Your name: ________________________________________________ Period: _____

# Measurement

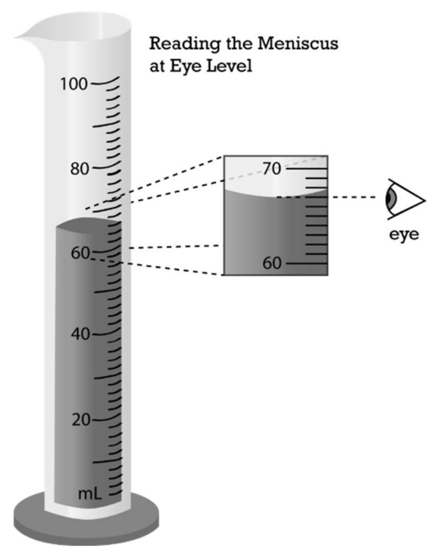
All of this is on Arbuiso.com, including the slide show we are using in class. If you are absent, you can follow along at home. You are responsible for getting the notes when you are absent. In addition to the notes, you are to read the included Measurement BASICS, which is basically everything you need to know about this topic.

You are expected to know this material, and the material that is reviewed in class and that you will learn about in the lab experiments. There is overlap, but some material is just from one of these three places.

Stay off the internet. New York State Regents Chemistry uses specific constants and numbers that are not interchangeable, even if your numbers are "better" or more accurate. You will use ONLY numbers and constants from our reference tables, our notes, and class notes. If you use the internet, it will be obvious, and your work will be marked with a big X.

If you need help, call me anytime. If I am available, I'll answer the phone. If I'm not, I'll call you back soon. Mr. Arbuiso cell: 607-727-3865. You can text too, but sometimes chemistry is hard to text, I may text you to call me. You won't be bothering me; I am your teacher until late June (and then some).


Aluminum is element #13
It has a density of 2.70 $g/cm^3$




We read our graduated cylinders by bending our knee and looking straight at it, and read to the NEAREST 10th milliliter.

Welcome! Let's start learning. Copy the objective of today's class in blank #1. Keep filling in the blanks.

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|---|---|
| 1 | Objective: |
| 2 | What is the chemistry formula for water? |
| 3 | The H stands for which happens to be element # |
| 4 | The O stands for which happens to be element # |
| 5 | The "little 2" means that there are |
| 6 | There is no "little 1" by the oxygen, why not? |
| 7 | Both hydrogen and oxygen are special elements in that they do not exist as single atoms when they are pure, not bonded to other atoms.<br><br>Their real formula would be and because they are DIATOMIC elements, which means that they are paired together when they are pure, when they are in their "elemental" state. |
| 8 | The arrow means |
| 9 | The skeleton formula for the synthesis reaction that combines $H_2 + O_2$ to form water is written this way |
| 10 | It's pretty obvious that we are missing one atom of oxygen on the right side of the arrow. Are we allowed to "lose" matter like this? |
| 11 | There is something called the LAW OF CONSERVATION OF MATTER which is |

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| 12 | There is more to the reaction, since we can't lose even a single atom. Watch how it's balanced, then copy the balanced chemical reaction |
| 13 | There never is a #13 in our class (except for the regents, or aluminum which is a great metal). |
| 14 | There are about 16 different kinds of chemical reactions that we will learn about this year. This one is called SYNTHESIS, which means combining 2 or more |
| 15 | In this reaction there are 2 reactants, which are |
| 16 | There is only one product, which has a science name of<br><br>but we can call it |

This reaction will release a LOT of energy, a FIREBALL of flame and heat, and NOISE.
It's safe, but it might surprise you too. Look at this photo, that's how you will protect your ears.
Let the sound hit your face, and blast past your hands. Let the sound bounce off of the wall behind you, and then catch it with your hands.
Try not to blink! This will look a lot like burning, which is casual talk for COMBUSTION,
but it's not that. The butane lighter is combusting. The candle is also a combustion reaction,
but the balloon will just be really fast and really loud SYNTHESIS.
When energy is released in a chemical reaction, we call it EXOTHERMIC.

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| 17 | The balanced "thermochemical reaction" will look like this |

<u>Tonight for Homework</u>

A. Read the whole 1st Day Handout; your parents NEED to look it over as well.

B. Fill in the Student Information Handout neatly. Numbers and letters need to be clear.

C. Get psyched, we are going to learn A LOT and we're going to have fun too. I can't wait.

Measurement Class #1 Percent Error and Density
Objective: Learning about Percent Error and how to make Density Calculations.
When we measure in chemistry we need to make the best measurements we can. We will try to be

