Chapter 1. Introduction: Matter, Energy, and Measurement

Media Resources

Important Figures and Tables:	Section:
Figure 1.4 The Three Physical States of Water —Water Vapor, Liquid Water, and Ice	1.2 Classifications of Matter
Figure 1.5 Molecular Comparison of Elements, Compounds, and Mixtures	1.2 Classifications of Matter
Figure 1.7 Electrolysis of Water	1.2 Classifications of Matter
Figure 1.9 Classification of Matter	1.2 Classifications of Matter
Figure 1.10 A Chemical Reaction	1.3 Properties of Matter
Figure 1.13 Distillation	1.3 Properties of Matter
Figure 1.18 The Scientific Method	1.5 Units of Measurement
Table 1.5 Prefixes Used in the Metric System and with SI Units	1.5 Units of Measurement
Figure 1.20 Volume Relationships	1.5 Units of Measurement
Figure 1.22 Precision and Accuracy	1.5 Units of Measurement

Interactive Sample Exercises:	Section:
1.1 Distinguishing among Elements, Compounds, and Mixtures	1.2 Classifications of Matter
1.2 Using SI Prefixes	1.5 Units of Measurement
1.5 Identifying and Calculating Energy Changes	1.5 Units of Measurement
1.6 Assigning Appropriate Significant Figures	1.6 Uncertainty in Measurement
1.8 Determining the Number of Significant Figures in a Calculated Quantity	1.6 Uncertainty in Measurement
1.13 Conversions Involving Density	1.7 Dimensional Analysis

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
Elementary My Dear Watson	1.2 Classifications of Matter
Elementary Riddles	1.2 Classifications of Matter
Cracks in the Periodic Table	1.2 Classifications of Matter
Elements Are Everywhere: A Crossword Puzzle	1.2 Classifications of Matter
What are Elements and Compounds?	1.2 Classifications of Matter
Classifying Matter: A Physical Model Using Paper Clips	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
Mentos and the Scientific Method	1.3 Properties of Matter
T-Shirt Chromatography: A Chromatogram You Can Wear	1.3 Properties of Matter
Visualizing Separations: How Shopping Can Be Useful for Introducing Chromatography	1.3 Properties of Matter
Having Fun with the Metric System	1.5 Units of Measurement
The Concept of Density	1.5 Units of Measurement
Method for Separating or Identifying Plastics	1.5 Units of Measurement
Meter Sticks in the Demonstrations of Error Measurement	1.6 Uncertainty in Measurement
Basic Skills of Scale Reading	1.6 Uncertainty in Measurement
Measuring with a Purpose: Involving Students in the Learning Process	1.6 Uncertainty in Measurement
Error, Precision, and Uncertainty	1.6 Uncertainty in Measurement
Precision and Accuracy in Measurements: A Tale of Four Graduated Cylinders	1.6 Uncertainty in Measurement
A Simple but Effective Demonstration for Illustrating Significant Figure Rules When Making Measurements and Doing Calculations	1.6 Uncertainty in Measurement
A Joke Based on Significant Figures	1.6 Uncertainty in Measurement
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Significant Figures: A Classroom Demonstration	1.6 Uncertainty in Measurement
The Box-and-Dot Method: A Simple Strategy for Counting Significant Figures	1.6 Uncertainty in Measurement
Expanded Dimensional Analysis: A Blending of English and Math	1.7 Dimensional Analysis
Appalachian Trail Problems	1.7 Dimensional Analysis
Learning Dimensional Analysis through Collaboratively Working Manipulatives	1.7 Dimensional Analysis

Live Demonstrations:	**Section:**
How to Make Chemistry Classroom Demonstrations and Experiments Safer	1.1 The Study of Chemistry
Making Water the Exciting Way: A Classroom Demonstration of Catalysis	1.1 The Study of Chemistry
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	1.5 Units of Measurement
Sugar in a Can of Soft Drink: A Density Exercise	1.5 Units of Measurement
The Mysterious Sunken Ice Cube	1.5 Units of Measurement
Densities and Miscibilities of Liquids and Liquid Mixtures	1.5 Units of Measurement
Bowling for Density	1.5 Units of Measurement
Whatever Floats (or Sinks) Your Can	1.5 Units of Measurement
The Sweeter Side of Density	1.5 Units of Measurement
Guessing the Number of Candies in the Jar—Who Needs Guessing?	1.7 Dimensional Analysis

Chapter 1. Introduction: Matter and Measurement

Common Student Misconceptions

- Students often confuse mass and weight.
- Students confuse power and energy.
- Students confuse heat with temperature.
- 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 pm = 1×10^{-12} m → 1 m = 1×10^{12} pm).
- Students often use precision and accuracy interchangeably.
- Students often use atoms and elements 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 measuring device.
- Students often cannot find exact numbers in calculations.
- Students confuse significant figures and decimal places in arithmetic manipulations.
- Students often round off too soon in calculations.
- Students often report results to as many figures as their calculator displays provide.
- 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. Thus they do not perform algebraic operations on both the number and the units.

