



Activity 3

Atoms and Their Masses



GOALS

In this activity you will:

- Explore the idea of atoms by trying to isolate a single atom.
- Measure how many times greater mass a copper atom has than an aluminum atom.
- Practice careful laboratory technique with measuring masses and filtration.
- Locate sources of the variation in the class's experimental results.
- Compare Dalton's experimental results to the masses of atoms known today.
- See that atoms react in definite proportions of mass when forming a compound.
- Relate the mole concept to real quantities.
- Use scientific notation in calculations.

What Do You Think?

Atoms are the smallest, indivisible part of an element.

- When did you first hear of atoms? What did they mean to you then, and what do they mean to you now?
- Use the following headings: Things I Know about Atoms, Things I Think I Know about Atoms, Things I Would Like to Know about Atoms. Include at least one entry under each heading.

Record your ideas in your *Active Chemistry* log. Be prepared to discuss your responses with your small group and the class.

Investigate

1. One way to think about an atom is to imagine trying to isolate a single atom from a large number of atoms.

Take a piece of aluminum foil and cut it in half. Take one of the resulting pieces and cut it in half again. Repeat this process with each successive half until you cannot make another cut.

- a) In your *Active Chemistry* log, record how many cuts you were able to make. How does the size of the smallest

What Do You Think?

The student's background should already support the concept of atoms and they probably have heard of atoms in primary school. Their concept of atoms at this point was probably related to science fiction and now they should associate atoms with compounds and elements. It would be interesting to have students discuss why they believe in atoms and not in the idea that everything is continuous. They may want to use some examples of why they believe in atoms.

Student Conceptions

In this activity, students develop an idea of the meaning of an atom and associate it with mass. The largest disparity between scientists' conception of atoms and that of high school students is that students often regard atoms as small pieces of solid having the same macroscopic-perceived properties (such as color and hardness) of the solid. This misconception arises because students have difficulty imagining how small an atom is. While atoms of an element do share properties with a macroscopic sample of the element, it is the chemical properties that arise from the internal structure of the atom that are shared. The macroscopic properties, chiefly those that can be perceived, are instead due to the arrangements of atoms in the material and how the atoms interact with each other. Students sometimes also confuse atoms with cells, since both are smaller than the eye can perceive and both are building blocks organized in definite structures. A second misconception students hold is that matter is continuous (not particulate), that is, that there are no spaces between particles. This may arise from the observation that solids and liquids are incompressible, coupled with a lack of understanding of the conservation of matter. While most high school students can imagine gases as being composed of particles, many do not attribute the same particulate composition to liquids and solids. A third misconception is that atoms do not have mass. This again points to a lack of understanding of the conservation of matter, for while gases are composed of particles which students believe to be atoms, many students believe gases to be massless. Students may also argue that because atoms are so tiny, they have no mass.

Investigate

Teaching Suggestions and Sample Answers

1. a) Try to keep the aluminum foil smooth throughout the cutting process. It will help in seeing that it maintains the same properties. Sharp scissors should allow students to cut a piece that is extremely small and it should lead them to the conclusion that it will still have the same properties as the original piece had. Exercise all necessary safety precautions when using sharp scissors.

piece of aluminum compare to the size of the original piece? Does the smallest piece of aluminum have the same properties as the original piece of aluminum foil?

2. An atom is the smallest part of an element. Since you can still cut the aluminum in pieces, you have not reached the size of a single atom.

Now imagine that there may be a way to cut the smallest piece of aluminum you have into even smaller and smaller pieces.

- a) How small can the smallest piece be and still retain the properties of aluminum? Could you cut the piece in half and half again 10 more times? 100 more times? 1000 more times?

Using your imagination of cutting and cutting, you will eventually get to one atom of aluminum.

3. Chemists combine elements to form new substances. By measuring the amounts of elements used and substances formed, they are able to draw conclusions about the properties of the elements involved. You will study the reaction between aluminum metal and a solution of copper (II) chloride.

Read through the procedure below.