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| 18 | ACCURATE |
| 19 | PRECISE |
| 20 | Mostly we'll try to be BOTH accurate and precise. |
| 21 | Our measurements are called the values. |
| 22 | The real measurements (the truth) come from science tables, and they are called the |
| 23 | Take out a ruler and measure this page top to bottom, in centimeters. It is cm |
| 24 | Measure it again to the nearest 10th of a centimeter now. cm |
| 25 | The actual length is 27.9 cm. How far "off" was your measurement in cm?<br>That boo-boo is called your ERROR. It's vocabulary - but we won't use error again. |
| 26 | What we will use to measure how close our measurements are to accurate is called percent error.<br>Write the formula for percent error below. Find it on the back of your reference table now.<br><br>Below that, we'll write it out in short hand. You must know both. Don't forget the % sign! |

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|---|---|
| 27 | Let's calculate your percent error on this measurement. If you happened to get 27.9 cm exactly (good for you) use 28.3 cm as your measured value instead. Write the formula again, shorthand style. |
| 28 | Let's use your eyes to measure the mass of your chemistry teacher in POUNDS.<br><br>I measure my teacher's mass to be _________________________ pounds (MV).<br><br>According to (the LOUSY) school scale, the actual value is ______________________ pounds (AV). |
| 29 | Calculate your percent error for the mass. Write the formula in shorthand first. |
| 30 | If your %E is positive, that means your measured value… |
| 31 | If your %E is negative, that means your measured value… |
| 32 | If you have no sign, positive or negative, that means… |
| 33 | Percent error ALWAYS gets a sign, or else… |

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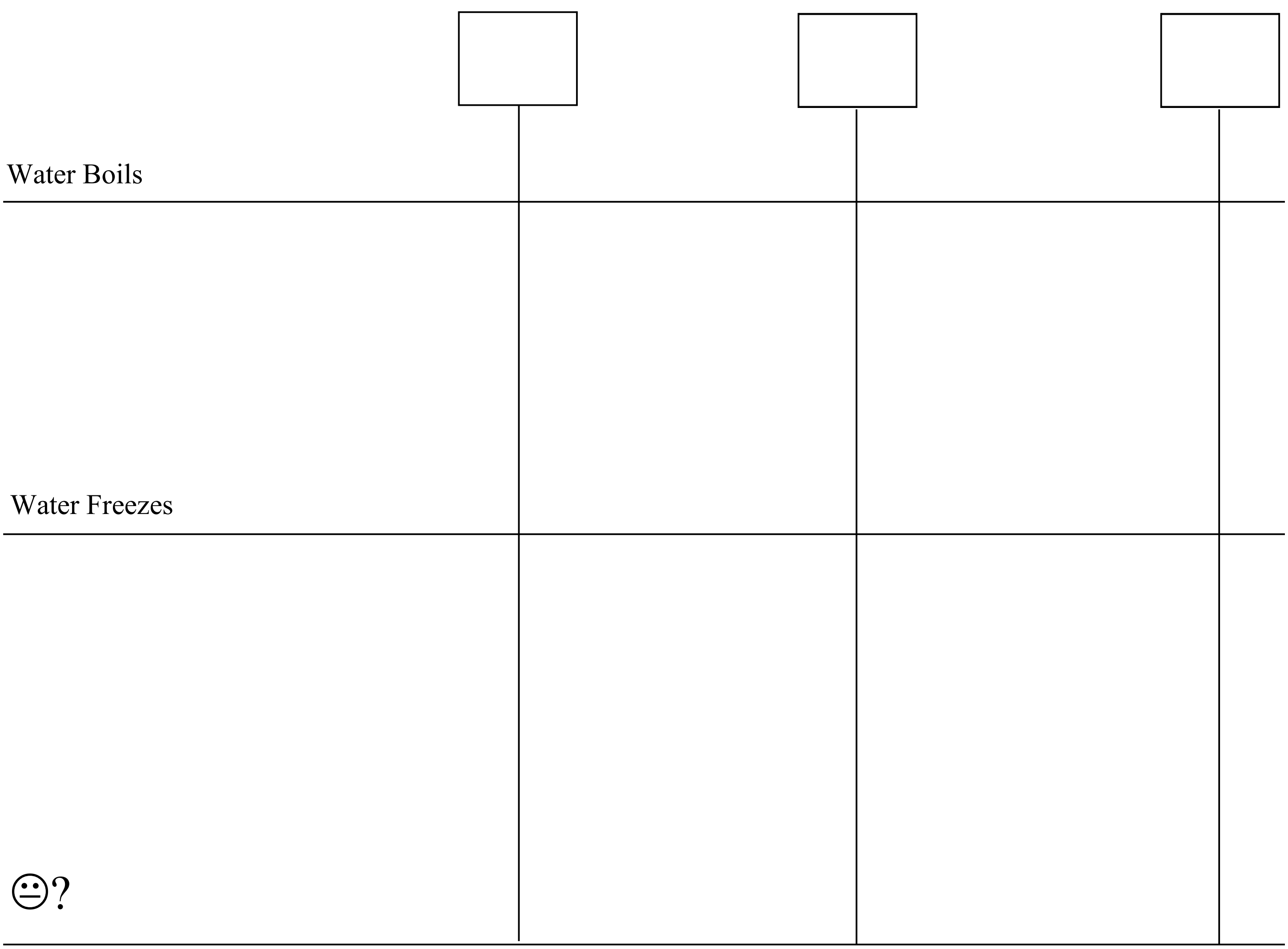
| 34 | You measure the density of copper to be 8.75 $g/cm^3$. What is the density of copper? (how do you know?) |
| 35 | Write the percent error formula in shorthand, calculate your percent error carefully. Get a SIGN too! |
| 36 | Is your measured value in this problem more than or less than the actual value?<br><br>Does your SIGN for percent error make sense? |
| 37 | The formula for Density is on the back of the reference table. Copy it then write it in short hand. |
| 38 | A bar of metal is 27.73 g and has volume of 4.70 $cm^3$. Is it gold? (start with density formula) |
| 39 | If it's not gold, what metal might it be instead? |
| 40 | You measure a hunk of aluminum to be 363 grams and have volume of 148 mL.<br>What is your measured density? What is your percent error? |

