A SourceBook Module

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$$6.02 \times 10^{23}$$

THE MOLE

ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

The mole concept pervades all of chemistry. Since most quantitative chemical calculations are based on the mole, an understanding of the mole is essential to the study of chemistry. An understanding of how the mole relates to mass, number of entities (atoms, molecules, ions, *etc.*) and volume of a gas is included in this module. Formulas are most often interpreted in terms of moles rather than atoms, molecules or formula units (see *Language of Chemistry* section) since chemical reactions are generally carried out on a macroscopic scale.

The SI definition of mole is: "... the amount of substance of a system that contains as many elementary entities as there are atoms in 0.012 kg of carbon 12. When the mole is used, the elementary entities must be specified and may be atoms, molecules, ions, electrons, other particles, or specified groups of such particles."

Chemists use the mole (symbolized *mol*) to express amount of substance (elements, compounds, ions, electrons, *etc.*). Avogadro's number (6.02×10^{23}) is the number of entities in one mole of a substance. Because it is easier to weigh large numbers of particles than to count them, chemists generally relate moles of a substance to mass rather than to the number of particles. The molar mass (mass of one mole) of a substance is the sum of the relative atomic masses (in grams) of the elements in a formula unit of the substance. Avogadro's number of formula units of a substance has a mass equal to the molar mass of the substance.

The subscripts in a formula can be interpreted as the number of moles of the element present in a mole of the compound. Calculations involving percent composition of a compound and empirical formula involve the mole concept. (The relationship of moles to stoichiometry is the subject of another *SourceBook* module.)

This topic lends itself to a historical presentation (see *History* section).

PLACE IN THE CURRICULUM

Since it is the basis for quantitative chemistry, the mole is usually introduced early in the course. Because it requires abstract reasoning, the mole is a challenging concept for many beginning students. If students have a poor mathematical background, it may be desirable to postpone introduction of the mole concept, perhaps until after a discussion of atomic structure and bonding.

CENTRAL CONCEPTS

1. One mole of a substance represents Avogadro's number (6.02×10^{23}) of units of that substance.
2. The amount of a substance (expressed in moles) is related to the mass of substance, and for a gaseous substance, to its volume.
3. The mole is the basis of quantitative chemistry (percent composition, empirical and molecular formula determinations).

RELATED CONCEPTS

1. Mass, volume
2. Formulas (empirical and molecular)
3. Atoms, molecules, formula units

4. Percent
5. Accuracy and precision
6. Gas behavior (particulate nature of matter, relationship of volume tonumber of molecules of gas)
7. Stoichiometry

RELATED SKILLS

1. Measurement
2. Exponential numbers
3. Significant figure manipulation
4. Problem solving strategies (dimensional analysis, *etc.*)
5. Algebra
6. Use of balance
7. Volume measurements
8. Use of burner
9. Use of tongs to handle crucibles

PERFORMANCE OBJECTIVES

After completing their study of the mole concept, students should be able to:

1. define mole in terms of mass and number of entities.
2. explain the difference between the terms *mole* and *molecule*.
3. explain why chemists find it convenient to express the quantity of a substance in moles (rather than mass, volume, number of entities, or some other unit).
4. calculate the molar mass of a substance, given its formula.
5. calculate moles, mass or number of entities (atoms, molecules, formula units, *etc.*), given appropriate data.
6. calculate moles, mass or volume of a gas sample at STP, given appropriate data.
7. determine percent composition of a substance, given its formula.
8. determine empirical formula, given the percent composition of a substance or the relative masses of the elements in the substance.
9. determine the molecular formula of a substance, given its empirical formula and other appropriate data.

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:
    a. Weigh the empty plastic cup again and record its mass.
    b. Determine the sum of the relative mass of one type bean and the mass of the plastic cup.
    c. 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.)
    d. Count the beans. Record this as the measured number of beans in one relative mass.
    e. Pour the beans into a pile. Retain your separate piles of relative masses of beans. You will answer questions about them later.
    f. 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 1.66×10^{-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	1.66×10^{-24}			
Carbon	2.00×10^{-23}			
Iron	9.30×10^{-23}			
Aluminum	4.49×10^{-23}			
Zinc	1.08×10^{-22}			
Lead	3.44×10^{-22}			
Copper	1.05×10^{-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 1.66×10^{-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.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Understanding the Mole

Major Chemical Concepts

Molar masses are derived from relative masses (see SI definition of mole in *Content in a Nutshell*). Molar masses of different substances have different masses and volumes. Each relative mass contains the same number of particles.

Level

This activity may be used with first-year high school chemistry students in basic, general, or honors levels.

Expected Student Background

Students should know how to use a balance. They should have a scientific calculator and be able to use and understand the calculator's exponential function. It would be helpful if students were acquainted with the concepts of significant figures, and knew that the last digit is an estimated number and that some variation is expected in this value.

Time

This is a two-day activity, assuming 45 to 55-min class periods. *Part I* is completed on the first day. *Part II* and *Part III* are done on the second day. Students turn in data tables, calculations, and answers to questions at the close of the second day.

Safety

1. No safety concerns are associated with this activity; students do not need to wear protective goggles, unless strict school regulations require them *whenever* students are at the laboratory bench.
2. The "one mole" of elements and compounds should be covered to prevent spilling. There could be a physical hazard of slipping on beans if many of them are spilled on the floor.

Materials (For 24 students working in pairs)

Nonconsumables

Covered containers for beans
12 Numbered plastic drink cups
48 Zip-closure bags (optional)
6 Beakers, 50-mL (with graduations to hold one-mole amounts of elements and water)
Beaker, 800-mL (with graduations; to hold one mole of sucrose)

Approximate quantities of beans:

Lentil beans, 1 bag (3 oz.)
Lima beans, 2 large bags (2 1/4 lb.)
Navy beans, 1 bag (8 oz.)
Pinto beans, 1 large bag (1 1/4 lb.)

One mole of each substance in powdered or granular form, except for water:

Water	18 mL (18 g)
Aluminum	27 g
Copper	63.5 g
Sucrose (table sugar)	342 g
Carbon	12 g
Tin	119 g
Zinc	65 g

Advance Preparation

First Day. Number the plastic cups. Cups may differ in mass by as much as one gram; each group should continue to use the same weighed cup. Place beans in a closable container that is easy for students to get to. If desired, approximately 120 beans of each type can be weighed into zip-closure bags (lentils - 8 g, navy beans - 208 g, pinto beans - 48 g, lima beans - 120 g per zip-closure bag). This will save student time.

Second Day. Place one mole of each of the specified elements, sucrose, and water in separate beakers. Label each one, cover each beaker with plastic wrap and put a rubber band on it. This prevents spillage and evaporation of water. Place these beakers around the laboratory. (These samples can be stored covered and reused indefinitely - except for water, which will evaporate.)

Pre-Laboratory Discussion

Draw students' attention to the Periodic Table. They will have seen the table in the classroom for several weeks—and in previous courses. They are usually curious about what it means and why it is arranged the way it is. You might ask them to spot any regularities in the chart. Among other things, they should be able to see that atomic numbers always increase across a row. Atomic *masses* usually increase—but not always. They also are curious about the decimal values in the atomic masses. Point out that a hydrogen atom is the lightest atom and that, for example, a gold atom weighs about 197 times more than a hydrogen atom. Students are curious how these numbers were determined, and how Avogadro's number was determined. Explain to students that by doing this activity, they will be able to understand, in part, how these things came about.

An analogy could also be discussed with students. A dozen golf balls would not have the same mass as a dozen ping pong balls. We could find the relative mass of the golf balls if we weighed a dozen ping pong balls and a dozen golf balls, and divided by the smaller mass. This would tell us not only how many more times massive a dozen golf balls were than a dozen ping pong balls, but also how many times more massive one golf ball is than one ping pong ball.

Teacher-Student Interaction

1. Move from group to group, checking to see that students understand the instructions. It is best to guide students by responding with questions that lead them to correct answers.
2. The concept of relative mass (Question 8 in *Part I Calculations*) may be a difficult concept for students. The concept should be clarified when the calculated value is verified experimentally (Question 10 in *Part I Calculations*).
3. Insure that students understand the relationship of *Part II* to *Part I*. (Since individual beans can be handled, they provide a concrete analogy. Since atoms and molecules cannot be seen, one must visualize them mentally. The calculations involving beans in *Part I* are analogous to those involving elements in *Part II*.)

Anticipated Student Results

Cup Number X

Mass of empty cup before activity 12.15 g

Mass of empty cup after activity 12.15 g

	Bean 1 (Lentils)	Bean 2 (Navy)	Bean 3 (Pinto)	Bean 4 (Lima)
Mass of 100 beans plus cup (g)	18.16	28.65	52.55	113.15
Mass of 100 beans (g)	6.01	16.50	40.40	101.00
Average mass of one bean (g)	0.0601	0.165	0.404	1.01
Relative mass of beans	1.00	2.75	6.72	16.80
Calculated number of beans in one relative mass	16.60	16.70	16.70	16.80
Measured number of beans in one relative mass	17	17	17	17

These values are typical student values. Lima beans vary greatly in size, thus having the largest uncertainty.

Answers to Implications and Applications

1. The calculated number of beans in one relative mass stayed the same at 16.7 ± 0.1 bean. The measured number stayed constant at 17 ± 1 bean.
2. The lima bean relative mass is about 17 times larger than the lentil bean relative mass. There are 17 beans in a relative mass. These values are the same.
3. The relative mass is the ratio of the mass of one type of bean to the mass of another type of bean. This mass ratio insures that when we weigh beans in this mass ratio, we must obtain the same number of beans. (Since a lentil bean is only 1/17 as massive as the most massive bean measured—the lima bean—there must be 17 beans in a relative mass.)
4. The volumes of the relative mass piles are not the same. Even though each pile has the same number of beans, they have different sizes.
5. The average mass of the least massive bean is 0.0601 g. The relative mass of the least-massive bean is 1.00.
6. If it is the least massive element, its relative mass should be 1.00.

Part II

Substance	Volume of one mole	Mass of one mole	Smallest unit of substance (atom or molecule?)
Sucrose ($C_{12}H_{22}O_{11}$)	450 mL	342.0 g/mol	Molecule
Carbon (C)	6 mL	12.01 g/mol	Atom
Tin (Sn)	23 mL	118.7 g/mol	Atom
Zinc (Zn)	20 mL	65.4 g/mol	Atom
Aluminum (Al)	18 mL	27.0 g/mol	Atom
Copper (Cu)	12 mL	63.5 g/mol	Atom
Water (H_2O)	18 mL	18.0 g/mol	Molecule

Note: These are not molar volumes, but the approximate volumes occupied by 1 mole of each substance, *eg.* crystalline sucrose.

7. Sucrose occupied the largest volume. Carbon occupied the smallest volume.
8. Sucrose had the largest molar mass and the most massive individual particles. These answers must be the same because all beakers contained the same number of individual units.
9. One mole of various entities have different masses because their individual particles have different masses, just as different kinds of beans have different masses.
10. One mole of various entities occupies different volumes because their individual particles have different volumes, just as piles of relative masses of beans have different volumes. (The volumes of solids that are not a single crystal, however, will not be directly related to the volume of the atoms or particles.)
11. The fastest way to obtain a relative mass of beans would be to count the beans. The fastest way to obtain a mole of beans would be to weigh them. (At least in principle. The mass of a mole of beans would be incredibly large—on the order of 10^{22} g.)

Part III

Atom	Mass of one atom (g)	Mass relative to hydrogen (calculate)	Atomic mass	Number of atoms in a relative mass (g) (calculate)*
Hydrogen	1.66×10^{-24}	1.00	1.008	6.02×10^{23}
Carbon	2.00×10^{-23}	12.0	12.01	6.03×10^{23}
Iron	9.30×10^{-23}	56.0	55.85	6.02×10^{23}
Aluminum	4.49×10^{-23}	27	26.98	6.01×10^{23}
Zinc	1.08×10^{-22}	65.1	65.45	6.03×10^{23}
Lead	3.44×10^{-22}	207	207.2	6.02×10^{23}
Copper	1.05×10^{-22}	63.3	63.45	6.03×10^{23}

* calculated using columns 2 and 3

12. All atomic masses agree with the relative masses to three significant figures. (Some variation in the last significant digit is always expected.)
13. Atomic masses are relative masses. They are calculated relative to some reference atom. *NOTE: Molar masses of elements are accurately known.* The number of carbon atoms in a unit cell is determined precisely by X-ray crystallography, and the density of carbon is determined. Using these precise measurements, all other atomic masses are taken relative to the carbon-12 isotope as exactly 12.
14. The number of atoms in a relative mass is constant at 6.02×10^{23}.
15. This is Avogadro's number.
16. There are 6.02×10^{23} atomic mass units in one gram.
17. A mole of these atoms would have a mass of 197 g. *(TEACHER NOTE: This is a gold atom.)*

Post-Laboratory Activities

At this point, students can begin calculations regarding molar masses of different molecules. Worksheets where "dozens" are discussed along with the concept of the mole should be provided. Throughout the year, when the mole concept or atomic masses of elements are discussed, remind students of this bean activity. Students usually make the connection easily and have a model to help them understand these ideas.

