurate" than your weakest measurement,
nor have more significant figures than your least accurate measurement.

## Percent Error

Percent Error is a way to compare your measurement to the ACTUAL measurement by percentage. It will ALWAYS have a + or - sign. A positive percent error means YOU measured larger than the actual.

A negative percent error means your measure is less than the actual.

Percent Error with NO sign means you did the math wrong and I will always deduct a point from your score for NOT PAYING ATTENTION TO DETAILS. Please take heed.

If you get a percent error of ZERO that means you measured perfectly (congratulations!) and you may forgo the sign for this one instance. Pay attention to SF (significant figures) in your percent error calculations.

$$\text{Percent Error} = \frac{\text{Measured value - Accepted value}}{\text{Accepted Value}} \times 100\% =$$

If you measure a piece of metal to be 23.5 grams but it's really 23.1 grams, your measurement is "off", and you can measure how far off using the Percent Error Formula.

$$\% E = \frac{23.5 \text{ grams} - 23.1 \text{ grams}}{23.1 \text{ grams}} \times 100\% = -1.731601732\%$$

Your percent error is –1.73% even though your calculator can figure out your "answer" to be –1.731601732%. You are limited to just three significant figures in your answer.

That last "2" is in the billionths place of that decimal. You didn't measure your metal to that sort of accuracy, so your percent error can't be that accurate either. Follow the rules about this (or else)

The next part of significant figures is the easiest part (or hardest if you think too much). Sometimes we have what are called equal values, say 454 grams is equal to 1 pound, or 212°F = 100°C. When two or more values are known to be equal, you understand that they are perfectly equal. You could just as easily state that 454.00000000 grams is equal to 1.000000000000 pounds because they are equal exactly. Equalities have what are called UNLIMITED significant figures. They are equal to the "nth" degree.

You could add as many zeroes after the decimal point as you want, to either or both sides of the equality, so they are going to be significant as you want them to be. When you use an equality in your conversion math, they do not limit your answer's significant figures in any way. They have UN SF, unlimited SF.

Practice: how many SF in each of these math expressions?

A. 34.56 grams ___________

B. 107 mL ___________

C. 0.4516 cm ___________

D. 34.25 g divided by 2.33 $cm^3$ = ___________

E. 1.000000 g ___________

F. 1000000 g ___________

G. 100.0000 g ___________

H. 0.000001 g ___________

*answers below*

A. 34.56 grams 4 SF B. 107 mL 3 SF C. 0.4516 cm 4 SF

D. 34.25 g divided by 2.33 $cm^3$ = 3 SF E. 1.000000 g 7 SF

F. 1000000 g 1 SF G. 100.0000 g 7 SF H. 0.000001 g 1 SF

# Scientific Notation

In chemistry we will talk a lot of atoms and molecules. Atoms are the smallest parts of an element (the limited number of pure substances that make up all matter, all listed on the Periodic Table of Elements). Molecules are the smallest parts of molecular compounds, which are made up of 2 or more atoms that are chemically combined into new substances, with new properties, such as water or carbon dioxide. Molecules are bigger than atoms that make them up, but both are sort of unimaginably small. It takes so many of them to amount to be measured. We talk about numbers of these particles in numbers larger than billions. To express these huge numbers (or tiny ones) we use scientific notation.

$10^2$ means 10 x 10 = 100 $10^3$ means 10 x 10 x 10 = 1000

$10^{23}$ is 10x10x10x10x10x10x10x10x10x10x10x10x10x10x10x10x10x10x10x10x10x10x10 =

100,000,000,000,000,000,000,000 which is a number I can't actually name in English. Numbers that big require exponents to become more easily written.

There are rules to follow using these numbers. We'll only use ten to a power. And we'll multiply a number in front of the ten, so 1000 would be written as $1 \times 10^3$. The number 66,500 would be $6.65 \times 10^4$.