| Measurement Class #2 Temperature Conversions, Centigrade and Kelvin (and NOT Fahrenheit) | |
|---|---|
| 41 | Centigrade is another way to say Celsius, but “centi” reminds us of cents, and there are __________ units of temperature from melting ice to boiling water temperature. It’s a Metric Temperature Scale. |
| 42 | Another scientific temperature scale is called the ____________________________________ scale.<br><br>It’s named after a guy named Lord _________________________, no relation to Lord Vader of Star Wars. |
| 43 | There are 100 units of temperature from melting ice to boiling water temperature.<br><br>It’s a Metric Temperature Scale too. This scale DOES NOT USE degrees, just |
| 44 | Let’s fill in this chart now (even with the F scale). |



| | Pros | Cons |
|---|---|---|
| Fahrenheit | 186 units (!) between melting ice + boiling water. There is less estimation, and almost everyone in the US knows this | Non-metric (wacky numbers) no scientists can use it, it won't work with our formulas at all. |
| Centigrade or Celsius | 100 units between melting ice + boiling water, totally metric. Almost everyone in the world except the US and England knows this | When it gets chilly in Vestal, you can get negative numbers which make the math go nuts. You can also have a zero temperature measurement on a relatively normal day. Zeros ruin the math. |
| Kelvin | 100 units between melting ice and boiling water, totally metric. Almost every scientist knows about it and it always works. It's the absolute temperature scale | Normal people have no clue. The numbers are weird. There is a conversion formula to learn because our thermometers are centigrade, but lots of our math is in Kelvin. |

To convert from Centigrade → Kelvin, or from Kelvin → Centigrade, we use the SAME formula.
It's on the back of the reference table, let's copy it now. Write it big, like you care.

| | |
|---|---|
| 45 | The boiling point for zinc metal is 693 Kelvin. Convert that to centigrade, use a formula. |
| 46 | The melting point of iron is ______________ Kelvin. Convert that to centigrade with a formula. |
| 47 | Chlorine gas boils at ___________________ Kelvin. Convert to centigrade with a formula.<br>Is this COLD or HOT? Explain that. |

| | |
|---|---|
| 48 | You have 416 $cm^3$ of iron. What is its mass? (hint, write density formula, fill it in neatly) |
| 49 | You measure the density of gold to be 19.7 $g/cm^3$. What is your percent error? |
| 50 | You measure the melting point of lead to be 615 Kelvin. What is your percent error? |
| 51 | You measure 18.25 grams of silver on the scale. What is the volume of this silver? |
| 52 | Convert the room temperature of 26.0°C into Kelvin. Round to three digits. |
| 53 | You measured the density of pure water to be 0.975 $g/cm^3$, but everyone knows that water's density is exactly 1.00 $g/cm^3$. What's your percent error? |

Measurement Class #3 Significant Figures

Significant figures are all of the numbers that you MEASURE that are important. There is a difference between what you measure and what you look up on a table, or even what you already "know".

They are the numbers that mean something that are not place holders. They are used to figure out how many places we are allowed to round our calculator answers to. Just because the calculator says something does not make it real. The SF will show you how many places your real answers are allowed to have. There are several easy rules, which you will have to master.

| | | | |
|---|---|---|---|
| 54 | 127.25 grams | 5 SF | All digits that are not zeros are significant |
| 55 | 107.25 grams | | All digits are significant, any zero that is IN BETWEEN SF is also significant |
| 56 | 0.625 meters | | "Leading" zeros are not SF. From the left, the first SF is a digit |
| 57 | 100. grams | | Because of the decimal point, that last zero is a SF.<br>The zero in the middle is between SF |
| 58 | 100 meters | | No decimal, the last zero is not a SF, the middle one is not between SF |
| 59 | $2.245 \times 10^4$ atoms | | we only have SF in the coefficient, or front part of scientific notation |
| 60 | 14.50 grams | | a zero at the "end" of a number and AFTER a decimal point is a SF |
| 61 | Density of water is 1.00 g/mL | | Unlimited SF means that in an equality, or numerical facts from tables will not limit your answer, but your measurements will. |