Extensions

1. Ask students to weigh out 100 grains of rice, and then recalculate their relative masses. They could then rethink some of their answers about why relative masses contain the number of items they do.
2. Students might also like to apply this same idea to other familiar objects—for example, game balls (baseballs, soccer balls, bowling balls, footballs, golf balls, ping pong balls, *etc.*) They could weigh individual balls or dozens of balls and prepare a table with names, symbols, and relative masses. Since bowling balls come in 12-pound, 16-pound sizes (and other sizes), the idea of isotopes could be discussed. The concept of a weighted average relative mass could be developed. The concept of isotopes (or relative masses) could be developed with paper clips, nails, styrofoam balls, or anything with relatively constant mass and several types of individual items.

Assessing Laboratory Learning

1. The laboratory data sheets should be collected and graded. Most students will identify any large errors in measurement of relative masses because they will not fit with other answers.
2. This is an opportunity to use cooperative learning and let groups of students discuss their ideas and give consensus answers to some of the more difficult questions. A group grade can then be given. Students appreciate this communication on discussion questions; cooperative learning helps foster a positive feeling about this process. It is far better than obtaining 30 papers with identical answers to discussion questions that represented original thoughts of only a few students.

3. Questions on worksheets as well as on tests focusing on concepts developed in this learning activity should be posed throughout this module and throughout the year. Clarity of thought on this pivotal concept should be evaluated in as many contexts as possible. (The worksheet on “dozens” uses familiar English units, so that students can work with calculations analogous to the mole concept using more familiar units. If desired, metric units could be substituted. A suitable worksheet follows.)

Answers to Worksheet on Dozens and Moles

Part I

1. (192 lb bowling balls) / (0.24 lb ping pong balls) = **8.0 x 10^2 times**
2. (1.44 lb golf balls) / (0.24 lb ping pong balls) = **6.0 times**
3. (0.24 lb/1 doz) x (1 doz/12 balls) = **0.020 lb**
   0.020 lb x (16 oz/1 lb) = **0.32 oz**
4. 1.44 lb x (1 doz/12 balls) = **0.120 lb**
   0.120 lb x (16 oz/1 lb) = **1.92 oz**
5. 0.040 lb x (1 doz/0.24 lb) x (192 lb/1 doz) = **32 lb**
6. 180 balls x (1 doz/12 balls) = **15 doz**
   15 doz x (1.44 lb/1 doz) = **21.6 lb**
7. 48 lb x (1 doz/192 lb) x (12 balls/1 doz) = **3.0 balls**
8. 20,000 balls x (1 doz/12 balls) x (1.44 lb/1 doz) = **2.40 x 10^3 lb**

Part II

Element	**Mass of one mole**
Helium	4.00 g/mol
Carbon	12.0 g/mol
Gold	197 g/mol

1. (197 g/mol gold) / (4.00 g/mol helium) = **49.3 times**
2. (4.00 g/1 mol) x (1 mol/6.02 x 10^{23} atoms) = **6.64 x 10^{-24} g/atom**
   6.64 x 10^{-24} g x (1 u/1.66 x 10^{-24} g) = **4.00 u**
3. (12.0 g/1 mol) x (1 mol/6.02 x 10^{23} atoms) = **1.99 x 10^{-23} g/atom**
   1.99 x 10^{-23} g x (1 u/1.66 x 10^{-24} g) = **12.0 u**
4. 0.50 g x (1 mol/4.00 g) x (197 g/1 mol) = **25 g**
   (Using solution to Question 1: 0.50 g x 49.3 = **25 g)**
5. 3.01 x 10^{24} atoms x (1 mol/6.02 x 10^{23} atoms) = **5.00 mol**
   5.00 mol x (12.0 g/1 mol) = **60.0 g**
6. 25 g x (1 mol/197 g) x (6.02 x 10^{23} atoms/1 mol) = **7.6 x 10^{22} atoms**
7. 3.01 x 10^{19} atoms x (1 mol/6.02 x 10^{23} atoms) x (197 g/1 mol) = **9.85 x 10^{-3} g**

Name ___________________________ Period ______ Date __________

Worksheet on Dozens and Moles

Part I. The following are masses of one dozen game balls

Type of Ball	Mass of one dozen
Ping pong	0.24 lb (110 g)
Golf	1.44 lb (654 g)
Bowling	192 lb (87.2 kg)

1. How many times more massive is a bowling ball than a ping pong ball?

2. How many times more massive is a golf ball than a ping pong ball?

3. How much does one ping pong ball weigh in pounds? In ounces?

4. How much does one golf ball weigh in pounds? In ounces?

5. You have 0.040 lb (18 g) of ping pong balls. How many pounds of bowling balls must you obtain to have the same number of each type of ball?

6. How many dozen golf balls are present in 180 golf balls?

   How much would these golf balls weigh in pounds?

7. If you had 48 pounds of bowling balls, how many bowling balls would you have?

8. You must ship 20,000 golf balls. You do not have time to count them. It is faster to weigh them. How many pounds of golf balls must you ship?

Part II

Look up the atomic masses for helium, carbon, and gold to answer the following questions.

Element	**Mass of one mole**
Helium	
Carbon	
Gold	

1. How many times more massive is a gold atom than a helium atom?

2. What is the average mass of one helium atom in grams?

   What is its mass in atomic mass units (u)? (1 u = 1.66×10^{-24} g)

3. What is the average mass (in grams) of one carbon atom?

   In atomic mass units?

4. You have 0.50 g of helium atoms. How many grams of gold atoms must you obtain to have the same number of atoms of each?

5. How many moles of carbon atoms are there in 3.01×10^{24} atoms of carbon?

   How many grams of carbon is this?

6. If you have 25 g of gold, how many atoms of gold do you have?

7. How many grams of gold would you have to obtain to have 3.01×10^{19} atoms of gold?

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Percent Composition of Magnesium Oxide

Introduction

The Law of Definite Composition is introduced early in general chemistry as a part of exploring differences and similarities between compounds. Proust, a French chemist, was one of the first to state this principle. Using this principle, chemists in laboratories and industry are able to determine with high precision the amounts of each substance needed to make a compound.

Purpose

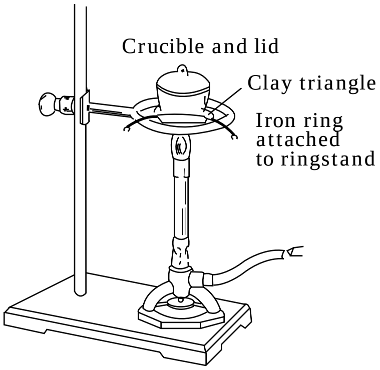
This activity is designed to support the concept of a definite percent by mass of each element in the binary compound magnesium oxide.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Do not touch the hot crucibles. Hot porcelain and cold porcelain look the same.
3. Follow proper procedures when lighting burners.
4. Keep your hair pulled back away from your face so it will not come close to the burner flame. Hair ignites easily, especially if hair spray is used.
5. Handle the Mg ribbon with tongs as much as possible. The ribbon is usually sharp along its edges from cleaning; it can cut easily. Do not look at the magnesium while it is burning.
6. When handling crucibles, hold a piece of wire gauze up close under the crucible while moving the crucible from the clay triangle to the balance, *etc.* This will save many crucibles and prevent hot porcelain from shattering and causing burns.
7. Dispose of materials as your teacher directs.
8. Thoroughly wash your hands before leaving the laboratory.

Procedure

1. Heat an empty crucible (with cover). Begin heating with a cool flame and gradually increase the temperature. Heat with a hot flame for 2-3 min. Allow to cool enough to handle with tongs. Weigh and record. Be careful not to place the crucible on counter tops, *etc.* that may melt or burn.
2. Place a coil of cleaned Mg ribbon provided for you in the crucible. (Your teacher may direct you to cut the Mg ribbon into small pieces to speed up the reaction.)
3. Weigh the crucible and Mg ribbon. Record.
4. Place the crucible and Mg ribbon in the clay triangle.
5. Heat the crucible and Mg ribbon in stages for 10-15 min. The clay triangle should glow red hot during this time.
6. Open the crucible with tongs to allow more air to enter while continuing to heat. Heat for 10 min more.



7. Remove the crucible and contents from the clay triangle and place it on a piece of wire gauze. Allow to cool about 5 min or until you can comfortably touch the crucible with a finger. *Be Careful!* Hold the back of your hand near the crucible to feel the heat.
8. Add 2 mL water to the crucible contents. (Use distilled water, if available.) Cover and return the crucible to the clay triangle.
9. Heat gently by moving the burner flame back and forth under the crucible for about 2 min. Then heat strongly for about 5 min to insure dryness.
10. Remove the crucible from the clay triangle and place it back on the wire gauze. Weigh when cool enough to handle. Record.
11. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

Data Chart

1. Mass of empty crucible and cover _________
2. Mass of crucible, cover, and Mg ribbon _________
3. Mass of Mg _________
4. Mass of crucible, cover, and contents after heating _________
5. Mass of contents (mostly MgO, we hope!) _________
6. Mass of oxygen combined with Mg _________

Observations

Be sure to note colors of the Mg ribbon and the crucible contents during the activity. Also note any odor you may have encountered when water was added.

Analysis of Data

1. Using the mass of the contents (mostly MgO) and the mass of the magnesium ribbon you used, calculate the percent magnesium in magnesium oxide. Using the mass of oxygen and mass of contents, calculate the percent oxygen in magnesium oxide.
2. Assume the formula for magnesium oxide is MgO. Calculate the theoretical mass percent of Mg and O. Molar masses are:

   Mg = 24.305 O = 15.9994 MgO = 40.304
3. Compare your data to theoretical values. Suggest reasons why your values might not match theoretical values.
4. Write a balanced equation for the primary reaction between Mg and O_2.

Implications and Applications

1. Why is it necessary to preheat the crucible?
2. Why was it necessary to open the crucible lid to allow more air to enter during the reaction?
3. Why was water added to the contents?
4. What changes in the procedure could have been made, if time allowed, to insure that the Mg ribbon had completely reacted?
5. How could you explain Mg mass percent values that are too high and those that are too low?
6. Explain how you could use this general laboratory procedure to identify a compound if you know the elements that compose it.
7. What effect would using unclean Mg ribbon have on your results?

*8. A student forgets to add water at the end of the procedure. What effect would this possibly have on the student's results? Do you think your equipment would be precise enough to detect this error?

*Optional

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Percent Composition of Magnesium Oxide

Major Chemical Concept

The activity illustrates the importance and usefulness of the Law of Definite Composition, emphasizes the mole concept, and involves calculations of empirical formulas.

Initial emphasis is on calculating percent composition; this is relatively easy for students to understand and calculations are straightforward. The determination of empirical formula and illustration of the utility of the mole concept are both clearer when pooled data for the entire class are used. Students calculate moles Mg and moles O from their experimental data. A plot of moles of Mg *vs.* moles of O for class data gives a graph with the expected slope of about 1.

Level

General and honors. Also possibly basic, with enough teacher guidance.

Expected Student Background

Students should be able to calculate simple percent values and deal with molar mass. They should be familiar with standard laboratory procedures. Additional background is optional, such as applications of oxidation-reduction concepts to this laboratory activity.

Time

This activity can be completed in one period if the pre-laboratory briefing is provided in a previous class period. If time becomes a concern, you can do the final weighing for students or save the final weighing until the next day, particularly if you have a drying oven to keep crucibles dry.

Safety

Read *Student Version* considerations.

Materials (For 24 students working in pairs)

Nonconsumables

Balance (to nearest 0.001 g if possible)
12 Burners
12 Clay triangles
12 Crucibles and lids
12 Crucible tongs
12 Graduated cylinders, 10-25 mL
12 Ringstands
12 Rings
12 Strikers (or other means to ignite the gas)
12 Wire gauze

Consumables

Distilled water, optional
Magnesium ribbon, Mg, 2.4 to 4.8 g

Advance Preparation

1. All equipment should be readily available to students. Having needed equipment in plain view at laboratory stations saves considerable student work time. In general, check equipment ahead of time.

2. Remove the oxide coating from the Mg ribbon by slightly sanding. Cut it in lengths with masses between 0.200 g and 0.400 g before the activity begins.
3. Check the balances to insure they work properly.
4. If distilled water is used, consider providing it in water bottles at each station.
5. You may wish to pose some problems (*e.g.*, percent composition, mole problems, *etc.*) for students to solve during the time samples are cooling.

