Experiment 2
Components of a Mixture – What Is That Stuff in the Bottom of the Cereal Box?

For the Instructor

The mixture per student should contain approximately 5 mg red #40, 100 mg Fe, 100 mg ZnO, 200 mg stearic acid, and 395 mg sucrose per gram. If the sucrose crystals are large, they might take a considerable time to dissolve, and grinding the sucrose before making the mixture is advisable.

The concentration of the methanol solution from 200 to 5-10 mL will take longer than a single lab period. This process can be facilitated by placing in a well-ventilated hood. During the following week, someone will need to check these and move them to a freezer at the appropriate time. Very little time will be required during the next laboratory period to complete the experiment.

This experiment can be performed much faster if gravity filtration is replaced with the use of a Buchner funnel, filter flask, and vacuum source.

Preparation Information – 24 students

Methanol	~6 L
Chloroform	<500 mL
(mixture – 25 g)	~25 g per section
Iron fillings	2.500 grams
Zinc oxide	2.500 grams
Sucrose	9.875 grams
Red #40	0.125 grams
Stearic acid	5.000 grams

Necessary Equipment – 24 students

Paper towels	
Disposable gloves	
Filter paper	
Weigh boats/weighing paper	
Funnel	1 per student
Beaker (20 mL)	2 per student
Beaker (400 mL)	2 per student
10-mL graduated cylinder	1 per student (optional)
100-mL graduated cylinder	1 per student (optional)
Ring stand	1 per student
Iron ring	1 per student

Clay triangle	1 per student
Stirring rod	1 per student
Wash bottle	2 per 1-4 students
Spatula	1 per student
Bar magnet	1 per 1-4 students
Rubber policeman	1 per student
Magnetic stirrer/hot plate	1 per 2-4 students (optional)
Balance	4-5 per room

Answers to Pre-Laboratory Questions

1. How could you determine (without tasting) whether a container of a colorless liquid contained ethanol or ethanol and sucrose?

   Allow the ethanol to evaporate; a white crystalline solid left would indicate the presence of sucrose.

2. Does this experiment demonstrate the law of conservation of matter?

   No. No chemical reactions are involved, only physical separations.

3. Explain the difference between filtration and decantation (see the "Experiment Equipment and Procedures" section of this manual). Why might one want to use filtration in this experiment rather than decantation?

   Decanting works if solid(s) stay at the bottom of the container. If they are likely to be disturbed, then filtration should be used.

4. How could one rapidly separate red #40 from zinc oxide? Indicate every step.

   Add water and agitate to dissolve the red #40. Filter and wash the solid with water. Solid remaining is zinc oxide.

5. Separation techniques are performed on a sample containing sand and salt. It was determined that there were 5.43 g of sand and 4.52 g of salt. The total sample weight was 10.50 g. What is the percent recovery of sand from the sample?

   $(5.43\text{ g} + 4.52\text{ g})/10.50\text{ g} \times 100\% = 94.8\%$

Chapter 2. Atoms and Elements

Student Objectives

2.1 Brownian Motion: Atoms Confirmed

- Describe Brownian motion
- Describe scanning tunneling microscopy (STM) and how atoms are imaged on surfaces.
- Define **atom** and **element**.

2.2 Early Ideas about the Building Blocks of Matter

- Describe the earliest definitions of atoms and matter (Greeks).
- Know that greater emphasis on observation and the development of the scientific method led to the scientific revolution.

2.3 Modern Atomic Theory and the Laws That Led to It

- State and understand the law of conservation of mass (also from Section 1.2).
- State and understand the law of definite proportions.
- State and understand the law of multiple proportions.
- Know the four postulates of Dalton's atomic theory.

2.4 The Discovery of the Electron

- Describe J. J. Thomson's experiments with the cathode ray tube and understand how they provide evidence for the electron.
- Describe Robert Millikan's oil-drop experiment and understand how it enables measurement of the charge of an electron.

