Your name: ___________________________________ Period: _____

Measurement

The Class Topic Notes are all on Arbuiso.com, and you can follow the slide show in class. If you are absent you can follow along at home. You are responsible to get the notes even when you are absent.

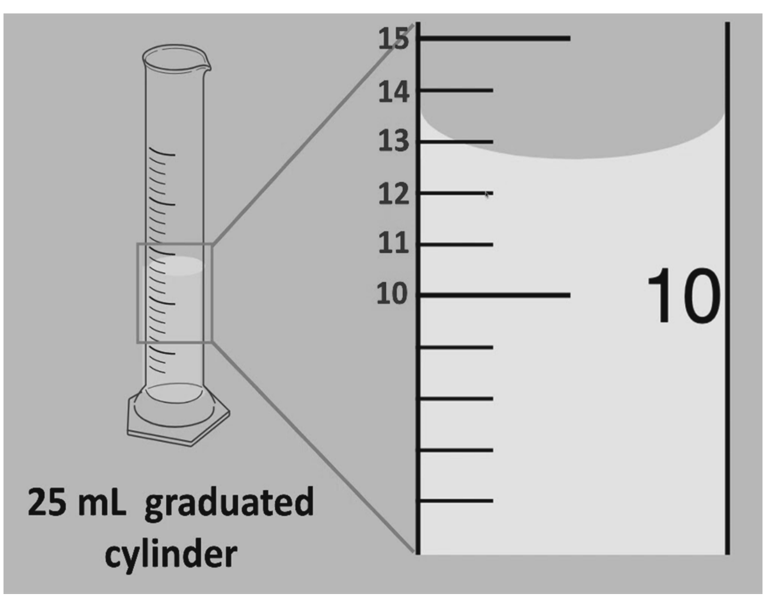
In addition to the notes you are to read the BASICS, which is basically everything you need to know about Measurement besides the notes. Both are required for your success.

This is 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.

Stay off of the internet. Sometimes online they use numbers we don't. We use our reference tables and our specific numbers. If you use the internet, I will know and your work will be incorrect. Our BASICS and our NOTES, and arbuiso.com only use NYS Regents values. Don't read the BASICS the night before your Celebration of Knowledge.

If you need help, call me, anytime… If I am available I'll answer. If not, I'll call you back.

Arbuiso cell: 607-727-3865 Text too, but sometimes chemistry is hard to text



Density of aluminum = 2.70 g/cm^3

AKA: 2.70 grams / centimeter cubed (cc)

Conversion Factors Allowed in Chem

(stay off of the internet)

Pressure

1 kilo-Pascal = 1000 Pascals

101.3 kPa = 760. mm of Hg = 1 atmosphere = 14.7 pounds per square inch

Volume

1 liter = 1000 mL

1 galloon = 4 quarts

1 mL = 1 cm3

2 pints = 1 quart

22.4 Liters = 1 mole of gas at STP

Energy

4.18 Joules = 1 calorie

1000 calories = 1 Calorie (kilo-Calorie)

1000 Joules = 1 kilojoule

Time

1 year = 365 days

1 day = 24 hours

1 hour = 60 minutes

1 minute = 60 seconds

Mass

1 pound = 16 ounces

1 pound = 454 grams

1 ton = 2000 pounds

1000 grams = 1 kilogram

Length

12 inches = 1 foot

3 feet = 1 yard

5280 feet = 1 mile

1 meter = 100 cm = 1000 mm

1 inch = 2.54 cm

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

1.

2. What is the chemistry formula for water? ___________________

3. The H stands for the element ______________________________ which happens to be element # _____

4. The O stands for the element ______________________________ 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. 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 skeleton formula for the synthesis reaction that combines $H_2 + O_2$ to form water is written this way:

   ____

9. The arrow means:

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 written this way:

12. So there must be a bit more to the reaction, since we can't lose even a single atom. Watch how it's balanced, then copy the balanced chemical reaction below

13. Skip this number, ok?

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 ___________________________ into the _____________________________

15. In this reaction there are 2 reactants, who's names are __________________ + ______________________

16. There is only one product, which has a science name of ____

    but we can call it ________________ too.

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, really loud SYNTHESIS. When energy is released in a chemical reaction, we call it EXOTHERMIC.

17. The balanced "thermochemical reaction" will look like this:

Let's watch, try not to blink, and promise me that you will tell your families about your first day of chemistry!

Tonight for Homework:

A. Read the whole1st 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 pencils, a calculator, and if you can a really big loose leaf binder to hold all of your old notes. We will retire everything once we are done with the topic. You will save them all, in order, to study for the 1st quarter cumulative celebration of knowledge, the Midterm, and the Regents Examination next June.
D. Get psyched, we are going to learn A LOT and we're going to have fun too. I can't wait.

