LHA Laboratory Guide

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LaurenHill Academy Chemistry Laboratory Guide

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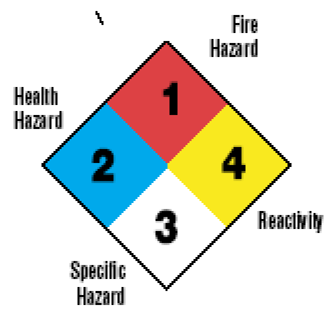
1. <u>Laboratory Safety</u>

A. Basic Rules

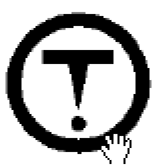
1. If long, loose hair styles cannot be avoided, tie up your hair, especially when using a Bunsen burner or strong solutions.
2. No food or beverages (including chewing gum) are allowed in the lab.
3. If you spill any chemical on your hands, wash it immediately.
4. Wash your hands before leaving the lab in case there is residue on them.
5. Students must keep all cell phones, pagers, CD players, etc. turned off and stowed in their bags during laboratory sessions.
6. Lab coats and safety glasses will be provided by us and must be worn by all students and teachers performing experiments, unless deemed unnecessary by the lab technician.
7. All equipment and reagents must be used only as directed.
8. Report spills and breakage immediately to the lab technician
9. Running and abrupt movements are prohibited in the laboratory.
10. Students may not enter the lab unless accompanied by a teacher or technician.

B. Hazardous Substances

The scale for each of the four categories for hazardous substances runs from 1 to 4, with 4 being the most dangerous.



SYMBOLS



Material causing chronic or ***long term*** **effects**. These include teratogens (cause birth defects), mutagens (cause mutations)and carcinogens (cause cancer), and others. Examples: benzene (carcinogen), mercury (Minimata disease)

Compressed, dissolved or liquefied **gas**—
Any of these materials under **pressure**.

Material causing ***immediate*** and serious **toxic** effects. Examples include cyanide(CN^-) and H_2S, both of which can kill immediately in low doses.

Oxidizing agent. Will cause flames if it comes into contact with a fuel such as dust or other reducing agent. Examples: $KMnO_4$ and NH_4NO_3

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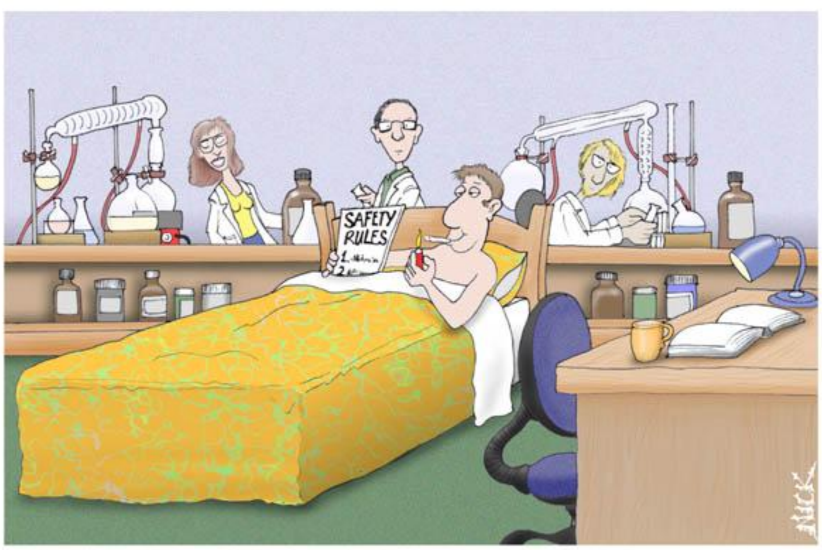
Dangerously reactive material. Some of these can lead to a violent reaction with just the addition of water. Examples of such compounds include all alkali metal compounds, calcium, calcium carbide(releases acetylene gas) , and acid halides which release halogens ifdumped in sink.

Corrosive material. Corrodes steel, pipes and especially flesh. Examples: NaOH, KOH, HCl, and other bases and acids.

Flammable materials. These will catch on fire either spontaneously or if placed near a heat source(whichdoes not have to be a flame). Examples: acetone, propane, carbon disulfide, powdered aluminum and ether.

Biohazardous and Infectious Materials These can cause disease and include viruses, some bacteria and fungi and samples of urine, blood and tissue.



Recent Advances in Science, #32:

Dr Ed Henderson demonstrates that it IS possible for someone to become too familiar with the safety rules.