Pre-Laboratory Discussion

1. Demonstrate any procedural skills students have not seen before, such as proper handling of crucibles with tongs and the lighting of a burner. It is important to show the set up so students will be able to quickly assemble it themselves.
2. Give students thorough instructions regarding safety. Warn students that cool and hot porcelain look the same; encourage caution when handling the crucibles. Holding a piece of wire gauze under the crucible when moving it from place to place will save some crucibles from rapid increases in entropy.
3. Explain that under the conditions of this activity magnesium reacts with nitrogen gas in air as well as with oxygen gas. The contents should be as free as possible of magnesium nitride, Mg_3N_2. This is accomplished by adding water to convert magnesium nitride into ammonia and magnesium oxide. A faint odor of ammonia, NH_3, may be detected by some students:

$$Mg_3N_2(s) + 3H_2O(l) \rightarrow 3MgO(s) + 2NH_3(g)$$

4. Students should be asked about procedure details such as opening the lid *[to allow more air in]*. They should also be asked why it is necessary to heat the Mg ribbon strongly until the clay triangle and possibly the crucible itself glow red hot *[to insure complete reaction and to dry the MgO after addition of water]*.
5. Insure that students understand how the collected data relate to the calculations they will perform.
6. You may decide to supply problems for students to solve or provide some other activity to keep students busy during the time samples are cooling. If so, explain the nature of the assignment during the pre-laboratory orientation.

Teacher-Student Interaction

1. Move from group to group, insuring proper set ups and safe lighting of burners. Do not hesitate to assist in the lighting of burners at first. This process should be monitored closely.
2. Since the crucibles will be heated for long periods of time, use this opportunity to further discuss the reactions with individual groups. Consider providing them with a list of questions you will expect them to answer after the activity so they can discuss them with each other or with you.
3. Circulate and be sure students make good observations when they open the crucible at different points in the procedure. They will note color changes.
4. Do not be alarmed if the final product is not pure white, as MgO should be. The gray coloration can be due to several reasons, but largely to impurities in the crucible and water. Students may note that this gray color is not the “expected” color of pure MgO. Invite them to suggest reasons. Actual laboratory data from this activity, however, usually do not indicate that this discoloration leads to any problems in the calculated results.

Anticipated Student Results

Here are actual data from one student's laboratory report.

1.	Mass of empty crucible and cover	27.225 g
2.	Mass of crucible, cover, and Mg ribbon	27.568 g
3.	Mass of Mg	0.343 g
4.	Mass of crucible, cover, and contents after heating	27.785 g
5.	Mass of contents (mostly MgO, we hope!)	0.560 g
6.	Mass of oxygen combined with Mg	0.217 g

% Mg = (0.343 g/0.560 g) x 100 = **61.3%**
% O = (0.217 g/0.560 g) x 100 = **38.8%**

Theoretical values:

% Mg = (24.3/40.3) x 100 = **60.3%**
% O = (16.0/40.3) x 100 = **39.7%**

Here are actual results from eight student groups:

% Mg	% O
74.0	26.0
63.1	36.9
60.9	39.1
60.8	39.2
69.9	30.1
61.1	38.9
60.7	39.3
98.3	1.70 *

* This group had continual trouble with their burner. At the close of the procedure the team noted that their Mg ribbon had hardly reacted. Do not be afraid to allow "poor" data to be collected and included in group analysis. Be prepared to lead the class in discussing what could have happened.

Answers to Implications and Applications

1. This removes water vapor and oils, which can affect masses if the balance is sensitive enough.
2. This is done to admit enough O_2 and N_2 to insure that the reaction of Mg is as complete as possible.
3. Water was added to convert magnesium nitride into magnesium oxide and ammonia, thus producing a purer MgO product.

$$Mg_3N_2(s) + 3H_2O(l) \rightarrow 3MgO(s) + 2NH_3(g)$$

4. Continuing to heat and reweigh until the masses are consistent would be one way to build confidence that the reaction was as complete as possible.
5. If the mass percent of magnesium is too high, it is likely that it did not react completely or some of the magnesium had oxidized to MgO before massing. If the mass percent is too low, it is most likely because not all of the water was evaporated. There is always a possibility of error in masses. (Also see the answer to Question 8, but do not expect students to include that in their answer to this question.)
6. If the elements in the compound were known and you could determine the mass percent composition, you could probably identify the compound by comparing the percent composition to that of known compounds.

7. If the Mg is not cleaned well enough at the start, some slight error would be introduced because the initial mass of Mg would already contain MgO. (This would lead to low results, as explained in the answer to Question 5.)

*8. The final products would consist of both magnesium nitride and magnesium oxide. Upon close mathematical examination, one would conclude that the products should have a lower mass. Mg_3N_2 has a molar mass of 100.9 g/mol, and although 3 MgO's contain the same mass of Mg it would have a molar mass of 120.9 g. Failure to add water would leave the 3 Mg in the nitride, not in the 3 MgO's. It is doubtful that most high school balances could detect this mass difference (probably < 0.05 g difference).

   (*NOTE: Very few students should be expected to display this much mathematical insight in answering the question.*)

* This question may be too difficult for students if stoichiometry has not been discussed at this point in the course.

Post-Laboratory Activities

1. Start by collecting class data for this activity. By posting results, a trend should emerge, revealing consistent results around 60% Mg and 40% O. Discuss how these results relate to the Law of Definite Composition.
2. Questioning about high or low percent values would be a natural follow up. Include as many students as possible in the discussion.
3. Ask such questions as "Would the amount of Mg used affect the results?" Also probe on the usefulness of knowing the percent composition of a compound. Propose that the students are chemists in charge of producing fertilizer for a company. Ask them to explain how this activity has a connection to their success in this company.

Extensions

1. A key part of this laboratory activity is to use data collected for the class (or several classes, if desired) to illustrate the importance of the mole concept and the determination of empirical formulas.
2. Ask students to calculate moles Mg and moles O from their experimental data. Plot moles Mg *vs.* moles O for the class. This usually gives a beautiful graph with the expected slope of about 1. Feel free to discard individual data points that are obviously far from the trend. These are decisions that scientists with guidance from statistics can make. Explain to students that many experimental trials are often needed to produce reliable results. With thousands of trials, data points far removed from the trend would have little effect on the overall statistical analysis of results.
3. This graph verifies the formula for magnesium oxide (MgO) that students used to calculate their theoretical mass percent values for Mg and O. Explain how the slope of the graph can be used to determine the empirical formula. (Alternately, extrapolate the graph to 1 mol Mg and show that the corresponding value for O is also 1 mol.)
4. Have students investigate common substances made up of the same elements that differ in their percent composition by mass, such as H_2O_2 and H_2O. Establish differences in their properties and uses.

5. Invite students to investigate a major industrial chemical process such as the Haber, Ostwald, Contact, or Hall process to identify major reactants and to reinforce the notion that knowledge of amounts of required substances is a component of being a successful industrial chemist or engineer.

Assessing Laboratory Learning

Laboratory Practical

Provide students with sample data on another compound composed of a metal and a nonmetal. Have them perform calculations to identify the compound by comparing their results to a list of "known" compound possibilities.

Examination Questions

1. Three metals are allowed to react as completely as possible to form their oxides. The possible oxides are Cu_2O, CuO, and ZnO. Which would have the highest percent oxygen by mass? *[CuO]*
2. If you were given 45.0 g of aluminum, what minimum mass of oxygen would you need to allow the aluminum to react completely to form Al_2O_3? *[40.0 g* O_2*]*

Other Laboratory Activities

1. A good activity is ask students to weigh a 1-in square of Al foil and calculate the number of such squares that must be stacked to provide 1 mol Al. A 1-in square of aluminum foil has almost exactly 10^8 Al atoms on each edge. Therefore, a 1-in cube of Al would approximate one mole (approximately 10^{24}) Al atoms.
2. *Determining the percent water in a hydrate.* This laboratory activity can also be used to illustrate the determination of percent composition. The calculations are not as clear to beginning students as the percent of an element in a compound (as in *Laboratory Activity 2*), but the results are usually good. (A typical laboratory experiment is given by Slowinski, 1985.)
3. *Determining Avogadro's number* (and molecular size) from measurements on a unimolecular layer of stearic or oleic acid. This activity provides a way to obtain an approximate value for Avogadro's number. The assumptions needed are usually beyond the grasp of beginning students. However, if you do not dwell on the assumptions, students can understand the calculations reasonably well and will obtain a value of approximately the correct order of magnitude (10^{22} to 10^{24}) for Avogadro's number. Corwin (1988) describes an experimental procedure.
4. *Determining an empirical formula from the reaction of Fe with* $CuSO_4$. This laboratory activity gives good results. A modified version of the activity is given in the *Stoichiometry* module. The data can be used to calculate the empirical formula for $CuSO_4$ (mole ratio of Cu^{2+} to SO_4^{2-}). The experimental procedure is the same.
5. *NOTE: We strongly discourage using the experimental determination of the empirical formula of a sulfide of copper, which you may find in older laboratory manuals.* Many individuals are extremely sensitive to the SO_2 formed in the activity. Persons with asthma can have severe reactions to SO_2.

DEMONSTRATIONS

Demonstration 1: A Molar Display

Purpose

To illustrate that one-mole samples of different substances have different volumes and masses.

Materials

One mole of various substances such as:

Water, H_2O, 18.0 g (18.0 mL)
Methanol, CH_3OH, 32.0 g (40.4 mL)
Copper (use copper dust), Cu, 63.6 g
Sodium chloride (table salt), NaCl, 58.5 g
Carbon (activated charcoal), C, 12.0 g
Sulfur, S, 32.1 g
Iron (iron filings), Fe, 55.8 g
Sucrose (table sugar), $C_{12}H_{22}O_{11}$, 342 g

Safety

No special precautions are needed if the above-listed substances are used.

Procedure

Display one mole of several different substances in identical containers, pointing out that one-mole samples of different substances have different volumes. Weigh one mole of several different substances to illustrate that moles of different substances have different masses. Suggestions of substances to use: water, methanol (or other alcohol), sucrose, salt, carbon, sulfur, iron filings, copper dust or other solids that are finely divided so the observed volume approximates the volume of the solid.

Demonstration 2: Molar Measurements

Purpose

To illustrate that the mole is a unit of measurement. (Extension of above demonstration.) This would also be an appropriate follow-up demonstration to the *Understanding the Mole* activity.

Materials

One mole and multiples and submultiples of a mole of a particular substance. For sodium chloride (table salt), NaCl, possible quantities are:

NaCl (0.100 mol), 5.85 g
NaCl (0.500 mol), 29.2 g
NaCl (1.00 mol), 58.5 g
NaCl (2.00 mol), 117 g
NaCl (5.00 mol), 292 g
NaCl (10.0 mol), 585 g

Safety

No special precautions are required if NaCl is used.

Procedure

Weigh one mole of a substance—for example sodium chloride or granular metal (Fe, Zn, *etc.*)—and arrange in a pile or heap. Label the pile "one mole NaCl" (or "one mole Zn,"*etc.*) Prepare several other piles of the same substance to represent different numbers of moles (1/2, 2, 3, 5, *etc.*); ask students to guess the number of moles of substance in each pile. Use this demonstration to discuss the concept that the mole is a way to measure substances—if one knows the amount that one mole represents, one can determine how many moles are present in quantities larger or smaller than one mole. (You could use sugar or salt packets and have students empty the contents and guess the number of moles. They could then weigh the substance and calculate the number of moles to check their answer. If desired, students could also calculate molecules of sugar, grams of carbon present, *etc.*)

Since "mole" means "pile" or "heap" and molecule means "smallest part of the pile or heap," you could also use this demonstration to highlight the difference between the terms "mole" and "molecule." Of course, students cannot see a molecule (or atom) of the substance, but if you use a granular substance, they can see that there are many "grains" in a pile and must therefore be many more molecules in a pile.

Demonstration 3: How High the Mole

Purpose

To help students visualize the size of Avogadro's number.

Materials

Mole Demo, a Seraphim computer program (see reference to AP 305, AR 301, PC 2301 in *Media* section)

Safety

No safety precautions are necessary

Procedure

The *Mole Demo* computer program illustrates the stacking of sheets of paper from earth to Pluto. Use this after introducing the mole and Avogadro's number to help students realize the magnitude of Avogadro's number.

Remarks

This program requires about 5 min. Since there is no dialog (but musical sound effects occur periodically), consider preparing your own running commentary to accompany the program.

Demonstration 4: Moles and Gases

This demonstration is on the Ring of Truth: Atoms *videotape and is also illustrated on CHEM Study video Gases and How They Combine (see references in* Media *section).*

Purpose

To illustrate the relation between gas volume and moles of gas.

Materials

Electrolysis apparatus (an inexpensive electrolysis apparatus is available from scientific supply vendors)
Distilled water
0.1 M Sulfuric acid, H_2SO_4, several milliliters (to serve as electrolyte to facilitate electrolysis)
Wooden splint (and match to ignite splint)

Safety

Observe usual safety precautions in handling the sulfuric acid. Follow directions supplied with electrolysis apparatus to set up the demonstration. Although the danger is slight for the small quantity of hydrogen produced in the electrolysis demonstration, the hydrogen could explode if ignited.

Procedure

An understanding of stoichiometry (or at least the meaning of formulas) is needed for students to understand this demonstration. Since the electrolysis of water produces two moles of hydrogen gas for every mole of oxygen gas, the volume of hydrogen formed in the electrolysis is twice the volume of oxygen. Consult directions accompanying the electrolysis apparatus to conduct the electrolysis demonstration. Collect 25-50 mL of the gases during electrolysis, or enough for the class to see that the volume of one gas is approximately twice the volume of the other. The gases collected can be tested to verify which is oxygen.