- a) Make a table in your *Active Chemistry* log for the data you will be collecting. You will need room for measurements (mass) and observations. You can use a table similar to the one provided below.
4. Check your balance to make sure that it reads zero with nothing on it. Then measure the mass of a 50-mL beaker.
- a) Record the mass in your *Active Chemistry* log.



Safety goggles and a lab apron must be worn during this activity.

Finding the mass of aluminum	
1. Mass of empty 50-mL beaker	g
2. Mass of beaker and aluminum foil	g
3. Calculate the mass of aluminum from 1 and 2 above.	g
Finding the mass of copper (II) chloride	
4. Mass of weighing paper	g
5. Mass of paper and copper (II) chloride	g
6. Calculate the mass of copper (II) chloride from 4 and 5 above.	g
Finding the mass of the product	
7. Mass of dry filter paper	g
8. Mass of filter paper with product, after drying	g
9. Calculate the mass of the product material from 7 and 8 above.	g

Chem Tip:

The density of aluminum is 2.70 g/cm^3 and 1 mole of aluminum has a mass of 27.0 g. Then the volume of 1 atom can be determined using the following calculations: $(27.0 \text{ g/mol})/(2.70 \text{ g/cm}^3) = 10.0 \text{ cm}^3/\text{mol}$
 $= 10.0 \text{ cm}^3/6.02 \times 10^{23} \text{ atoms}$
 $= 1.66 \times 10^{-23} \text{ cm}^3$

2. a) Theoretically, if we had 1 mole of aluminum and it has a mass of 27.0 g, then 1 atom of aluminum would have a volume of $1.66 \times 10^{-23} \text{ cm}^3$.

This is extremely small and of course we cannot see it. Do not expect the students to be able to give this size, as they do not have the background at this time to mathematically figure out what the value should be. The students should realize that it will be extremely small in size.

3. Word of caution: the early scientists did call the combining of atoms the compound atom. We do not want the students to think that every compound will be called a molecule. Ionic compounds differ from molecular compounds. Water happens to be a molecular compound.
- a) The data table shown in the text is provided as a **Blackline Master** in this Teacher's Edition at the end of the chapter.

The mass measurements were based on using a balance that measures to the 100th place. Sample data are provided on page 55. This activity provides good practice on working with significant digits.

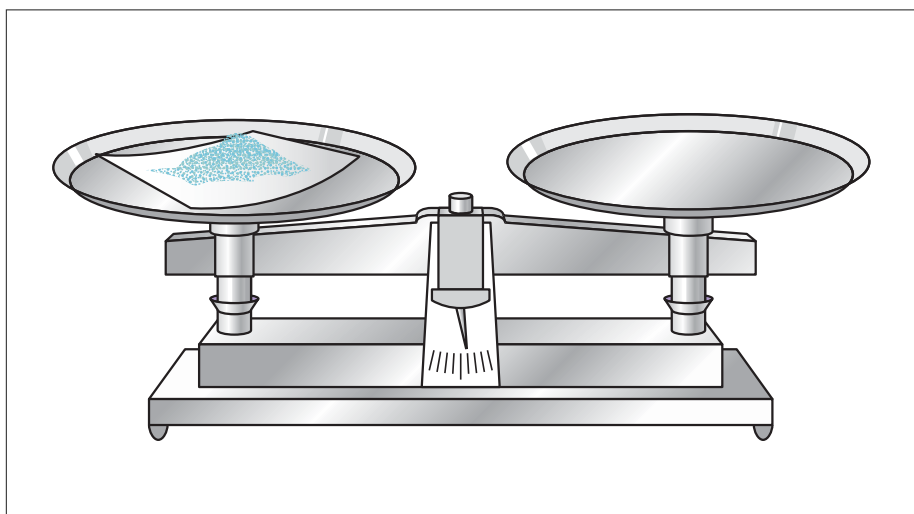
4. See page 55 for sample data.

