Chapter 1. Introduction: Matter and Measurement

Common Student Misconceptions

- Students often confuse mass and weight.
- Students have difficulty with algebraic manipulation. Conversion of temperatures between Celsius and Fahrenheit scales is particularly problematic.
- Students tend to equate density with mass.
- Students often either are unfamiliar with the prefixes used in the metric system or cannot use them properly (e.g. $1\ \text{pm} = 1 \times 10^{-12}\ \text{m} \rightarrow 1\ \text{m} = 1 \times 10^{12}\ \text{pm}$).
- Students often use precision and accuracy interchangeably.
- Students often do not appreciate that in chemistry, measurement yields numbers determined with certain precision and in certain units; both depend on the type of the measuring device.
- Students often cannot find exact numbers in calculations.
- Students confuse significant figures and decimal places in arithmetic manipulations.
- Students often either round off too soon in calculations or they report the result to as many figures as their calculators produce.
- Some students do not understand the use of a conversion factor of exactly one.
- In dimensional analysis problems students do not see that a physical quantity is a multiplication of value and units. Therefore, they do not perform algebraic operations on both the number and units.

Teaching Tips

- Many students have problems using dimensional analysis ("from physics") in chemistry. This text bases the whole of stoichiometry on dimensional analysis: students should be encouraged to embrace the concept as soon as possible.
- Emphasize the importance of quickly mastering the use of common metric prefixes and scientific notation.
- Conversion involving commonly encountered metric prefixes need frequent reinforcement; for example, $10^6\ \mu\text{g} = 10^3\ \text{mg} = 1\ \text{g} = 10^{-3}\ \text{kg}$

Lecture Outline

1.1 The Study of Chemistry

- Chemistry:
  - is the study of properties of materials and changes that they undergo.
  - can be applied to all aspects of life (e.g., development of pharmaceuticals, leaf color change in fall, etc.).

The Atomic and Molecular Perspective of Chemistry[1,2,3,4,5,6,7]

Chemistry involves the study of the properties and the behavior of matter.

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[1] "The First Demonstration: Proof that Air Is a Substance" from Live Demonstrations
[2] "Oxygen" 3-D Model from Instructor's Resource CD/DVD
[3] "Water" 3-D Model from Instructor's Resource CD/DVD
[4] "Hydrogen" 3-D Model from Instructor's Resource CD/DVD
[5] "Ethanol" 3-D Model from Instructor's Resource CD/DVD
[6] "Ethylene Glycol" 3-D Model from Instructor's Resource CD/DVD
[7] "Aspirin" 3-D Model from Instructor's Resource CD/DVD

- **Matter**:
  - is the physical material of the universe.
  - has mass.
  - occupies space.
  - A **property** is any characteristic that allows us to recognize a particular type of matter and to distinguish it from other types of matter.
  - About 100 **elements** constitute all matter.
- **Elements**:
  - are made up of unique **atoms**, the building blocks of matter.
  - Names of the elements are derived from a wide variety of sources (e.g., Latin or Greek, mythological characters, names of people or places).
- **Molecules**:
  - are combinations of atoms held together in specific shapes.
  - *Macroscopic* (observable) properties of matter relate to *submicroscopic* realms of atoms.
  - Properties relate to composition (types of atoms present) and structure (arrangement of atoms) present.

Why Study Chemistry?[8,9,10]

We study chemistry because:

- it has a considerable impact on society (health care, food, clothing, conservation of natural resources, environmental issues, etc.).
- it is part of your curriculum! Chemistry serves biology, engineering, agriculture, geology, physics, etc.. **Chemistry is the *central science***.

1.2 Classifications of Matter

- Matter is classified by *state* (solid, liquid, or gas) or by *composition* (element, compound or mixture).

States of Matter[11]

- Solids, liquids and gases are the three forms of matter called the **states of matter**.
- Properties described on the macroscopic level:
  - **gas** (vapor): no fixed volume or shape, conforms to shape of container, compressible.
  - **liquid**: volume independent of container, no fixed shape, incompressible.
  - **solid**: volume and shape independent of container, rigid, incompressible.
- Properties described on the molecular level:
  - **gas**: molecules far apart, move at high speeds, collide often.
  - **liquid**: molecules closer than gas, move rapidly but can slide over each other.
  - **solid**: molecules packed closely in definite arrangements.