Ask the class to decide which gas is oxygen on the basis of the stoichiometry for the reaction (or the formula for water). Test only that gas. (Insert a glowing splint into a sample of the gas. If the splint bursts into flame, the gas is oxygen.) Warning: Do not attempt to ignite hydrogen as a proof that the other gas is hydrogen. Students will accept that the other electrolysis product must be hydrogen if water is electrolyzed.

GROUP AND DISCUSSION ACTIVITIES

1. Ask students to construct (from poster board) a 22.4 L cube so they can visualize the volume of one mole of gas at STP. (A typical 5 gal waste paper basket contains about 20 L, also a good approximation for the molar volume of a gas.)
2. Challenge students to calculate the number of molecules (formula units) present in various volumes of water or amounts of other substances (*i.e.*, molecules/moles in a grain of salt or sugar, drop of water, sugar cube, sugar or salt packet, *etc.*)
3. Students could set up *Demonstration 2* (each group using a different substance), organized as a game in which other students guess the number of moles.

Key Questions

1. Why do chemists often refer to moles of a substance rather than molecules (or atoms or formula units)? *[The mole is a convenient size unit to express the amounts of substances involved in laboratory-scale reactions.]*
2. How are moles, mass, and number of entities for a sample of matter related? *[1 mol of entities = formula mass (in grams) x (the molar mass) =* 6.02×10^{23} *entities]*
3. What is the relationship of moles to volume of a gas at STP? *[For an ideal gas at STP, 1 mol = 22.4 L]*
4. How is percent composition related to mole ratios? *[The percent of an element (A) in a compound is given by:*

$$\%A = \frac{(\text{moles A in compound}) \times (\text{molar mass A})}{(1 \text{ mol compound}) \times (\text{molar mass compound})} \times 100$$

5. How is the empirical formula related to the mole ratios of elements in a compound? *[The empirical formula relates to the relative numbers of moles of each element in the compound. The relative numbers of moles expressed as integers are the subscripts in the empirical formula. For example, if a compound contains 1.5 mol A for every mole of B, the ratio expressed as integers is 3:2—the empirical formula for the compound is* A_3B_2*.]*
6. Explain why equal masses of different substances do not contain equal numbers of molecules or moles. *[Since different substances usually do not have the same molar mass, one mole or Avogadro's number of entities (molecules,* etc.*) will have different masses. Therefore, equal masses must contain different numbers of moles or molecules.]*

Counterintuitive Example

1. 200 g of lead has fewer atoms than 1 g of hydrogen. (Other examples can be given.) One mole of lead (208 g) has the same number of atoms (Avogadro's number, or 6.02×10^{23}) as one mole of hydrogen (1 g). However, 200 g of lead is less than one mole of lead and has fewer than Avogadro's number of atoms, thus fewer than 1 g of hydrogen.

Analogies

1. A mole (6.02×10^{23}) is a counting number like dozen (12), gross (144), ream (500), *etc.* used to count objects such as eggs, pencils, or sheets of paper. Because atoms are so very small, measurements using ordinary laboratory-measuring equipment can only be made on very large numbers of them—on the order of Avogadro's number.
2. Astronomers estimate that there is a mole (6.02×10^{23}) of stars in the universe.
3. Analogy to mole/mass/molecule calculations: Counting rolls of coins and determining the number of coins and their value is analogous to mole/mass/molecule calculations, as the following activity suggests.
4. **Roll-Mole Analogy**(by Caroline L. Ayers). As anyone who has taught chemistry knows, the mole concept is one of the most difficult in chemistry. The roll analogy was devised to show students that calculations involving moles are similar to some they can easily do—calculations involving money. Any student can do relatively complex calculations if they involve money.

   The Roll-Mole Worksheet (which follows) is designed to be used by answering Problems 1, 2, 1', and 2' first. Then have students fill in the blanks on the equalities. After this point, students are told to "forget" that a nickel is worth 5¢ and a quarter is worth 25¢. They must not use that information in solving Problems 3 through 6 or the purpose of the exercise is defeated. (Nobody remembers the mass of an atom of an element or a formula unit of a compound!) The worksheet points out the information students are allowed to use to solve the problems. To use the worksheet most effectively, work Problem 3 and 3' side-by-side on the board so students see that the reasoning is identical.
   Students should be made aware of the reasoning (use a set up) for each problem so that they will recognize that moles are not really impossible to work with—Avogadro's number is just too large to picture. Students resist setting up the money problem. They say they do not know how they do it, they just do it. That is because such calculations are so familiar and thus easy to do. If they cannot tell what the set up should be, tell them and write it down, including units so they can see the units cancel out just as in the corresponding mole problem. Work Problems 4 and 4' together, *etc.* In the analogy, value ($) is like mass (g), number of coins is like number of atoms, and rolls of coins are like moles of atoms. (Note that "roll" rhymes with "mole".)

$$3 \text{ rolls nickels} \times \frac{40 \text{ nickels}}{1 \text{ roll nickels}} = 120 \text{ nickels}$$

$$3.00 \text{ mol S} \times \frac{6.02 \times 10^{23} \text{ atoms S}}{1 \text{ mol S}} = 1.81 \times 10^{24} \text{ atoms S}$$

Roll-Mole Analogy

Given the following information: There are 40 nickels in a roll.
There are 40 quarters in a roll.

Answer these questions. Show Set Up!

(1) What is the value of a roll of nickels?	(1') What is the mass of a mole of S?
(2) What is the value of a roll of quarters?	(2') What is the mass of a mole of Br?

_____ nickels = 1 roll nickels = $_______

_____ quarters = 1 roll quarters = $_______

______________ atoms S = 1 mole S = ________ g S

______________ atoms Br = 1 mole Br = ________ g Br

Use only the above equalities to answer the following questions:

(3) How many nickels are in 3 rolls of nickels?	(3') How many atoms of S are in 3.00 mol S?
(4) What is the value of 3 rolls of nickels?	(4') What is the mass of 3.00 mol S?
(5) How many rolls of quarters is $500?	(5') How many moles are in 53.8 g Br?
(6) How many nickels are in $500 of nickels?	(6') How many atoms are in 1.5 g Br?

Answers: (1) $2 (2) $10 (3) 120 (4) $6 (5) 50 (6) 10,000 (1') 32.1 g S (2') 79.9 g Br (3') 1.81 x 10^{24} atoms S (4') 96.3 g S (5') 0.673 mol Br (6') 1.1 x 102^{2} atom Br

4. *Empirical Formula Analogy* (by Kerro Knox, 1980)

   Given the average mass (weight) of boys and girls in a class and the percent boys by mass and percent girls by mass, calculate the ratio of boys to girls and the "formula" for the class (B_xG_y).

Example:	Average mass of boys	72 kg (160 lb.)
	Average mass of girls	54 kg (120 lb.)
	Percent boys in class	66.6%
	Percent girls in class	33.3%

   Assume 100 kg of the class. The class would be 66.7% x 100 kg = 66.7 kg boy (B) and 33.3% x 100 kg = 33 kg girl (G).

$$66.7 \text{ kg B} \times \frac{1\text{B}}{72 \text{ kg B}} = 0.93 \text{ B}$$

$$33.3 \text{ kg G} \times \frac{1 \text{ G}}{54 \text{ kg G}} = 0.62 \text{ G}$$

$$\frac{0.93 \text{ B}}{0.62 \text{ G}} = \frac{1.5 \text{ B}}{1.0 \text{ G}} \times \frac{2}{2} = \frac{3.0 \text{ B}}{2.0 \text{ G}}$$

   Formula for the class = B_3G_2

5. Analogies to illustrate the size of Avogadro's number

   a. One mole of high school textbooks would cover the USA to a depth of about 320 km (200 miles).

   b. Water flows over Niagara Falls at about 650,000 kL (172,500,000gallons) per minute. At this rate it would take 134,000 years for one mole of water drops (6.02×10^{23} drops) to flow over Niagara Falls.

   c. Avogadro's number (6.02×10^{23}) is the approximate number of milliliters of water in the Pacific Ocean (7×10^8 km^3 or 7×10^{23} mL). (Analogy by M. Dale Alexander, Gordon J. Ewing, and Floyd T. Abbott, 1984.)

   d. If one mole of particles the size of sand grains were released by the eruption of Mount St. Helens, they would cover the entire state of Washington to the depth of a ten-story building. (Analogy attributed to Glenn Crosby.)

   e. Analogies by A. Truman Schwartz and Lucy T. Pryde:

      One mole of marshmallows would cover the USA to a depth of 105,000 km (6500 miles). *NOTE: The volume of a marshmallow is estimated as 16* cm^3 *(1.0* in^3*). The area of the USA is* 9.32×10^6 km^2 *or* 3.6×10^6 mi^2.

      If an Avogadro number of pennies were distributed evenly among the 4.9×10^9 human inhabitants of earth, each man, woman, and child would have enough money to spend a million dollars every hour—day and night—and still have over half of it unspent at death.

      One guacamole is the amount of taco chip dip that can be made from an Avogadro number of avocados—plus appropriate quantities of tomatoes, onions, and chili. A train stretching to the North Star and back 2-1/2 times would be required to transport one guacamole. *NOTE: This assumes that the volume of one standard avocado (pit removed) is 278* cm^3 *and that other ingredients make up 25% of total volume. The average coal car has a capacity of 110 kL (4000* ft^3*) and is 16 m (53 ft) long. The North Star is 680 light years distant.*

Suppose the Greek god Zeus, after observing the Big Bang 15 billion years ago, became bored and decided to count one mole of atoms. Zeus is omnipotent. He can count very fast (one million atoms per second) and, of course, never sleeps. He has currently completed over three-fourths of the task, and will be finished in just another four billion years.

One mole of moles (animal-type), placed head to tail, would stretch 11 million light years and weigh 9/10 as much as the moon. *NOTE: Each mole is assumed to be 17 cm long with a mass of 100 g. Speed of light = 3.0×10^8 m/s. Mass of moon = 6.7×10^{22} kg.*

One mole of marbles, each 2 cm in diameter, would form a mountain 116times higher than Mount Everest. The base of the marble mountain would be slightly larger than the area of the USA. *NOTE: Marbles are assumed to have hexagonal closest packing and the mountain has a cone of angle 30°. Area of USA = 9.32×10^6 km^2. Volume of cone = $\pi r^2 h/3$, h = height = $r \tan 30°$.*

TIPS FOR THE TEACHER

Language of Chemistry

1. The term "formula unit" is used in this module to refer to the smallest unit of a substance of a defined formula. For example, one formula unit of CO_3^{2-} refers to one CO_3^{2-} ion. One formula unit of Ne would be one Ne atom. This terminology is especially useful when referring to ionic compounds because "molecules" of solid ionic compounds do not exist. One formula unit of NaCl, consists of one Na^+ ion and one Cl^- ion. Likewise, a formula unit of $CuSO_4 \cdot 5H_2O$ contains one Cu^{2+} ion, one SO_4^{2-} ion, and 5 H_2O molecules. Because the molar mass of any substance is defined on the basis of its formula, the use of the term "formula unit" is not necessary. It is used here to avoid referring to atoms, molecules, ions, *etc.* each time we refer to moles.
2. Due to the SI definition of mole, there is no official basis for using the terms "gram moles," "pound moles," "ton moles," *etc.* even though these are sometimes still encountered. The molar mass is the mass of one mole of any species expressed in grams.
3. "Formula weight," "gram formula weight," "molecular weight," "gram molecular weight," and "atomic weight" are terms that refer to moles of specific categories of substances. These are used synonymously with "molar mass," which remains the preferred term.
4. The meaning of formulas in terms of moles (rather than atoms) should be stressed so students begin to think "moles" rather than "atoms," "molecules" or "formula units" when they see a formula or interpret its meaning. This will greatly facilitate calculations encountered later in this module (percent composition and empirical formula) and stoichiometry problems, which follow soon after this topic in most texts.
5. The empirical formula is the simplest formula for a substance. In the empirical formula the subscripts are reduced to the smallest possible integers. The formulas for many ionic substances are empirical formulas; however, molecules of covalent substances are frequently a multiple of the empirical formula. The formula for a molecule is called the molecular formula. The empirical formula for a covalent compound is found by dividing all the subscripts by the largest common factor (*e.g.*, the empirical formula for $C_6H_{12}O_6$ is CH_2O).

Common Student Misconceptions

1. **"Equal volumes or equal masses of different substances contain the same number of moles (or atoms/molecules/formula units)."**

   It *is* true that equal volumes of gases contain the same number of moles of gas molecules (assuming equal temperatures and pressures), but such an assertion is wide of the mark for liquids or solids. No such generalizations are possible for the masses of substances—whether in gaseous state or not.

2. **"The term 'mole' refers only to 'molecules'."**

   This is particularly problematic if the mole concept is only applied to molecules (or formula units). It is advisable to apply the concept of moles also to ions, electrons, and/or other entities so that students will not associate "mole" exclusively with the term "molecule."

