Assignment #3 – In Class & After Class

Wooster Honor Chemistry

Name:______________________________

Date:_________________ Period:________

Scientific Measuring Lab

1. Get a graduated cylinder from your instructor. If you are carrying it through the room, walk carefully back to your desk/table. Write how you would measure the volume in this liquid in the space below. You may **carefully** pass the graduated cylinder around your group so everyone has a chance to examine the liquid up close. **Do not ask anyone else in your group how they would/have record/ed this measurement.**

2. Once each member of your group has made his or her own judgement, then compare. Make note of all your measurements below, writing them exactly as they were shown originally.

3. Discuss together how you measured. What would you be willing to claim about the true value of this liquid volume? Create a statement such as, "We are 95% certain that the true value lies between 5.5 mL and 6.0 mL." Write the statement below.

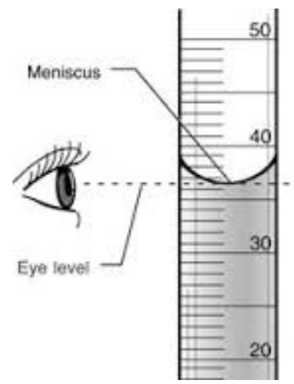
4. One way to investigate the true value is by finding the mass and dividing by the density of water, which is conveniently 1.0 g/mL. "Tare" or zero the mass of a small cup and pour all the liquid from your graduated cylinder into the cup—**pouring must be done over a sink as it could damage the digital scale.** Record the mass below. Then write the same number but switch the units to mL; this is the true value of your liquid volume.

5. If your true value falls within the range described in #3 above, you may take a minute to give a super fancy fist bump to your group members.

6. Pour the liquid back from the small cup into your graduated cylinder. You may use a funnel. Dry the small cup and return it to its stack. Take your graduated cylinder back to the table; the value is probably a little smaller as some liquid was lost during the transfer process.

7. What were some things that might have gone wrong with your group's original guesses in question 1? (Another way of thinking about this might be, what are some common measuring mistakes that could cause you to get quite a different value from the true value?)

8. To read liquid volumes accurately, follow the convention of placing your eye exactly at the level of the liquid, and read to the middle of the "meniscus" or curve of the liquid; in water (and most liquids), this is the lowest spot. Try reading the volume when your eye is above and below the meniscus; how does that affect your readings?



Eye above the meniscus:

Eye below the meniscus:

9. Return the graduated cylinder to your instructor.

Scientific measurement is all about ethically, precisely, and accurately communicating information.

To avoid miscommunication/misrepresenting the data, scientists follow certain conventions or rules.

- **Every scientific measurement must include units.** Deduce the correct units from hints—inspect the glassware you are using, etc. (Ask if you need help.)

- **Every scientific measurement must also include an estimated uncertainty.** As scientists-in-training, you are expected to convey information about how certain you are of your measurements alongside your reported value for each measurement. The idea behind estimating uncertainty is that we recognize all measurements are imprecise; that is, it is impossible for us to measure anything absolutely perfectly. Maybe you think you weigh 140 pounds, but your true weight is 139.873901524…pounds. An infinite number of digits would be required for a perfect representation of the actual, true value of your weight, but what the scale gives you is an acceptable level of information, even if it is off by 0.126098475…pounds. The scientific community has established some conventions for how large your estimated uncertainty should be, as follows:
  - For all digital readings, the uncertainty is a one in the last place value of the reported value.
  - Thus, a scale that reads 23.452**9** g would therefore have an uncertainty of ± 0.000**1** g, and a scale that reported 2**3** g would have an uncertainty of ± **1** g.
  - One exception is a digital timer, which often reports to the hundredths of a second; you can't really read as fast as the numbers are moving; therefore, all time measurements should use an uncertainty of ± 1 s due to human reaction time as the limiting factor.
- Glassware/rulers are on a case-by-case basis:
  - Metric rulers with divisions every 1 mm – uncertainty is ± 0.5 mm
  - 10.00 mL Pipet – uncertainty is ± 0.02 mL
  - 25.00 mL Pipet – uncertainty is ± 0.03 mL
  - Buret - uncertainty is ± 0.02 mL
  - 10 mL graduated cylinder with divisions shown every 0.1 or 0.2 mL – uncertainty is ± 0.05 mL
  - 100 mL graduated cylinder with divisions shown every 1 mL – uncertainty is ± 0.5 mL
  - Thermometer with divisions shown every 1 °C – uncertainty is ± 0.5 °C
  - Beakers and Erlenmeyer flasks – uncertainty is a personal best-guess for how big a range do you think you need to be 95% sure your estimated value ± your uncertainty includes the true value.
- Notice that uncertainty includes a unit.

- The final unit of the measurement and the final unit of the uncertainty must have the same place value.
    - Good examples / Bad examples

Good examples		Bad examples
54.32 g $\pm$ 0.01 g	/	54.32 g $\pm$ 0.0001 g
34.55 mL $\pm$ 0.05 mL		35 mL $\pm$ 0.05 mL
43 s $\pm$ 1 s		43.24 s $\pm$ 1 s

10. Practice correct lab safety throughout this lab; anyone engaging in horseplay, needing more than one reminder to wear safety goggles, or otherwise off task will lose the opportunity to complete the lab.

