

Chapter 1: Chemistry and Measurement

Problems: 2, 4-5, 7-10, 14, 16, 17, 24-26, 27a, 28-44, 47-52, 55-56, 59-77, 79-88, 93-96, 105-112, 115-132

science: methodical study of nature followed by a logical explanation of the observations

Chemistry: study of matter and its properties, the changes that matter undergoes, and the energy associated with those changes

1.1 MODERN CHEMISTRY

In what fields is chemistry used today? And how is it used?

1.2 EXPERIMENT AND EXPLANATION

Consider the following:

On the first day of school, you get in your car and turn the key on the ignition, but nothing happens. What could be the problem?

The Scientific Method

1. Perform **experiments** and **record observations** on system
2. Analyze the data, and propose a **hypothesis** to explain observations
3. Conduct additional experiments to test hypothesis. If initial hypothesis holds up to extensive testing, the hypothesis becomes a **theory**.

theory: a tested qualitative explanation of basic natural phenomena

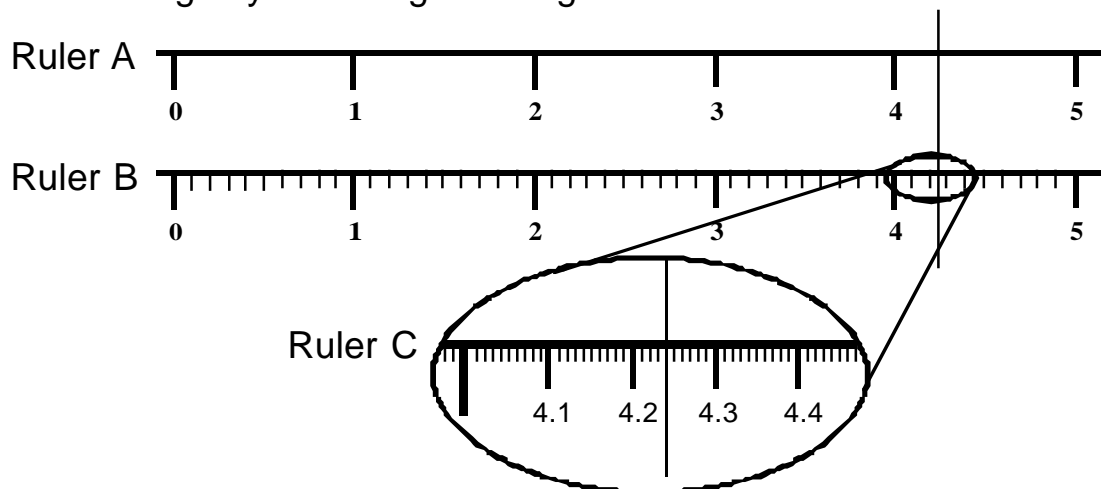
If all or part of the hypothesis does not hold up to testing, then it is adjusted or a new hypothesis is proposed to explain the observations.

law: a simple statement or mathematical equation about a fundamental relationship

1.5 MEASUREMENT AND SIGNIFICANT FIGURES

measurement: a number with attached units

When a measurement is recorded, all the numbers known with certainty are given along with the last number, which is estimated. All the *digits* are *significant* because removing any of the digits changes the measurement's uncertainty.



Ruler	Measurement/quantity	# of sig figs
A	_____	_____
B	_____	_____
C	_____	_____

Which ruler gives **the most exact measurement**? _____

Guidelines for Sig Figs (if measurement is given):

Count the number of digits in a measurement from **left to right**:

1. When there is a decimal point:

- For measurements **greater than 1**, count all the digits (even zeros).
 - 62.4 cm has 3 sig figs, 5.0 m has 2 sig figs, 186.000 g has 6 s.f.
- For measurements **less than 1**, start with the first nonzero digit and count all digits (even zeros) after it.
 - 0.011 mL and 0.00022 kg each have 2 sig figs

2. When there is no decimal point:

- Count all non-zero digits and zeros between non-zero digits
 - 125 g has 3 sig figs, 107 mL has 3 sig figs

How many significant digits do the following numbers have?