3. **"One mole of various substances may contain different numbers of units of those substances."**

   Although Avogadro's number (6.02×10^{23}) is such a large number that it cannot be meaningfully visualized, it is important that students conceive of one mole of a substance as containing the *same number of units* as one mole of any other substance. However, you must work hard to destroy the idea that equal masses or equal volumes of two substances contain the same number of moles, as was highlighted in Misconception 1.

Problem Solving

1. Problem solving is essential to understanding the mole concept. It is important to illustrate a variety of permutations (moles to grams, volume of gas at STP to mass, molecules to moles, *etc.*) when solving mole problems. Practice should be given, initially with your assistance and later as independent assignments. After students master mole-mass and mass-mole conversions, proceed to percent composition problems. When percent composition problems are understood, the concept of empirical *vs.* molecular formula can be introduced and empirical formula problems can be solved. The determination of molecular formulas from empirical formulas and molar masses can be introduced either before or after problems dealing with calculating empirical formula.

2. In demonstrating how to solve mole-mass-entity problems, emphasize the relationship of mass and number of entities (molecules/atoms/formula units/ions, *etc.*) to moles. Suggest that students devise an equality that relates mole/ mass/molecules for the substance in the question. For example:

   $$1 \text{ mol HCl} = 36.5 \text{ g HCl} = 6.02 \times 10^{23} \text{ molecules HCl}$$

   The mass (grams) of the substance can be converted to moles by using this relationship to construct a "unit conversion factor" and the unit-label method. Since *mole* is the unifying idea, if the number of molecules is desired and mass is given, mass should be converted to moles and then moles to molecules, using two factors (see transparency master).

   Example: How many molecules are in 146 g HCl?

$$146 \text{ g} \times \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} \times \frac{6.02 \times 10^{23} \text{ molecules HCl}}{1 \text{ mol HCl}} = 2.41 \times 10^{24} \text{ molecules HCl}$$

3. If you prefer an algebraic approach, a generalized form of the relationship among mass, moles, and entities is:

$$\frac{\text{g X}}{\text{molar mass X}} = \text{mol X} = \frac{\text{number of X's}}{N_A}$$

X = formula; N_A = Avogadro's number

4. In percent composition problems, it is useful to illustrate the calculation using a compound having more than one atom of some element (*e.g.*, K_2O) to illustrate that the total mass (2 x molar mass K atoms, not molar mass of K atoms) is used to calculate percent composition. It is a very common error to use the mass of 1 mol K (rather than 2 mol K) in calculating the percent K in K_2O. The resulting mass percent potassium will be incorrect (too low by a factor of two). Remind students that the sum of the percents should be 100 (or close to 100).

5. One procedure for solving percent composition problems can be outlined using the mole concept:

$$\frac{\text{Total moles of element} \times \text{Formula mass of element}}{\text{1 mol compound} \times \text{Formula mass compound}} \times 100 = \text{\% of element in compound}$$

6. When illustrating empirical formula problems, start with a binary compound containing subscripts that are simple multiples of each other (*e.g.*, CO_2, PCl_3, CH_4). Then proceed to a binary compound where the mole relationship is more complex (2 to 3 or 3 to 4), requiring the additional step of multiplying by an integer to convert relative numbers of moles to integers. In solving empirical formula problems emphasize that subscripts give the relative numbers of moles of elements in the compound; if we know the relationship of the moles of the elements we know the subscripts. This type of problem is difficult for most students. Provide needed practice, both supervised and independent.

7. If students set up empirical formula problems "upside down" (*i.e.*, divide atomic mass by mass of the element rather than calculate moles correctly), their subscripts will be opposite to what they should be (*i.e.*, they will get the correct subscripts but for the wrong elements). Watch for this and reemphasize that students calculate moles of the elements. Of course, the units will be incorrect if students set up the problem this way, but that does not seem to bother many students. This would be a good place to emphasize how to use units to check that the problem solution is reasonable.

   The procedure for solving empirical formula problems can be outlined using the mole concept:

   [Percent composition] → Masses of elements →
   Ratio of moles of elements → Empirical formula for compound

HISTORY: ON THE HUMAN SIDE

1. The mole concept is attributed to Amedeo Avogadro, who postulated that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules. He based this hypothesis on Gay-Lussac's Law of Combining Volumes (when two gases react, their volumes are in whole number ratios) and Dalton's Atomic Theory (substances combine in integral ratios of atoms). Avogadro proposed that gases exist as molecules that can be composed of more than one atom of an element and that these molecules can break up during reaction. Avogadro, however, never determined what we know today as Avogadro's number.

2. Because Dalton and other influential scientists of his day did not accept his hypothesis, Avogadro’s idea was not accepted for many years. After his death, Avogadro’s hypothesis was “rediscovered” and accepted. When the number of molecules in a mole was estimated by Perrin during the early years of this century, he suggested that the number be called Avogadro’s number. Since that time, Avogadro’s number has been measured by several different methods, including a measurement by Rutherford using alpha particles and the measurement by Millikan in his famous oil-drop experiment in which he determined the charge on an electron. More recently, Avogadro’s number has been measured by X-ray crystallographic measurements.

HUMOR: ON THE FUN SIDE

1. A poem by Doris Kolb (from *J. Chem. Educ., 1978*):

A mole is a particular amount
Of a substance–just its formulary weight
Expressed in grams, with Avogadro’s count
Of units making up the aggregate.
A mole is a specific quantity:
Its volume measures twenty-two point four
In liters (for a gas at STP).
A mole’s a counting unit, nothing more.
A mole is but a single molecule
By Avogadro’s number multiplied;
One entity, extremely miniscule,
A trillion trillion times intensified.
A mole is an expedient amount,
For molecules are just too small to count.

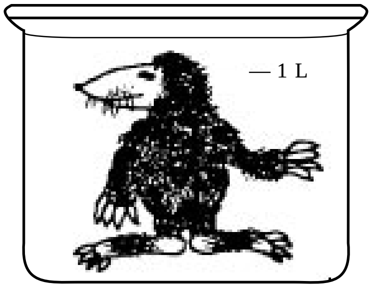
2. Mole Song (to “Deck the Halls”)

Avogadro’s Number makes one
mole of atoms or of molecules.
Changing it to grams is easy
If you know what things to use.
Multiply it by a fraction
Molar mass over one mole.
Then you calculate your answer
Canceling like terms, you’ll reach your goal.

3. A mole stuffed animal can be obtained from American Bear, Inc., 401 N. Wabash, Suite 500, Chicago, IL 60611 (The Hole in the Ground Gang Mole #1407. Phone number: 312-329-0020.)
4. Mole posters and shirts available from Mole Reproduction, 315 Lincoln Street, Verona, WI 53593.
5. Classified ad:

FOR MOLE CONTROL
CALL
∞ 602 – 1023 ∞
A. vAogadr

6. A one-molar solution:



7. Message on a T-shirt: A Chemist is a Mole's Best Friend.
8. Sign on a bumper sticker: Everyone has...Avogadro's Number—6.023 x 10^{23}.
9. **THE MOLE**

You say a mole is brown and furry
with eyes developed wrong,
You ask a chemist his description
which is quite a different song.
A chemists mole is not of fur,
nor is it very brown,
It comes in many varieties—square, oblong,
or round.

The mole can stand for numbers, formulas
or items of everyday,
A mole of girls, a mole of boys, whatever
you wish to say.
Chemists indulge jubilantly,
solving problems it creates,
My brains find problems no fun at all
pondering them till it aches.

CHEM 13 NEWS, February 1972, p. 393

10. The Mole Joke Book (CHEM 13 NEWS, February 1983, p. 12 and October 1983, p. 3):

QUESTIONS

1. What did Avogadro give his ex-wife every month?
2. What did Avogadro have on his pancakes?
3. On which American mountain was Avogadro's face carved in stone?
4. What song did Avogadro's family sing on New Year's Eve?
5. Where did Avogadro send his CARE packages?
6. Why was there only one Avogadro?
7. What was Avogadro's favorite board game?
8. What kept Avogadro in bed for two months?
9. Who was Avogadro's favorite singing group?
10. Who was Avogadro's favorite composer?
11. Avogadro loved to watch M.A.S.H. Which character did he like the most?
12. What did Avogadro call his church services?
13. How would you have described Avogadro's room while he was a teenager?
14. Which was Avogadro's favorite Indian tribe?
15. Which recent movie would Avogadro have really liked?
16. What did Avogadro think of hemorrhoids?
17. What did Avogadro put into the pockets of his tweed suit?
18. Can you name two movies that Avogadro really liked?
19. What was Avogadro's best day in golf?
20. How did Avogadro send a secret message with his walkie-talkie?

21. What is Avogadro's favorite kind of music?
22. Which Walt Disney characters was Avogadro fond of?
23. What is Avogadro teaching his astronomy class about?
24. What do you get when you have a bunch of moles acting like idiots?
25. How did Avogadro help his team win the soccer playoffs?
26. Which is one of Avogadro's best songs?
27. What did Avogadro invent for his wife to use as a night cream?
28. What was Avogadro's favorite drink?
29. Why did people say Avogadro was lazy?
30. What is Avogadro's favorite chocolate bar?
31. What did Avogadro get when he mixed ice cream, chocolate syrup, and milk?
32. What kind of shell-fish did Avogadro like to eat?
33. Which tooth did Avogadro have pulled out?
34. What religion did Avogadro belong to?
35. What did Avogadro's bird do when it was time for him to send his feathers?
36. What was the mad bomber's favorite drink?
37. Where did Mrs. Avogadro do her shopping?
38. What did Avogadro teach his students in math class?
39. What are moles made of?
40. Which team lost the World Series in 1982?
41. What brand of cigarettes did Avogadro smoke?
42. What are mammoles?
43. What do you sue to flatten hot asphalt?
44. How did Avogadro get through the desert?
45. What do chemists do in a math class?
46. Which part of the universe did Avogadro like to study?
47. What kind of fruit did Avogadro eat in the summer?
48. How much money did Avogadro make from being a chemist?
49. What is Avogadro's favorite sport?
50. What is Avogadro's favorite drink?
51. What happened to Avogadro when he had to drive over a bridge?
52. What area did Avogadro explore?
53. What were Avogadro's houses made of?
54. How much does Avogadro exaggerate?
55. What did J. J. Thomson have for dessert?
56. What happens to Avogadro's dog in the summer?
57. What did Avogadro do when he lost his job?
58. Where does Avogadro plant his trees?
59. What was Avogadro's best subject?
60. What was Avogadro dressed as when he went to the masquerade ball?
61. Where did Avogadro go sailing?
62. What was the problem Avogadro had with his shoes?
63. At what time was Avogadro at his romantic best?
64. What are Avogadro's favorite places?
65. What scientist was a member of the pig family?
66. What is Avogadro's favorite character?
67. What was Avogadro full of?
68. What is Avogadro's favorite song?
69. What did Avogadro collect at the seashore?
70. Where do students graduate?

ANSWERS to the Mole Joke Book

1. Alimoley
2. Moleasses
3. Mount Rushmole
4. “Mole Lang Syne”
5. Moleasia
6. When they made him, they broke the moled.
7. Moleopoly
8. Moleonucleosis
9. The Moleing Stones
10. Molezart
11. Father Molecahy
12. Molar mass
13. A Molar mess
14. The Molehawks
15. “On Moleton Pond”
16. They’re a pain in the mass.
17. Molth balls
18. “Mole Miner’s Daughter” and “Moleby Dick”
19. When he got a mole-in-one.
20. He used the Molse Code.
21. Rock + Mole
22. Mickey and Minnie Mole
23. The molar system
24. A bunch of moleasses.
25. He scored the winning mole.
26. “There’s a mole in my bucket, dear Liza, dear Liza...”
27. Oil of Molay
28. Moleson Export Ale
29. They said he was as slow as moleasses.
30. Molted milk
31. A chocolate molted
32. Mole-uscs
33. One of his molars
34. Molslem
35. It moleted
36. Moletov cocktails
37. In a shopping mole
38. Moletiplication
39. Molecules
40. The Molewaukee brewers
41. DuMolier
42. Four-legged animoles
43. A steam-moler
44. On a camole
45. Moletiply
46. The Moleky Way
47. Watermolens
48. Enough to make him into a moleionnaire.
49. Lawn moles
50. Coca molar
51. He had to pay at the mole gate.
52. The South Mole
53. Moleskin
54. He makes mountains out of mole hills.
55. Plum pudding
56. It molets.
57. He joined the mole queue.
58. Moles in the ground.
59. Mass
60. An avocado
61. The Moleucca Strait
62. He wore the mole out too quickly.
63. On Mole-lit nights
64. Moland, Massachusetts
65. Niels Bohr
66. Masswell Smart
67. Molar energy or 6 x 10^{23} particles
68. Molar Kintyre
69. Mole-uscs
70. At moleage

11. In the U.S., Mole Day is celebrated on Oct. 23 (10/23) from 6:02 a.m. to 6:02 p.m. In Europe, Mole Day is celebrated on Feb. 6 (6/02) from 10:23 a.m. to 10:23 p.m. (In the U.S., dates are written month/date; in Europe, date/month.)