The number in the front is the coefficient. The coefficient is multiplied by a power of ten. The rules for significant figures apply to the coefficients only.

If you have one million atoms, you would write $1 \times 10^6$ atoms.

If you have 1,300,000 atoms, it would be written as $1.3 \times 10^6$ atoms. That has just 2 significant figures (the 1 and the 3)

When doing math, the answers are limited to the lowest number of significant figures are the lowest number in the coefficients in the math problem.

Multiplying Scientific Notation Math

To multiply in scientific notation, say $(2.0 \times 10^5)(3.0 \times 10^4) =$

You multiply the coefficients, 2.0 x 3.0 = 6.0 (with two sig figs only in answer)
Then add the powers of ten (5 + 4 = 9) The answer would be $6.0 \times 10^9$.

Dividing Scientific Notation Math

To divide in scientific notation, say → $\frac{9.00 \times 10^8}{3.0 \times 10^5}$

You divide the coefficients, 9.00/3.0 = 3.0
then subtract the powers of ten. 8—5 = 3 The answer would be $3.0 \times 10^3$

Note that 9.00 has 3 sig figs and 3.0 has 2 sig figs, your answer must have two SF as well—the same as the least number of sig figs in your math problem.

The Addition Rule take an additional get ready step, that is getting the exponents to match.

$$\begin{array}{r} 2.35 \times 10^{7} \\ \underline{+1.34 \times 10^{6}} \end{array}$$

This can't be done until you match the exponents to either both being 7th or 6th power (doing it either way will give the same answer). Then just work with the coefficients.

Change the powers of ten to $10^6$ or to $10^7$

| | |
|---|---|
| $23.5 \times 10^6$<br>$\underline{+1.34 \times 10^6}$<br>$24.84 \times 10^6 = 2.48 \times 10^7$ with 3 SF | $2.35 \times 10^7$<br>$\underline{+0.134 \times 10^7}$<br>$2.484 \times 10^7$<br><br>this changes to $2.48 \times 10^7$ with 3 sig figs |

Subtracting powers of ten rules are the same as for addition, except you subtract the coefficients instead of adding them.

The last "rule" for us in chemistry is that we always make our coefficients between 1.00 and 9.99.

It is true that $1.00 \times 10^9$ is the same as $10 \times 10^8$, or $100 \times 10^7$, we will always adjust our exponents so that our coefficients are more than one, but less than ten. We will always have to adjust our answers to comply with this rule.

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# Temperature Scales

We live in America, we use the Fahrenheit scale of temperature almost everywhere but science class. We'll almost never use it in chemistry, except in comparison to the other scales. Centigrade is the same as Celsius, but your teacher almost always will say centigrade. The third scale we'll learn is Kelvin, named after the famous chemist Lord Kelvin.

What ever the temperature is outside, or in the room, that temperature can be measured on different scales, but it's still the same temperature.

Room temperature is about 20°C, but it's closer to 293 in Kelvin. The numbers are really different, but the temperatures are the same. Kelvin NEVER uses degrees, they are just Kelvin.

As shown in this chart, the 3 scales are related as follows…

|  | F | C | K |
|---|---|---|---|
| water boils | 212 | 100 | 373 |
| water freezes | 32 | 0 | 273 |
| absolute zero |  | -273 | 0 |

Water freezes at STANDARD TEMPERATURE.
On table A of your reference tables this is pointed out.

To convert from centigrade to Kelvin, or vice versa, use this formula:

$$K = C + 273$$

That formula is also on your reference table, table T on the back page.

We will not need to convert to or from Fahrenheit in our class.

Example: What temperature in Kelvin is steam at 105°C?
$K = C + 273$
$K = 105 + 273 = 378$ Kelvin.

Kelvin units are Kelvins, NOT DEGREES. No little circles indicating degrees as is for °C or °F.

What is a really hot day outdoors in Kelvin, if the outside temperature is 36.0°C? (answer at bottom)
Always, always start with a formula, use units, watch your SF.