62. Let's look over these measurements and determine the number of SF present. The first one is an example

| | | | | |
|---|---|---|---|---|
| 200. grams of Mg<br>3 SF | 35.66 grams Cu | 100 cm | 100. mm | 4,005,033 atoms |
| 0.552°C | 1.552°C | 10.552°C | 23.00552°C | $6.02 \times 10^{23}$ atoms |
| 1.00 g/mL | 1.000 g/mL | 1.00000 g/mL | 0.0000005 grams | $3.550 \times 10^{-17}$ grams |
| The density of metal with 125 g mass and volume of 35 mL | Calculate density with mass of 1025 gm & volume of 650 mL | The answer to 5.000 g divided by 3.88 $cm^3$ | the answer to 2.5 $cm^3$ x 5.6788 $g/cm^3$ = | |

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|---|---|
| 63 | You measure the density of nickel metal to have density of 9.1 g/cm$^3$.<br>What is your % Error with the correct number of SF? |
| 64 | You measured your floor to be 14.5 feet X 15 feet and you want a rug. How many square feet of rug do you need to buy? (SF count) |
| 65 | You measure the mass of metal to be 74.35 grams and its volume to be exactly 12.0 mL.<br>What is the density of this metal with correct SF? |
| 66 | You know that 12 inches = 1 foot. How many inches are in 8,375 feet? (SF count!) |

| | | |
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| 67. How many SF are in these measurements? | 0.0000164 g/cm$^3$ (density of a gas) | 7180 K (melting point) |
| 303 K (a melting point) | 3560. K (a boiling point) | 640 grams (mass of a liquid) |

# Measurement Class #4 Dimensional Analysis or Math Conversions

Converting from one unit to another is going to happen a lot in chemistry. We already have done conversions between Kelvin and centigrade and converted mass and volume into density. Other conversions will need to be made as well. Some are easy, they are one step conversions, but some are multiple step, and require this "technique" of dimensional analysis, which is just scary talk for unit conversion math. Here goes.

Nothing says LOVE more than a dozen roses.

How many roses are a dozen? _________ There, you did dimensional analysis in your head.

How many roses are in 3 dozen roses? _________ Ha! again. You converted dozens of roses → roses.

If 12 inches = 1 foot, how many inches are in 4.0 feet? ___________ Another conversion, feet → inches.

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| 68 | Let's look at what your brain was doing. First you decide your "starting point" and put that number "over 1" as a fraction.<br><br>Your starting point X Your conversion factor = Your answer with proper SF<br><br>2 SF Unlimited SF Limited to 2 SF |
| 69 | Convert 1.24 kilograms into grams (1 kg = 1000 g). (watch SF) |
| 70 | Convert 56,750 mL into liters (l L = 1000 mL). (watch SF) |
| 71 | Assume you are EXACTLY 16.33 years old <u>right NOW</u>. Convert that into minutes. |

72. Convert the 400. meter race into yards so the football players can easily compare that length to their field.
1 inch = 2.54 centimeters

73. Below are a set of equalities that are meaningless. This problem is to practice finding your "starting point" and cancelling out units properly, to get the right answer. This is mental exercise, and you should take it as a personal challenge.
How many blinks are in 244 winks? (round to the nearest whole blink)

6.75 zinks = 1.09 blinks
14 winks = 3.4 jinx
0.95 pinks = 2.0 zinks
7.0 jinx = 2.11 pinks

74. It's exactly 7.10 miles to Binghamton High School from Vestal High School according to mapquest.com
Convert that distance into meters. 1 inch = 2.54 cm 1 mile = 5280 feet 100 cm = 1 m

# Measurement Class #5 Scientific Notation for fun and personal enjoyment!

Really big and really small numbers are often presented in scientific notation, and look like this:
$6.02 \times 10^{23}$ atoms is one mole of atoms; the density of helium is $1.64 \times 10^{-4}$ grams/mL

The front part of the number is called the coefficient, and the back part is called the exponent.
In our class the RULE for scientific notation is that the coefficient MUST BE at least 1, but less than 10.
If your answers work out weirdly, you must change your coefficient and exponent to abide by this rule.

## Converting big and small numbers into scientific notation first.

| | |
|---|---|
| 75. 17,000,000,000 ants | 76. 6,374,000 meters |
| 77. 0.034 grams | 78. 0.000000000154 meters |
| 79. 0.0000083 meters | 80. 4,500,000,000,000 years |

## Convert scientific notation into numbers

| | |
|---|---|
| 81. $6.02 \times 10^{23}$ molecules | 82. $3.5 \times 10^{4}$ grams |
| 83. $1.25 \times 10^{-7}$ meters | 84. $2.290 \times 10^{3}$ Kelvin |

85. Convert 36.8 kilograms into ounces, your answer to be given in scientific notation.
Hints: 454 g = 1 pound = 16 ounces 1kg = 1000 grams Round to correct SF

