Hey Future Advanced Chemistry Students!

My name is Stephanie Meadows, and I am looking forward to being your Chemistry Teacher for the 2024-25 school year at Hargrave High School!

I hope that you fill find that Chemistry is SUPER FUN!!! And, it is, at times, SUPER CHALLENGING!!! You will have homework. You will have to study. You will do a lot of math. You will be forced to adhere to our late work policy and you will be expected to work hard all year long. In return, you will learn more this year than probably any other class you've had up to this point! You will get to do some cool labs along the way. Are you ready???

If you feel like you are correctly placed academically, and you are up to the challenges ahead, then here is where our adventure begins! Advanced Chemistry students are expected to address certain selected topics independently during the summer in order to prepare for the upcoming school year.

Your assignment consists of the following:

1. Work through the attached notes and worksheets (1) and (2) and be prepared to turn these in on **Monday, August 12th**.

*Feel free to work the problems on your own paper for more room and box in your final answers. Try your best to apply the significant figure rules when writing your final answer. However, the answers will not be graded on significant figures on the summer assignment.

*No credit will be given without work shown.

*There will be a QUIZ on Tuesday, August 13th, so that you can show me that you have mastered this unit of study.

*I will be grading select problems from the assignment for a daily grade on the first nine-weeks grading period. You will have tutorial time before and after school to ask questions and get clarification. Form study groups and work on the problems together!

If you need help over the summer, and you have watched the video do not hesitate to contact me for help! If you choose to not work through the summer assignment, you will find the pace of the course somewhat overwhelming when school starts, as I will move through the material rather quickly. Thank you and good luck! Can't wait to see you in August!

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Summer Assignment 2024-2025 Advanced Chemistry

Matter, Measurements, and Calculations

I. **Chemistry**: The branch of science that deals with the identification of the substances of which matter is composed; the investigation of their properties and the ways in which they interact, combine, and change; and the use of these processes to form new substances. (Matter = anything that has mass and takes up space)

II. **SI (System of International) Units of Measurements**: adopted in 1960 by the General Conference on Weights and Measures.

a. Metric System – must know this

i. Mass is measured in kilograms (other mass units: grams, milligrams)
ii. Volume in liters
iii. Length in meters

b. **Prefixes** are added to the stem or base unit to represent quantities that are larger or smaller than the stem or base unit. You must know the following:

Prefix	Value/Meaning	Abbreviation	Example
Milli	10^{-3} / 0.001	m	mm, mg
Centi	10^{-2} / 0.01	c	cl, cg
Deci	10^{-1} / 0.1	d	dl, dg
Deka	10^{1} / 10	da	dag, dal
Hecto	10^{2} / 100	h	hl, hm
Kilo	10^{3} / 1000	k	kl, kg

Examples:
1Mm=1,000,000m
1km=1000m
1hm=100m
1dam=10m
1m=1m
1dm=0.1m
1cm=0.01m
1mm=0.001m
1μm=0.000001m

"When solving problems, I will always put a 1 with the prefix."

c. **Derived Units**: combinations of quantities:
area (m^2), Density (g/cm^3), Volume (cm^3 or mL) 1 cm^3 = 1 mL

d. **Temperature**: Be able to convert between degrees Celsius and Kelvin. Absolute zero is 0 K, a temperature where all molecular motion ceases to exist. Has not yet been attained, but scientists are within thousandths of a degree of 0 K. No degree sign is used for Kelvin temperatures.

i. Celsius to Kelvin: K = C + 273

Convert 98° C to Kelvin: 98°C + 273 = 371 K

Ex: New materials can act as superconductors at temperatures above 250 K.
Convert 250 K to degrees Celsius: 250 K = _____°C

III. **Density**: relationship of mass to volume D = m/V
Density is a derived unit (from both mass and volume)
For solids: D = grams/cm^3
Liquids: D = grams/mL **Know these units.**
Gases: D = grams/ liter

Density is a conversion factor. Water has a density of 1 g/mL which means 1g = 1mL!

IV. **Dimensional Analysis**: When you finish this section, you will be able to: convert between English and metric units; convert values from one prefix to another. We will use dimensional analysis all year…it gets easier with practice, but it can be very confusing at first.

Dimensional analysis is the **single most valuable mathematical technique** that you will use in general chemistry. The method involves using conversion factors to cancel units until you have the proper unit in the proper place. A conversion factor is a ratio of equivalent measurements, so a conversion factor is equal to one. Example conversion factors:

4 quarters = $1.00 1 kg = 1000 g 1 kg = 2.2 lbs

What is the mass in kilograms of a 125 pound box?

$$?\text{kg} \rightarrow \frac{125\text{lbs}}{1} \times \frac{1\text{ kg}}{2.2\text{ lbs}} = 56.8\text{ kg}$$

Notice that the unit "lbs" cancel out and your answer is in "kg."

