Chemistry S338
Relative Atomic Mass

Introduction:
   The relative mass of an object is how many times more massive the object is than a standard object. The purpose of the following experiment is to become familiar with how to determine relative masses. Part I involves determining the relative masses of a variety of seeds. Part II extends the idea of relative masses to atoms.

Procedure: (Part I)

1. Count an equal number of grains of orzo, of black beans, and of garbanzo beans into each of three plastic cups. All the orzo goes in one container, all the black beans in another, and so on. Your teacher will assign you a number of seeds to use, so that all the groups have a different number between 40 to 60 seeds. **You must have the same number of each.** Try and pick individual seeds that look pretty close in size, and discard any that are broken.
2. Find the mass of the seeds in each container. (Tare the empty cup on the balance as instructed by your teacher to determine the mass of just the seeds).
3. Copy/complete the following data table in the data section of your lab notebook.

Type of Seed	Mass of Seeds (g)	Relative Mass Compared to the Lightest Seed (See # 4 below)	"Atomic Mass" (See # 5 below)	Number of Seeds Needed to Equal the "Atomic Mass" (See # 7 below)
Orzo				
Black Beans				
Garbanzo Beans				

4. To calculate the "Relative Mass" column, identify the lightest of all seeds (it should probably be orzo). Divide the "Mass of Seeds" for each type by this lightest mass, and enter the result in the Relative Mass column. (The lightest seed should have a relative mass of 1, and the others will be heavier).
5. All of the atomic masses that you see on the periodic table <u>are</u> relative masses. (They have all been compared to something – you will find out what that something is in part two of this lab)! Since atomic masses *are* relative masses, your atomic mass column should be the exact same as your relative mass column. The units for your answers should be in grams.
6. Compare your "atomic mass" column with other students who did not originally use the same number of seeds as you.
7. Add seeds of the lightest grain one by one to the balance, until you have as close to its "atomic mass" number that as you can get. Count them. Repeat this procedure for the other grains, except each time count the number of seeds needed to make up the "atomic mass" in grams for *that* seed.

Questions for Discussion (Part I)
1. How did your atomic mass compare with students that used a different numbers of seeds? Explain your finding.
2. How did your "Number of Seeds Needed To Equal The Atomic Mass" column compare with students that used a different number of seeds? Explain your finding.
3. Maybe some of you already did this, maybe some of you didn't…explain <u>how</u> you could have mathematically solved for the "Number of Seeds Needed to Equal the Atomic Mass" column (rather than adding the grains one by one on the balance until you got as close as you could get to its atomic mass). What numbers would you need? What kind of math would you do?

Procedure: (Part II)

1. Below is a chart with the average masses of individual atoms. Reproduce this table in Microsoft Excel. (Enter 1.674 x 10⁻²⁴ as 1.674E-24).

Atom	Mass of one atom (g)	Mass relative to hydrogen	Average Atomic Mass (Look on periodic table)	Number of atoms in a relative mass	Percent Abundance of *This Particular Isotope*
Hydrogen-1	1.674×10^{-24}				99.985%
Carbon-12	1.993×10^{-23}				98.89%
Cobalt-59	9.786×10^{-23}				~100%
Aluminum-27	4.480×10^{-23}				~100%
Gold-197	3.271×10^{-22}				~100%
Bismuth-209	3.470×10^{-22}				~100%
Chlorine-35	5.807×10^{-23}				75.73%
Copper-65	1.078×10^{-22}				30.9%

2. Your goal is to create formulas that will allow you to quickly calculate the "relative mass" and "# of atoms…" columns. Your teacher may need to help you if you have never written formulas on a spreadsheet.
3. Complete the table comparing the mass of each atom to hydrogen. (Hint: Use the table you used for part one as a model).
4. **Create and complete a new table, this time using** <u>carbon</u> **as the standard atom for comparison rather than hydrogen.** You may want to copy the previous table and simply alter the formulas. (**Three important changes** - In this table, include an extra column called "Adjusted Relative Mass" inbetween the "Mass relative to…" and the "Average Atomic Mass…" columns. Be sure to change the title of column three to be "Mass relative to *Carbon*." Finally, change the "# of atoms…." column to be "Number of atoms in an *adjusted* relative mass").
5. In order to calculate the "adjusted" relative mass column, each of the values in the "mass relative to carbon" column should be multiplied by 12.
6. Have your teacher check your table. After he/she has checked it, print out a copy of each table and tape them in to your lab notebook.

Questions for Discussion (Part II) - You may find your textbook useful when answering these questions.

1. Why were you asked to adjust the masses relative to carbon, but not those relative to hydrogen?
2. How do the relative masses for hydrogen and the adjusted relative masses carbon compare to the atomic masses from the periodic table?
3. Which standard (hydrogen or carbon) gives values the closest to the actual atomic masses on the periodic table? Each of these isotopes has been used as a mass standard at one time in "history." Which is used today, and why was it chosen over others? (We will discuss this question in class if you can't find the information easily in your textbook or other resources).
4. When comparing relative to atomic masses, did you notice any isotopes that didn't fit the pattern? Why is this the case? (Hints: Think of the differences between the masses you were asked to solve for and the masses on the periodic table. Also look at the percent abundance column to try to solve this question).
5. What is the number of atoms of each element in one relative mass? What does this number represent? (Look at page 83 in your book for the second part of this question).

Reference: The idea for Part II of the lab came from a lab found at
http://intro.chem.okstate.edu/ChemSource/Moles/mole6.htm