Pure Substances[12]

- **Pure substances**:
  - are matter with distinct properties and fixed composition..
  - are **elements** (cannot be decomposed into simpler substances; i.e. only one kind of atom) or **compounds** (consist of two or more elements).

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[8] "Chemistry in the Real World" from Further Readings
[9] "Science Demonstrations, Experiments, and Resources. A Reference List for Elementary through College Teachers Emphasizing Chemistry with Some Physics and Life Science" from Live Demonstrations
[10] "Chemicals in Everyday Life" from Further Readings
[11] "Phases of Water" Activity from Instructor's Resource CD/DVD
[12] Figure 1.5 from Transparency Pack

- **Mixtures**:
  - are a combination of two or more pure substances.
  - Each substance retains its own identity.

Elements[13,14,15,16,17,18,19,20,21]

- There are 117 known elements.
- They vary in abundance.
  - Oxygen, silicon, aluminum, iron, and calcium make up over 90% of the Earth's crust (including oceans and atmosphere).
  - Oxygen, carbon, and hydrogen make up over 90% of mass of the human body.
- Each is given a unique name and is abbreviated by a chemical *symbol*.
- They are organized in the *periodic table*.
- Each is given a one- or two-letter symbol derived from its name.

Compounds[22,23]

- **Compounds** are combinations of elements.
  Example: The compound H_2O is a combination of elements H and O.
- The opposite of compound formation is decomposition.
- Compounds have different properties than their component elements (e.g., water is liquid, hydrogen and oxygen are both gases at the same temperature and pressure).
- **Law of Constant Composition or the Law of Constant (Definite) Proportions** (Proust): A compound always consists of the same combination of elements (e.g., water is always 11% H and 89% O).

Mixtures[24,25,26]

- A **mixture** is a combination of two or more pure substances.
  - Each substance retains its own identity, each substance is a *component* of the mixture.
  - Mixtures have variable composition.
  - *Heterogeneous* mixtures do not have uniform composition, properties and appearance, e.g., sand.
  - *Homogeneous* mixtures are uniform throughout, e.g., clean air or vinegar; they are **solutions**.

FORWARD REFERENCES

- States of matter will be essential in properly writing chemical reaction equations, including net ionic equations, in Chapter 4, as well as equilibrium constant expressions in Chapters 15, 19, and 20.
- Solutions will be further discussed in Chapters 4 and 13.
- Mixtures of gases will be discussed in Chapter 10.

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[13] "What's the Use?" from Further Readings
[14] "Phases of the Elements" Activity from Instructor's Resource CD/DVD
[15] "It's Elementary" from Further Readings
[16] "Element ZOO" from Further Readings
[17] "Important Elements" from Further Readings
[18] "Origin of the Names of Chemical Elements" from Further Readings
[19] "Connecting Element names with the Names of U.S. Towns" from Further Readings
[20] "Elementary My Dear Watson" from Further Readings
[21] "Elementary Riddles" from Further Readings
[22] "Electrolysis of Water" Animation from Instructor's Resource CD/DVD
[23] Figure 1.7 from Transparency Pack
[24] Figure 1.9 from Transparency Pack
[25] "Classification of Matter" Activity from Instructor's Resource CD/DVD
[26] "Classifying Matter: A Physical Model Using Paper Clips" from Further Readings

1.3 Properties of Matter

- Each substance has a unique set of physical and chemical properties.
  - **Physical properties** are measured without changing the substance (e.g., color, density, odor, melting point, etc.).
  - **Chemical properties** describe how substances react or change to form different substances (e.g., hydrogen burns in oxygen).
  - Properties may be categorized as intensive or extensive.
    - **Intensive properties** do not depend on the amount of substance present (e.g., temperature, melting point etc.).
    - **Extensive properties** depend on the quantity of substance present (e.g., mass, volume etc.).
    - Intensive properties give an idea of the composition of a substance whereas extensive properties give an indication of the quantity of substance present.

Physical and Chemical Changes[27,28,29,30]

- **Physical change**: substance changes physical appearance without altering its identity (e.g., **changes of state**).
- **Chemical change** (or **chemical reaction**): substance transforms into a chemically different substances (i.e. identity changes, e.g., reaction of hydrogen and oxygen gases to produce water).

Separation of Mixtures[31,32,33,34]

- Key: separation techniques exploit differences in properties of the *components*.
  - Filtration: remove solid from liquid.
  - Distillation: boil off one or more components of the mixture.
  - Chromatography: exploits differing abilities of substances to adhere to the surfaces of solids..