Absolute zero is a theoretical temperature. It is the temperature so low that all atomic motion stops. Scientists have gotten close to, but cannot ever get to absolute zero, but explaining that involves a lot of talk that is not part of our course, especially now. If all motion "stops" (and it does), time stops as well, at least where the temperature is 0K (okay?). Also, if you could get to this temperature, just being close enough to observe it would impart some energy on it, raising above absolute zero.

What is a really hot day outdoors in Kelvin, if the outside temperature is 36.0°C? (do this answer below) Always, always start with a formula, use units, watch your SF.

$K = C + 273$ $K = 36.0 + 273 = 309$ K (3 SF)

# Dimensional Analysis

In science, or math, you can label different measurements with different units, but all of them measuring the exact same thing properly. Different measures of the same thing. Different units.

I might be five feet eight inches tall. And I am 68 inches tall. You might measure my height in meters, centimeters, millimeters, or even miles! Each of the numbers would be different and each would have a different unit. All of them are equal to each other (with their proper units).

To convert from one unit to another mathematically is called unit conversion, or dimensional analysis. It's actually sort of fun, but requires you write every single unit or else you will make big mistakes in the math. With the units, you really can't make a mistake.

If you multiply any number by one, you get the same number.

12 x 1 = 12 10000 x 1 is still 10000 234 x 1 = 234

But one can be written in many different ways

$\frac{2}{2}$ Is the same thing as 1

$\frac{157}{157}$ Is the same thing as 1

$\frac{12 \text{ inches}}{1 \text{ foot}}$ Is the same thing as 1

$\frac{60 \text{ seconds}}{1 \text{ minute}}$ Is the same thing as 1

When we create a sort of fraction, with equivalent units in the numerator as in the denominator, we are essentially creating a new way to write "1". All equalities can create 2 conversion factors. If 12 bagels = 1 dozen bagels, then this is true (it is).

$$\frac{12 \text{ bagels}}{1 \text{ dozen bagels}} = 1 = \frac{1 \text{ dozen bagels}}{12 \text{ bagels}}$$

$$\frac{12 \text{ inches}}{1 \text{ foot}} = 1 = \frac{1 \text{ foot}}{12 \text{ inches}}$$

Since these fractions are equal to exactly one, we can multiply by them and change units, but not the actual value. For example, how do you convert from inches to feet? How many feet is 5700 inches? Most students could figure this out, but there is an easy way to convert that many inches to feet, just convert using dimensional analysis.

Think of a good conversion factor...

$$\frac{1 \text{ foot}}{12 \text{ inches}} = 1$$

How do we convert 5700 inches in to feet?

$$\frac{5700 \text{ \sout{inches}}}{1} \quad \frac{1 \text{ foot}}{12 \text{ \sout{inches}}} = \mathbf{475 \text{ feet}}$$

Convert 1.50 pounds to grams.
Pounds in the numerator will cancel pounds in the denominator of the conversion factor, do the math, keep the unit you need in your answer, check sig figs.

To do this you need to know other conversion factors. Some you must know, some you should know. We'll practice many of them all year.

This is an easy, one step conversion. The conversion factor is on every can of corn in your life.

$$\frac{1.50\ \cancel{\text{pounds}}}{1} \times \frac{454\ \text{grams}}{1\ \cancel{\text{pound}}} = 681\ \text{grams}$$

An elephant weighs in at 6.5 tons. Convert to grams in scientific notation. Some problems have multiple steps, multiple conversions, to go from tons to grams for instance. There is a conversion factor of 1 ton = some number of grams, but I don't know it. I do know others, but it will take multiple steps to do the math

$$\frac{6.5\ \cancel{\text{tons}}}{1} \times \frac{2000\ \cancel{\text{pounds}}}{1\ \cancel{\text{ton}}} \times \frac{454\ \text{grams}}{1\ \cancel{\text{pound}}} = 5902000\ \text{grams}$$

$$5902000\ \text{grams} = 5.9 \times 10^{6}\ \text{grams}$$

6.5 tons has only 2 significant figures. Both conversion factors have numerators equal to their denominators, so they both have UNLIMITED significant figures. Your answer is limited to have just 2 SF. That is how to "round" in chemistry class. You can't be more accurate than two significant figures here.