Chem Tip:

The pieces of aluminum that were used in **Step 1** and **2** could be used in this part of the activity and would make it easier for the students to come close to the 0.20 g of aluminum.



Active Chemistry The Periodic Table



5. Measure out approximately 0.20 g of aluminum foil into the empty beaker. Try to get your mass measurement close to the assigned value.

If you have a centigram balance, you'll need to adjust the balance to read 0.20 more grams than the beaker alone. Then add pieces of aluminum foil until it rebalances.

If you have an electronic balance, simply add pieces of aluminum foil until the display indicates 0.20 g more than the empty beaker.

- a) Record the value that you obtain, even though you might not hit the target value.

6. Measure the mass of a piece of weighing paper.

Place approximately 2.00 g of copper (II) chloride on the weighing paper. (The chemical elements you would expect to find in copper (II) chloride

are copper and chlorine.) Again, remember that your target value is 2.00 g and that you may be slightly over or under this value.

- a) Record the masses in your *Active Chemistry* log.

7. Add the copper (II) chloride to the beaker.

Next add water to the beaker until it is approximately half full.

- a) Record your observations in your *Active Chemistry* log. Consider including the following: how the beaker feels when you touch it; what you hear when you listen closely to the beaker; what you see happening in the beaker.

- b) What color forms on the aluminum foil? What do you think is responsible for this color? Where is the color coming from?



Wipe up any spills immediately.

5. The aluminum is the limiting reagent in this reaction. If you use dihydrate copper (II) chloride you need a minimum of 1.9 g. Some of the aluminum may not react and will be one of the factors that causes the value of copper recovered to be a little light. Theoretically, 0.2 g of aluminum will produce 0.7 g of copper.
- a) A sample data table is provided later in this activity.

Chem Tip:

The pieces of aluminum that were used in **Step 1** could be used in **Step 5**. It would make it easier for the students to come close to the 0.20 g of aluminum required.

6. a) A sample data table is provided later in this activity.
7. a) The solution color becomes blue and the aluminum starts to react with the cupric ions to form aluminum ions. Students should be able to hear the reaction taking place and they should be able to detect a temperature change.
- b) The copper appears to be forming on the aluminum and the color of the solution is becoming lighter.

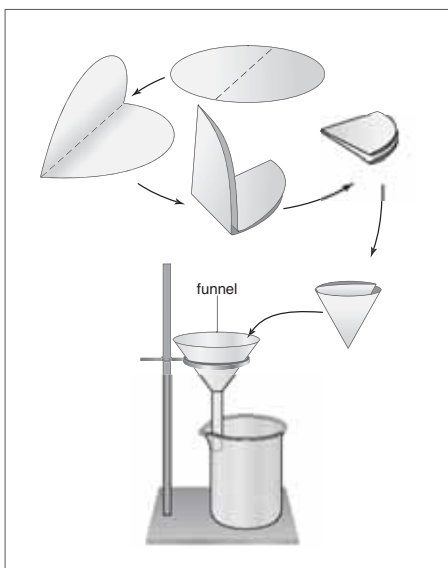
Chem Tip:

1 mole of copper has a mass of 63.5 g. We see from the equation that 3 mol of copper will be produced when 2 mol of aluminum metal reacts. Then the total mass of copper produced is 3 mol (63.5 g/mol) = 190.5 g = 190 g or 1.90×10^2 g (final answer must have 3 significant digits, 190 would be read as though it only has 2 significant digits). 2 mol of aluminum would have a mass of 2 mol (27.0 g/mol) = 54.0 g. Then we find that the mass of copper is 3.5 times greater than the mass of aluminum.

8. You will now need to find the mass of the substance formed in the chemical reaction. You can filter out this substance and then find its mass.

First measure the mass of a piece of dry, clean filter paper.

- a) Record the mass in your *Active Chemistry* log.
9. Set up a filtration system, as shown in the diagram below.
10. Wait until you no longer see or hear any reaction between the aluminum foil and copper (II) chloride and the liquid begins to clear up. Pour the contents of the 50-mL beaker into the funnel. You should rinse the beaker a couple of times with some water to be sure that all of the contents of the beaker are transferred.