The Scientific Method[35,36,37,38,39]

- **The scientific method** provides guidelines for the practice of science.
  - Collect data (observe, experiment, etc.).
  - Look for patterns, try to explain them and develop a **hypothesis** *or tentative explanation*.
  - Test hypothesis, then refine it.
  - Bring all information together into a **scientific law** (*concise statement or equation that summarizes tested hypotheses*).
  - Bring hypotheses and laws together into a theory. A **theory** should explain general principles.

1.4 Units of Measurement

- Many properties of matter are quantitative, i.e., associated with numbers.

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27 "Changes of State" Animation from Instructor's Resource CD/DVD
28 Figure 1.10 from Transparency Pack
29 "Ira Remsen's Investigation of Nitric Acid" from Live Demonstrations
30 "Sodium and Potassium in Water" Movie from Instructor's Resource CD/DVD
31 "Mixtures and Compounds" Movie from Instructor's Resource CD/DVD
32 Figure 1.16 from Transparency Pack
33 "Paper Chromatography of Ink" Movie from Instructor's Resource CD/DVD
34 "T-Shirt Chromatography: A Chromatogram You Can Wear" from Further Readings
35 "The Inquiry Wheel, an Alternative to the Scientific Method: A View of the Science Education Research Literature" from Further Readings
36 "An Experiment to Demonstrate the Application of the Scientific Method" from Live Demonstrations
37 "Why Does Popcorn Pop? An Introduction to the Scientific Method" from Further Readings
38 "Mentos and the Scientific Method: A Sweet Combination" from Further Readings
39 "Using History to Teach the Scientific Method: The Role of Errors" from Further Readings

- A measured quantity must have BOTH a number and a unit.
- The units most often used for scientific measurement are those of the **metric system**.

SI Units[40,41,42]

- 1960: All scientific units use **S**ystème **I**nternational d'Unités (**SI** Units).
- There are seven base units.
- Prefixes are used to indicate smaller and larger units obtained by decimal fractions or multiples of the base units.

Length and Mass

- SI base unit of length = meter (1 m = 1.0936 yards).
- SI base unit of mass (not weight) = kilogram (1 kg = 2.2 pounds).
  - **Mass** is a measure of the amount of material in an object.

Temperature[43]

- *Temperature* is the measure of the hotness or coldness of an object.
  - Physical property that determines the direction of heat flow.
  - Heat flows spontaneously from a substance of higher temperature to one at lower temperature.
- Scientific studies use Celsius and Kelvin scales.
- **Celsius scale**: water freezes at 0 °C and boils at 100 °C (sea level).
- **Kelvin scale** (SI Unit):
  - Water freezes at 273.15 K and boils at 373.15 K (sea level).
  - is based on properties of gases.
  - Zero is lowest possible temperature (*absolute zero*).
  - 0 K = –273.15 °C.
- Fahrenheit (not used in science):
  - Water freezes at 32 °F and boils at 212 °F (sea level).
  - Conversions:

$$°F = (9/5)°C + 32 \qquad °C = (5/9)(°F - 32)$$

$$°C = K - 273.15 \qquad K = °C + 273.15$$

Derived SI Units

- These are formed from the seven base units.
- Example: velocity is distance traveled per unit time, so units of velocity are units of distance (m) divided by units of time (s): m/s.

Volume[44]

- Units of *volume* = (units of length)^3 = m^3.
- This unit is unrealistically large, so we use more reasonable units:
  - cm^3 (also known as mL (*milliliter*) or cc (*cubic centimeters*))
  - dm^3 (also known as *liters*, L).
- **Important: the liter is not an SI unit.**

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[40] "A Simple Demonstration for Introducing the Metric System to Introductory Chemistry Classes" from Live Demonstrations
[41] "Having Fun with the Metric System" from Further Readings
[42] Table 1.5 from Transparency Pack
[43] "Temperature" Activity from Instructor's Resource CD/DVD
[44] Figure 1.18 from Transparency Pack

Density[45,46,47,48,49,50,51]

- **Density** is defined as mass divided by volume.
- Units: g/cm^3 or g/mL (for solids and liquids); g/L (often used for gases).
- Was originally based on mass (the density was defined as the mass of 1.00 mL of pure water at 25˚C).