2.5 The Structure of the Atom

- Define **radioactivity**, **nucleus**, **proton**, and **neutron**.
- Understand Thomson's plum-pudding model and how Ernest Rutherford's gold-foil experiment refuted it by giving evidence for a nuclear structure of the atom.

2.6 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms

- Define **atomic mass unit**, **atomic number**, and **chemical symbol**.
- Recognize chemical symbols and atomic numbers on the periodic table.
- Define **isotope**, **mass number**, and **natural abundance**.
- Determine the number of protons and neutrons in an isotope using the chemical symbol and the mass number.
- Define **ion**, **anion**, and **cation**.
- Understand how ions are formed from elements.

2.7 Finding Patterns: The Periodic Law and the Periodic Table

- Define the **periodic law**.
- Know that elements with similar properties are placed into columns (called groups) in the periodic table.
- Define and distinguish between metals, nonmetals, and metalloids.
- Identify main-group and transition elements on the periodic table.
- Know the general properties of elements in some specific groups: noble gases, alkali metals, alkaline earth metals, and halogens.
- Know and understand the rationale for elements that form ions with predictable charges.

2.8 Atomic Mass: The Average Mass of an Element's Atoms

- Calculate atomic mass from isotope masses and natural abundances.
- Define **mass spectrometry** and understand how it can be used to measure mass and relative abundance.

2.9 Molar Mass: Counting Atoms by Weighing Them

- Understand the relationship between mass and count of objects such as atoms.
- Define **mole** and **Avogadro's number**.
- Calculate and interconvert between number of moles and atoms.
- Calculate and interconvert between number of moles and mass.

Section Summaries

Lecture Outline

- Terms, Concepts, Relationships, Skills
- Figures, Tables, and Solved Examples

Teaching Tips

- Suggestions and Examples
- Misconceptions and Pitfalls

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples
2.1 Brownian Motion: Atoms Confirmed • Description of scanning tunneling microscopy (STM) • Introduction to macroscopic and microscopic perspectives. • Definitions of atom and element.	• Figure 2.1 Imaging Atoms
2.2 Early Ideas about the Building Blocks of Matter • History of chemistry from antiquity (~450 bc) • Scientific revolution (1400s-1600s)	

2.3 Modern Atomic Theory and the Laws That Led to It
• Law of conservation of mass
  o Matter is neither created nor destroyed.
  o Atoms at the start of a reaction may recombine to form different compounds, but all atoms are accounted for at the end.
  o Mass of reactants = mass of products.
• Law of definite proportions
  o Different samples of the same compound have the same proportions of constituent elements independent of sample source or size.
• Law of multiple proportions
• John Dalton's atomic theory
	• unnumbered figure: models and photos of Na and Cl_2 forming NaCl
	• Example 2.1 Law of Definite Proportions
	• unnumbered figure: models of CO and CO_2 illustrating the law of multiple proportions
	• Example 2.2 Law of Multiple Proportions
	• Chemistry in Your Day: Atoms and Humans

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.1 Brownian Motion: Atoms Confirmed • Other STM images can be found readily on the Internet. • It is useful to reiterate the analogies about size; for example, comparing an atom to a grain of sand and a grain of sand to a large mountain range.	• STM is not actually showing images of atoms like one might imagine seeing with a light microscope. • Atoms are not colored spheres; the images use color to distinguish different atoms.
2.2 Early Ideas about the Building Blocks of Matter • The view of matter as made up of small, indestructible particles was ignored because more popular philosophers like Aristotle and Socrates had different views. • Leucippus and Democritus may have been proven correct, but they had no more evidence for their ideas than Aristotle did. • Observations and data led scientists to question models; the scientific method promotes the use of a cycle of such inquiry.	• Theories are not automatically accepted and may be unpopular for long periods of time. • Philosophy and religion can be supported by arguments; science requires that theories be testable and therefore falsifiable.
2.3 Modern Atomic Theory and the Laws That Led to It • That matter is composed of atoms grew from experiments and observations. • Conceptual Connection 2.1 The Law of Conservation of Mass • Investigating the law of definite proportions requires preparing or decomposing a set of pure samples of a compound like water. • Investigating the law of multiple proportions requires preparing or decomposing sets of pure samples from related compounds like NO, NO_2, and N_2O_5. • Conceptual Connection 2.2 The Laws of Definite and Multiple Proportions	• Measurements to establish early atomic theories were performed at the macroscopic level. The scientists observed properties for which they could collect data (e.g., mass or volume).