Remove the filter from the funnel and place it on a piece of folded paper towel and allow it to dry overnight. Label the paper towel so that you can identify your filter paper.

Clean and put away your equipment and dispose of your chemicals as directed by your teacher. Wash your hands.

11. When the filter paper is dry, measure the mass of the filter paper and its contents. Dispose of the filter paper and its contents as directed by your teacher. Wash your hands.
- a) Record the mass in your *Active Chemistry* log. Determine and record the mass of the contents of the filter paper.
 - b) What element is inside the filter paper?
12. The reaction you witnessed is called a **single-displacement reaction**. In this reaction, a single element (aluminum) replaces another element (copper) in its combined form (copper (II) chloride). As a result of this reaction, the copper leaves its combined form to become an uncombined, or free element. The aluminum leaves its uncombined form to join with the chlorine to form a new compound. It's time to look at your data and see if you can make some sense of the numbers.
- a) How many grams of aluminum did you start with? How many grams of copper did you end up with (contents of the dry filter paper)?
 - b) How many times as great is the mass of the copper as compared

8. a) A sample data table is provided below.
9. Draw the students' attention to the diagram showing how to fold and insert the filter paper into the funnel. A better seal of the filter paper on the funnel can be sometimes obtained if the outside corner of the filter paper is torn off. Also, you may wish to suggest that students wet the entire filter paper prior to inserting it into the funnel.
10. The solution can be poured down the drain.
11. b) Copper metal is collected on the filter paper. All of the aluminum should have reacted in the reaction. The aluminum metal is the limiting reagent in this reaction.

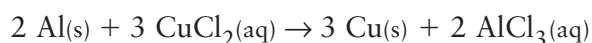
Sample data for the activity is provided below. The mass measurements were based on using a balance that measures to the 100th place.

1. Mass of empty 50-mL beaker	0.97 g
2. Mass of beaker with Al foil in it	1.17 g
3. Mass of Al foil alone = (#2) – (#1)	0.20 g
4. Mass of weighing paper	0.65 g
5. Mass of weighing paper with copper (II) chloride	2.65 g
6. Mass of copper (II) chloride alone = (#5) – (#4)	2.00 g
7. Mass of dry filter paper before using it	0.66 g
8. Mass of filter paper with product material, after drying	1.37 g
9. Mass of product material = (#8) – (#7)	0.71 g

Or, if you want to use zinc in place of aluminum then the values obtained would be:

1. Mass of empty 50-mL beaker	0.97 g
2. Mass of beaker with Zn in it	1.17 g
3. Mass of Zinc alone = (#2) – (#1)	0.20 g
4. Mass of weighing paper	0.65 g
5. Mass of weighing paper with copper (II) chloride	2.65 g
6. Mass of copper (II) chloride alone = (#5) – (#4)	2.00 g
7. Mass of dry filter paper before using it	0.66 g
8. Mass of filter paper with product material, after drying	0.85 g
9. Mass of product material = (#8) – (#7)	0.19 g

The overall reaction for aluminum with copper (II) chloride is:



Cu : Al mass ratio is 190 g : 54 g or about 3.5 : 1



Active Chemistry The Periodic Table

to the mass of aluminum you originally used? (What is the ratio of mass of copper to mass of aluminum?)

- c) If one atom of copper is released for each atom of aluminum that becomes combined (with chlorine), the masses of copper and aluminum determined in (a) should contain the same number of atoms. What does this tell you about the relative masses of copper and aluminum atoms? The number of atoms in 0.20 g of aluminum EQUALS the number of atoms in 0.71 g of copper. How many times more massive is a copper atom than an aluminum atom? (You may wish to compare

objects that you are familiar with: a dozen bowling balls of mass 60.0 kg can be compared to a dozen eggs of mass 2.4 kg. Since there are a dozen bowling balls and a dozen eggs, you can find that the bowling balls are 25 times as heavy as the eggs (60.0 kg/2.4 kg).