FORWARD REFERENCES

- Prefixes (Table 1.5) will be heavily used in future chapters: kJ (Chapters: 5-8, 14, 19-21); nm, pm and MHz (Chapter 6).
- Density will be a quantity commonly used in Chapter 10 (for gases), Chapter 11 (to compare liquid and solid water), Chapter 13 (for solutions).
- Conversions of temperature units will be commonplace in Chapters 5, 14-20.

1.5 Uncertainty in Measurement[52,53,54,55]

- There are two types of numbers:
  - *exact numbers* (known as counting or defined).
  - *inexact numbers* (derived from measurement).

Precision and Accuracy

- **Precision**: how well measured quantities agree with each other.
- **Accuracy**: how well measured quantities agree with the "true value".
- Figure 1.23 is very helpful in making this distinction.

Significant Figures[56,57,58,59,60,61,62]

- All measurements have some degree of *uncertainty* or *error* associated with them.
- In a measurement it is useful to indicate the exactness of the measurement. This exactness is reflected in the number of **significant figures**.
- Guidelines for determining the number of significant figures in a measured quantity are:
  - The number of significant figures is the number of digits known with certainty plus one uncertain digit. (Example: 2.2405 g means we are sure the mass is 2.240 g but we are uncertain about the nearest 0.0001 g.)

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[45] "Sugar in a Can of Soft Drink: A Density Exercise" from Live Demonstrations
[46] "The Mysterious Sunken Ice Cube" from Live Demonstrations
[47] "Method for Separating or Identifying Plastics" from Further Readings
[48] "Densities and Miscibilities of Liquids and Liquid Mixtures" from Live Demonstrations
[49] "The Concept of Density" from Further Readings
[50] "Bowling for Density!" from Live Demonstrations
[51] "Whatever Floats (or Sinks)" from Live Demonstrations
[52] "Uncertainty in Measurement" Activity from Instructor's Resource CD/DVD
[53] "Meter Sticks in the Demonstration of Error Measurement" from Further Readings
[54] "Basic Principles of Scale Reading" from Further Readings
[55] "Measuring with a Purpose. Involving Students in the Learning Process" from Further Readings
[56] "Error, Precision, and Uncertainty" from Further Readings
[57] "Precision and Accuracy in Measurements: A Tale of Four Graduated Cylinders" from Further Readings
[58] "A Simple but Effective Demonstration for Illustrating Significant Figure Rules When Making Measurements and Doing Calculations" from Further Readings
[59] "Significant Figures" Activity from Instructor's Resource CD/DVD
[60] "A Joke Based on Significant Figures" from Further Readings
[61] "Significant Figures" from Further Readings
[62] "Significant Figures: A Classroom Demonstration" from Further Readings

- Final calculations are only as significant as the least significant measurement.
- Rules:
  1. Nonzero numbers and zeros between nonzero numbers (ie. imbedded zeros) are always significant.
  2. Zeros before the first nonzero digit (ie. leading zeros) are not significant. (Example: 0.0026 has two significant figures.)
  3. Zeros at the end of the number after a decimal point are significant.
  4. Zeros at the end of a number before a decimal point are ambiguous (e.g., 10,300 g). Exponential notation, such as scientific notation, eliminates this ambiguity.
- Method:
  1. Write the number in scientific notation.
  2. The number of digits remaining is the number of significant figures.
  3. Examples:
     2.50×10^2 cm has 3 significant figures as written.
     1.03×10^4 g has 3 significant figures.
     1.030×10^4 g has 4 significant figures.
     1.0300×10^4 g has 5 significant figures.

Significant Figures in Calculations

- In calculations, the least certain measurement limits the certainty of the calculated result
  - The answer is reported with only 1 uncertain digit.
- Guidelines for keeping track of significant figures:
  - Addition and Subtraction:
  - Report to the least number of decimal places (e.g., 20.4 g – 1.322 g = 19.1 g).
  - Multiplication and Division:
  - Report to the least number of significant figures (e.g., 6.221 cm × 5.2 cm = 32 cm^2).
- In multiple step calculations always retain an extra significant figure until the end to prevent rounding errors.

FORWARD REFERENCES

- Working in scientific notation will be required in Chapters: 6 (wavelength, frequency, energy of levels), 14 (rate constant), 15 (equilibrium constant calculations), 16 and 17 (acid or base ionization constants, concentrations), 19 ($\Delta G°$ vs K calculations), 20 ($\Delta G°$ vs K vs $E°$ calculations).
- Rules for significant figures will appear whenever calculations are to be performed; rules for significant figures in calculations with logarithms will be needed in Chapters 14, 16, 17, 19, and 20.