When you are setting up problems using dimensional analysis, you are more concerned with units than with numbers.

Example: How many atoms of copper are present in a pure copper penny? The mass of the penny is 3.2 grams.

Needed conversion factors:
6.02×10^{23} atoms = 1 mole copper
1 mole copper = 63.5 grams

PROBLEM SOLVING STEPS

1. List the relevant conversion factors.
2. Rewrite the problem as follows

$$?\text{ atoms} \rightarrow \frac{3.2\text{ g}}{1} \times \frac{1\text{ mole}}{63.5\text{ g}} \times \frac{6.02\times10^{23}\text{ atoms}}{1\text{ mole}} =$$

3. Multiply all the values in the numerator and divide by all those in the denominator.
4. Double check that your units cancel properly. If they do, your numerical answer is probably correct. If they don't, your answer is certainly wrong.

Density as a Conversion Factor

- Density is a conversion factor that relates mass and volume.

Example Problem:

The density of mercury is 13.6 g/mL. What would be the mass of 0.75 mL of mercury?

$$?\text{g} \rightarrow \frac{0.75\text{mL}}{1} \times \frac{13.6\text{ g}}{1\text{ mL}} =$$

Solve using dimensional analysis when applicable.

1. A gas has a density of 0.824 g/L and occupies a volume of 3.00 liters. What is the mass in grams?

2. An unknown metal having a mass of 287.8 g was added to a graduated cylinder that contained 31.47 mL of water. After the addition of the metal, the water level rose to 58.85 mL. Determine the volume of the metal. Calculate the density of the metal.

3. A solid with dimensions of 3.0 cm X 4.0 cm X 2.0 cm has a mass of 28 g. Will this solid float in water? (water has a density of 1.00 g/mL)

Complete the following using dimensional analysis:

1. Convert the following metric units:

   a. 62.9 kg to g

   b. 49.8 ml to liters

   c. 33.9 cm to m

2. Convert the following units:

   a. 7.51 miles to meters

   b. 38 feet to cm

3. How many atoms of carbon are present in a 56 gram sample of charcoal (carbon)?
   (1 mole = 12.01 grams, 1 mole = $6.02X10^{23}$atoms)

A. **Counting Significant Figures**

When you report a measured value it is assumed that all the numbers are certain except for the last one, where there is an uncertainty of ±1.

Example of nail that is 6.36 cm long. The 6.3 are certain values and the final 6 is uncertain! There are 3 significant figures in the value 6.36 cm (2 certain and 1 uncertain). All measured values will have one (and one only) uncertain number (the last one) and all others will be certain. The reader can see that the 6.3 are certain values because they appear on the ruler, but the reader has to estimate the final 6.

The rules for counting the number of significant figures in a value are:

1. **If the number has a decimal, start from the LEFT (Front) side of the number, go to the first number that is NOT zero and count the rest of the numbers**
   → 0.0<u>30400</u> 5 sig figs

2. **If the number does NOT have a decimal, start from the RIGHT (End) side of the number, go to the first number is NOT zero and count the rest of the numbers**
   <u>309</u>0000 ← 3 sig figs

Give the number of significant figures in the following values:

a. 38.4703 mL
b. 0.00052 g
c. 0.05700 s
d. 500 g

If your value is expressed in proper scientific notation, all of the figures in the pre-exponential value are significant, with the last digit being the least significant figure.

7.143×10^{-3} grams contains 4 significant figures.
If that value is expressed as 0.007143, it still has 4 significant figures. Zeros, in this case, are placeholders. If you are ever in doubt about the number of significant figures in a value, write it in scientific notation.

Give the number of significant figures in the following values:
a. 6.19×10^{1} years
b. 7.4×10^{6} years
c. 3.802×10^{-19} J

<u>Helpful Hint:</u> Convert to scientific notation if you are not certain as to the proper number of significant figures.

When solving multiple-step problems, DO NOT ROUND OFF THE ANSWER UNTIL THE VERY END OF THE PROBLEM.

D. **Significant Figures in Calculations**

(Pay attention because the rules change depending on WHAT TYPE of calculation you are doing! Are you watching the video? If not, you should be…)

1. In <u>addition</u> and <u>subtraction</u>, your answer should have the same number of decimal places as the measurement with the <u>least number of decimal places</u>.

   Example: 12.734 mL - 3.0 mL = _____

   Solution: 12.734 mL has 3 figures past the decimal point. 3.0 mL has only 1 figure past the decimal point. Therefore, your final answer should be rounded off to one figure past the decimal point.