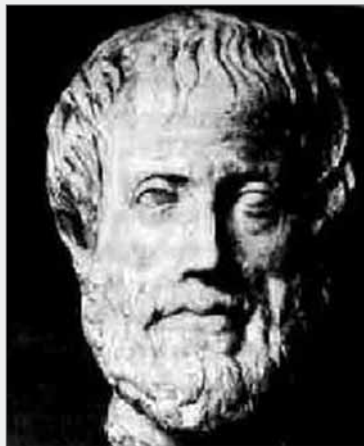
13. Recall that every group in your class reacted the same mass of aluminum with the same mass of copper (II) chloride. Discuss the similarities and differences in the data and calculations among the groups in the class.
- a) Record your thoughts on how and why the results are similar and/or different.

ChemTalk

ATOMIC MASS

Atoms

In **Activity 2**, you defined the term element and explored the properties of some common elements. In this activity, you focused on atoms. An **atom** is the smallest representative part of an element. The ancient Greek philosopher Aristotle did not believe in the existence of atoms. In his thinking, if atoms did exist, there would have to be empty space between them.



Aristotle

Chem Words

atom: the smallest representative part of an element

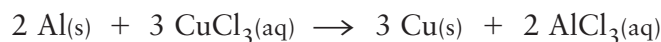
12. Using the data table provided on the previous page:

- a) 0.20 g of aluminum and it produced 0.71 g of copper.

Chem Tip:

The aluminum is the limiting reagent in this reaction. An excess of copper chloride is used in this reaction. Only 1.5 g of copper (II) chloride is needed to react with 0.20 g of aluminum. About 0.5 g of copper (II) chloride remains in the solution and this also explains why the solution is still slightly blue.

- b) 0.71 g of copper/0.20 g of aluminum. The copper is 3.5 times greater than the mass of aluminum.
- c) The mass of a copper atom is 3.5 times greater than the mass of an aluminum atom. Note that the values that we are using are the correct values. The students' values will vary from this and you could discuss errors in measurements and what percent error students have in their measurements. Also, remember that we are working with Dalton's law and he assumed that elements reacted atom-to-atom ratio. Some of the students should question this when they check the periodic table. We know that the equation is actually:

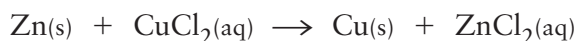


From the equation we see that 54 g of aluminum will produce 191 g of copper and the ratio is $191/54 = 3.5$.

13. The results should be similar since the reaction of copper ions with aluminum metal should always be the same. If it is not the same then laboratory technique needs to be reviewed.

If you use the data for the reaction of zinc with cupric chloride, then:

The overall reaction for zinc with copper (II) chloride is:



Cu : Zn mass ratio is 63.5 g : 65.4 g or about 1 : 1

0.20 g of copper will react with 0.19 g of zinc.

Aristotle did not believe it was possible to have empty space. Not everyone agreed with Aristotle. Another ancient Greek named Democritus believed that matter was made up of tiny particles that could not be broken down further. He called the particles atoms, from the Greek word *atomos*, meaning indivisible.

If you could have continued cutting the aluminum foil until it could no longer be cut, by any means, you would have reached one atom of aluminum. A mind-expanding fact is that if you started with 27 g of aluminum, you would find that there are 6.02×10^{23} atoms of aluminum. Nobody has ever counted this nor could they. Scientists have determined this number by other means and are very confident that it is correct.

Masses of Elements and Compounds in a Reaction

By the turn of the 19th century, chemists were combining elements to form new substances. The new substance was called a **compound**, because the atoms of the elements were believed to combine to form what they called a compound atom. The chemists were also particularly interested in measuring the amounts of elements used and substances formed. Their first attempts in determining masses were wrong, possibly due to the equipment that they had available at that time.