1.6 Dimensional Analysis[63,64]

- **Dimensional analysis** is a method of calculation utilizing a knowledge of units.
- Given units can be multiplied and divided to give the desired units.
- Conversion factors are used to manipulate units:
  - desired unit = given unit × (conversion factor)
- The **conversion factors** are simple ratios:
  - conversion factor = (desired unit) / (given unit)
    - These are fractions whose numerator and denominator are the same quantity expressed in different units.
    - Multiplication by a conversion factor is equivalent to multiplying by a factor of 1.

[63] "Expanded Dimensional Analysis: A Blending of English and Math" from Further Readings
[64] "Appalachian Trail Problems" from Further Readings

Using Two or More Conversion Factors

- We often need to use more than one conversion factor in order to complete a problem.
- When identical units are found in the numerator and denominator of a conversion, they will cancel. The final answer MUST have the correct units.
- For example:
  - Suppose that we want to convert length in meters to length in inches. We could do this conversion with the following conversion factors:
    - 1 meter = 100 centimeters and 1 inch = 2.54 centimeters
- The calculation would involve both conversion factors; the units of the final answer will be inches:
    - (# meters) (100 centimeters / 1 meter) (1 inch / 2.54 centimeters) = # inches

Conversions Involving Volume

- We often will encounter conversions from one measure to a different measure.
- For example:
  - Suppose that we wish to know the mass in grams of 2.00 cubic inches of gold given that the density of the gold is 19.3 g/cm^3.
  - We could do this conversion with the following conversion factors:

    $$2.54 \text{ cm} = 1 \text{ inch and } 1 \text{ cm}^3 = 19.3 \text{ g gold}$$
  - The calculation would involve both of these factors:

    $$(2.00 \text{ in.}^3)\ (2.54 \text{ cm / in.})^3\ (19.3 \text{ g gold / } 1 \text{ cm}^3) = 633 \text{ g gold}$$
  - Note that the calculation will NOT be correct unless the centimeter to inch conversion factor is cubed!! Both the units AND the number must be cubed.

Summary of Dimensional Analysis

- In dimensional analysis always ask three questions:
  1. What data are we given?
  2. What quantity do we need?
  3. What conversion factors are available to take us from what we are given to what we need?

FORWARD REFERENCES

- Solving problems using dimensional analysis can be found virtually in each chapter.

Other Resources

Further Readings:	Section:
Chemistry in the Real World	1.1 The Study of Chemistry
Chemicals in Everyday Life	1.1 The Study of Chemistry
What's the Use?	1.2 Classifications of Matter
It's Elementary	1.2 Classifications of Matter
Element Zoo	1.2 Classifications of Matter
Important Elements	1.2 Classifications of Matter
Origin of the Names of Chemical Elements	1.2 Classifications of Matter
Connecting Element names with the Names of U.S. Towns	1.2 Classifications of Matter
The Inquiry Wheel, an Alternative to the Scientific Method: A View of the Science Education Research Literature	1.3 Properties of Matter
Why Does Popcorn Pop?	1.3 Properties of Matter
Using History to Teach The Scientific Method: The Role of Errors	1.3 Properties of Matter
Having Fun with the Metric System	1.4 Units of Measurement
Method for Separating or Identifying Plastics	1.4 Units of Measurement
Error, Precision, and Uncertainty	1.5 Uncertainty in Measurement
Precision and Accuracy in Measurements: A Tale of Four Graduated Cylinders	1.5 Uncertainty in Measurement
A Simple but Effective Demonstration for Illustrating Significant Figure Rules When Making Measurements and Doing Calculations	1.5 Uncertainty in Measurement
A Joke Based on Significant Figures	1.5 Uncertainty in Measurement
Significant Figures	1.5 Uncertainty in Measurement
Significant Figures: A Classroom Demonstration	1.5 Uncertainty in Measurement
Expanded Dimensional Analysis: A Blending of English and Math	1.6 Dimensional Analysis
Appalachian Trail Problems	1.6 Dimensional Analysis