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One more "real" problem, convert 2.5 years into seconds, then convert to scientific notation.

$$\frac{2.50\ \cancel{\text{years}}}{1} \times \frac{365\ \cancel{\text{days}}}{1\ \cancel{\text{year}}} \times \frac{24\ \cancel{\text{hours}}}{1\ \cancel{\text{day}}} \times \frac{60\ \cancel{\text{minutes}}}{1\ \cancel{\text{hour}}} \times \frac{60\ \text{seconds}}{1\ \cancel{\text{minute}}} =$$

This time it takes four different conversion factors to convert all the way from years to seconds. It could take just one step, but you would need to know the number of seconds in one year! I don't know that conversion factor, but I do know the rest are all equal to zero. The number of steps doesn't matter, as long as you multiply by "one" over and over, it's all the same.

Doing the math, watching for significant figures = 78840000 seconds

Since we're limited to 3 significant figures (from 2.50 years) your answer is 78,800,000 seconds, *(rounded correctly)* or written in scientific notation $7.88 \times 10^{7}$ seconds

Remember, all conversion factors have equal numerators and denominators, so they have UNLIMITED SF.

Imagine if you made a crazy error and put the 1 day = 24 hour conversion factor upside down above, so it looked like this:

$$\frac{1 \text{ day}}{24 \text{ hours}}$$

This equality does = one, but your units will NOT CANCEL OUT.

Your answer will be zany. Unless you "fake" the units, then you will just be WRONG. USE your units, make them cancel out so you KNOW that all the numbers are in the right place.

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Sometimes to see if you're really thinking the Regents will test you in dimensional analysis using make believe units. The units are there to set up the math, to cancel each other out, and to get the proper answer, with proper sig figs. Don't sweat the strangeness of some problems. It's a math *game*, but an *excellent learning tool* to solving bigger chemistry problems, as we'll see.

Last problem:

1.0 pigs equal 1.6 dogs
2.2 dogs is equal to 0.95 cats
1.9 cats is equal to 3.1 birds
1.0 bird is the same as 11.0 spiders
3.7 spiders is the same as 8.5 bugs

If this is true, how many bugs make up 1.0 pig? Convert to the nearest whole number of bugs.

$$\frac{1.0 \text{ } \cancel{\text{pig}}}{1} \times \frac{1.6 \text{ } \cancel{\text{dog}}}{1.0 \text{ } \cancel{\text{pig}}} \times \frac{0.95 \text{ } \cancel{\text{cat}}}{2.2 \text{ } \cancel{\text{dogs}}} \times \frac{3.1 \text{ } \cancel{\text{birds}}}{1.9 \text{ } \cancel{\text{cats}}} \times \frac{11.0 \text{ } \cancel{\text{spiders}}}{1.0 \text{ } \cancel{\text{birds}}} \times \frac{8.5 \text{ bugs}}{3.7 \text{ } \cancel{\text{spiders}}} =$$

Do the math, cancel all units in order, make sure you watch out for SF (you're limited to the 2 SF in the 1.0 pigs from the question. All other significant figures in the conversion factors are unlimited.

So, $$\frac{1.0 \times 1.6 \times 0.95 \times 3.1 \times 11.0 \times 8.5}{1 \times 1.0 \times 2.2 \times 1.9 \times 1.0 \times 3.7} = \frac{440.572}{15.466} = 28.48648649 \text{ bugs} = 28 \text{ bugs}$$

This is clearly a wacky problem, but is shows a proper dimensional analysis set up, proper use of units, proper cancelling of units, and proper significant figures. If you can follow this, dimensional analysis will be easy!