John Dalton, an early 19th century chemist who did much to advance the belief in the existence of atoms, expected that atoms combined in the simplest possible relationship. He reported that seven pounds of oxygen reacted with one



John Dalton

Chem Words

compound: a material that consists of two or more elements united together in definite proportion



Active Chemistry The Periodic Table

pound of hydrogen to form water. Accurate modern experiments give eight pounds to one. We will use modern values rather than historical ones to avoid confusion. Dalton reported that five pounds of nitrogen reacted with one pound of hydrogen to form ammonia. He also reported that seven pounds of oxygen reacted with five pounds of nitrogen to form a compound he called nitrous gas.

In 1809, Joseph Gay-Lussac reported that the hydrogen reacting with oxygen to form water occupied twice as much volume as the oxygen. He also noted that the hydrogen reacting with nitrogen to form ammonia occupied three times as much volume as the nitrogen. Furthermore, he found that equal volumes of nitrogen and oxygen reacted to form nitrous gas (now known as nitric oxide or nitrogen monoxide, NO).

Gay-Lussac's data was inconsistent with Dalton's assumption that water, ammonia, and nitrous gas are formed from one atom of each of the combining elements. This inconsistency was subsequently resolved by Amadeo Avogadro, who furthered the understanding of the correct chemical formulas and atomic masses.

Relative Mass of Atoms

Eventually, chemists determined a scale of relative masses of atoms through the systematic study of chemical reactions. By measuring the masses of two elements reacting with each other and knowing the formula for the compound that was formed, the relative mass of the two elements was determined. In this way, chemists were able to determine, for example, that one element has twice the mass of a second element. Relative mass does not tell you the exact mass measured in kilograms. It does provide a relative scale. Comparison of many reactions resulted in a scale of relative masses. Atoms of carbon were found to have a mass 12 times greater than the mass of hydrogen atoms, whereas oxygen atoms were found to have a mass 16 times greater than the mass of hydrogen. The units for this scale are called **atomic mass units**, defined in such a way that the mass of one type of carbon (carbon-12) is exactly 12 atomic mass units. The average mass of an atom of a given element in atomic mass units is known as the **atomic mass**. Atoms of hydrogen have an atomic mass of one unit. In addition to the physical and chemical properties of

Chem Words

atomic mass unit (amu): a unit of mass defined as one-twelfth of the mass of a carbon-12 atom

atomic mass: atomic mass is determined by the mass of the protons and neutrons of the atom

law of definite proportions: the composition of a pure substance is always the same or the elements of the compound always combine in the same proportion by mass

Element	Relative atomic mass
Aluminum	26.98
Copper	63.55
Iodine	126.90
Iron	55.85
Magnesium	24.31
Silicon	28.09
Sulfur	32.06
Zinc	65.38

elements, the relative mass (incorrectly called weight) of each element was known and used by Mendeleev as he organized his table. The atomic mass is still one of the most prominent pieces of information provided for each element on the periodic table. The table gives the relative atomic masses of the eight elements that you observed in **Activity 2**.

Checking Up

1. What is the difference between an element and a compound?
2. What is an atom?
3. How is an atomic mass unit defined?
4. How can the existence of atoms help to explain the Law of Definite Proportions?

The Law of Definite Proportions

Chemists at the beginning of the 19th century noted that eight pounds of oxygen always reacted with one pound of hydrogen to form nine pounds of water. This observation is an example of **The Law of Definite Proportions**.

This law, first articulated by Joseph Proust in 1799, states that whenever two elements combine to form a compound, they do so in a definite proportion by mass. Proust based this statement on his observations that 100 pounds of copper, dissolved in nitric acid and precipitated by carbonates of soda (sodium) or potash (potassium), invariably gave 180 pounds of green carbonate. The Law of Definite Proportions is not direct proof of the existence of atoms. However, if you believe in the existence of atoms it does make it easier to explain why the Law of Definite Proportions should hold. The existence of atoms can also help explain why a given mass of aluminum reacting with sufficient copper (II) chloride in solution should always produce the same mass of copper.



