The concept of atoms dates back to the Greeks when philosophers argued the merits of matter that was either continuous or discontinuous. The discontinuous theory advocated that all substances were collections of tiny particles – the smallest possible pieces of matter. Although the Greeks never pursued this idea experimentally, they are credited with its origin.

In the 1700’s scientists began to investigate matter in more quantitative ways, and several important scientific laws ensued due to the collecting, recording and analysis of data. A French chemist, Joseph Proust, studied samples of copper ore called malachite. Malachite is composed of atoms of copper, carbon, and oxygen. He collected malachite from various countries around the world and analyzed them.

<table>
<thead>
<tr>
<th>Origin of Sample</th>
<th>Mass (g)</th>
<th>Cu Mass (g)</th>
<th>C Mass (g)</th>
<th>O Mass (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Africa</td>
<td>23.91</td>
<td>12.30</td>
<td>2.32</td>
<td>9.29</td>
</tr>
<tr>
<td>South America</td>
<td>1.206</td>
<td>0.620</td>
<td>0.117</td>
<td>0.469</td>
</tr>
<tr>
<td>Asia</td>
<td>7.909</td>
<td>4.069</td>
<td>0.767</td>
<td>3.073</td>
</tr>
</tbody>
</table>

One pattern that appears in studying the data is that the total mass of the sample is equal to the sum of its components. This seems obvious to us today, but this concept was a major step forward at that time. For instance:

\[ g \text{ of malachite} = g \text{ copper} + g \text{ carbon} + g \text{ oxygen} \]
\[ 23.91 \quad = \quad 12.30 \quad + \quad 2.32 \quad g \quad + \quad 9.29 \quad g \]

The idea that the weight of the products should be equal to the weight of the reactant(s) is known as the Law of Conservation of Mass. Chemists make use of this law in many ways, sometimes to calculate the % purity of an impure sample, or to help them confirm the % composition of a compound.

The % composition of a pure substance can be calculated from the above data also. If we calculate the % copper in the African malachite we would compare the amount of copper to the weight of the total sample (this is assuming that the sample is pure).

\[
\% \text{ Cu} = \frac{\text{wt copper}}{\text{wt sample}} \times 100\% = \frac{12.30 \text{ g}}{23.91} \times 100\% = 51.44\%
\]

If we do the same for the sample from Asia, we would find the % Cu to be 51.45%. The two are virtually the same within experimental error! So...in other words, % composition of a pure substance is independent of origin. This is known as the Law of Constant Composition or the Law of Definite Proportions and can be stated as ‘all samples of a pure substance have the same % composition.’ This means that to confirm that we have a given substance, one test we can perform is that of determining the % composition. If two samples have the same % composition and the same intrinsic properties, they are the same substance!
In 1803 John Dalton proposed his atomic theory based on his own observations, as well as those of many others. He also relied on the Law of Conservation of Mass and the Law of Constant Composition to propose his hypotheses.

1. Matter is composed of tiny, indivisible particles called atoms.
2. All atoms of an element are alike in weight, size, etc.
3. Atoms of different elements have different properties from each other.
4. Atoms of different elements combine in simple, whole-number ratios to form compounds.
5. The same set of elements can combine in different ratios, but each ratio corresponds to a different compound.

Postulate 4 is a reflection of the Law of Constant Composition. Postulate 5, interestingly, adds a wrinkle to the combination of specific atoms. This means, for instance, that hydrogen and oxygen can combine to form water, H₂O, but on the other hand – can also combine to form hydrogen peroxide, H₂O₂.

In our laboratory exercise today, we will investigate the reaction between zinc and hydrochloric acid. The reaction produces a compound between zinc and chlorine (from the acid). Our purpose today will be to determine whether or not the same compound is produced if the initial amount of zinc is varied from reaction to reaction. In other words, will some combinations produce ZnCl while others produce ZnCl₂, ZnCl₃, or Zn₂Cl, etc? The compound that is produced in the reaction is soluble in water and so it will be unseen, but dissolved in the liquid in the beaker.

For purposes of mathematical calculations, we will use the word formula:

\[
\text{zinc} + \text{chlorine} \Rightarrow \text{compound}
\]

The amount of zinc in the reaction will be varied from 1 – 4 g as indicated below.