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| 86 | Convert 300. seconds into years, your answer must be in scientific notation<br>(hint: your answer will be a small fraction of years, your exponent will be negative) |

## Rules to use scientific notation in math problems…

Multiplication Rule for Scientific Notation

| | | |
|---|---|---|
| 87<br>88 | $(3 \times 10^5)(2 \times 10^2) =$ | $\begin{array}{r} 5.0 \times 10^4 \\ \underline{\text{X } 3.0 \times 10^2} \end{array}$ |

Division Rule for Scientific Notation

| | | |
|---|---|---|
| 89<br>90 | $\frac{3.0 \times 10^4}{2.0 \times 10^2}$ | $\frac{9.0 \times 10^5}{3.0 \times 10^3}$ |

| | |
|---|---|
| 91 | Addition Rules for scientific notation |

| | | |
|---|---|---|
| 92<br>93 | $\begin{array}{r} 6.5 \times 10^7 \\ \underline{+ 2.2 \times 10^7} \end{array}$ | $\begin{array}{r} 6.2 \times 10^8 \\ \underline{+ 1.5 \times 10^6} \end{array}$ |

Subtraction Rules for scientific notation

| | |
|---|---|
| 95. $$\begin{array}{r} 8.5 \times 10^{3} \\ \underline{-\ 2.4 \times 10^{3}} \end{array}$$ | 96. $$\begin{array}{r} 7.1 \times 10^{5} \\ \underline{-\ \ 1.6 \times 10^{4}} \end{array}$$ |
| 97. $$\begin{array}{r} 8.72 \times 10^{11} \\ \underline{+\ 1.72 \times 10^{10}} \end{array}$$ | 98. $$\begin{array}{r} 4.65 \times 10^{14} \\ \underline{-\ 2.25 \times 10^{15}} \end{array}$$ |
| 99. $$\begin{array}{r} 6.02 \times 10^{23} \\ \underline{\times\ 1.50 \times 10^{2}} \end{array}$$ | 100. $(9.05 \times 10^{19}) \div (3.2 \times 10^{16}) =$ |

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| 101 | Convert 2450 mL into gallons. Show all units (3 SF) (1.06 Qt = 1 L) |
| 102 | How many millimeters are in 1000. yards? Put answer into scientific notation. |

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| 103 | You have a special moment and discover a hunk of metal in your yard in the dirt. It's stamped "pure osmium" and "100.0 grams" as well. It looks pretty new and you believe it is real.<br>What is the volume of this hunk of metal in $cm^3$? Show a formula and all your work. Use SF. | | |
| 104 | REVIEW: determine how many significant figures are in each of these measurements | | |
| 5,600 g | | 5.600 kg | 4.305 mL |
| 0.678°C | | 0.00065 moles | 1.400 seconds |
| 105 | Calculate the quotient: 4.569 g ÷ 2.0 $cm^3$ = | | |
| 106 | In each set of temperatures, decide which is the coldest, which is the hottest.<br>SET A: 10 K 10 C 10 F SET B: 280 K 32°F 6.0°C | | |
| 107 | Convert 125 grams into kilograms | | |
| 108 | Convert 34.75 liters into mL | | |
| 109 | You measure some pure niobium metal to have a density of 8.00 $g/cm^3$. What is your percent error?<br>(hint, write the formula first) | | |

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| 110 | $(3.5 \times 10^6) \times (2.0 \times 10^2) =$ |
| | $(8.0 \times 10^8) \div (4.0 \times 10^{12}) =$ |
| | $(3.3 \times 10^8) + (1.2 \times 10^7) =$ |
| | $(5.64 \times 10^5) - (2.33 \times 10^4) =$ |
| 111 | Look up the boiling point of aluminum, covert that into scientific notation. |
| | Convert the boiling point of aluminum into centigrade. (use a formula, watch out for SF!) |

I cannot stress how important it is to read the BASICS.
The BASICS is a sort of textbook that your teacher has put together over many years, that highlights basically all that you need to know for this topic.

There is a BASICS for each of the 19 topics we cover in chemistry.
It's like your textbook; you are required to read it.

Sometimes there are questions on the celebration that are NOT spoken aloud in class, but are included in the reading. If you don't read, you will NOT get a 100%.

It starts on the next page, and the sooner you read it the smarter you'll get.

# 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: the solutions is blue, or the iron bar is cold.

## 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 is always perfect, if we can be careful enough in lab, we'll be able to prove that to ourselves.

When you make a measurement that is very close or perfectly correct, that measurement is said to be accurate. An accurate measurement is right—it is the same as the ACTUAL or ACCEPTED VALUE.
This is good.

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, or they might be inaccurate, but they are close together. If your measures are close together but inaccurate, you are measuring properly, but your tool is broken.

In chemistry class the plan is to be both accurate & precise every time we measure.


In this close up of two rulers, top metric (mm and cm), the bottom in English (eighths of inches & 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.

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 m 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 3.67cm x 2.0cm =?