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples

2.4 The Discovery of the Electron
• Thomson’s cathode-ray tube experiments
  o High voltage produced a stream of particles that traveled in straight lines.
  o Each particle possessed a negative charge.
  o Thomson measured the charge-to-mass ratio of the electron.
• Millikan’s oil-drop experiments
  o Oil droplets received charge from ionizing radiation.
  o Charged droplets were suspended in an electric field.
  o The mass and charge of each oil drop was used to calculate the mass and charge of a single electron.
	• Figure 2.2 Cathode Ray Tube
	• unnumbered figure: properties of electrical charge
	• Figure 2.3 Thomson’s Measurement of the Charge-to-Mass Ratio of the Electron
	• Figure 2.4 Millikan’s Measurement of the Electron's Charge

2.5 The Structure of the Atom
• Thomson’s plum-pudding model: negatively charged electrons in a sea of positive charge
• Radioactivity
  o Alpha decay provides the alpha particles for Rutherford’s experiment.
• Rutherford’s experiment
  o Alpha particles directed at a thin gold film deflect in all directions, including back at the alpha source.
  o Only a concentrated positive charge could cause the alpha particles to bounce back.
• Rutherford’s nuclear theory
  o most mass and all positive charge contained in a small nucleus
  o most of atom by volume is empty space
  o protons: positively charged particles
  o neutral particles with substantial mass also in nucleus
	• unnumbered figure: plum-pudding model
	• Figure 2.5 Rutherford’s Gold Foil Experiment
	• Figure 2.6 The Nuclear Atom
	• unnumbered figure: scaffolding and empty space

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.4 The Discovery of the Electron • Review the attraction, repulsion, and additivity of charges. • Discuss the physics of electric fields generated by metal plates. • A demonstration of a cathode ray tube will help students better understand Thomson's experiments. • Demonstrate how Millikan's calculation works and why he could determine the charge of a single electron.	• Millikan did not measure the charge of a single electron; he measured the charge of a number of electrons and deduced the charge of a single electron.
2.5 The Structure of the Atom • It may be useful to give a brief description of radioactivity. Rutherford's experiment makes more sense if one knows some properties of the alpha particle and from where it comes. • Thomson identified electrons and surmised the existence of positive charge necessary to form a neutral atom. The plum-pudding model is the simplest way to account for the observations. • The figure about scaffolding supports discussion about an atom being mostly empty space but still having rigidity and strength in the macroscopic view. This is another example of apparent differences between the microscopic and macroscopic properties.	• Students often don't understand the *source* of alpha particles in Rutherford's experiments.

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples

2.6 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms
• Properties of subatomic particles
  o atomic mass units (amu)
    ➢ proton, neutron: ~1 amu
    ➢ electron: ~0.006 amu
  o charge
    ➢ relative value: –1 for electron, +1 for proton
    ➢ absolute value: 1.6×10^{-19} C
• Atomic number (number of protons): defining characteristic of an element
• Isotope: same element, different mass (different number of neutrons)
• Ion: atom with nonzero charge
  o anion: negatively charged (more electrons)
  o cation: positively charged (fewer electrons)
	• unnumbered figure: baseball
	• Table 2.1 Subatomic Particles
	• unnumbered figure: lightning and charge imbalance
	• Figure 2.7 How Elements Differ
	• Figure 2.8 The Periodic Table
	• unnumbered figure: portrait of Marie Curie
	• Example 2.3 Atomic Numbers, Mass Numbers, and Isotope Symbols
	• Chemistry in Your Day: Where Did Elements Come From?