Live Demonstrations:	Section:
The First Demonstration: Proof that Air Is a Substance	1.1 The Study of Chemistry
Science Demonstrations, Experiments, and Resources. A Reference List for Elementary through College Teachers Emphasizing Chemistry with Some Physics and Life Science	1.1 The Study of Chemistry
Ira Remsen's Investigation of Nitric Acid	1.3 Properties of Matter
An Experiment to Demonstrate the Application of the Scientific Method	1.3. Properties of Matter
A Simple Demonstration for Introducing the Metric System to Introductory Chemistry Classes	1.4 Units of Measurement
Sugar in a Can of Soft Drink: A Density Exercise	1.4 Units of Measurement
The Mysterious Sunken Ice Cube	1.4 Units of Measurement
Densities and Miscibilities of Liquids and Liquid Mixtures	1.4 Units of Measurement

Further Readings:

1. Martin B. Jones and Christina R. Miller, “Chemistry in the Real World,” *J. Chem. Educ.*, Vol. 78, **2001**, 484–487.

2. Raymond B. Seymour, “Chemicals in Everyday Life,” *J. Chem. Educ*., Vol. 64, **1987**, 63–68.

3. Each of the “What’s the Use?” articles, written by Alton Banks, focuses on the uses of a specific element. See *J. Chem. Educ*., Vols. 66 (**1989**), 67 (**1990**), 68 (**1991**) and 69 (**1992**).

4. Milton J. Wieder, “It’s Elementary,” *J. Chem. Educ.*, Vol. 78, **2001**, 468–469.

5. Terry L. Helser, “Element ZOO”, *J. Chem. Educ.*, Vol. 80, **2003**, 409–410.

6. P. G. Nelson, “Important Elements,” *J. Chem. Educ*., Vol. 68, **1991**, 732–737.

7. Vivi Ringes, “Origin of the Names of Chemical Elements,” *J. Chem. Educ*., Vol. 66, **1989**, 731–738.

8. Nicholas C. Thomas, “Connecting Element names with the Names of U.S. Towns,” *J. Chem. Educ*., Vol. 86, **2009**, 181-184. A sampling of towns in the US with names based on chemical elements.

9. Terry L. Helser, “Elementary My Dear Watson,” *J. Chem. Educ*., Vol. 66, **1989**, 980. Puzzles involving element names and symbols.

10. Doris Eckey, “Elementary Riddles”, *J. Chem. Educ*., Vol. 71, **1994**, 1051. A series of riddles involving the names of the elements.

11. Bob Blake, Lynn Hogue, and Jerry L. Sarquis, “Classifying Matter: A Physical Model Using Paper Clips,” *J. Chem. Educ.*, Vol. 83, **2006**, 1317–1318.

12. Jeanne M. Buccigros, “T-Shirt Chromatography: A Chromatogram You Can Wear,” *J. Chem. Educ*., Vol. 69, **1992**, 977–978.

13. Frederick C. Sauls, “Why Does Popcorn Pop? An Introduction to the Scientific Method,” *J. Chem. Educ*., Vol. 86, **1991**, 415–416.

14. Carmen J. Guinta, “Using History to Teach Scientific Method: The Case of Argon,” *J. Chem. Educ*., Vol. 75, **1998**, 1322–1325. The scientific method is introduced using a case-study approach involving the discovery of argon.

15. Jack F. Eichler, Heather Patrick, Brenda Harmon and Janet Coonce, “Mentos and the Scientific Method: A Sweet Combination,” *J. Chem. Educ.*, Vol. 84, **2007**, 1120–1123.

16. William R. Robinson, “The Inquiry Wheel, an Alternative to the Scientific Method: A View of the Science Education Research Literature,” *J. Chem. Educ.*, Vol. 81, **2004**, 791–792.

17. Mark L. Campbell, “Having Fun with the Metric System,” *J. Chem. Educ*., Vol. 68, **1991**, 1043. A word game.

18. Kenneth E. Kolb and Doris K. Kolb, “Method for Separating or Identifying Plastics,” *J. Chem. Educ*., Vol. 68, **1991**, 348. A demonstration of the concept of density.

19. Stephen J. Hawkes, “The Concept of Density,” *J. Chem. Educ.*, Vol. 81, **2004**, 14–15.

20. Robert Suder, “Meter Sticks in the Demonstration of Error Measurements*,*” *J. Chem. Educ*., Vol. 66, **1989**, 437.

21. Gavin D. Peckhan, “Basic Principles of Scale Reading,” *J. Chem. Educ*., Vol. 71, **1994**, 423–424.

22. Patricia A. Metz and Jeffrey R. Pribyl, “Measuring with a Purpose. Involving Students in the Learning Process,” *J. Chem. Educ*., Vol. 72, **1995**, 130–132.