<u># of sig figs</u>		<u>of sig figs</u>		<u># of sig figs</u>	
a. 165.3	_____	c. 90.40	_____	e. 0.19600	_____
b. 105	_____	d. 100.00	_____	f. 0.0050	_____

SCIENTIFIC NOTATION

Some numbers are very large or very small \Rightarrow difficult to express.

For example,

Avogadro's number = 602,000,000,000,000,000,000

an electron's mass = 0.000 000 000 000 000 000 000 000 91 kg

Also, it's not clear how many sig figs there are in some measurements.

For example,

Express 100.0 g to 3 sig figs: _____

Express 100.0 g to 2 sig figs: _____

Express 100.0 g to 1 sig fig: _____

To handle such numbers, we use a system called **scientific notation**.

All numbers can be expressed in the form: $N \times 10^n$

where **$N = \text{digit term} = \text{a number between 1 and 10}$** , so there can only be one number to the left of the decimal point: **$\#.####$**

$n = \text{an exponent} = \text{a positive or a negative integer (whole \#)}$.

To express a number in scientific notation:

Count the number of places you must move the decimal point to get N between 1 and 10.

Moving decimal point to the **right (if $\# < 1$) \Rightarrow **negative exponent**.**

Moving decimal point to the **left (if $\# > 1$) \Rightarrow **positive exponent**.**

Express the following numbers in scientific notation (each with **3 sig figs**):

555,000 \Rightarrow _____

0.000888 \Rightarrow _____

602,000,000,000,000,000,000,000 \Rightarrow _____

ROUNDING OFF NONSIGNIFICANT DIGITS

How do we eliminate **nonsignificant digits**?

- If first nonsignificant digit < 5 , just drop **ALL nonsignificant digits**
- If first nonsignificant digit ≥ 5 , **raise the last sig digit by 1 then drop ALL nonsignificant digits**

For example, express **72.58643** with 3 sig figs:

last significant digit

72.58643 g

first nonsignificant digit

72.58643 $\xrightarrow{\text{to 3 sig figs}}$ _____

Express each of the following with the number of sig figs indicated:

a. 376.276 $\xrightarrow{\text{to 3 sig figs}}$ _____

b. 500.072 $\xrightarrow{\text{to 4 sig figs}}$ _____

c. 0.00654321 $\xrightarrow{\text{to 3 sig figs}}$ _____

d. 1,234,567 $\xrightarrow{\text{to 5 sig figs}}$ _____

e. 2975 $\xrightarrow{\text{to 2 sig figs}}$ _____

Express measurements in scientific notation when necessary to make it clear how many sig figs there are in the measurement.

ADDING/SUBTRACTING MEASUREMENTS

When adding and subtracting measurements, your final value is limited by *measurement with the largest uncertainty*—i.e. the number with the fewest **decimal places**.

Ex 1: $106.61 + 0.25 + 0.195 = 107.055$

⇒ 107.055 to the correct number of sig figs: _____

Ex 2: $725.50 - 103 = 622.50$

⇒ 622.50 to the correct # of sig figs: _____

MULTIPLYING/DIVIDING MEASUREMENTS

When multiplying or dividing measurements, the final value is limited by the *measurement with the least number of significant figures*.

Ex 1: $106.61 \times 0.25 \times 0.195 = 5.1972375$

⇒ 5.1972375 to the correct # of sig figs: _____

Ex 2: $106.61 \times 91.5 = 9754.815$ w/ correct sig figs: _____

MULTIPLYING/DIVIDING WITH EXPONENTIAL NUMBERS:

When multiplying or dividing measurements with exponents, use the digit term (N in " $N \times 10^n$ ") to determine number of sig figs.

Ex. 1: $(6.02 \times 10^{23})(4.155 \times 10^9) = 2.50131 \times 10^{33}$

How do you calculate this using your scientific calculator?

Step 1. Enter " 6.02×10^{23} " by pressing:

6.02 then **EE** or **EXP** (which corresponds to " $\times 10$ ") then **23**

⇒ Your calculator should read or something similar.