3.67 cm x 2.0 cm = 7.34 cm$^2$ **but that is INCORRECT** (your answer can have just 2 SF)
3.67 cm x 2.0 cm = 7.3 cm$^2$ 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

% Error is a way to compare your measurement to the ACTUAL measurement by percentage. It will ALWAYS have a + or - sign. A POSITIVE % ERROR means YOU measured more than the actual.

A negative % Error means your measurement is less than the actual value.

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.

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

This drawing represents darts shot at a target.
On the 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 (precise but not accurate).

At the far right the darts are all very close to the center, which makes them both 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 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 an accurate measurement.

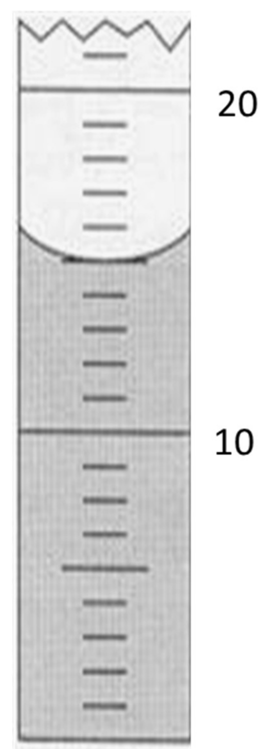
If you thought it was 15.00 mL that would be wrong too. Your eyes can never "see" that accurately. 15.000 mL would be an even worse measure, because you just can't really see, or estimate properly, to the thousandth of a mL unit.

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 places 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 there are rules for doing math with SF as well.

Once you measure properly, you must keep your sig figs consistent in all calculations you do with these measurements. You are not allowed to get more accurate (nor less) than the original measurements were made to.

Sig Figs are not free, nor can you give them away. They will be right, or they will be incorrect. Learn them ASAP.



| |
|---|
| 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. |

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

Your percent error is really just –1.73% even though your calculator said your "answer" was –1.731601732%. Your answer must be limited to just 3 SF.

That last "2" above 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 = 1.000000000000 pounds because they're 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 unlimited SF.

# Scientific Notation

In chemistry we 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 an amount we can feel. 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 name in English.
Numbers that are big like this one usually get written with exponents – in this case $1 \times 10^{23}$.

There are rules to follow using these numbers. The front number is the coefficient and it has to be at least 1 and less than 10. So 1000 would be written as $1 \times 10^3$

The number 66,500 would be $6.65 \times 10^4$

The rules for significant figures apply to the coefficients only.

If you have 349 million atoms, you will write that as $3.49 \times 10^6$ atoms.

If you have 1,300,000 atoms, it would be written as 1.3 x $10^6$ atoms. That has just 2 SF (the 1 and 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.

<table>
<tr><th colspan="2">Multiplying Scientific Notation Math</th></tr>
<tr><td>To multiply (2.0 x $10^5$) x (3.0 x $10^4$) =<br>Multiply the coefficients<br>Add the exponents</td><td>2.0 x 3.0 = 6.0 (2 SF each, so 2 SF in answer)<br>(5 + 4 = 9)<br>The answer is 6.0 x $10^9$.</td></tr>
<tr><th colspan="2">Dividing Scientific Notation Math</th></tr>
<tr><td>To divide (9.00 x $10^6$) ÷ (3.0 x $10^5$) =<br><br>You divide the coefficients<br>then subtract the exponents</td><td>9.00/3.0 = 3.0<br>8 - 5 = 3<br>The answer is 3.0 x $10^3$<br><br>Note that 9.00 has 3 sig figs and 3.0 has 2 sig figs, your answer must have two SF, the same as the least number of sig figs in your math problem.</td></tr>
<tr><th colspan="2">Adding with Scientific Notation</th></tr>
<tr><td>To add these two expressions<br>2.35 x $10^7$<br>+1.34 x $10^6$<br><br>first you must change them (temporarily) so they have the same exponent (six or seveeeeen, sorry).</td><td>Once both exponents match,<br>Just add the coefficients and keep the exponents.<br>If your coefficient is less than 1 or more than 10,<br>adjust your answer.</td></tr>
<tr><td>This becomes…<br>2.35 x $10^7$ → 23.5 x $10^6$<br>+1.34 x $10^6$ → +1.34 x $10^6$<br>24.84 x $10^6$<br>which must change to 2.48 x $10^7$ with 3 SF</td><td>Or it becomes this…<br>2.35 x $10^7$ → 2.35 x $10^7$<br>+1.34 x $10^6$ → +.134 x $10^7$<br>2.848 x $10^7$<br>Which must change to 2.84 x $10^7$ with 3 SF</td></tr>
<tr><th colspan="2">Subtraction with Scientific Notation</th></tr>
<tr><td colspan="2">With subtraction, the rules are the same for addition, but you SUBTRACT the coefficients.<br><br>Once both exponents match, Just add the coefficients and keep the exponents.<br>If your coefficient is less than 1 or more than 10, adjust your answer.</td></tr>
</table>

# Temperature Scales

We live in America; we use the Fahrenheit scale of temperature here, almost everywhere but science class. We'll almost never use it in chemistry.