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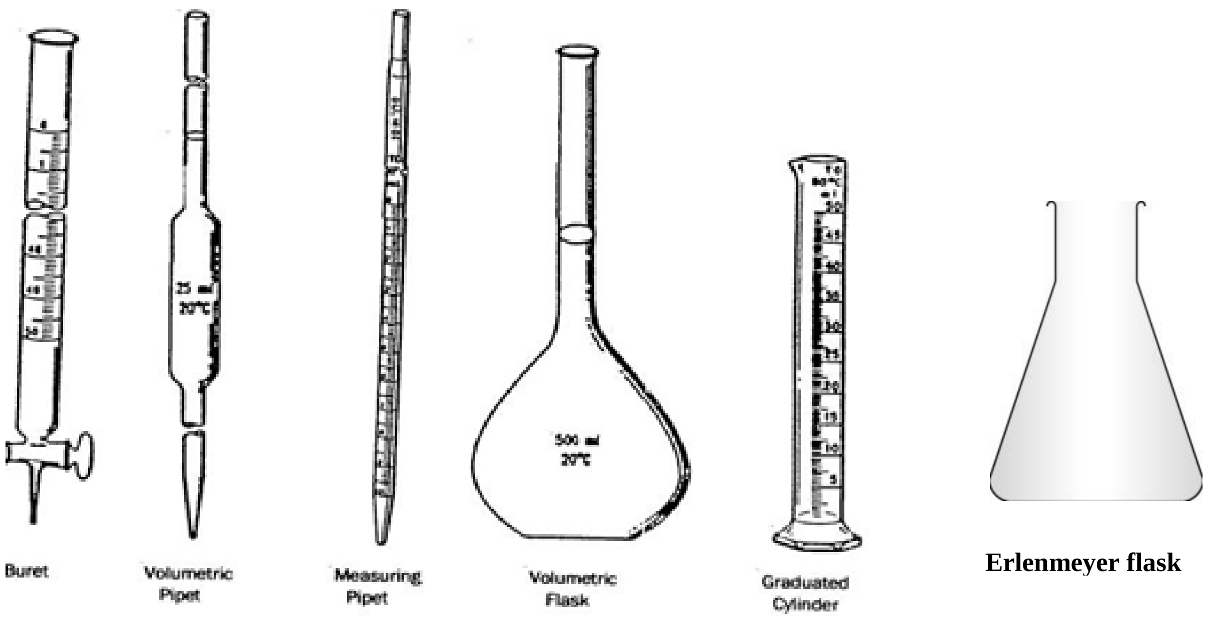
C. Knowing Your Equipment



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2. Guidelines for Writing a Lab Report

(1) **Procedure** Write the procedure without any recorded data, mathematical steps or discussion of formulas. Only write what you actually performed in the lab, and make sure you mention all the equipment.

NOTE:

- The procedure has to be written in numbered format.
- *If the procedure is given to you, there is no need to recopy it. The procedure will not be given for problem-solving labs and for lab exams.*

(2) **Data and observations** Record your data in a table.

- All data must include 1 estimated figure. If only 1 decimal can be estimated, for example, and you think the temperature is exactly twenty degrees, you must write 20.**0** ^{o}C.
- Include all units for all measurements

(3) **Analysis** Explain your reasoning and show how you arrived at your answer mathematically. Usually you will just solve the given problems in this section.

(4) **Conclusion**

- In two or three sentences, mention the answer to your problem and what equipment and method (very briefly!---the details will be found in your analysis) were used to solve the problem. If the experiment includes an "aim", that will be a major clue for writing the conclusion.

- Also include any equations for reactions that took place.

- Avoid the common mistake of writing as if the reader has already done the lab. Help your reader understand what you learned from the lab.

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3. <u>Significant Figures</u>

What are significant figures?

They are digits that are obtained or meaningfully derived from a measurement. ***They include the single estimated figure that is part of every measurement.*** For example suppose that you weigh some sodium chloride. The electronic balance reads 2.0 g. It's important that you record the mass as 2.0 g. The balance is telling you (at the very best; the uncertainty can be even greater.) that the actual mass is anywhere between 1.95 and 2.04 grams. But if you record 2g, you are telling your reader that the mass lies anywhere between 1.5 and 2.4 grams: an injustice to the balance!

When performing calculations, significant figures must also be respected. For instance if the 2.0 g of NaCl is dissolved in 60.0 mL of water, the calculator yields a concentration of 2.0/0.060 L = 33.3333333g /L. Where do you round off? The correct answer is 33g/L. When dividing or multiplying, the answer must have as many significant figures as contained in the original measurement *with the least number of significant figures*.