<u>Measurement Class #1 Percent Error and Density</u>

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

18. ACCURATE

19. We will also try to be 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, which we usually get from science tables are called the

____ values.

23. Take out a ruler and measure this page top to bottom in centimeters. My measure 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? ______________

That boo boo is called your ERROR, which is vocabulary but we won't use it much ever again.

26. What we really want to use to measure how close we measured to accurate - THAT is called percent error.
Write the formula for percent error in the box at left. It's on the back of your reference table.
Below, we'll write it out in short hand. You must know both. Don't forget the % sign after the 100!

% Error =

%E =

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 instead. Write the formula again, shorthand style.

28. Let's use your eyes to measure the mass of the teacher in POUNDS. My measure his mass to be

_____________ pounds. According to the LOUSY school scale, the actual value is __________ pounds.

29. Calculate your percent error, 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…

34. You measure the density of Cu to be 8.75 g/cm^3. What element is this, what is the density of this element (and how did you figure that out?)

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? ____________

Does your percent error sign make sense? _________________

37. The formula for Density is on the back of the reference table. Copy it now, then in short hand.

Density = ______________________________ D = ___________

38. A bar of metal is 27.73 g and has volume of 4.70 cm^3. Is it gold? (start with a formula, look at table S)

39. What metal might it be instead? ____________________________

40. You measure a hunk of aluminum to be 363 grams and have volume of 148 mL. 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. (good)

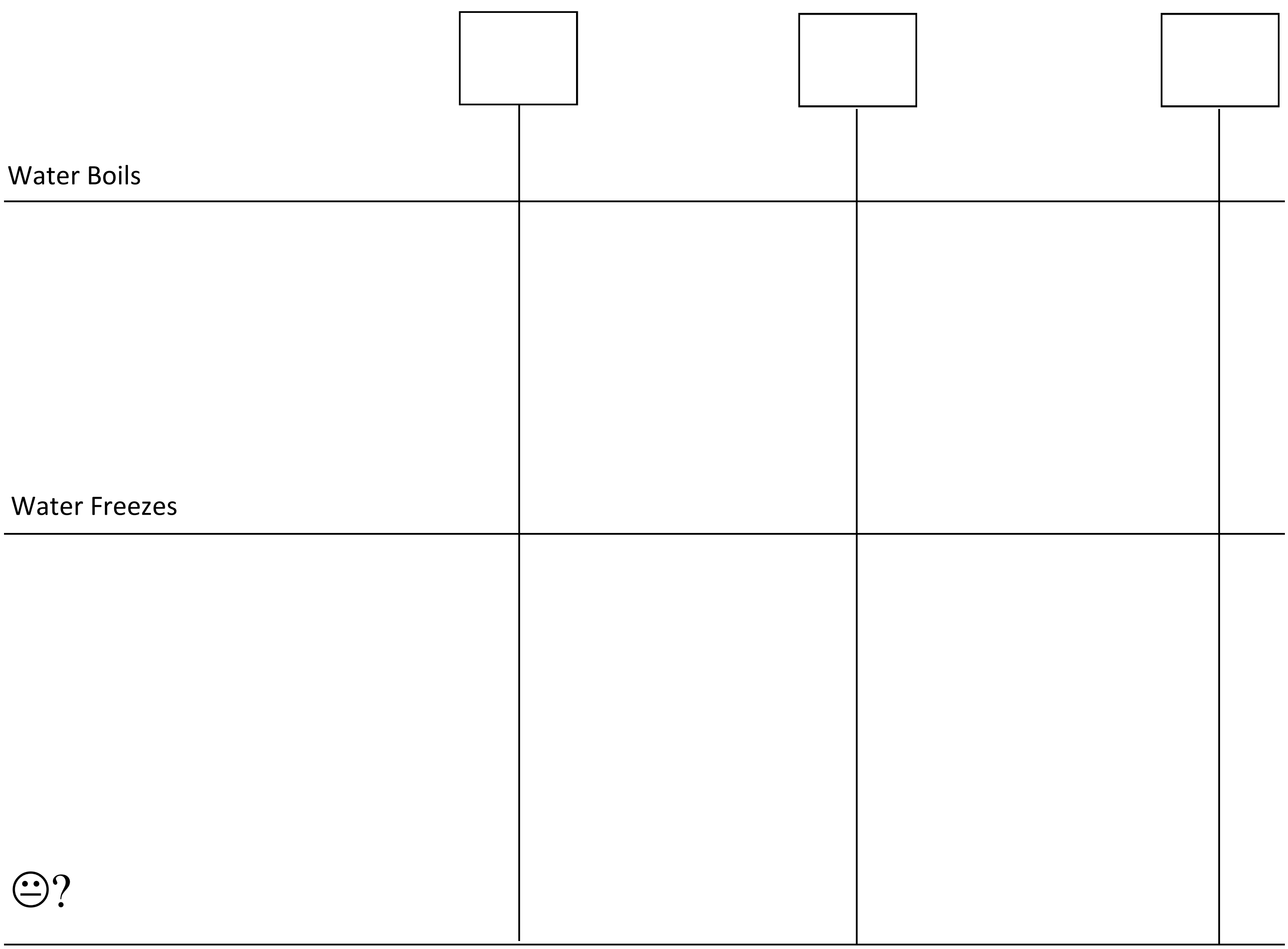
42. Another scientific temperature scale is called the ______________ scale. It's named after a guy named