Teaching Tips

- Incorporate "Strategies for Success" placed throughout the chapter into your lecture presentations.
- Practice makes perfect! Encourage students to do as many end-of-chapter exercises as possible, starting with exercises recommended under the "Learning Outcomes".
- 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.
- Stress the importance of quickly mastering the use of common metric prefixes and scientific notation. Incorporate various prefixes and scientific notation into as many problems as possible to give students continuous opportunity to practice these skills. Continue this practice throughout the course.
- Conversion involving commonly encountered metric prefixes needs frequent reinforcement; for example, 10^6 μg = 10^3 mg = 1 g = 10^{-3} kg. Emphasize that unit conversions will be widely used throughout the course and beyond.
- Emphasize that atoms and elements are not to be used interchangeably.
- Weave energy arguments into discussions of what drives chemical and physical processes.
- Expose students to the micro- and macro- representations of matter (e.g., Figure 1.4).
- Encourage students to perform simple calculations with numbers in scientific notation *without* the aid of a calculator.

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.

The Atomic and Molecular Perspective of Chemistry[1,2,3]

- Chemistry involves the study of the properties and the behavior of matter.
- **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).
  - For some elements, the basic unit is not an atom, but a molecule (e.g., O_2, S_8.)
- **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?[4,5,6]

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[7]

- Solids, liquids and gases are the three forms of matter called the **states of matter (phases)**.
- 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.

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1 "How to Make Chemistry Classroom Demonstrations and Experiments Safer" from Live Demonstrations
2 "The First Demonstration: Proof That Air Is a Substance" from Live Demonstrations
3 "Making Water the Exciting Way: A Classroom Demonstration of Catalysis" from Live Demonstrations
4 "Chemistry in the Real World" from Further Readings
5 "Science Demonstrations, Experiments, and Resources: A Reference List for Elementary through College Teachers Emphasizing Chemistry with Some Physics and Life Science" from Live Demonstrations
6 "Chemicals in Everyday Life" from Further Readings
7 Figure 1.4

Pure Substances[8]

- **Pure substances**:
  - are matter with distinct properties and fixed composition.
  - are **elements** (cannot be decomposed into simpler substances, i.e., consist of only one kind of atom) or **compounds** (consist of two or more elements).
- **Mixtures**[9]:
  - are a combination of two or more pure substances.
  - Each substance retains its own identity.

Elements[10,11,12,13,14,15,16,17,18,19,20]

- There are 118 known elements.
- They vary in abundance.
  - Hydrogen constitutes about 74% of the mass in the Milky Way galaxy and helium constitutes 24%.
  - Oxygen, silicon, aluminum, iron and calcium make up over 90% of 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*.
- Symbols may or may not correlate with the *English* names of elements.
- Each is given a one- or two-letter symbol derived from its name. The first letter is always capitalized, the second letter, if present, is always in lower case.
- They are organized in the *periodic table*.

Compounds[21,22]

- **Compounds** are combinations of elements.
  Example: The compound H_2O is a combination of the elements H and O.
- The opposite of compound formation is decomposition.
- Compounds have different properties than their component elements (e.g., water is liquid, but 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).

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8 Figure 1.5
9 Interactive Sample Exercise 1.1
10 "What Do Chemists Mean When They Talk about Elements?" from Further Readings
11 "What's the Use?" from Further Readings
12 "It's Elementary" from Further Readings
13 "Element ZOO" from Further Readings
14 "Important Elements" from Further Readings
15 "Origin of the Names of Chemical Elements" from Further Readings
16 "Connecting Element Names with the Names of U.S. Towns" from Further Readings
17 "Elementary My Dear Watson" from Further Readings
18 "Elementary Riddles" from Further Readings
19 "Cracks in the Periodic Table" from Further Readings
20 "Elements Are Everywhere: A Crossword Puzzle" from Further Readings
21 "What Are Elements and Compounds?" from Further Readings
22 Figure 1.7

Mixtures[23,24]

- A **mixture** is a combination of two or more pure substances.
  - Each substance retains its own identity, and 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 to proper writing of chemical reaction equations (including net ionic equations) in Chapter 4 and 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.