- A: 1.000 g zinc
- B: 2.000 g zinc
- C: 3.000 g zinc
- D: 4.000 g zinc

**PROCEDURE**

1. Clean a 50 mL beaker with your towel. Put on the balance and hit the TARE control.
2. If you have been assigned amount A, place a star at the top of that column on the data sheet to distinguish your data from others in the class. Add zinc pieces to the beaker until you come close to 1.000 g Zn as you can. For instance you might record 1.054 g or 0.962 g; this will not affect the experiment as long as you record the amount of Zn correctly on your data sheet.
3. Measure 10.0 mL 6M HCl (hydrochloric acid) in your graduated cylinder (the unit M is a concentration unit for the acid). Add to the beaker and record your observations on the data sheet.
4. Allow the reaction to proceed for at least 20 minutes.
5. Weigh a dry, clean 100 mL beaker and record the weight on the data sheet.
6. After the reaction has stopped, pour the liquid from your sample into the 100 mL beaker. If there is any metallic zinc left, carefully pour out the solution, keeping the zinc in the beaker.
7. Wash the 50 mL beaker (and zinc, if any) with 5 mL distilled water. Then pour this wash water carefully into the 100 mL beaker.
8. Place any remaining zinc on a paper towel to dry for 10 minutes. Weigh a weighing boat, then scrape the zinc into the boat and reweigh. Record these two weights on the data sheet. Put this left-over zinc in the Zinc Waste Container.

9. Place the 100 mL beaker on a hot plate. This is to evaporate the solution down to the dry compound. Do not splatter the solution!! If there are signs of splattering, turn the hot plate down.

10. Continue heating until the solid in the dish is dry. Be sure that there is no liquid water left in the bottom of the beaker. The top of the compound should appear dry with no moist areas apparent.

11. Allow the beaker to cool for five minutes, then test for coolness by cupping your hands over the beaker. If warmth still noted, allow to cool further. The bottom and sides of the beaker must be cool to the touch before it is weighed.

12. Weigh the beaker and its contents and record this data.

13. Find the number of grams of chlorine in the compound by mathematics. Fill in the table using your data.

14. Obtain data from others in the class and complete the other columns in the table.

15. You may wash the compound out with lots of water and put down the sink.
2. Record the data below.

- **Original wt of zinc**
- **Wt. 100 mL beaker**
- **Wt. Weighing boat**
- **Wt. Zinc + weighing boat**
- **Wt leftover zinc**
- **Wt beaker + compound**
- **Wt compound**

3. Complete Table I for your data. Include all units. Show your work as indicated on the next page.
   
   *Be sure to use the rules for significant figures in your calculations!*

<table>
<thead>
<tr>
<th></th>
<th>SAMPLE A</th>
<th>SAMPLE B</th>
<th>SAMPLE C</th>
<th>SAMPLE D</th>
</tr>
</thead>
<tbody>
<tr>
<td>Wt zinc</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Wt left-over Zn</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Wt Zn reacted</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Wt chlorine reacted</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Wt compound</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ratio Zn/Cl</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>% Zn</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>% Cl</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

a. Weight of zinc in the compound (zinc reacted):
b. Weight of the compound:

c. Weight of chlorine in the compound (chlorine reacted):

d. Ratio Zn/Cl. Often instead of reporting the % composition, chemists calculate the ratio of one element to another in a compound. This is merely the decimal representation of dividing the weight of the Zn by the weight of the Cl.

e. % Zn and % Cl:

**QUESTIONS:**

1. a. When washing out the 50 mL beaker, the wash water contains the dissolved compound. How would your results for % Zn be changed if this wash water was thrown away? (Consider the equations you use to calculate % Zn.) Explain your answer.

   b. If you did not completely evaporate all the water from the compound in the 100 mL beaker, how would your % Zn change? Be specific and explain your answer.

2. Emma performed the experiment as written; she had no zinc left when the reaction was complete. After she finished the experiment she checked her data. How did Emma know that she had made an error somewhere? Explain fully.
3. Use the numbers for % Zn in all the columns to calculate the % deviation of your sample. Show your work.

\[
\text{% deviation} = \frac{\text{your # - average}}{\text{average}} \times 100 =
\]

4. Do you believe that the same compound was produced in each sample? Why or why not? Use your numerical calculations to justify your answer.

5. Explain how you used the two laws of chemistry in this experiment.
   a) Law of Conservation of Mass
   b) Law of Constant Composition