Lord _____________________, no relation to Lord Vader from Star Wars.

43. There are 100 units of temperature from melting water to boiling water temperature. 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		
Centigrade or Celsius		
Kelvin		

To convert from Centigrade to Kelvin, or from Kelvin to 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 water is 100°C. Convert that to Kelvin, 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.
Is this COLD or HOT? Explain what you mean

48. You have 416 cm^3 of iron. What is it's 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 control how many places your real answers are allowed to have. There are several easy rules, which you will have to master.

For each of these measures, we will write how many SF are present, and what rule applies.

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

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. 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	2.245×10^4 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 with an equality that you will use in a math problem, or numerical facts from tables will not limit your answer, but your measurements will.

200. grams of Mg 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	6.02×10^{23} atoms
1.00 g/mL	1.000 g/mL	1.00000 g/mL	0.0000005 grams	3.550×10^{-17} grams
The answer to a density problem with mass of 125 grams and volume of 35 mL	Calculate density with mass of 1025 gm + volume of 350 mL	The temperature in K when you convert 24.5°C → Kelvin.	the answer of $2.5\ cm^3 \times 5.6788\ g/cm^3 =$	
If there are 454 grams = 1 pound, the number of grams in 3.750 pounds when you convert it.		The equality of 1000 grams = 1 kilogram		12.00 inches

63. You measure the density of nickel metal to have density of 9.1 g/cm^3. What is your % Error with the correct number of SF?

64. You measure your floor to be 14.5 feet X 15 feet and you want a rug. How big of a rug do you need in square feet? (SF count)

65. You measure the mass of metal to be 74.35 grams and it's volume to be exactly 12.0 mL. 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!)

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)

Sig Figs, or SF are the highest value topic of the year. SF are in all lab reports from now on, they're on every celebration and midterm, most HW and most classwork assignments. You will be measuring all sorts of stuff, and doing all types of calculations all year long. Figuring out the RIGHT answer requires you to follow the rules of SF. Learn SF or else.

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. But other conversions will need to be done as well. Some are easy, 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 into roses.

If 12 inches = 1 foot, how many inches are in 4.0 feet? ___________ another conversion, from feet to inches.

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.

Your starting point	X	Your conversion factor	=	Your answer with proper SF
☐	X	☐	=	☐
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 (1 L = 1000 mL). (watch SF)

71. Assume you are EXACTLY 16.33 years old right NOW. 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:

6.02×10^{23} atoms is one mole of atoms or the density of helium is 1.64×10^{-4} grams/mL

The front part of the number is called the coefficient, and the back part is called the power of ten.

In our class the RULE for scientific notation is that the coefficient MUST BE greater than 1 but less than 10.

If your math works out differently, you must adjust your answer to an equivalent answer in the proper form.

Converting big and small numbers into scientific notation first.

75. 17,000,000,000 ants ____________________ 76. 6,374,000 meters ____________________

77. 0.034 gram ________________________ 78. 0.000000000154 meters _________________

79. 0.0000083 meters ____________________ 80. 4,500,000,000,000,000,000,000,000 years ______________

Convert scientific notation into numbers

81. 6.02×10^{23} molecules _____________________________ 82. 3.5×10^{4} grams _____________________

83. 1.25×10^{-7} meters _______________________________ 84. 2.290×10^{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

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

Rules to use scientific notation in math problems…

Multiplication Rule for Scientific Notation: ____

87. $(3 \times 10^5)(2 \times 10^2) =$ ____________________________________

88.
$$\begin{array}{r} 5.0 \times 10^4 \\ \underline{\times\ 3.0 \times 10^2} \end{array}$$