Step 2. Multiply by pressing: **×**

Step 3. Enter "4.155 × 10⁹" by pressing:

4.155 then EE or EXP (which corresponds to "× 10") then 9

⇒ Your calculator should now read

Step 4. Get the answer by pressing: =

⇒ Your calculator should now read: 2.50131³³ or 2.50131 E 33
or something similar indicating 2.50131×10³³

Thus, the answer with the correct sig figs = _____

Be sure you can do exponential calculations with your calculator. Many of the calculations we do in chemistry involve very large and very small numbers with exponential terms.

Ex. 2: $(3.75 \times 10^{15})(8.6 \times 10^4) = 3.225 \times 10^{20}$
with the correct sig figs: _____

Ex. 3: $(3.75 \times 10^{15}) \div (8.605 \times 10^4) = 4.357931435 \times 10^{10}$
with the correct sig figs: _____

EXACT NUMBERS

Although measurements can never be exact, we *can* count an exact number of items. For example, we can count exactly how many students are present in a classroom, how many M&Ms are in a bowl, how many apples in a barrel.

1.8 UNITS AND DIMENSIONAL ANALYSIS (FACTOR-LABEL METHOD)

UNIT EQUATIONS AND UNIT FACTORS

Unit equation: Simple statement of two equivalent values

Unit conversion factor = unit factor = equivalents:

- Ratio of two equivalent quantities

Unit equation

1 dollar = 10 dimes

Unit factor

$\frac{1 \text{ dollar}}{10 \text{ dimes}}$ or $\frac{10 \text{ dimes}}{1 \text{ dollar}}$

Equivalents are **exact** if we can count the number of units equal to another or the units are in the same system (metric or English). For example, the following unit factors and unit equations are exact:

$$\frac{365 \text{ days}}{1 \text{ year}} \quad \frac{7 \text{ days}}{1 \text{ week}} \quad \frac{12 \text{ inches}}{1 \text{ foot}} \quad \text{and} \quad 1 \text{ yard} \equiv 3 \text{ feet}$$

Exact equivalents have an *infinite* number of sig figs
⇒ **never** limit number of sig figs!

Note: When the relationship between two units or items is exact, the “≡” (meaning “equals exactly”) is used instead of the basic “=” sign.

Other equivalents are **inexact** or **approximate** because they are measurements or approximate relationships, such as

$$\frac{1.61 \text{ km}}{1 \text{ mile}} \quad \frac{65 \text{ mi}}{\text{hour}} \quad \text{speed of light} = \frac{3.00 \times 10^8 \text{ m}}{\text{s}}$$

Approximate equivalents **do limit** the **sig figs for the final answer**.

UNIT (DIMENSIONAL) ANALYSIS PROBLEM SOLVING

1. Write the units for the answer.
2. Determine what information to start with.
3. Arrange all other unit factors—showing them as fractions—with correct units in the numerator and denominator, so all units cancel except for the units needed for the final answer.
4. Check for correct units and number of sig figs in the final answer.

Example 1: If a marathon is 26.2 miles, how many inches are in the marathon?
(1 mile = 5280 feet)

Example 2: The speed of light is about 3.00×10^8 meters per second. Express this speed in miles per hour. (1.609 km = 1 mile, 1000 m = 1km)

1.6 SI Units (from French "le Système International d'Unités")

– standard units for measurement

Metric system: A unified *decimal system of measurement* with a basic unit for each type of measurement

quantity	basic unit	symbol
length	meter	m
mass	gram	g
volume	liter	L
time	second	s

Metric Prefixes

Multiples or fractions of a basic unit are expressed as a prefix

⇒ Each prefix = power of 10

⇒ *The prefix increases or decreases the base unit by a power of 10.*

Prefix	Symbol	Multiple/Fraction
mega	M	10^6
kilo	k	10^3
deci	d	$0.1 = 10^{-1}$
centi	c	$0.01 = 10^{-2}$
milli	m	$0.001 = 10^{-3}$
micro	μ (Greek "mu")	10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}

KNOW these metric units above (highlighted in Table 1.3)!