Centigrade is the same as Celsius, but your teacher almost always says centigrade.

The third scale we'll learn is Kelvin, named after the famous chemist Lord Kelvin.

Whatever 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 20ºC, that's close to 293 in Kelvin. The numbers are really different, but the temperatures are the same. Kelvin NEVER uses degrees; they are just Kelvin.

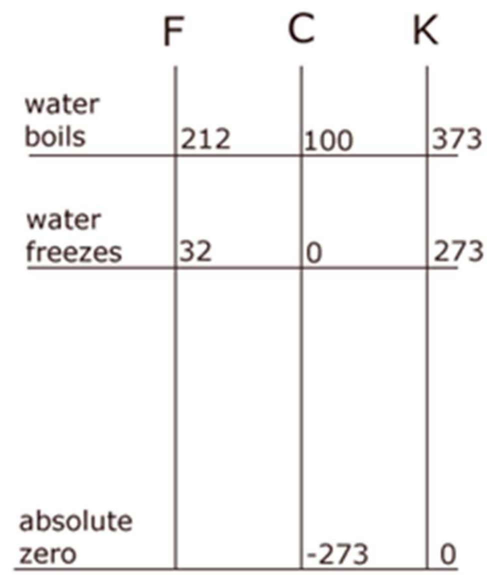
As shown here, the 3 scales are related as follows… Water freezes at <u>STANDARD TEMPERATURE</u>.
On table A of your reference tables this is pointed out.

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 do any math with Fahrenheit in high school chem.




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 like in °C or °F.

Absolute zero is a theoretical temperature. It's a temperature so low that all atomic motion stops.

Scientists have gotten close to it but they cannot reach absolute zero temperature.
Explaining that involves some fancy science talk that is not part of our course, especially now, but here goes.

If you could reach this temperature (say in a lab), just being close enough to observe it would impart some energy upon it, raising the temperature above absolute zero. Also, if all motion stops, even the movement of electrons spinning around the nucleus, theoretically time would stop there as well, creating a paradox that can't be explained even in a Star Wars movie. It remains just a theoretical idea, not a real temperature.

# Dimensional Analysis

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

I am five feet eight inches tall. I am also 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, but all of these measurements are equal (with the units).

To convert from one unit to another mathematically is called unit conversion math, or dimensional analysis. It's sort of fun but requires you write every single unit or else you will make silly mistakes in the math. Using your units correctly means you won't make any mistakes.

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

$$\frac{2}{2} = 1 \qquad \frac{12\text{ inches}}{1\text{ foot}} = 1$$

But one can be written in many different ways.

$$\frac{60\text{ seconds}}{1\text{ minute}} = 1 \qquad \frac{1000\text{ grams}}{1\text{ kilogram}} = 1$$

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, and 1000 mL = 1 liter, then this is possible too…

$$\frac{12\text{ bagels}}{1\text{ dozen bagels}} = 1 = \frac{1\text{ dozen bagels}}{12\text{ bagels}} \quad \text{or} \quad \frac{1000\text{ mL}}{1\text{ liter}} = 1 = \frac{1\text{ Liter}}{1000\text{ mL}}$$

If the equality of 1000 mL = 1 Liter is in your working math knowledge, these two problems would be easy for you to handle (this way) using dimensional analysis.

Convert 4.54 liters into milliliters.

$$\frac{4.54\text{ Liters}}{1} \text{ X } \frac{1000\text{ mL}}{1\text{ Liter}} = 45{,}400\text{ mL}$$

Note, the liters in the numerator of the first expression cancel the liter units in the denominator of the second The answer is given the "only units" left. You converted liters into milliliters using a conversion factor = 1.

Convert 865 mL into liters.

$$\frac{865\text{ mL}}{1} \text{ X } \frac{1\text{ Liter}}{1000\text{ mL}} = 0.865\text{ Liters}$$

Here, you used the same equality but used the "other" conversion factor since you wanted to change from milliliters to Liters. Conversion factors always equal 1, they just change units.

Convert 1.50 pounds to grams.
You must know that 1 pound = 454 g to do this.

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

An elephant weighs 6.53 tons. Convert that to grams, then change your answer into scientific notation (wow).

$$\frac{6.53 \text{ tons}}{1} \times \frac{2000 \text{ pounds}}{1 \text{ ton}} \times \frac{454 \text{ grams}}{1 \text{ pound}} = \begin{array}{c}5{,}929{,}240 \text{ grams} \\ 5{,}930{,}00 \text{ g 3 SF}\end{array}$$

Which changes to $5.93 \times 10^6$ grams with 3 SF
This is a multistep conversion, first the tons cancel out, then the pounds cancel.
Both conversion factors are equal to 1.