Division Rule for Scientific Notation: ____

89. $\dfrac{3.0 \times 10^4}{2.0 \times 10^2}$

90. $\dfrac{9.0 \times 10^5}{3.0 \times 10^3}$

91. Addition Rules for scientific notation: ____.

92.
$$\begin{array}{r} 6.5 \times 10^7 \\ \underline{+\ 2.2 \times 10^7} \end{array}$$

93.
$$\begin{array}{r} 6.2 \times 10^8 \\ \underline{+\ 1.5 \times 10^6} \end{array}$$

94. Subtraction Rules for scientific notation: ____

95. 8.5×10^{3}
$\underline{- 2.4 \times 10^{3}}$

96. 7.1×10^{5}
$\underline{- \ 1.6 \times 10^{4}}$

97. 8.72×10^{11}
$\underline{+ 1.72 \times 10^{10}}$

98. 4.65×10^{14}
$\underline{- 2.25 \times 10^{15}}$

99. 6.02×10^{23}
$\underline{\times 1.50 \times 10^{2}}$

100. $(9.05 \times 10^{19}) \div (3.2 \times 10^{16}) =$

PRACTICE MATH for measurement. Do all of these in PENCIL please.

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.

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 even believe this is real. 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 grams_____ 5.600 kilograms_____ 4.305 mL_____

0.678°C_____ 0.00065 moles Hg_____ 1.400 seconds_____

105. Calculate the quotient: 4.569 g $\div$ 2.0 cm^3 = ______________________

106. In each set of temperatures, decide which is the coldest, which is the hottest.

SET A: 10 K or 10 C or 10 F SET B: 280 K or 32°F or 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? (hint, write the formula first)

110. Do what the math says to do: $(3.5 \times 10^6) X (2.0 \times 10^2) =$

$(8.0 \times 10^8) \div (4.0 \times 10^{12}) =$

$(3.3 \times 10^8) + (1.2 \times 10^7) =$

$(5.64 \times 10^5) - (2.33 \times 10^4) =$

111. Look up the boiling point of aluminum, covert that into scientific notation. Also, convert the Kelvin into centigrade. (watch out for SF!)

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 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 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 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 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.

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: 3.67mm 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 <u>least</u> accurate measurement.

<u>Percent Error</u>

% 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.

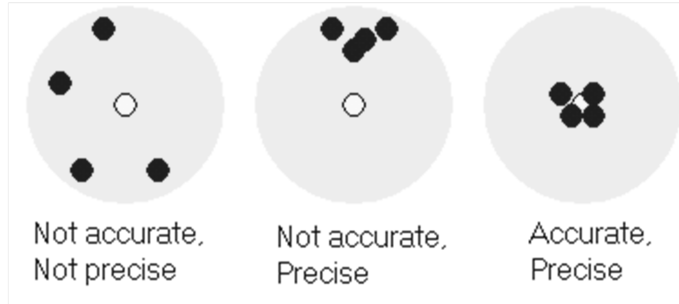
Percent Error with NO sign means you did the math wrong and I will always deduct a point from your score for NOT PAYING ATTENTION TO DETAILS. Please take heed.

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. 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).

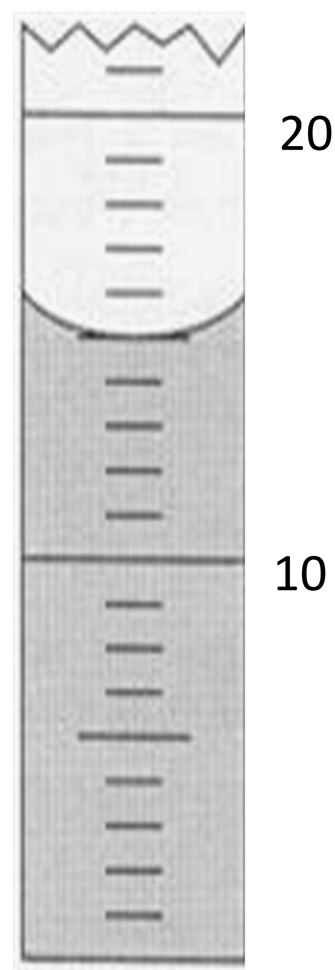


At the far right the darts are all very close to the center, which makes them (precise + accurate).

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.

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 = 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 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 $\qquad$ 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 1×10^3.

The number 66,500 would be 6.65×10^4.

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 1×10^6 atoms.

If you have 1,300,000 atoms, it would be written as 1.3×10^6 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, $2.0 \times 3.0 = 6.0$ (with two sig figs only in answer)
Then add the powers of ten $(5 + 4 = 9)$ The answer would be 6.0×10^9.