Metric Conversion Factors

Ex. 1 Complete the following unit equations:

a. $1 \text{ kg} = \underline{\hspace{2cm}} \text{ g}$ c. $1 \text{ cm} = \underline{\hspace{2cm}} \text{ m}$

b. $1 \text{ s} = \underline{\hspace{2cm}} \text{ ns}$ d. $1 \text{ L} = \underline{\hspace{2cm}} \text{ mL}$

Ex. 2 Write 2 unit factors for each of the unit equations above.

3.3 Metric-Metric Conversions: Solve the following using dimensional analysis.

Ex. 1 Convert 75 miles per hour into units of meters per second.

Ex. 2 Convert 12.0 kilograms into milligrams.

Metric-English Conversions

English system: Our general system of measurement.

Scientific measurements are exclusively metric. However, most Americans are more familiar with inches, pounds, quarts, and other English units.

⇒ A method of conversion between the two systems is necessary.

Know these conversions!

Quantity	English unit	Metric unit	English–Metric conversion
length	1 inch (in)	1 cm	1 in. = 2.54 cm (exact)
mass	1 pound (lb)	1 g	1 lb = 453.6 g (approximate)
volume	1 quart (qt)	1 mL	1 qt = 946.4 mL (approximate)

All other metric-English conversions will be provided on quizzes and exams.

Ex. 1 What is the mass in kilograms of a person weighing 215 lbs?

Ex. 2 What is the volume in cups for a 2.0-L bottle?

Ex. 3 A light-year ($\sim 5.88 \times 10^{12}$ miles) is the distance light travels in one year. Calculate the speed of light in meters per second. (1 mile=1.61 km approx.)

Temperature:

— A measure of the average energy of a single particle in a system.

The instrument for measuring temperature is a ***thermometer***.

Temperature is generally measured with these units:

Fahrenheit degree (°F) Celsius degree (°C)

References	English system	Metric system
freezing pt for water	32°F	0°C
boiling pt for water	212°F	100°C
Phoenix Summer day	113°F	45°C

Conversion between Fahrenheit and Celsius scales:

$$^{\circ}\text{C} = \frac{(^{\circ}\text{F} - 32)}{1.8}$$

$$^{\circ}\text{F} = (^{\circ}\text{C} \times 1.8) + 32$$

- Ex. 1 In Europe, temperature is generally reported in the Celsius scale. A European tourist asks you how hot it is. What is the equivalent Celsius temperature if on that cool Summer day it's only 95°F?

Kelvin Temperature Scale

There is a third scale for measuring temperature: **the Kelvin scale**.

The unit for temperature in the Kelvin scale is **Kelvin (K, NOT °K!)**.

The Kelvin scale assigns a value of zero kelvins (0 K) to the lowest possible temperature, which we call **absolute zero** and corresponds to **-273.15°C**.

(The term **absolute zero** is used because this is the theoretical lowest limit.)

Conversion between °C and K:

$$K = ^\circ C + 273.15$$

$$^\circ C = K - 273.15$$

- Ex. 1 The average temperature in Reykjavik, Iceland during the summer is about 15°C. What is the equivalent Kelvin temperature? What is the equivalent Fahrenheit temperature?
- Ex. 2 In the movie *Terminator 2*, a tanker crashes and pours out liquid nitrogen. This freezes the T-1000 because liquid nitrogen has a temperature of 77 K. What is the equivalent temperature in degrees Celsius?

1.7 DERIVED UNITS

Volume: Amount of space occupied by a solid, gas, or liquid.

- measured using graduated cylinder, a buret, a pipet, a volumetric flask, etc.
- generally in units of liters (L), milliliters (mL), or cubic centimeters (cm³)
- **Know the following:**

$$1 \text{ L} = 1 \text{ dm}^3 \qquad 1 \text{ mL} = 1 \text{ cm}^3$$

These are both exact!