11. On a later day, we will share our data across the class, and you will need it for some of the post lab questions. Please work efficiently but take care to have quality measurements.

Data Table: (8 points)

Station	*Description of Item Being Measured*	*Measured Value ± Uncertainty w/Units*
Ex. a	Mass of small, dry beaker, empty	*29.48 g ± 0.01 g*
Ex. b	Mass of small, dry beaker plus one 10.00 mL pipet-full of antifreeze	*40.70 g ± 0.01 g*
1	Volume of liquid in beaker	
2	Volume of liquid in 10 mL graduated cylinder	
3	Temperature of water in Erlenmeyer flask	
4a	Length of aluminum foil rectangle	
4b	Width of aluminum foil rectangle	
4c	Mass of aluminum foil rectangle	
5	Mass of a nickel	
6a	Mass of ethanol	
6b	Time (per stopwatch) at which ethanol mass measurement is taken	
7	Voltage of a battery	
8	Volume of liquid in 100 mL graduated cylinder	

Homework Part A = Post-Lab Questions (2 points; ALL students)

12. Calculate the mass of antifreeze present for the class example, Ex. b – Ex. a. Since this value involved subtracting two measurements that each had an uncertainty of ± 0.01 g, the uncertainty of the difference increases to ± 0.02 g. (Perhaps the first measurement was 0.01 g too high, and the second measurement was 0.01 g too low, or vice versa.) (Write your calculated value plus or minus this increased uncertainty below; 2 points.)

Read this! For Ex. a, we would say that 0.01 g is the absolute uncertainty, which we could show as: **Δm** = 0.01 g (where Δ stands for “uncertainty in” and **m** stands for mass).

**We may also show uncertainty as a percentage, where the percentage uncertainty is:

Percentage uncertainty for any measurement of x = $\frac{\Delta x}{x} \times 100\%$

Thus, the percentage uncertainty of measurement Ex. a is $\frac{0.01\text{ g}}{29.48} \times 100\% = 0.03\ \%$
Note that percentage uncertainty values are only reported to the first non-zero digit.

Homework Part B = Differentiated Post-Lab Questions (10 points; choose either 13 & 14 <u>OR</u> 15 & 16.)

13-14. Regular Honors Chemistry Desired (10 points):

13. Calculate the percentage uncertainty for every measurement in stations 1 – 8, writing the values in the space to the right of your Data Table. In the space below, label "smallest percentage uncertainty = ____" and "largest percentage uncertainty = _____", and show your work for each calculation.

14. What vocabulary was new to you during this lab? Chose 2 items that weren't very familiar, and write the words and a brief description OR picture below. (Use additional paper if desired; 2 points.)

15-16. For Students Desiring the Accelerated Path (10 points + 2 bonus points possible):

15. Calculate the percentage uncertainty for your measurement in station 1, your measurement in station 2, and your measurement in station 8. Label and show your work for each calculation. (4 points)

16. Density is defined as mass ÷ volume and commonly has units of g/mL (for liquids), g/cm^3 (for solids), or g/L (for gases). ***Free example = Calculate the density of antifreeze.***

Mass of antifreeze = 11.22 g ± 0.02 g (or 0.2%)
Volume of antifreeze-delivered by a 10.00 mL pipet = 10.00 mL ± 0.02 mL (or 0.2%)

Calculate density normally, temporarily ignoring uncertainty values
$D = \frac{m}{V}$: Density = $\frac{11.22\ g}{10.00\ mL} = 1.122$ g/mL

Since the density calculation involves dividing two values, we add the two percentage uncertainties together to determine the percentage uncertainty of the density.
Percentage uncertainty of density = 0.2% + 0.2% = 0.4%

Next, we multiply the percentage uncertainty of density to the calculated value for density to determine the absolute uncertainty of density, ΔD:
ΔD = (Percentage uncertainty of density) × (**calculated density)**
ΔD = 0.4% × 1.122 g/mL = 0.004 × 1.122 g/mL = 0.004488 g/mL
→ round to first non-zero digit
→ ΔD = 0.004 g/mL

Finally, as noted in (5) at the top of page 3, "The final unit of the measurement and the final unit of the uncertainty must have the same place value." *Therefore, we check that this is the case…and it is good as is. (If there is a mismatch, it usually means we need to round the calculated value to match the place value of the uncertainty.)*

We report the density of antifreeze as 1.122 g/mL ± 0.004 g/mL.

Your turn: Calculate the thickness of your aluminum foil, given that the density of aluminum is $2.70\ g/cm^3 \pm 0.01\ g/cm^3$. (Volume = length × width × thickness.)
Show complete calculations, including those needed to determine the uncertainty of your calculated thickness. (Use additional paper if desired.; 6 points w/ 2 bonus points possible.)