Convert 2.50 years into seconds. Answer must be in scientific notation.
This is a multi-step problem because there is not a conversion factor to go from years directly to seconds. Each conversion factor must equal 1, all units must cancel out.

| | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|
| $\frac{2.50 \text{ years}}{1}$ | X | $\frac{365 \text{ days}}{1 \text{ year}}$ | X | $\frac{24 \text{ hours}}{1 \text{ day}}$ | X | $\frac{60 \text{ minutes}}{1 \text{ hour}}$ | X | $\frac{60 \text{ seconds}}{1 \text{ min}}$ | = 78840000 seconds | |

78,840,000 seconds with 3 SF (to match the 2.50 years)

then convert to scientific notation → 78,800,000 seconds → $7.88 \times 10^7$ seconds (also 3 SF)

# Density

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

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$ Which is often abbreviated to... $$D = \frac{m}{v}$$ or $$\frac{D}{1} = \frac{m}{v}$$

Because the mass and volume have this relationship, it needs to be clear that for any pure substance (element or compound) the more mass of matter you have, there is a proportional volume increase.

This happens because density is a CONSTANT.

The density of pure water is 1.00 g/mL.

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

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

If you have 0.000356 grams water, its volume is 0.000356 mL, density is a constant.

## For any form of pure matter, density is constant.

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

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

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

| | | | | | | |
|---|---|---|---|---|---|---|
| An unknown substance has mass of 89.0 grams and volume of 10.05 cm$^3$.<br>What element could it be from Table S in your reference table? | | | | | | |
| $\frac{D}{1}$ | = | $\frac{m}{v}$ | → | $\frac{89.0\ g}{10.05\ cm^3}$ | = | 8.85572139 g/cm$^3$ = 8.86 g/cm$^3$ (3 SF) |
| The element on table S that has this exact density is #27 Cobalt | | | | | | |

| | | | | | | | |
|---|---|---|---|---|---|---|---|
| Your piece of copper has a volume of 25.00 cm$^3$. What is the mass of this copper? | | | | | | | |
| $\frac{D}{1}$ | = | $\frac{m}{v}$ | → | $\frac{8.96\ g/cm^3}{1}$ | = | $\frac{m}{25.00\ cm^3}$ | Here you cross multiply and solve for "m" |
| | | | | | | | |
| 8.96 g/cm$^3$ x 25.00 cm$^3$ = 224.0 grams (units cancel, 4 SF) | | | | | | | |

| | | | | | | | |
|---|---|---|---|---|---|---|---|
| Another piece of copper with an irregular shape has a mass of 923.40 grams.<br>What is the volume of this metal? | | | | | | | |
| | | | | | | | |
| $\frac{D}{1}$ | = | $\frac{m}{v}$ | → | $\frac{8.96\ g/cm^3}{1}$ | = | $\frac{923.40\ g}{v}$ | Here you cross multiply and solve for "v" |
| | | | | | | | |
| 8.96 g/cm$^3$ x v = 923.40 g → v = 103.058 cm$^3$ = 103.06 cm$^3$ 5 SF | | | | | | | |
| For equalities like density, they have unlimited SF, they do not affect the SF in the answers. | | | | | | | |

# Conversion Factors Allowed in Chem

| | | | | |
|---|---|---|---|---|
| volume | 1 liter = 1000 mL | 1 gal = 4 qts | 1 mL = 1 $cm^3$ | 2 pints = 1 quart |
| energy | 1000 calories = 1 Calorie (kilo-Cal) | 1000 Joules = 1 kilojoule | 4.18 Joules = 1 calorie | |
| time | 1 year = 365 days | 1 day = 24 hours | 1 hour = 60 min | 1 min = 60 sec |
| mass | 1000 g = 1 kilo | 1 pound = 16 oz | 1 pound = 454 g | 1 ton = 2000 pounds |
| length | 1 meter = 100 cm = 1000 mm<br>12 inches = 1 foot | 1 cm = 10 mm<br>3 feet = 1 yard | 1 inch = 2.54 cm<br>5280 feet = 1 mile | |
| pressure | 1 kilo-Pascal = 1000 Pascals<br>101.3 kPa = 760. mm of Hg = 1 atmosphere = 14.7 pounds per square inch | | | |

On Tables C and D on your reference tables are basic science units, and their prefixes.

Of all the prefixes, only kilo is larger than the base unit. Kilo = $10^3$ or 10x10x10 the unit

1 gram is small, 1 kilogram is 1000x BIGGER.
1 meter is one large step; 1 kilometer is a long walk 1000x longer than a meter.
1 joule is a tiny amount of energy, but a kilojoule is 1000x bigger

All other prefixes are smaller, and have negative exponents, which means they're decimals.

Centimeters are $10^{-2}$ of one meter or $1/100^{th}$ the size. There are 100 cm in one meter.

Milliliters are $10^{-3}$ of one liter or $1/1000^{th}$ the size. There are 1000 mL in one liter.

Micrometers are $10^{-6}$ of one meter or $1,000,000^{th}$ the size.
There are 1,000,000 micrometer = 1 meter

## Grasping this table's function is important for you.