1.3 Properties of Matter

- Each substance has a unique set of physical and chemical properties.
  - **Physical properties** can be observed or described 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 *amount* of substance present.

Physical and Chemical Changes[25,26]

- **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 substance (i.e., identity changes, e.g., reaction of hydrogen and oxygen gases to produce water).

Separation of Mixtures[27,28,29]

- 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.

FORWARD REFERENCES

- Changes in the states of matter will be covered in Chapter 11 (phase diagrams, heating curves).
- The effects of intermolecular interactions on the properties of substances in different states of matter will be discussed in detail in Chapter 13.

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[23] Figure 1.9
[24] "Classifying Matter: A Physical Model Using Paper Clips" from Further Readings
[25] Figure 1.10
[26] "Ira Remsen's Investigation of Nitric Acid" from Live Demonstrations
[27] Figure 1.16
[28] "T-Shirt Chromatography: A Chromatogram You Can Wear" from Further Readings
[29] "Visualizing Separations: How Shopping Can Be Useful for Introducing Chromatography" from Further Readings

1.4 The Nature of Energy

- To understand chemistry, we need to also understand energy and the changes that happen during chemical processes.
- Definitions:
  - **Energy** is the capacity to do work or to transfer heat.
  - **Work** is energy used to cause an object with mass to move.

$$w = F \times d$$

  - **Heat** is the energy used to cause the temperature of an object to increase.
  - A **force** is any kind of push or pull exerted on an object.
    - The most familiar force is the pull of gravity.

Kinetic Energy and Potential Energy

- **Kinetic energy** is the energy of motion:

$$E_k = \frac{1}{2}mv^2$$

  - where E_k is the magnitude of the kinetic energy, m is the mass, and v is the velocity.
- **Potential energy** is the energy an object possesses by virtue of its position or composition.
  - *Electrostatic potential energy* (E_{el}) is an example.
    - It arises from interactions between charged particles.
    - The strength increases as the magnitude of the charges increases and decreases as the distance between charges increases.
- Many substances release energy when they react.
  - An example is a fuel.
  - The *chemical energy* of a fuel is due to the potential energy stored in the arrangement of its atoms.
  - Chemical energy is released when bonds between atoms are formed and is consumed when bonds between atoms are broken.

FORWARD REFERENCES

  - Energy transformations and the first law of thermodynamics will be introduced in Chapter 5.
  - The second law and the third law of thermodynamics will be covered in Chapter 19.

The Scientific Method[30,31,32,33,34,35]

- **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 that hypothesis, and 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.

1.5 Units of Measurement

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

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[30] "The Inquiry Wheel, an Alternative to the Scientific Method: A View of the Science Education Research Literature" from Further Readings
[31] "An Experiment to Demonstrate the Application of the Scientific Method" from Live Demonstrations
[32] "Why Does Popcorn Pop? An Introduction to the Scientific Method" from Further Readings
[33] "Mentos and the Scientific Method: A Sweet Combination" from Further Readings
[34] "Using History to Teach the Scientific Method: The Role of Errors" from Further Readings
[35] Figure 1.18

SI Units[36,37,38,39]

- 1960: All scientific units use **S**ystème **I**nternational d'Unités (**SI Units**).
- There are seven *base units*.
- Prefixes are used to indicate smaller (or larger) units obtained by taking decimal fractions (or multiples) of the base units.
    - For example, "*milli*-" represents a 10^{-3} fraction, so a milligram (mg) is 10^{-3} gram.
    - Common metric system prefixes are show in Table 1.4

Length and Mass[40]

- 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

- **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 of 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

- A **derived unit** is formed by multiplication or division of one of the seven base units.
- Example: *speed* is distance traveled per unit time, so the derived SI unit for speed is distance (m) divided by units of time (s): m/s or "meters per second."

Volume[41]

- 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 (*milliliters*) or cc (*cubic centimeters*))
    - dm^3 (also known as *liters*, L).
- Note**:** the liter is <u>not</u> an SI unit.