12. Word Search (see *Appendix* for master copy)

~~A V O G A D R O S N U M B E R~~ O T
G T F V C M C L S S D S A Y K K S
I V I O B D D X S M R L J J V X T
A V I O W S Y V A R U H U E N E N
O U M D S V N O M M W B D F W J E
Y ~~G R A M~~ O H W R J U T S A M B I
S Q W O B X X O A V X T D A G O C
T V N R R T F T L Z B P F D G L I
E D A ~~Y R T E M O I H C I O T S~~ F
D C F Z B G T I M S N W X Y R D F
M C ~~M O L E C U L E S~~ A X A L P E
S T W X T I Z Y L U ~~E L O M~~ J K O
D X R I I O O C D R S F W U G K C

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. A ____ of a substance represents Avogadro's number of units of that substance.
2. 6.02 x 10^{23} (2 words).
3. Mass, in grams, of 1 mole of a substance (2 words).
4. SI unit of mass ÷ 1000
5. __________ describes the quantitative relationships among elements and compounds as they undergo chemical changes.
6. Symbolism showing a substance's chemical composition.
7. Smallest units of compounds.
8. Numerical multipliers showing ratios between reactants and products in a chemical equation.
9. One mole of a substance in this state at STP occupies a volume of 22.4 L.
10. One mole of this element weighs 12.0 g.

Answers: 1. MOLE 2. AVOGADRO'S NUMBER 3. MOLAR MASS 4. GRAM 5. STOICHIOMETRY 6. FORMULA 7. MOLECULES 8. COEFFICIENTS 9. GAS 10. CARBON

13. See cartoons at end of module.

MEDIA

1. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
    a. For the Apple II computer running on ProDOS: AR 301, AR 302, AR 305
    b. For the Apple II computer: AP 304, AP 305, AP 306, AP 307
    c. For IBM PCs and PC-compatibles: PC 2201, PC 2301
2. "The Atom" videotape from the *Ring of Truth* series by Philip Morrison includes about a 3-min segment on the mole idea (reaction of different volume ratios of gases). Videotapes are available through PBS (800-424-7963). (To appreciate the entire tape, students probably need a broader understanding of chemistry than is held by the average student at the point in which development of the mole concept generally occurs in the high school chemistry course.) Essentially the entire text of the tape and all important visual images are also included in the book *The Ring of Truth: An Inquiry Into How We Know What We Know* by Philip and Phyllis Morrison (1987). PBS address (Ring of Truth):

    PBS Video
    1320 Braddock Place
    Alexandria, VA 22314
    (703) 739-5000

3. The video *Powers of Ten* goes from galaxy size down to nuclear size. It may help students visualize the size of atoms and the comparison of Avogadro's number in terms of powers of ten (Morrison, 1982). Video available from Video Catalog, P.O. Box 64428, St. Paul, MN 55164-0428 or High Bridge Co., 1000 Westgate Drive, St. Paul, MN 55114-1631.
4. CHEM Study Video: Vol. 12, *Gases and How They Combine*. (1989). Illustrates Avogadro's Hypothesis (equal volumes of gases contain equal numbers of molecules). Available from Ward's Natural Science Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9978.
5. Videodiscs published by *JCE:* Software, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

    "A Generic Spectroscopic Instrument" and "The Mole – From Counting to Weighing," two chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.

6. The *World of Chemistry* videotape "Number 2: The Mole." World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.

EQUIPMENT AND INSTRUMENTATION

A simple electrolysis apparatus suitable for the electrolysis of water to illustrate the relation of the mole to gas volume is available from laboratory equipment vendors.

Links/Connections

WITHIN CHEMISTRY

1. Since the mole is the central concept in quantitative chemistry, it is related to almost all topics supported in high school chemistry. Here are some examples:
   a. **Chemical Reactions and Stoichiometry** The law of conservation of mass requires that the same numbers of atoms of all species be present before and after reaction. Because the small size of atoms makes it impractical to count atoms, chemists use moles of atoms. In balancing chemical equations, the coefficients are related to the moles of substances present. The relationship of moles of reactants and products in balanced equations is central in determining mass relationships in chemical reactions (stoichiometry).

   b. **Solutions** Chemists generally express solution concentration in terms of molarity (moles of solute per liter of solution) or mole fraction (moles of solute per total moles in solution). Because chemical reactions are based on moles, molarity is convenient in determining the quantities of solutions needed for chemical reactions. Colligative properties of solutions (freezing point depression, boiling point elevation, osmotic pressure, vapor pressure lowering) are related to the moles of solute particles in solution.

   c. **Rates of Reaction** Rates of reaction are often expressed in terms of moles of reactant per liter (mol/L) disappearing per unit time (or some quantity related to molarity, such as pressure of gas, absorbance/transmittance of light, *etc.*).

   d. **Equilibrium** Since coefficients in a balanced equation relate moles of reactants and products, quantities of substances present at equilibrium are usually expressed in terms of molarity (moles/liter) for reactions in solution or pressure (which is proportional to moles per liter) for gases.

   e. **Acids and Bases** A monoprotic acid is a compound that can furnish one mole of hydrogen (hydronium) ions per mole of compound. A diprotic acid can furnish two moles of hydrogen ions per mole of compound; a triprotic acid three. When soluble acids and bases react, the net ionic equation involves one mole of hydrogen (hydronium) ion and one mole of hydroxide ion. The pH of a solution is related to the molarity (moles/liter) of hydrogen (hydronium) ion present in solution.

   f. **Electrochemistry and Oxidation-Reduction** In oxidation-reduction reactions, transfer of electrons accompanies chemical change. Electrochemical reactions can be described by indicating the number of moles of electrons (Faradays) transferred per mole of reactant species oxidized or reduced. In electrolysis reactions, current, time, and amount of product are related to the moles of electrons involved.

   g. **Gases**. The amount (moles) of a gas sample as well as its temperature and pressure determine its volume. (The ideal gas law relates these four quantities.) Since the volume of a gas at constant temperature and pressure is proportional to the amount of gaseous substance, the relative volumes of reactant and product gases provide a convenient method for determining or monitoring the stoichiometry of gaseous reactions.

h. **Thermochemistry** In a thermochemical equation, the physical states of all reactants and products as well as the mole relationships among reactants and products determine the quantity of energy involved. The thermal energy involved in a reaction is proportional to the amounts (moles) of reactants used and products formed.

i. **Condensed States** The molar heat of fusion (thermal energy required to melt one mole of substance), the molar heat of vaporization (thermal energy required to evaporate one mole of substance), and the molar heat capacity (thermal energy required to raise the temperature of one mole of substance by one degree) are all commonly used quantities.

2. When chemists prepare new compounds (particularly organic compounds), one way the compound can be characterized is from the elemental analysis. A sample of the compound is sent to a laboratory that provides the mass percent of each element (*e.g.*, % C, % H, % O, % N, *etc.*). From this information, the chemist can derive the empirical formula for the compound. Generally, spectral data (infrared, ultraviolet, nuclear magnetic resonance, mass spectrum, *etc.*) are used to predict the structure and to deduce the arrangement of the atoms in the molecule. From spectral data and knowledge of chemistry (prediction of the reaction products based on analogous reactions with known products), the chemist proposes a formula and structure for the compound. If the empirical formula of the proposed compound and the empirical formula predicted from elemental analysis agree, the proposed structure is possible. If they do not agree, the proposed structure is incorrect and additional studies must be made to identify the compound.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Since the mole is central to chemistry, related areas such as medicine, biochemistry, biology, geology, physics, agronomy, and environmental science use the mole concept with reference to chemical processes important in those disciplines. A few examples are given below.

1. The mole (particularly millimoles and molarity) is a commonly used unit in medicine. For example, concentrations of intravenous solutions and drugs are expressed in terms of moles.
2. For the complete oxidation of glucose using the glycolytic pathway and the Krebs cycle, 38 moles of ATP are produced per mole of glucose metabolized.

   $$C_6H_{12}O_6 + 36ADP + 36H_3PO_4 + 6O_2 \rightarrow 6CO_2 + 42H_2O + 38ATP$$

   The amount of ATP formed in a biochemical reaction is related to the energy since the hydrolysis of each mole of ATP corresponds to 31 kJ of free energy. For the metabolism of glucose in the Krebs cycle, energy production efficiency is roughly 40% (approximately 1.2×10^3 kJ/mol glucose in the Krebs cycle compared to approximately 2.9×10^3 kJ/mol glucose for combustion of glucose to CO_2 and H_2O). This is higher efficiency than in the automobile engine (about 20%) and most other heat engines (maximum efficiency about 35%).
3. Formation of stalactites and stalagmites involves the dissolution and reprecipitation of calcium carbonate. Dissolved CO_2 causes limestone ($CaCO_3$) to dissolve according to the reaction $CaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ca^{2+}(aq) + 2HCO_3^{-}(aq)$. The reverse reaction results in reprecipitation of $CaCO_3$. The concentration (moles/liter) of dissolved CO_2 in water seeping through limestone caves is one factor responsible for formation of stalactites and stalagmites.

4. One source of “fixed” nitrogen (nitrogen that can be used by plants) is the combination of nitrogen and oxygen during highly energetic conditions such as during a lightning storm. The reaction is $N_2(g) + O_2(g) \rightarrow 2\,NO(g)$. The NO is further oxidized to NO_2, which dissolves in water to form nitric and nitrous acids.

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

$$2NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + HNO_2(aq)$$

The quantities of nitrogen “fixed” is related to the moles of nitrogen reacting with oxygen.

TO THE CONTEMPORARY WORLD

Community

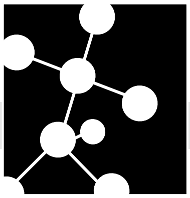
Knowledgeable Individuals

A pharmacist or clinical laboratory scientist (medical technologist) could explain how the mole concept is used in medicine.

Societal

1. Commercially available materials sometimes list the percent composition, particularly with respect to important or active components in the formulation. For example, the label on cleaning agents (such as dishwasher detergent) may include the percent phosphorus (by mass).
2. The composition of soil fertilizers is generally expressed in terms of % P_2O_5 (or H_3PO_4), % N (or NH_3) and % K_2O (see module on *Industrial Inorganic Chemistry*).
3. The relationship of moles to chemical reactions opens an inexhaustible supply of applications such as combustion of gasoline in automobile engines, reaction of SO_2 (from the burning of sulfur-containing fossil fuels) with $CaCO_3$ (marble, limestone, *etc.* in statuary and building materials), as well as the production of “acid rain” from SO_2. (For additional examples see module on *Stoichiometry*.)

Extensions



1. The percent water in popcorn can be determined. *Activity 2: Analyzing the "Pop" in Popcorn* in the *Gases* module provides directions.
2. A procedure for determining the percent fat and water in hot dogs is given by Tocci (1984).
3. Students could investigate methods used for determining the elemental composition of various compounds. The use of spectra in compound identification is usually beyond the scope of first-year high school chemistry; however, an outstanding student may be interested in learning what kinds of information can be obtained from various spectral techniques (*e.g.*, how atoms are connected from infrared and nuclear magnetic resonance spectra, how atoms are grouped based on how they "fall apart" from mass spectra).

References

Module developed by Caroline L. Ayers, Philip Dail, and Carolyn Morse, the North Carolina team.

Alexander, M. D., Ewing, G. J., and Abbott, F. T. (1984). Analogies that indicate the size of atoms and molecules and the magnitude of Avogadro's number. *Journal of Chemical Education*, *61*, 591.

Corwin, C. (1988). Avogadro's number. In W. S. Seese and G. W. Daub (Eds.), *Laboratory experiments for basic chemistry* (5th Ed.) (pp. 111 - 115). Englewood Cliffs, NJ: Prentice Hall.

This is one example of a laboratory activity to estimate Avogadro's number.

Dickson, T. R. (1983). Empirical formula. In *Laboratory experiments for introduction to chemistry* (4th Ed.) (pp 47-56). New York, NY: Wiley.

This experiment is typical of those illustrating the determination of the empirical formula of MgO.

Knox, K. (1980). Empirical formulas—a ratio problem. *Journal of Chemical Education*, *57*, 879.

This article describes an exercise in using the female-male composition of a class to determine an empirical formula.

Kolb, D. (1978). The mole. *Journal of Chemical Education, 55,* 728-732.

This article is a review of the mole concept. It also describes methods for measuring Avogadro's number.

Morrison, P., and Morrison, P. (1987). *The ring of truth: An inquiry into how we know what we know*. New York, NY: Random House.

This book contains pictures from the videotape "The Atom" from the *Ring of Truth* series by Philip Morrison.

Morrison, P., Morrison, P., and the Office of Charles and Ray Eames. (1982). *Powers of ten.* Redding, CT: Scientific American Library.

This book illustrates orders of magnitude from 10^{24} m (larger than our galaxy) to the size of nuclei. It contains pictures taken from the film *Powers of Ten: A Film Dealing with the Relative Size of Things in the Universe and the Effect of Adding Another Zero* produced in the office of Charles and Ray Eames for IBM in 1977.