2.7 Finding Patterns: The Periodic Law and the Periodic Table
• Periodic law and the periodic table
  o generally arranged by ascending mass
  o recurring, periodic properties; elements with similar properties arranged into columns: groups (or families)
• Major divisions of the periodic table
  o metals, nonmetals, metalloids
  o main-group elements, transition elements
• Groups (families)
  o noble gases (group 8A)
  o alkali metals (group 1A)
  o alkaline earth metals (group 2A)
  o halogens (group 7A)
• Ions with predictable charges: based on stability of noble-gas electron count
  o group 1A: 1+
  o group 2A: 2+
  o group 3A: 3+
  o group 5A: 3–
  o group 6A: 2–
  o group 7A: 1–
	• unnumbered figure: discovery of the elements
	• Figure 2.9 Recurring Properties
	• Figure 2.10 Making a Periodic Table
	• unnumbered figure: stamp featuring Dmitri Mendeleev
	• Figure 2.11 Metals, Nonmetals, and Metalloids
	• Figure 2.12 The Periodic Table: Main-Group and Transition Elements
	• unnumbered figure: the alkali metals
	• unnumbered figure: the halogens
	• Example 2.4 Predicting the Charge of Ions
	• Figure 2.13 Elements That Form Ions with Predictable Charges
	• Chemistry and Medicine: The Elements of Life
	• Figure 2.14 Elemental Composition of Humans (by Mass)

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.6 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms • The analogy of the baseball and a grain of rice to a proton and an electron is meant to illustrate the difference in mass but not size. • Electrical charge can be demonstrated with static electricity. Two balloons charged with wool or human hair will repel each other. • Names of elements come from various sources. Tom Lehrer's "Element Song" can be found on the Internet. • Isotopic abundances are invariant in typical lab-sized samples because of such large numbers of atoms. • Conceptual Connection 2.5 The Nuclear Atom, Isotopes, and Ions • The history of chemistry involves considerable cultural and gender diversity. Examples include both Lavoisiers (French), Dalton (English), Thomson (English), Marie Curie (Polish/French), Mendeleev (Russian), Millikan (American), Robert Boyle (Irish), Amedeo Avogadro (Italian). • The Chemistry in Your Day box gives a broad description of the origin of atoms.	• Students sometimes confuse the mass number as being equal to the number of neutrons, not the number of neutrons plus the number of protons. • Students logically (but mistakenly) presume that the mass of an isotope is equal to the sum of the masses of the protons and neutrons in that isotope.

2.7 Finding Patterns: The Periodic Law and the Periodic Table • Other displays of the periodic table can be found in journals (Schwartz, *J. Chem. Educ*. **2006**, *83*, 849; Moore, *J. Chem. Educ*. **2003**, *80*, 847; Bouma, *J. Chem. Educ.* **1989**, *66*, 741), books, and on the Internet. • Periodic tables are arranged according to the periodic law but can compare many features, e.g. phases of matter, sizes of atoms, and common ions. These are presented as a series of figures in the text. • Chemistry and Medicine: The Elements of Life provides an opportunity to relate the topics to everyday life. Some of the other elements in the figure and table represent trace minerals that are part of good nutrition. The periodic law accounts for why some are necessary and others are toxic.	• The periodic table is better at predicting microscopic properties, though macroscopic properties are also often illustrated.