# Density

Density is the relationship between the mass and the volume of matter.
The formula is…

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \qquad \text{Which is often abbreviated to...} \qquad D = \frac{m}{v}$$

Because the mass and volume are in a particular relationship, it needs to be clear that the more mass of matter you have, the same proportion of volume increase is needed. Because this is true, no matter how much mass you have, your proportional volume will work out in the formula to a CONSTANT.

The density of pure water is 1.00 g/mL.

If you have 57 grams of water it will have 57 mL volume, density = 1.00 g/mL

If you have 9,825 grams water, it's volume is 9,825 mL, density = 1.00 g/mL

If you have 0.000356 g water, the volume is 0.000356 mL, density is the same.

For any kind of matter, density is constant.

Units for density are either grams/milliliter (g/mL) or grams/centimeter cubed (g/cm$^3$)

Since these volumes, 1 mL = 1 cm$^3$, we can interchange them whenever we want to.

You will be required to use the formula above to solve for density, mass or volume.

1. An unknown substance has mass of 89.0 grams and your volume is 10.05 cm$^3$.
   What element could it be from the chart below?

$$D = \frac{m}{v} \qquad D = \frac{89.00\text{g}}{10.05\ \text{cm}^3} \qquad D = 8.85572139\ \text{g/cm}^3$$

8.86 g/cm$^3$ with 3 SF

It's cobalt!

Elements and their density values

| | | |
|---|---|---|
| copper | Cu | 8.96 g/cm$^3$ |
| nickel | Ni | 8.90 g/cm$^3$ |
| cobalt | Co | 8.86 g/cm$^3$ |
| tin | Sn | 7.29 g/cm$^3$ |
| iron | Fe | 7.87 g/cm$^3$ |

2. Your piece of copper has a volume of 25.00 $cm^3$. What is the mass of this copper?

$$D = \frac{\text{mass}}{\text{volume}} \qquad 8.96\ \text{g/cm3} = \frac{\text{mass}}{25.00\ \text{cm}^3}$$

Solve for "m" by multiplication, 8.96 x 25.00 = 224 grams (3 SF) (units omitted, but they do cancel out

3. Another piece of copper has an irregular shape too big for a graduated cylinder. It has a mass of 923.4 grams. What is the volume of this metal?

$$D = \frac{\text{mass}}{\text{volume}} \qquad 8.96 = \frac{923.4}{V}$$

Do the algebra to solve for "x" volume -cross multiply, 8.96 x V = 923.4

Solving for V, V = 923.4 g / 8.96 = which becomes 103.058 $cm^3$, which is 103.1 $cm^3$ with 4 SF.
Here, the density of copper has UNLIMITED SF, it's from a table, not a measurement.
The answer must round to 4 SF.

Density is a physical constant. Every element you will need to know about has the density listed in Table S. Water, when pure has a density of 1.00 g/mL. Ice, which of course is solid water has a density slightly less than that, therefore ice floats in water. Only rarely does a solid float in its own liquid phase.

When you have more than one liquid, the denser one goes to the bottom of a container, while the less dense one floats above. Oil floats on vinegar in Italian restaurants, gasoline floats on water. Hopefully you will never see that water can float on mercury (really neat, but mercury is dangerous to your health!)

If there are multiple liquids, they will arrange into layers, most dense at the bottom, least on top.

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Challenge problem.

You find a hunk of shiny metal that is in a box that says pure silver (atom #47). The mass is 485 grams and the volume measures out to be 39.2 $cm^3$ - could it be silver?

Solve for the density = mass/volume

Density = 485 grams/39.2 cm3 = 13.4 g/cm3. The density of silver is on table S, and it's 10.5 g/cm3

Hafnium has density of 13.3 $g/cm^3$, depending upon how well you measured that seems more likely.

It's probably not silver.