Schwartz, A. T., and Pryde, L. T. *Mole poster*. Available from Education Division, American Chemical Society, 1155 16th Street, N.W., Washington, DC 20036.

Many mole analogies presented in this module are taken from this poster.

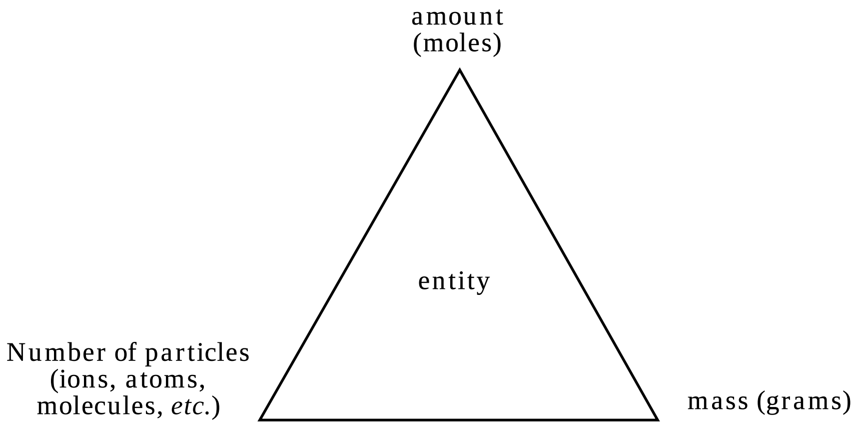
Slowinski, E. J., Wolsey, W. C., and Masterton, W. L. (1985). Water of hydration. In *Chemical principles in the laboratory* (4th Ed.) (pp. 37-42). New York, NY: Saunders.

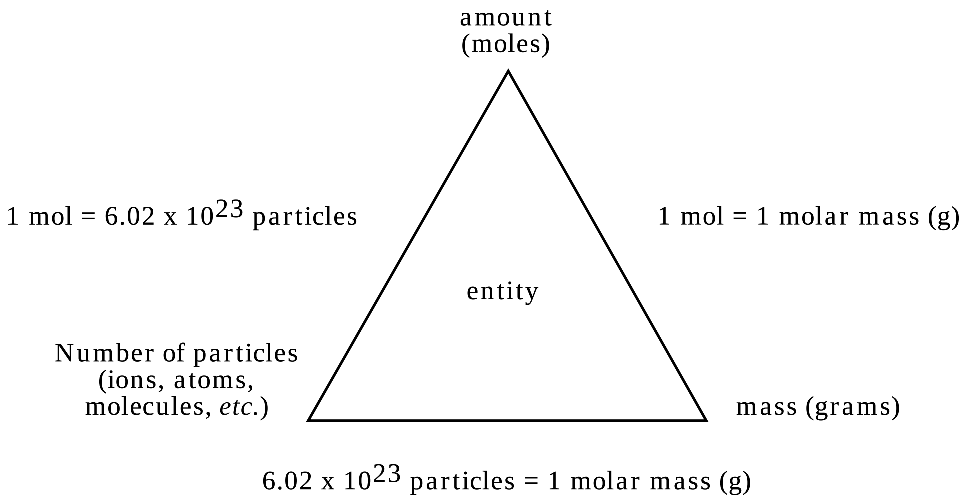
This is an example of a typical water of hydration experiment.

Tocci, S. (1984). How much meat in a hot dog? *The Science Teacher*, *51*(3), 62.

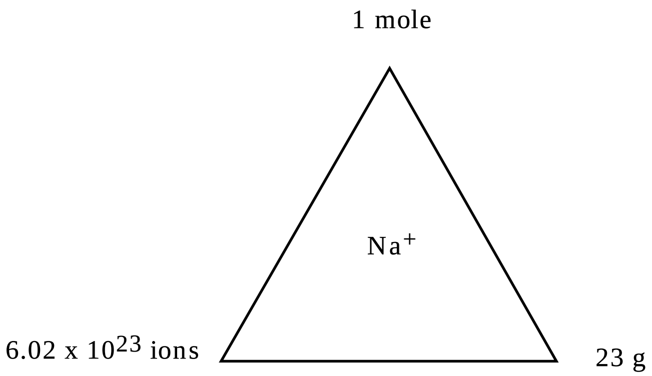
A brief description is given of a method for determining the percent fat and percent water in a hot dog.

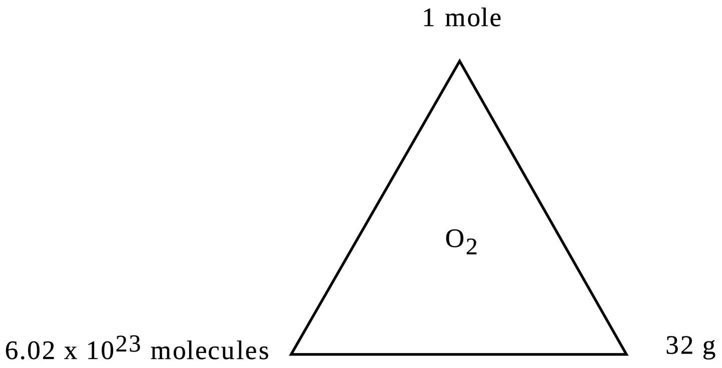
Triangular Relationship among Amount, Mass, and Number of Particles for an Entity

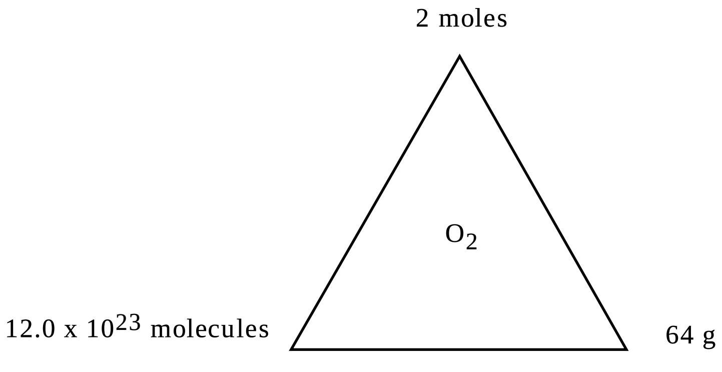




Examples of Triangular Relationship







Word Search

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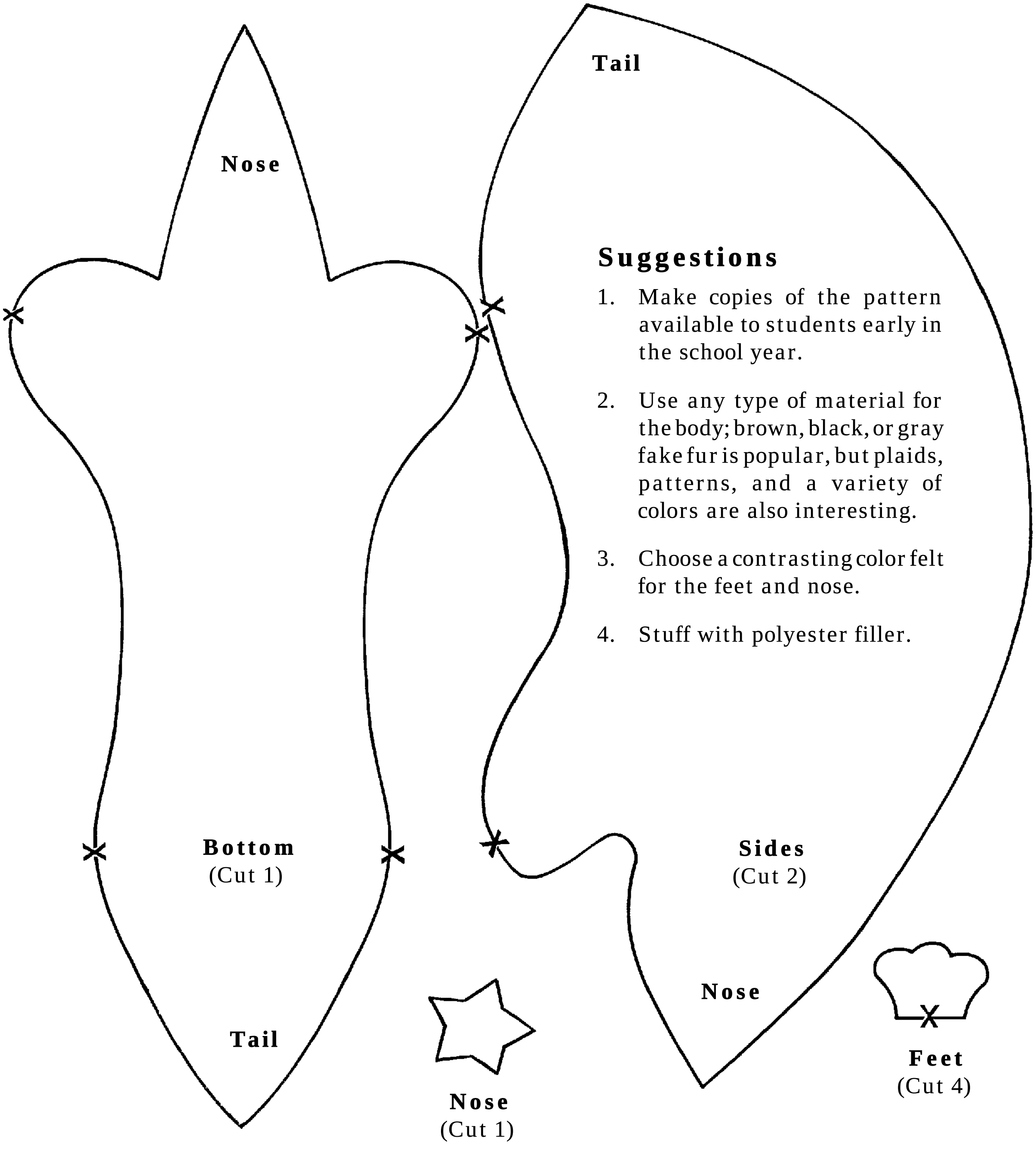
A V O G A D R O S N U M B E R O T
G T F V C M C L S S D S A Y K K S
I V I O B D D X S M R L J J V X T
A V I O W S Y V A R U H U E N E N
O U M D S V N O M M W B D F W J E
Y G R A M O H W R J U T S A M B I
S Q W O B X X O A V X T D A G O C
T V N R R T F T L Z B P F D G L I
E D A Y R T E M O I H C I O T S F
D C F Z B G T I M S N W X Y R D F
M C M O L E C U L E S A X A L P E
S T W X T I Z Y L U E L O M J K O
D X R I I O O C D R S F W U G K C
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. A ____ of a substance represents Avogadro's number of units of that substance.
2. 6.02 x 10^{23} (2 words).
3. Mass, in grams, of 1 mole of a substance (2 words).
4. SI unit of mass ÷ 1000
5. __________ describes the quantitative relationships among elements and compounds as they undergo chemical changes.
6. Symbolism showing a substance's chemical composition.
7. Smallest units of compounds.
8. Numerical multipliers showing ratios between reactants and products in a chemical equation.
9. One mole of a substance in this state at STP occupies a volume of 22.4 L.
10. One mole of this element weighs 12.0 g.

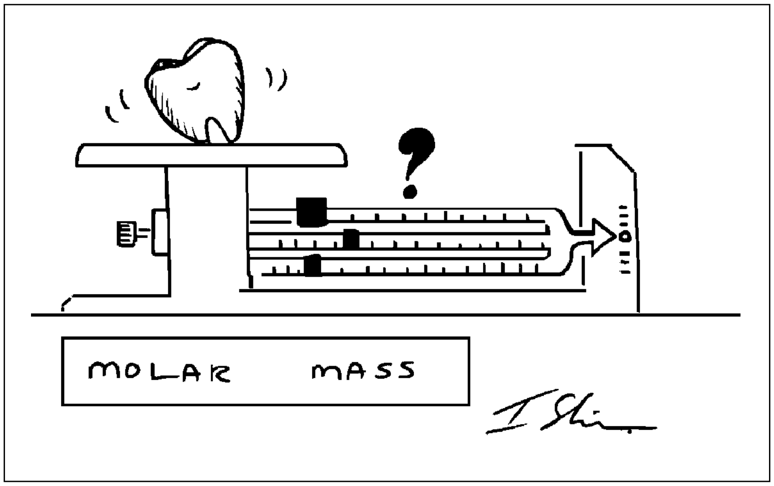
Mole Pattern

(Adapted by Ray Bruzan from a pattern obtained from Claire Smith [retired], Mills College, who discovered the original mole in an art store in Maine. Ready-made mole price list available from DeWayne Lieneman, 2509 59th St., Woodridge, IL 60517; (708) 810-0335.)



Suggestions

1. Make copies of the pattern available to students early in the school year.
2. Use any type of material for the body; brown, black, or gray fake fur is popular, but plaids, patterns, and a variety of colors are also interesting.
3. Choose a contrasting color felt for the feet and nose.
4. Stuff with polyester filler.