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples

2.8 Atomic Mass: The Average Mass of an Element's Atoms
• Average atomic mass is based on natural abundance and isotopic masses.
• Mass spectrometry
  o atoms converted to ions and deflected by magnetic fields to separate by mass
  o output data: relative mass *vs*. relative abundance
	• unnumbered figure: periodic table box for Cl
	• Example 2.5 Atomic Mass
	• Figure 2.15 The Mass Spectrometer
	• Figure 2.16 The Mass Spectrum of Chlorine

2.9 Molar Mass: Counting Atoms by Weighing Them • Mole concept and Avogadro's number • Converting between moles and number of atoms • Converting between mass and number of moles	• unnumbered figure: pennies containing ~1 mol of Cu • unnumbered figure: 1 tbsp of water contains ~1 mol of water • Example 2.6 Converting between Number of Moles and Number of Atoms • unnumbered figure: relative sizes of Al, C, He • unnumbered figure: balance with marbles and peas • Example 2.7 Converting between Mass and Amount (Number of Moles) • Example 2.8 The Mole Concept–Converting between Mass and Number of Atoms • Example 2.9 The Mole Concept

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.8 Atomic Mass: The Average Mass of an Element's Atoms • The masses of isotopes must be reconciled with an element having only whole number quantities of protons and neutrons; the values should be nearly integral since the mass of electrons is so small. • Mass spectrometry is an effective way to demonstrate where values of natural abundance are obtained.	• Students are tempted to calculate average atomic mass by adding together isotopic masses and dividing by the number of isotopes. • Atomic mass on the periodic table is usually not integral even though elements have only whole numbers of protons and neutrons.
2.9 Molar Mass: Counting Atoms by Weighing Them • Review the strategy for solving numerical problems: sort, strategize, solve, check. • Estimating answers is an important skill; the number of atoms will be very large (i.e. some large power of ten) even from a small mass or small number of moles. • Conceptual Connection 2.7 Avogadro's Number • Conceptual Connection 2.8 The Mole	• Many students are intimidated by estimating answers in calculations involving powers of ten.

Additional Problem for Converting between Number of Moles and Number of Atoms (Example 2.6)	**Calculate the number of moles of iron in a sample that has 3.83 $\times$ 10^{23} atoms of iron.**
Sort You are given a number of iron atoms and asked to find the amount of iron in moles.	**Given** 3.83×10^{23} Fe atoms **Find** mol Fe
Strategize Convert between number of atoms and number of moles using Avogadro's number.	**Conceptual Plan** atoms $\rightarrow$ mol $\dfrac{1\text{ mol Fe}}{6.022 \times 10^{23}\text{ Fe atoms}}$ **Relationships Used** $6.022 \times 10^{23} = 1$ mol (Avogadro's number)
Solve Follow the conceptual plan. Begin with 3.83 $\times$ 10^{23} Fe atoms and multiply by the ratio that equates moles and Avogadro's number.	**Solution** $3.83 \times 10^{23}\text{ Fe atoms} \times \dfrac{1\text{ mol Fe}}{6.022 \times 10^{23}\text{ Fe atoms}} = 0.636\text{ mol Fe}$
Check	The sample was smaller than Avogadro's number so the answer should be a fraction of a mole. The value of the sample has 3 significant figures, and the answer is provided in that form.

Additional Problem for Converting between Mass and Number of Moles (Example 2.7)	**Calculate the number of grams of silver in an American Silver Eagle coin that contains 0.288 moles of silver.**
Sort You are given the amount of silver in moles and asked to find the mass of silver.	**Given** 0.288 mol Ag **Find** g Ag
Strategize Convert amount (in moles) to mass using the molar mass of the element.	**Conceptual Plan** mol Ag $\rightarrow$ g Ag $\frac{107.87\text{ g Ag}}{1\text{ mol Ag}}$ **Relationships Used** 107.87 g Ag = 1 mol Ag
Solve Follow the conceptual plan to solve the problem. Start with 0.288 mol, the given number, and multiply by the molar mass of silver.	**Solution** $0.288\ \cancel{\text{mol Ag}} \times \frac{107.87\text{ g Ag}}{1\ \cancel{\text{mol Ag}}} = 31.07\text{ g Ag}$ $31.07\text{ g} = 31.1\text{ g Ag}$
Check	The magnitude of the answer makes sense since we started with an amount smaller than a mole. The molar amount and answer both have 3 significant figures.

