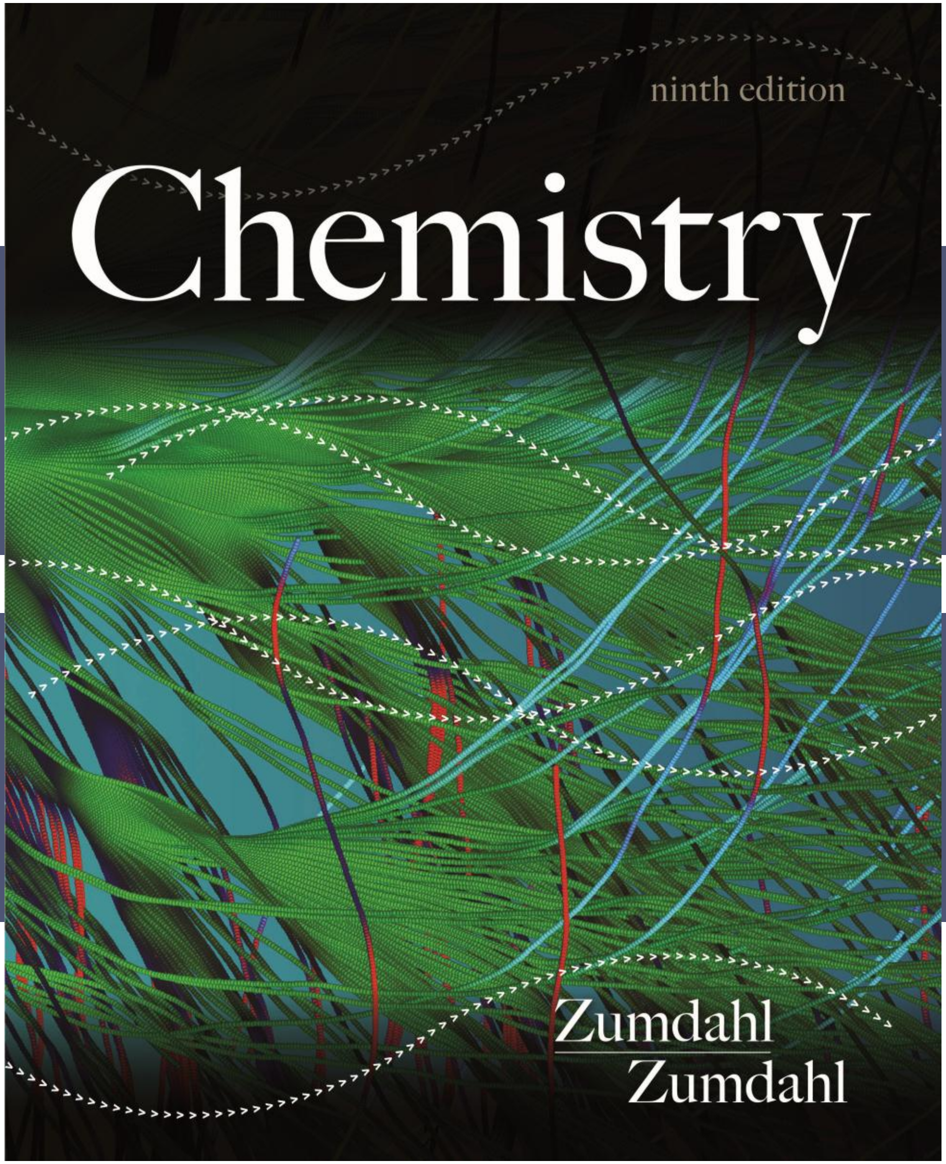


Chapter 1

Chemical Foundations

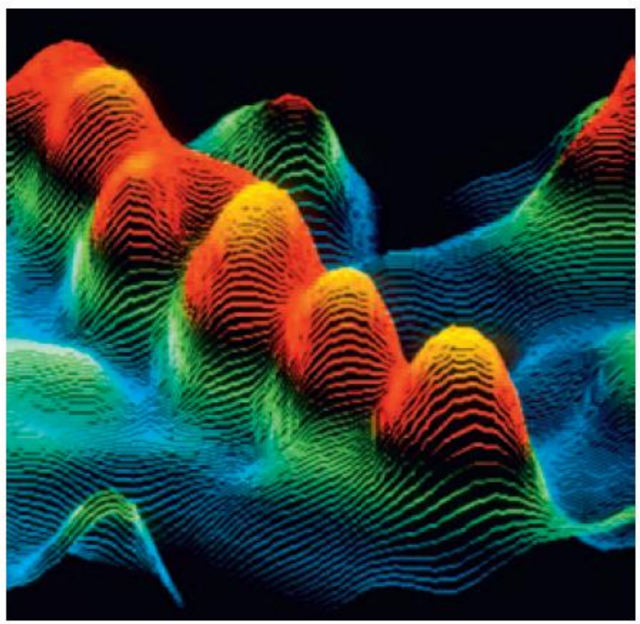
Table of Contents

Chemistry - Introduction

- Matter is believed to be composed of atoms
- Recent development
  - One can view individual atoms using a scanning tunneling microscope (STM)
    - STM - Uses an electron current from a tiny needle to probe the surface of a substance