ChemTalk

Checking Up

1. Elements contain only one kind of atom and a compound contains two or more different kinds of atoms in some definite ratio.
2. An atom is the smallest particle of an element that retains the same chemical reactivity as a larger sample of the element.
3. The atomic mass unit is defined as $1/12$ th the mass of a carbon-12 atom.
4. The existence of atoms leads us to believe that regardless of the amount of elements that we use, the ratio in combining or reacting is always the same and leads to the Law of Definite Proportions.



Active Chemistry The Periodic Table

Reflecting on the Activity and the Challenge

In this activity you learned how chemists measured elements in chemical reactions to determine the relative masses of atoms and how these masses were assembled into a scale of atomic masses. The atomic mass is one of the most important pieces of information listed for each element on the periodic

table. How will you incorporate it into your game about the periodic table? Will you test players' ability to simply identify the atomic mass from a periodic table, or will you require that players understand how the relative scale was determined?

Chemistry to Go

1. John Dalton believed that water was formed from the simplest combination of hydrogen and oxygen atoms—one of each. Observations today show that eight pounds of oxygen react with one pound of hydrogen to form water. We now know that Dalton's values were wrong but his attempt was based on the data that he had available at the time.
 - a) Based on these two statements, what conclusion could Dalton draw about the relative masses of oxygen and hydrogen atoms? How many times more massive is an oxygen atom than a hydrogen atom?
 - b) The atomic mass of oxygen is 16 and the atomic mass of hydrogen is 1. How do the current atomic masses of oxygen and hydrogen compare to Dalton's?
 - c) You know that water molecules are not made from one atom of hydrogen and one atom of oxygen. A water molecule is made up of 2 atoms of hydrogen and one atom of oxygen. If the pound of hydrogen reacting with eight pounds of oxygen is due to twice as many hydrogen atoms as there are oxygen atoms, how many times more massive is an oxygen atom than a hydrogen atom?
 - d) Are the values of these revised masses closer to the current atomic masses of oxygen and hydrogen atoms?
2. In Dalton's time, ammonia was observed to be formed when nitrogen reacted with hydrogen. Today's values show that fourteen pounds of nitrogen react with three pounds of hydrogen.
 - a) If ammonia were formed from Dalton's simplest formula of one atom of each element, what would he have concluded about the relative masses of nitrogen and hydrogen atoms?
 - b) Ammonia molecules are made not from one atom each of hydrogen and nitrogen but from three atoms of hydrogen and one atom of nitrogen. If the

Chemistry to Go

1.
 - a) The mass of oxygen was 7 times that of the hydrogen mass based on Dalton's observations. We now know that water is 2 hydrogen atoms to 1 oxygen atom and that the mass of oxygen to hydrogen is 8:1. Dalton assumed that it was one atom of hydrogen to 1 atom of oxygen.
 - b) Referring to **Part A**, we can see that currently the ratio of oxygen to hydrogen is 8:1 and Dalton's was 7:1.
 - c) A water molecule consists of 2 hydrogen atoms and 1 oxygen atom. This gives you a total mass of 18 grams or in Dalton terminology we would say 18 pounds. Then oxygen must weigh 16 pounds or the ratio of oxygen to hydrogen is 16:1.
 - d) Yes.
2.
 - a) If we use the value of 14 pounds for nitrogen and 3 pounds for hydrogen, the ratio of nitrogen to hydrogen is 14:3.
 - b) Since 3 pounds of hydrogen from **Part A** is really for 3 atoms of hydrogen, then the ratio of a nitrogen atom to a hydrogen atom is 14:1.

three pounds of hydrogen reacting with fourteen pounds of nitrogen is due to three times as many hydrogen atoms as there are nitrogen atoms, how many times more massive is a nitrogen atom than a hydrogen atom?