Additional Problem for the Mole Concept—Converting between Mass and Number of Atoms (Example 2.8)	**What mass of iron (in grams) contains 1.20×10^{22} atoms of Fe? A paperclip contains about that number of iron atoms.**
Sort You are given a number of iron atoms and asked to find the mass of Fe.	**Given** 1.20×10^{22} Fe atoms **Find** g Fe
Strategize Convert the number of Fe atoms to moles using Avogadro's number. Then convert moles Fe into grams of iron using the molar mass of Fe.	**Conceptual Plan** Fe atoms $\rightarrow$ mol Fe $\rightarrow$ g Fe $\dfrac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}}$ $\dfrac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}$ **Relationships Used** $6.022 \times 10^{23} = 1$ mol (Avogadro's number) $55.85 \text{ g Fe} = 1 \text{ mol Fe}$
Solve Follow the conceptual plan to solve the problem. Begin with 1.20×10^{22} atoms of Fe, multiply by the ratio derived from Avogadro's number, and finally multiply by the atomic mass of Fe.	**Solution** $1.20 \times 10^{22} \cancel{\text{Fe atoms}} \times \dfrac{1 \cancel{\text{mol Fe}}}{6.022 \times 10^{23} \cancel{\text{Fe atoms}}} \times \dfrac{55.85 \text{ g Fe}}{1 \cancel{\text{mol Fe}}}$ $= 1.11 \text{ g Fe}$
Check	The units and magnitude of the answer make sense. The sample is smaller than a mole. The number of atoms and mass both have 3 significant figures.

Additional Problem for the Mole Concept (Example 2.9)	**An iron sphere contains 8.55×10^{22} iron atoms. What is the radius of the sphere in centimeters? The density of iron is 7.87 g/cm^3.**
Sort You are given the number of iron atoms in a sphere and the density of iron. You are asked to find the radius of the sphere.	**Given** 8.55×10^{22} Fe atoms $d = 7.87$ g/cm^3 **Find** radius (r) of a sphere
Strategize The critical parts of this problem are density, which relates mass to volume, and the mole, which relates number of atoms to mass: (1) Convert from the number of atoms to the number of moles using Avogadro's number; (2) Convert from the number of moles to the number of grams using the molar mass of iron; (3) Convert from mass to volume using the density of iron; (4) Find the radius using the formula for the volume of a sphere.	**Conceptual Plan** Fe atoms → mol Fe → g Fe → V (cm^3) $\dfrac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}}$ $\dfrac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}$ $\dfrac{1 \text{ cm}^3}{7.87 \text{ g Fe}}$ V (cm^3) → r (cm) $V = \frac{4}{3}\pi r^3$ **Relationships Used** $6.022 \times 10^{23} = 1$ mol (Avogadro's number) 55.85 g Fe = 1 mol Fe d (density of Fe) = 7.87 g/cm^3 $V = 4/3\ \pi r^3$ [volume of a sphere with a radius of r]
Solve Follow the conceptual plan to solve the problem. Begin with 8.55×10^{22} Fe atoms and convert to moles, then to grams and finally to a volume in cm^3. Solve for the radius using the rearranged equation.	**Solution** $8.55 \times 10^{22} \cancel{\text{atoms}} \times \dfrac{1 \cancel{\text{ mol Fe}}}{6.022 \times 10^{23} \cancel{\text{ atoms}}} \times \dfrac{55.85 \cancel{\text{ g Fe}}}{1 \cancel{\text{ mol Fe}}} \times \dfrac{1 \text{ cm}^3}{7.87 \cancel{\text{ g Fe}}} = 1.00757 \text{ cm}^3$ $r = \sqrt[3]{\dfrac{3\text{ V}}{4\ \pi}} = \sqrt[3]{\dfrac{3 \times 1.00757 \text{ cm}^3}{4\ \pi}} = 0.622 \text{ cm}$
Check	The units (cm) are correct and the magnitude of the answer makes sense compared with previous problems.