Dividing Scientific Notation Math

To divide in scientific notation, say

$$\frac{9.00 \times 10^8}{3.0 \times 10^5} = 3.0 \times 10^3$$

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

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 exponents to 10^6	Or Change the exponents to 10^7
23.5×10^6 $\underline{+1.34 \times 10^6}$ $24.84 \times 10^6 = 2.48 \times 10^7$ with 3 SF	2.35×10^7 $\underline{+0.134 \times 10^7}$ 2.484×10^7 this changes to 2.48×10^7 with 3 SF

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 1.00×10^9 is the same as 10×10^8, or 100×10^7, 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.

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. 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 20ºC, 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 here, the 3 scales are related as follows… 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 into Fahrenheit or from Fahrenheit in our class.

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

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.

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.

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.

$12 \times 1 = 12$ 10000 x 1 is still 10000 $234 \times 1 = 234$

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? (do the math)

$$\frac{5700 \text{ in}\cancel{\text{ches}}}{1} \quad \frac{1 \text{ foot}}{12 \text{ in}\cancel{\text{ches}}} = 475 \text{ feet}$$

Convert 1.50 pounds into 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 = X of grams, but I don't know it. I do know others, but it will take multiple steps to do the math

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

Then, into scientific notation… 5902000 grams = 5.9×10^6 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 2 significant figures here.

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

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

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 SF = 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 7.88×10^7 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.

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

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

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

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 <u>wacky</u> 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…

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

Which is often abbreviated to...

$$D = \frac{m}{v} \quad \text{or} \quad \frac{D}{1} = \frac{m}{v}$$

Because the mass and volume are in a particular relationship, it needs to be clear that for any pure substance (element or compound) 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.

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

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

It's cobalt!

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

$$D = \frac{\text{mass}}{\text{volume}} \qquad \frac{8.96\ g/cm^3}{1} = \frac{\text{mass}}{25.00\ cm^3}$$

Solve for "m" by: 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?

$$D = \frac{\text{mass}}{\text{volume}} \qquad \frac{8.96\ g/cm^3}{1} = \frac{923.4}{V}$$

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.

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 cm^3 = 12.4 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.

The Significance of Significant Figures

Significant figures are how we keep track of our numbers in science classes. We always try to measure as perfectly as we can, but we are limited by our own abilities and by our instruments. We strive to make the best measures, and use these measurements in math formulas. The significant figures limit our mathematical answers, we're not allowed to get "more accurate" with a calculator, to somehow measure better by math, nor should we ever give away our exact measures by rounding casually.

There are rules to follow, they are sometimes easy to forget, but you must memorize them. In truth, you will likely lose more points to significant figures than any other one thing in our class. Unless you learn them.

THE RULES for SIGNIFICANT FIGURES

1. Any digit 1 to 9 will always be a significant figure.
2. All zeros between significant figures will be significant.
3. Zeros before significant figures are not significant.
4. Zeros before a "missing" decimal point are not significant, but if the decimal point is included, then the zeros before it are significant.
5. Zeros at the far right of a decimal are significant.
6. Unlimited significant figures happen with Equalities, such as 12 inches = 1 foot, both have unlimited SF because these are not measures, we understand they are equal to the "nth" degree.
   12.00... inches = 1.000… foot with as many zeros (as many SF) as you like.
7. With math, answers must have the same number of significant figures as the LEAST number of SF in the measurements you are calculating with.
8. With Scientific Notation, SF are counted only in the co-efficient portion of the scientific notation.

23 grams has 2 SF 23.5 grams has 3 SF 23.54 grams has 4 SF 23.543 grams has 5 SF	100 grams has 1 SF 1,000 grams has 1 SF 1,001 cm has 4 SF 1,001,001inches has 7 SF	100 has 1 SF 100. has 3 SF 90 has 1 SF 90. has 2 SF
0.005 kJ has 1 SF 0.4 grams has 1 SF 1.005 has 4 SF 0.000000000001 has just 1 SF	12.00 inches has 4 SF 9.0000000 grams has 8 SF 1.00 g/cm^3 has 3 SF	Equalities have unlimited SF, meaning when you use them in a conversion, they do not limit your answer. 12 inches = 1 foot 1000 grams = 1 kilogram Any atomic mass or density
454 grams = 1 pound 22.4 liters of gas = 1 mole of gas with equalities, both have unlimited SF	6.02×10^{23} atoms has 3 SF 3.550×10^{-4} moles has 4 SF 7.75043×10^{12} m has 6 SF	2.5 cm x 5.6788 cm = ? the answer has just 2 SF 4,550 grams divided by 1,255 cm^3 = ? the answer has to have 3 SF