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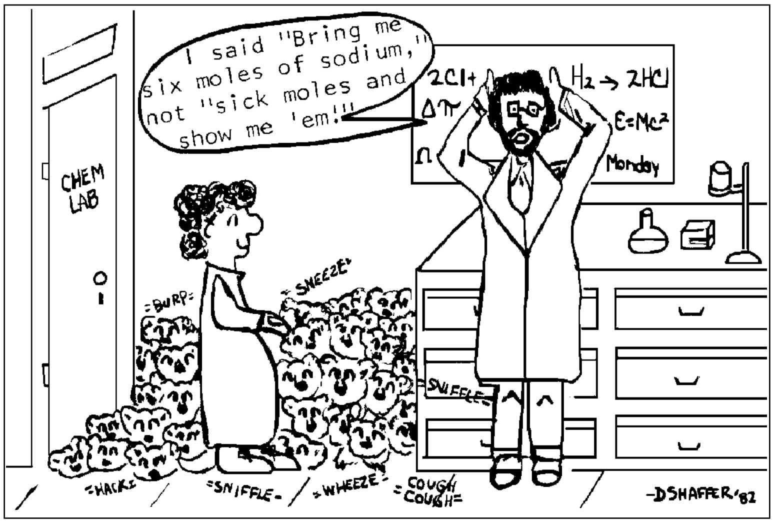

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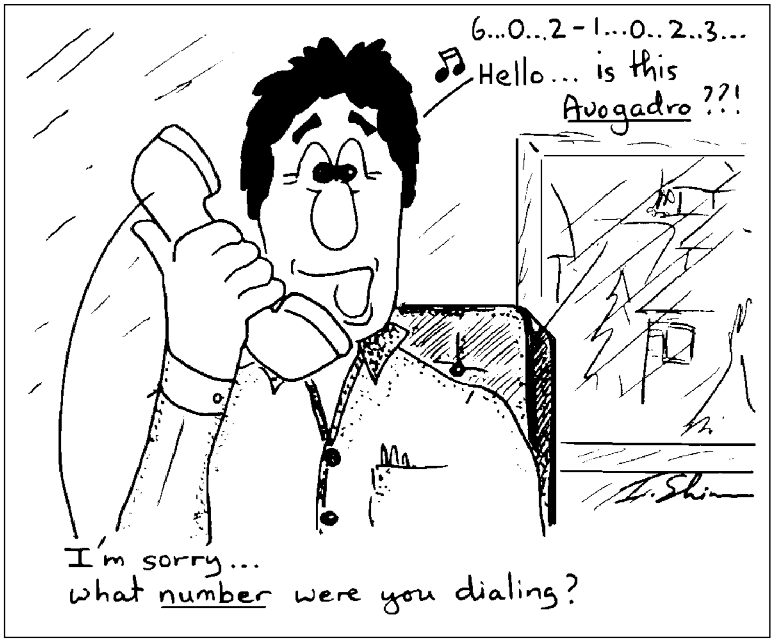

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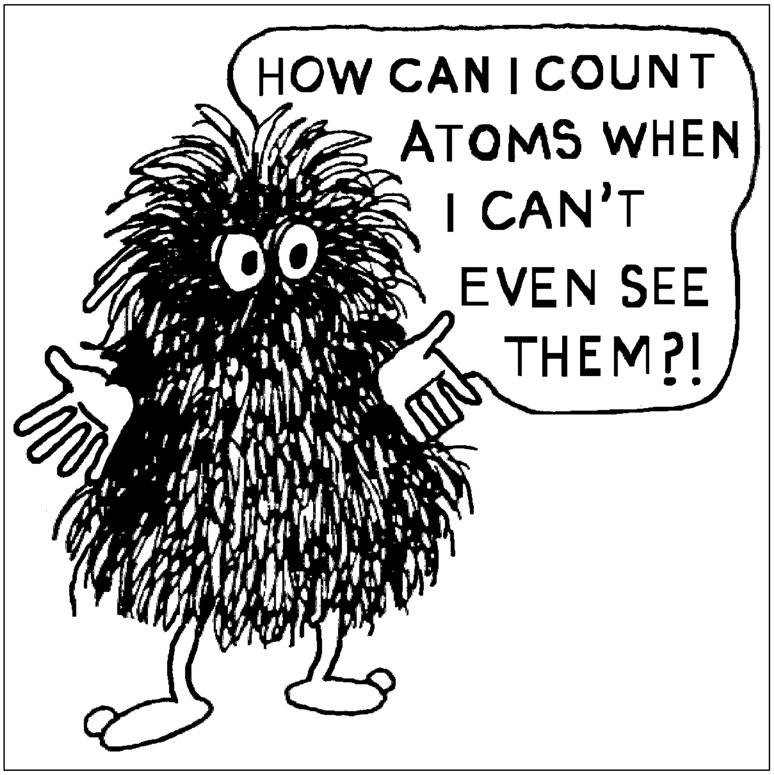

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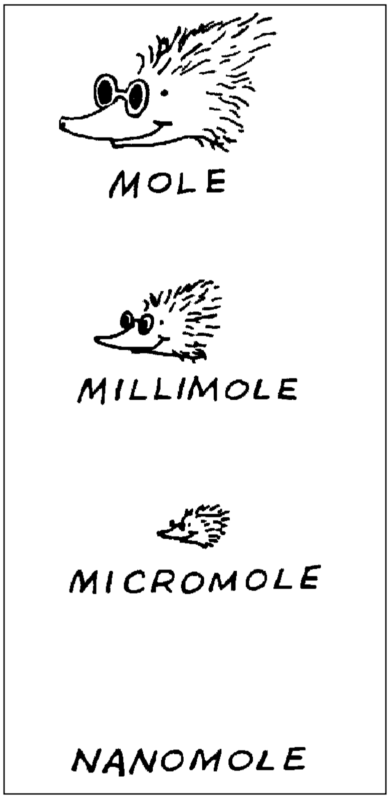

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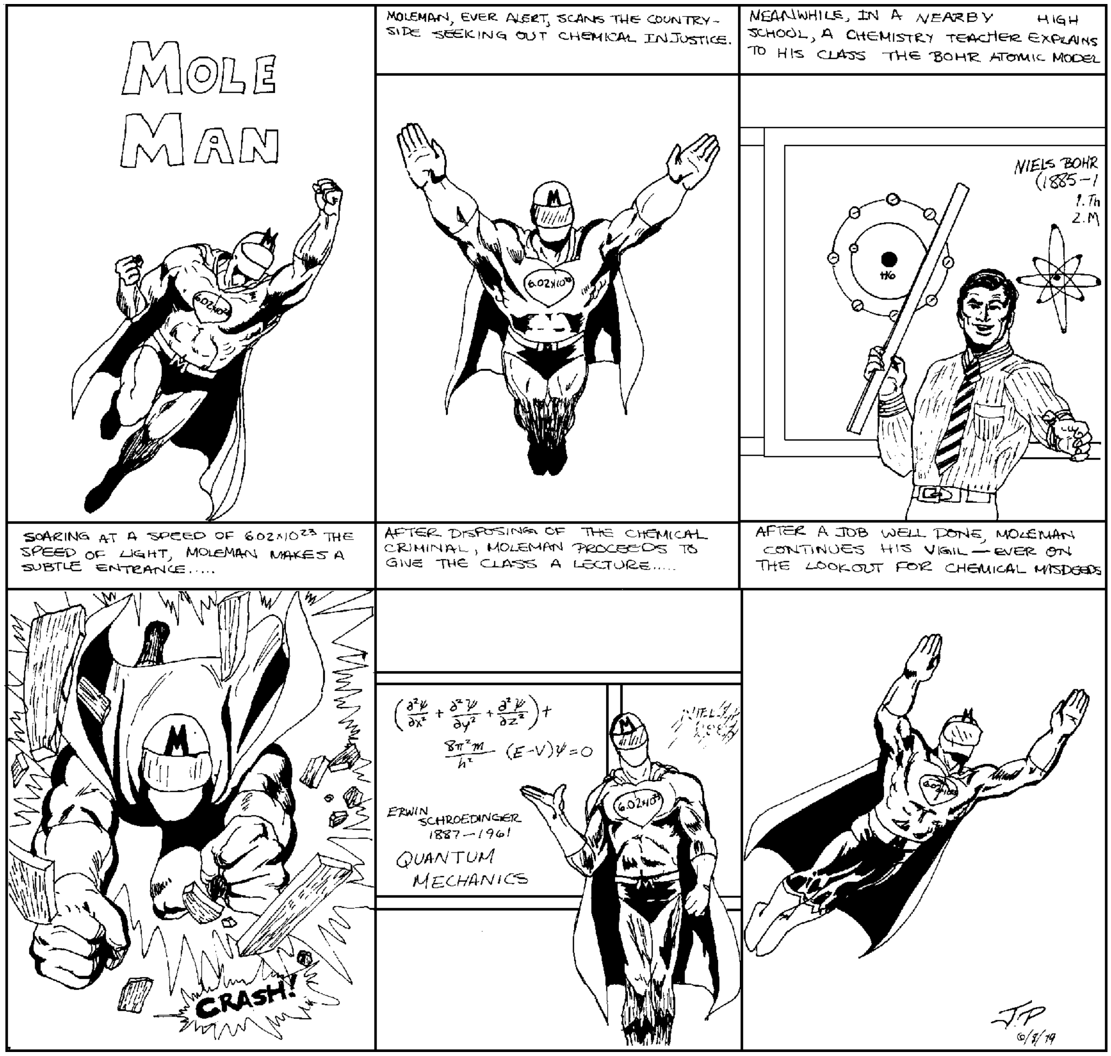

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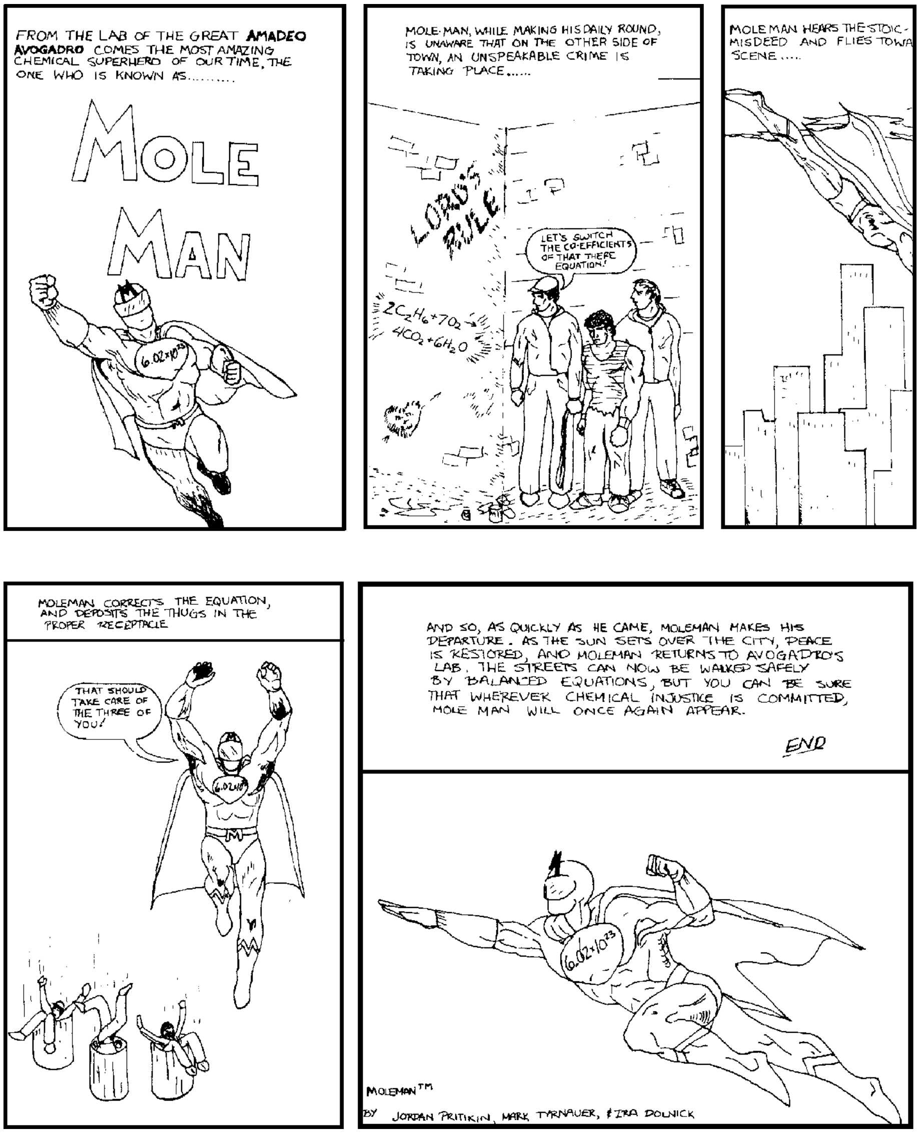

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