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[36] "A Simple Demonstration for Introducing the Metric System to Introductory Chemistry Classes" from Live Demonstrations
[37] "Having Fun with the Metric System" from Further Readings
[38] Table 1.5
[39] Interactive Sample Exercise 1.2
[40] "What Is a Kilogram in the Revised International System of Units (SI)?" from Further Readings
[41] Figure 1.20

Density[42,43,44,45,46,47,48]

- **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).

Units of Energy[49]

- SI unit is the **joule**, J.
- Joules are a derived unit:

$$E_k = \frac{1}{2}mv^2, \;\; 1 \text{ J} = 1 \text{ kg} \times \frac{\text{m}^2}{\text{s}^2}$$

- Traditionally, we use the **calorie** (cal) as a unit of energy.
  - 1 cal = 4.184 J (exactly)
- The nutritional Calorie, Cal = 1,000 cal = 1 kcal.

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 water and ice), and Chapter 13 (for solutions).
- Conversions of temperature units will be commonplace in Chapters 5 and 14–20.
- Energy in J will be used in Chapter 6 to express energy of a hydrogen atom and the energy of an emitted.
- Energy in kJ/mol will be used in Chapter 7 for ionization energies and electron affinities and in Chapter 8 to measure lattice energy and average bond enthalpies and to calculate enthalpies of reactions.
- Energy in J (and kJ) to energy of thermodynamic calculations throughout Chapters 19, 14 (Arrhenius equation in section 14.5), and Chapter 20 (Gibbs free energy vs. cell potential in section 20.5).

1.6 Uncertainty in Measurement[50,51,52]

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

Precision and Accuracy[53]

- **Precision**: how well-measured quantities agree with each other.
  - Often expressed in terms of *standard deviation*: how much the individual measurements differ from the average.
- **Accuracy**: how well-measured quantities agree with the "true value."
- Figure 1.22 is very helpful in making this distinction.

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[42] "Sugar in a Can of Soft Drink: A Density Exercise" from Live Demonstrations
[43] "The Mysterious Sunken Ice Cube" from Live Demonstrations
[44] "Method for Separating or Identifying Plastics" from Further Readings
[45] "Densities and Miscibilities of Liquids and Liquid Mixtures" from Live Demonstrations
[46] "The Concept of Density" from Further Readings
[47] "Bowling for Density!" from Live Demonstrations
[48] "Whatever Floats (or Sinks) Your Can" from Live Demonstrations
[49] Interactive Sample Exercise 1.5
[50] "Meter Sticks in the Demonstration of Error Measurement" from Further Readings
[51] "Basic Principles of Scale Reading" from Further Readings
[52] "Measuring with a Purpose: Involving Students in the Learning Process" from Further Readings
[53] Figure 1.22

Significant Figures[54,55,56,57,58,59,60]

- 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.)
  - Final calculations are only as significant as the least significant measurement.
- Rules:
  1. Nonzero numbers and zeros between nonzero numbers (i.e., imbedded zeros) are always significant.
  2. Zeros before the first nonzero digit (i.e., 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[61,62]

- 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 $\times$ 5.2 cm = 32 cm^2).
- In multiple-step calculations, always retain an extra significant figure until the end to prevent rounding errors.

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[54] "Error, Precision, and Uncertainty" from Further Readings
[55] "Precision and Accuracy in Measurements: A Tale of Four Graduated Cylinders" from Further Readings
[56] "A Simple but Effective Demonstration for Illustrating Significant Figure Rules When Making Measurements and Doing Calculations" from Further Readings
[57] "A Joke Based on Significant Figures" from Further Readings
[58] "Significant Figures" from Further Readings
[59] "Significant Figures: A Classroom Demonstration" from Further Readings
[60] Interactive Sample Exercise 1.6
[61] "The Box-and-Dot Method: A Simple Strategy for Counting Significant Figures" from Further Readings
[62] Interactive Sample Exercise 1.8

FORWARD REFERENCES

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

1.7 Dimensional Analysis[63,64,65,66]

- **Dimensional analysis** is a method of calculation utilizing a knowledge of units.
- Given units can be multiplied and divided to obtain the desired units.

Conversion Factors[67]

- **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.

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 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.

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[63] "Expanded Dimensional Analysis: A Blending of English and Math" from Further Readings
[64] "Appalachian Trail Problems" from Further Readings
[65] "Learning Dimensional Analysis through Collaboratively Working Manipulatives" from Further Readings
[66] "Guessing the Number of Candies in the Jar—Who Needs Guessing?" from Live Demonstrations
[67] Interactive Sample Exercise 1.13

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 in virtually every chapter.

Further Readings:

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

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