3. A student uses magnesium instead of aluminum in this activity and obtains the following data:

Mass of beaker: 30.20 g

Mass of beaker + magnesium strip: 30.40 g

Mass of beaker + magnesium + copper (II) chloride: 32.40 g

Mass of beaker + magnesium + copper (II) chloride + water: 57.40 g

Mass of dry filter paper: 0.67 g

Mass of new beaker: 30.50 g

Mass of beaker + wet filter and residue + solution: 58.37 g

Mass of dry filter + residue: 1.19 g

- How many grams of magnesium did the student use in this experiment?
- How many grams of copper did the student measure in the dry filter paper?
- How many times as great is the mass of the copper as the mass of magnesium the student originally used? (What is the ratio of mass of copper to mass of aluminum?)
- If one atom of copper was released for each atom of magnesium that becomes combined (with chlorine), the masses of copper and magnesium determined in (a) and (b) should contain the same number of atoms. What does this tell you about the relative masses of copper and magnesium atoms? How many times more massive is a copper atom than a magnesium atom?



3.

Mass of beaker	30.20 g
Mass of beaker + magnesium strip	30.40 g
Mass of magnesium used	0.20 g
Mass of beaker + magnesium + copper (II) chloride	32.40 g
Mass of beaker + magnesium + copper (II) chloride + water	57.40 g
Mass of dry filter paper	0.67 g
Mass of new beaker	30.50 g
Mass of beaker + wet filter and residue + solution	58.37 g
Mass of dry filter + residue	1.19 g
Mass of residue (copper)	0.52 g

- a) 0.20 g
- b) 0.52 g
- c) 2.6:1 (mass of copper/mass of aluminum = $0.52 \text{ g}/0.20 \text{ g} = 2.6/1$)
- d) 2.6 times heavier.

Note that we did not use all of the data that was collected. Students need to know what data is important. We could have also given them room temperature as well.

- 4. They now will probably arrange the elements in order of increasing masses and they will probably keep the metals separated from the nonmetals.



Active Chemistry The Periodic Table

4. Look at the table in the **ChemTalk** reading section that gives the atomic masses of the eight elements that you observed in **Activity Two**. Add this data to your element cards. Can you now improve upon the way you sorted the cards in the previous activity taking this new information about relative masses into account?
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Preparing for the Chapter Challenge

In a paragraph, explain how the relative scale of atomic masses is determined.

Inquiring Further

Avogadro's number and a mole

Chemists are interested in keeping track of quantities of particles. However, the particles are very small so chemists use a particular quantity that is convenient for counting particles. The quantity is called a **mole**. The quantity of particles in a mole is 602,000,000,000,000,000,000,000. The mole can be represented more easily in scientific notation as 6.02×10^{23} . This is a very large number because many, many small particles (atoms or molecules) make up a mole. The number 6.02×10^{23} is sometimes called Avogadro's number.

Research to find the significance of Avogadro's number and a mole. Record your findings in your *Active Chemistry* log.

Then, to appreciate how huge a mole is, answer the following question:

Imagine you have 7 billion people (7×10^9 people, which is approximately the human population of the world) and they are given the task of dropping \$1-bills once every second into a large hole. How long will it take 7×10^9 people to drop 1 mole of dollar bills into the hole? How old will you be when they complete this task?

Preparing for the Chapter Challenge

Students may struggle with this question. They know the definition of atomic mass unit and that the mass of an element can be determined if they know the ratio of the atom to one atom of hydrogen or mass of the element to the mass of one gram of hydrogen. They can then use different relationship that they know to determine the other masses of the elements that had not been associated with hydrogen.

Inquiring Further

Avogadro's number and a mole

$$\frac{6.02 \times 10^{23} \text{ dollars}}{9 \times 10^{13} \text{ s}} = 9 \times 10^{13} \text{ s}$$

$$\frac{9 \times 10^{13} \text{ s}}{3 \times 10^6 \text{ years}} = 3 \times 10^6 \text{ years or 3 million years old}$$

Note that the 3600 s/hr, 24 hr/day are significant to as many places as necessary. The 365 day/yr has 3 significant digits. The 7×10^9 dollars/s only has 1 significant digit.