23. Richard S. Treptow, “Precision and Accuracy in Measurements: a Tale of Four Graduated Cylinders,” *J. Chem. Educ*., Vol. 75, **1998**, 992–995.

24. Charles J. Guare, “Error, Precision, and Uncertainty,” *J. Chem. Educ*., Vol. 68, **1991**, 649–652.

25. Kenton B. Abel and William M. Hemmerlin, “Significant Figures,” *J. Chem. Educ*., Vol. 67, **1990**, 213.

26. H. Graden Kirksey and Paul Krause, “Significant Figures: A Classroom Demonstration,” *J. Chem. Educ*., Vol. 69, **1992**, 497–498.

27. Ben Ruekberg, “A Joke Based on Significant Figures,” *J. Chem. Educ*., Vol. 71, **1994**, 306.

28. Arden P. Zipp, “A Simple but Effective Demonstration for Illustrating Significant Figure Rules When Making Measurements and Doing Calculations,” *J. Chem. Educ.*, Vol. 69, **1992**, 291.

29. Ronald DeLorenzo, "Expanded Dimensional Analysis: A Blending of English and Math," *J. Chem. Educ*., Vol. 71, **1994**, 789–791.

30. Brian N. Akers, “Appalachian Trail Problems,” *J. Chem. Educ*., Vol. 75, **1998**, 1571–1572. This short article provides some unit analysis problems.

31. W. Kirk Stephenson, “The Box-and-Dot Method: A Simple Strategy for Counting Significant Figures,” *J. Chem. Educ*., Vol. 86, **2009**, 933-935.

Live Demonstrations:

1. David A. Katz, “Science Demonstrations, Experiments, and Resources. A Reference List for Elementary through College Teachers Emphasizing Chemistry with Some Physics and Life Science,” *J. Chem. Educ*., Vol. 68, **1991**, 235–244. An extensive listing of various demonstrations, experiments and resources for the science educator.

2. Lee R. Summerlin, Christie L. Borgford, and Julie B. Ealy, “Ira Remsen's Investigation of Nitric Acid,” *Chemical Demonstrations. A Sourcebook for Teachers*, Volume 2 (Washington: American Chemical Society, **1988**), pp. 4–5.

3. Irving R Tannenbaum, “An Experiment to Demonstrate the Application of the Scientific Method,” *J. Chem. Educ*., Vol. 66, **1989**, 597. A simple experiment to heat water to boiling.

4. David A. Franz and David Speckhard, "Densities and Miscibilities of Liquids and Liquid Mixtures," *J. Chem. Educ*., Vol. 68, **1991**, 594. Demonstrations involving densities and miscibility of liquids.

5. Lee. R. Summerlin, Christie L. Borgford, and Julie B. Ealy, “The First Demonstration: Proof That Air Is a Substance,” *Chemical Demonstrations, A Sourcebook for Teachers*, Volume 2 (Washington: American Chemical Society, **1988**), p. 3.

6. Clarke W. Earley, “A Simple Demonstration for Introducing the Metric System to Introductory Chemistry Classes,” *J. Chem. Educ.*, Vol. 76, **1999**, 1215–1216.

7. Lee. R. Summerlin, Christie L. Borgford, and Julie B. Ealy, “Sugar in a Can of Soft Drink: A Density Exercise,” *Chemical Demonstrations, A Sourcebook for Teachers*, Volume 2 (Washington: American Chemical Society, 1988), p. 126–127.

8. Lee. R. Summerlin, Christie L. Borgford, and Julie B. Ealy, “The Mysterious Sunken Ice Cube,” *Chemical Demonstrations, A Sourcebook for Teachers*, Volume 2 (Washington: American Chemical Society, 1988), p. 15–16.

9. Kathleen Holley, Diana Mason, and Kirk Hunte, “Bowling for Density!”, *J. Chem. Educ.*, Vol. 81, **2004**, 1312A.

10. Michael J. Sanger, “Whatever Floats (or Sinks) Your Can,” *J. Chem. Educ.*, Vol. 83, **2006**, 1632A.

11. Michael Davis and Charles Henry, “The Sweeter Side of Density,” *J. Chem. Educ.*, Vol. 85, **2008**, 1088A-1088B.