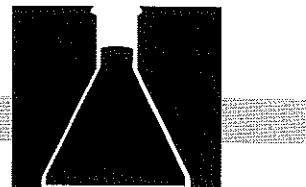


# Concept/Skills Development



## LABORATORY ACTIVITY: STUDENT VERSION

### Activity 1: Understanding the Mole

[Adapted from: Laboratory Manual accompanying Chemical Principles by Masterton and Slowinski. Modifications by Norma Mackenzie of Charlotte Country Day School, Charlotte, North Carolina.]



#### Introduction

The relative mass of an object is how many times more massive the object is than a standard object. The atomic masses of atoms are all relative masses. They can be considered relative to any particular element. Historically, both oxygen and carbon have served as the reference standard. For our purposes we can also consider atomic masses relative to the least massive element—hydrogen, with an atomic mass of approximately one. Fluorine, with a relative mass of 19, is 19 times more massive than hydrogen, *etc.* In this laboratory exercise you deal with the relative masses of beans. Then you will be asked to draw a parallel to the atomic masses of elements.

#### Purpose

To develop an understanding of the mole concept and molar masses of elements through an analogy with a model system.

#### Safety

Just don't spill the beans!

#### Procedure

##### Part I

1. Obtain a numbered plastic cup. Record the number of the cup; be sure to use the same cup during the entire activity.
2. Record the mass of your cup to the nearest 0.01 g.
3. Count out exactly 100 beans of one type. Discard any beans that differ greatly from an average bean. If you fail to do this, your results will not be accurate.
4. Weigh the beans; record their mass to the nearest 0.01 g.
5. Repeat Steps 3 and 4 for each type of bean provided.

#### Calculations

6. Calculate the mass of 100 beans by subtracting the mass of the plastic cup from the mass of the plastic cup + 100 beans.
7. Calculate (do not weigh) the mass of one bean of each type. Record the value in the data table. [NOTE: "Calculate" means to take the total mass of 100 beans and divide by 100 rather than weighing one particular bean.]
8. Determine the relative mass of each type of bean.

$$\text{Relative Mass} = \frac{\text{Average mass of one bean}}{\text{Average mass of the lightest type of bean}}$$

9. Calculate the number of beans in one relative mass of each bean. Do this by dividing the average mass of one bean into the relative mass.

$$\text{Number of beans in one relative mass} = \frac{\text{Relative mass}}{\text{Average mass of one bean}}$$

10. Check your calculated results in Step 9 by following these steps:
- Weigh the empty plastic cup again and record its mass.
  - Determine the sum of the relative mass of one type bean and the mass of the plastic cup.
  - Place the plastic cup on the balance pan and add beans of that type until the balance contains one relative mass of that type bean. (The mass should be that calculated for the cup plus one relative mass.)
  - Count the beans. Record this as the measured number of beans in one relative mass.
  - Pour the beans into a pile. Retain your separate piles of relative masses of beans. You will answer questions about them later.
  - Repeat for each type bean.

### Data Analysis and Concept Development

Cup Number \_\_\_\_\_

Mass of empty cup  
before activity \_\_\_\_\_ g

Mass of empty cup  
after activity \_\_\_\_\_ g

	Bean 1	Bean 2	Bean 3	Bean 4
Mass of 100 beans plus cup (g)				
Mass of 100 beans (g)				
Average mass of one bean (g)				
Relative mass of beans				
Calculated number of beans in one relative mass				
Measured number of beans in one relative mass				



## Implications and Applications

1. What did you find out about the number of beans in one relative mass? How do your calculated values compare to your measured values?
2. How many times larger is the lima bean relative mass than the lentil bean relative mass? How does this relate to the number of beans in the relative mass?
3. In a very logical statement, explain why there are always the same number of beans in one relative mass. Think very hard about this one.
4. Compare the volume of relative mass piles. Are they the same? Why or why not?
5. a. What is the average mass of the lightest bean?  
b. What is the relative mass of the lightest bean?
6. Among the elements, hydrogen has the least massive atoms—an atom of hydrogen has an average mass of  $1.66 \times 10^{-24}$  g. This is very small, but remember it is only *one* atom! What is the relative mass of hydrogen if it is the least massive element?

### Part II

Examine the beakers of elements and compounds on display. Each beaker contains one mole of a particular substance. Record information on the chart below.

Substance	Volume of one mole	Mass of one mole	Smallest unit of substance (atom or molecule?)
Sucrose ( $C_{12}H_{22}O_{11}$ )			
Carbon (C)			
Tin (Sn)			
Zinc (Zn)			
Aluminum (Al)			
Copper (Cu)			
Water ( $H_2O$ )			

7. a. Which substance occupies the largest volume?  
b. Which substance occupies the smallest volume?
8. a. One mole of which substance has the largest mass?  
b. Which substance has the most massive individual particles?  
c. Why *must* the answers to Questions 14a and 14b be the same?
9. Why do equal amounts (in moles) of different elements and compounds have different masses even though they have the same number of particles?
10. Why do equal amounts (in moles) of elements or compounds occupy different volumes?
11. Describe the fastest way to measure one relative mass of beans.
12. Describe the fastest way to measure out one mole of beans.

**Part III**

Below is a chart reporting the average masses of individual atoms. Calculate the relative mass of each element and record it in the chart. Then look up the molar mass (atomic mass) of each element on a Periodic Table and record it in the table.

Atom	Mass of one atom (g)	Mass relative to hydrogen (calculate)	Atomic mass	Number of atoms in a relative mass (g) (calculate)*
Hydrogen	$1.66 \times 10^{-24}$			
Carbon	$2.00 \times 10^{-23}$			
Iron	$9.30 \times 10^{-23}$			
Aluminum	$4.49 \times 10^{-23}$			
Zinc	$1.08 \times 10^{-22}$			
Lead	$3.44 \times 10^{-22}$			
Copper	$1.05 \times 10^{-22}$			

\* Although this calculation works with a relative mass of hydrogen of 1.00 as its basis, atomic masses are actually calculated as relative to a particular type of carbon atom viewed as exactly 12.00.

13. Comment on how the atomic masses found on the Periodic Table compare to the relative masses you calculated.
14. What are atomic masses, and how are they determined?
15. What did you find out about the number of atoms of each element in one relative mass?
16. Whose name is given to the number of atoms in the relative mass?
17. One atomic mass unit (u) is the approximate mass of a proton or a neutron. This mass is equal to  $1.66 \times 10^{-24}$  g. How many atomic mass units are in one gram?
18. If an atom has a mass of 197 u, find the mass (in g) of a mole of these atoms.