   12.734 mL - 3.0 mL = 9.734 □ 9.7 mL

   a. 32.3 mL – 25.993 mL = _____

   b. 84 g + 34.99 g = _____

   c. 43.222 mL – 38.12834 mL = _____

2. In <u>multiplication</u> and <u>division</u>, your answer should have the same number of significant figures as the least precise measurement (or the measurement with the fewest number of SF). (SF = Sig Figs, a term we will use all year long!)

   61 cm x 0.00745 cm = 0.45445 cm^2 = 0.45cm^2 (2 SF)

   a. 32 m x 0.00003987 m = _____

   b. 5 cm x 1.882 cm = _____

   c. 47. 8823 g ÷ 9.322 mL = _____

In multiple step problems, if addition or subtraction AND multiplication or division is used the rules for rounding are based off of multiplication and division (it "trumps" the addition and subtraction rules).

3. There is no uncertainty in a conversion factor; therefore, they do not affect the degree of certainty of your answer. The answer should have the same number of SF as the initial value.

   a. Convert 25.0 meters to millimeters.

   b. Convert 0.12L to mL.

Measurements and Calculations Worksheet #1

Mixed Conversions: Show ALL work using dimensional analysis.

1. 4.305 liters to milliliters

2. 3.80 km to meters

3. 3.88 miles to meters

4. 2.994 ounces to milligrams

5. 926 tons to Kilograms

6. How many centimeters are there in 7.88×10^2 feet?

7. Convert 6.775 yards to millimeters.

8. Convert 5.47×10^{-2} hectograms to ounces

9. How many centimeters are there in 51.004 miles?

1 mile = 5280 feet
1 inch = 2.54 cm
1 gram = 0.0353 ounces
2000 pounds = 1 ton
1 pound = 454 grams
1 yard = 3 feet
1 km = 0.621 miles
1 yard = 0.914 m
1 ton = 907.2 kg
1 ft = 12 in
1 gallon = 3.785 L
1 m = 39.37 inch
1 m^3 = 35.31 ft^3
1 grain = 0.00229 ounces
1 grain = 0.0648 grams
*****Look to notes for all metric conversions.**

Density: Show ALL work.

10. If an unknown solid weighs 84.0 grams and occupies 30.0 cm^3 of space, what is its density?

11. What is the mass, in grams, of a liquid having a density of 1.50 g/ml and a volume of 3.5 liters?

12. What volume, in cm^3, would a 200 gram sample of gold have if its density is known to be 20.5 g/cm^3?

13. A solid block of substance is 74.0 cm by 55.0 cm by 29.0 cm and it weighs 625 kg. Assuming that it did not chemically react with water nor dissolve in it, would it float in water? The density of water is 1 g/mL. Show your work.

14. A gas has a volume of 7.0 liters and a mass of 4.44×10^5 milligrams. What is its density in g/L?

15. A certain liquid has a density of 0.855 g/mL. How many LITERS would have a mass of 1.00 kg?

Measurements and Calculations Worksheet #2

Dimensional Analysis (Exercising Problem Solving Skills)

Show ALL work, using dimensional analysis, Report answers using the correct number of significant figures.

1. The record long jump is 349.5 inches. Convert this to meters.

2. A car traveling 55.0 miles per hour. Convert this to meters per second.

3. How many milligrams are there in a 5.00 grain aspirin tablet?

4. What is the volume, in mL, of 100. grams of mercury? Mercury has a density of 13.54 g/mL (which means 1 mL = 13.54 g).

5. In 1980, the US produced 18.4 billion pounds (1.84×10^{10} lbs) of phosphoric acid to be used in the manufacture of fertilizer. The average cost of the acid is $318/ton. What was the total value of the phosphoric acid produced?

6. On planet Zizzag, city Astric is 35.0 digs from city Betrek. The latest in teenage transportation is a Zeka which can travel a maximum of 115 millidigs/zip. On Zizzag their time system divides each dyne into 25.0 zips. How many dyne will it take Pezzi to get from Astric to Betrek to see his girlfriend?

7. While prospecting in the North Woods, Joe found a gold nugget which had a density of 19.2 g/cm3. Joe dropped the gold into water in a graduated cylinder, the water level increased by 15.0 mL. How many grams of gold did he have?

8. After removing 68.1 kilograms of old copper tubing from air conditioning units Mark takes his load to a recycling yard. There he is paid $2.50 per pound.

    a. How much money did he make?

    b. If 1 mole copper = 63.45 grams and $6.02X10^{23}$ atoms = 1 mole, how many atoms of copper did Mark recycle?

YAY! You've reached the end…of the very beginning! I am proud of you for working hard to get to this point! I know this wasn't an easy assignment, but this stuff will eventually be the "easy" stuff. Practice helps! Enjoy your summer! ☺