DETERMINING VOLUME

Volume is determined in three principal ways:

1. Volume of any liquid can be measured directly using calibrated glassware (graduated cylinder, pipets, burets, etc.)
2. **Volume** of a solid with a regular shape (rectangular, cylindrical, uniformly spherical or cubic, etc.) can be determined **by calculation**.
 - e.g. volume of rectangular solid = length x width x thickness
3. Volume of an irregular solid is found indirectly by the amount of liquid it displaces. This technique is called **volume by displacement**.

Volume By Displacement

- a. Fill a graduated cylinder halfway with water, and record the initial volume.
- b. Carefully place the object into the graduated cylinder so as not to splash or lose water.
- c. Record the final volume.
- d. **Volume of object = final volume – initial volume**

density: The amount of mass in a unit volume of matter

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad \text{or} \quad d = \frac{m}{V} \quad \text{generally in units of } \mathbf{g/cm^3} \text{ or } \mathbf{g/mL}$$

For water: 1.00 g of water occupies a volume of 1.00 cm³

$$d = \frac{m}{V} = \frac{1.00 \text{ g}}{1.00 \text{ cm}^3} = \mathbf{1.00 \text{ g/cm}^3}$$

Applying Density as a Unit Factor

Given the density for any matter, you can always write two unit factors. For example, the density of ice is 0.917 g/cm³.

Two unit factors would be:

$$\frac{0.917 \text{ g}}{\text{cm}^3} \quad \text{or} \quad \frac{\text{cm}^3}{0.917 \text{ g}}$$

Ex. 1 Aluminum has a density of 2.70 g/cm³. What is the mass of a piece of aluminum with a volume of 0.525 cm³?

Ex. 2 Ethanol is used in alcoholic beverages and has a density of 0.789 g/mL. What volume of ethanol (in liters) would have a mass of 500.0 mg?

1.4 MATTER: PHYSICAL STATE AND CHEMICAL COMPOSITION

Matter exists in one of three **physical states**: solid, liquid, gas (See Fig. 1.11)

solid: Has definite shape, fixed volume

- At molecular level, particles are packed closely, in a fixed position

liquid: Fixed volume, but shape can change—takes shape of container

- Particles are packed closely together but can move past each other

gas (or vapor): Volume is variable, particles are widely spaced

⇒ If volume expands, particles move apart

 If volume decreases, particles, move closer together

⇒ Takes shape of container

ELEMENTS, COMPOUNDS, AND MIXTURES

PHYSICAL & CHEMICAL PROPERTIES

Physical Properties:

physical state (solid, liquid, gas)

color

density

melting and boiling points

electrical & heat conductivity

solubility

hardness

odor

Chemical Properties: how a substance reacts with other substances

PHYSICAL & CHEMICAL CHANGES

physical change:

- a process that does not alter the chemical composition of a substance
 - eg. changing shape, changing physical state, dissolving
 - eg. boiling water, melting ice, hammering gold into foil

chemical change or reaction:

- a process that changes the chemical composition (and thus the chemical formula) of starting materials (reactants)
 - eg. oxidation of matter (burning or rusting)
 - release of gas bubbles (fizzing)
 - formation of insoluble solid (precipitation)
 - release of heat or light

pure substances:

- Matter having constant composition, definite and consistent properties

Two types of pure substances:

elements:

- consist of only one type of atom
- cannot be broken down by chemical reaction
- eg. carbon (C), hydrogen (H₂), sulfur (S₈), copper wire (Cu)

compounds:

- consist of more than one type of atom and *has a specific formula*
- can be broken down by chemical reaction
- eg. ethanol (C₂H₅OH) can be broken down to C, H, & O

Two or more pure substances can combine to form **mixtures:**

mixtures:

- consist of various compounds and/or elements, with *no specific formula*
- Matter having variable composition with either definite or varying properties depending on the sample
- can be broken down into individual components
- eg. Any alloy like brass, steel, 10-K to 18-K gold;
course mixtures like sea water, carbonated soda, salt and iron fillings; air consists mostly of nitrogen and oxygen gases

Also, be able to identify elements, compounds, and mixtures given images at the atomic/molecular level. (See Concept Check 1.1 on p. 14)