Not convinced? Remember when we measure 2.0, we don't know what the third decimal really is. Similarly, for 60.0 mL the actual volume may be anywhere from 59.95 to 60.04. So let's look at some possibilities.

mass(g)	**/olume(ml**	**volume (L)**	**concentration (g/L)**
2.01	60.04	0.06004	33.47768155
2.01	60.01	0.06001	33.4944176
2.01	59.95	0.05995	33.52793995
2.04	60.04	0.06004	33.97734843
2.04	60.01	0.06001	33.99433428
2.04	59.95	0.05995	34.02835696
1.96	60.04	0.06004	32.6449034
1.96	60.01	0.06001	32.66122313
1.96	59.95	0.05995	32.69391159

The results range from 32.6 to 34.0. Clearly the decimal place is meaningless. 33 g/L is the answer that reflects the accuracy of our measurements.

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Rules for Determining Significant Figures

1. All non-zero digits are significant. In any measurement, the last significant figure is an estimate but is still significant.

Example	**Number of Significant Figures**
2365	4
1.3365	5
11	2

2. Leading zeros (0's before the nonzero digits) are **NOT** significant.

Examples	**Number of Significant Figures**
0.005333	4
0.3	1
0.000010	2

If you find this illogical, remember that the above numbers can be converted into scientific notation. For example 0.3 = 3 X10^{-1} has one significant figure.

3. Captive zeros (those between nonzero digits) are significant.

Examples	**Number of Significant Figures**
0.005003	4
0.301	3
3005	4

4. Trailing zeros (those after nonzero digits) are significant only if the measurement contains a decimal.

Examples	**Number of Significant Figures**
50000	1
50000.	5
7.000	4

5. Exact numbers and irrational numbers have an infinite number of significant numbers.

Examples	**Number of Significant Figures**
π	infinite
The 2 in C = 2 πr	infinite
3 people—no doubt about measurement	infinite

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6. After carrying out a multiplication or division, the answer must have as many significant figures as contained in the original measurement *with the least number of significant figures*.

Examples	**Answer**	**Number of Sig Figs in Answer**
100./25	4.0	2
622.3 *3.5	2.2 X 10^3	2
3.65 per 4 people	0.912 per person	3(note the 4 people have an infinite number of sig figs and do not act as the limiting measurement if there is no doubt that there are 4 people

7. For adding and subtracting, the answer must have the same number of decimal places as the least precise measurement (the one with the **least** decimal places.)

Examples	**Answer**
23.91 + 11.999	35.91
11110.2 + .1333	11110.3

<u>Important:</u> The rules for significant figures should be applied only to the final answer. That implies that all decimal places should be retained in the middle of a calculation.

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Siginificant Figures Exercise

1. What is the correct number of significant figures in the following measurements?

   a. 2.0004 cm
   b. 300. cm(there is a deliberate point after the last zero)
   c. 300 m
   d. There are about 6 billion people on earth.
   e. 1.9900 ml
   f. 2.00×10^4 kg
   g. 0.00403 g

2. The mass of an empty can is 61 ± 1 grams. Then 30 ± 1 grams of water are added to the can. What is the lowest possible total mass for the can and water? The highest?

3. A student observed that the temperature of 100.0 ml of water with a known density of 1.0 g/ml increased from 10.5 ^oC to 22.8 ^oC. Express the amount of heat absorbed by the water in kJ with the correct number of significant figures. Use c = 4.19 J/(g^{o}C).

4. A gas sample contains 1000.233 moles of He and 10.35 moles of H_2. What is the total number of moles of gas in the sample, expressed with the correct number of sig- figs?

5. H_2 with a molar mass of 2(1.00797) g/mole consumes 8.0 grams of sodium, according to the following reaction:

   $$2\ Na_{(s)} + H_{2(g)} \rightarrow 2\ NaH_{(s)}$$

   How many grams of sodium hydride, NaH will be produced? Express with the correct number of significant figures.

6. The initial concentration of ozone, O_3 is 7.0 moles/L. If the equilibrium concentration of oxygen is 3.0 moles/L, what is the value of K, expressed with the correct number of sig figs?

7. In a lab, our measurements for the height of a tube ranged from 5.5 cm to 6.8 cm. The least accurate concentrations of the solutions used was 0.0010 M. Assuming that the value for K(an equilibrium constant) should have been expressed with just as many significant figures as in the above numbers, how should Peter have expressed the average of the following 4 values for K?

   96.37754
   87.05914
   126.1661
   230.3015