Scanning Tunnelling Microscope (STM) Images

- Depict bridges (electrons) that connect atoms



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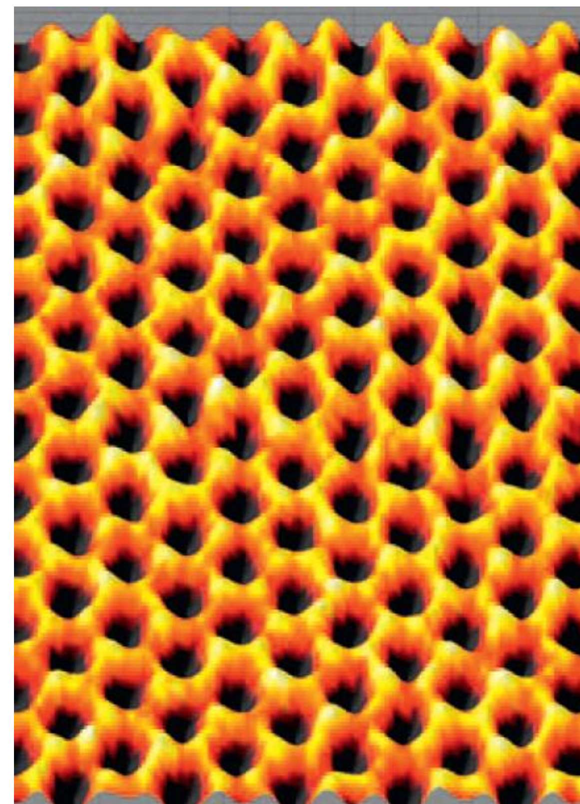

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Chemistry - Complexity and Challenges

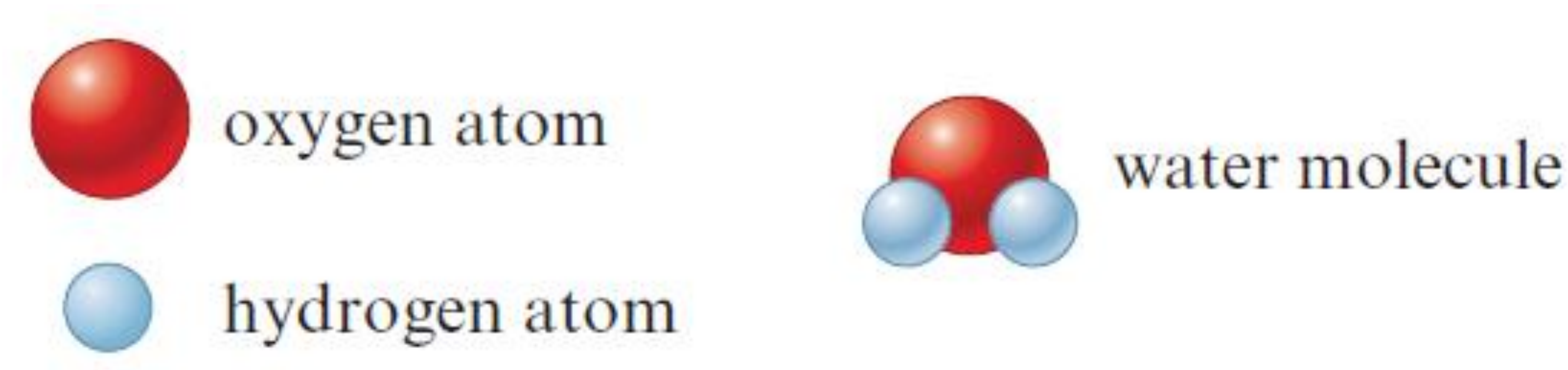
- Nature of atoms is complex
  - Atoms do not behave like objects one views in the macroscopic world
- Challenge
  - To understand the connection between the macroscopic world and the microscopic world of atoms and molecules
    - Requires an individual to learn to think on the atomic level

Critical Thinking

- The scanning tunneling microscope allows us to see atoms
  - What if you were sent back in time before the invention of the scanning tunneling microscope?
  - What evidence could you give to support the theory that all matter is made of atoms and molecules?

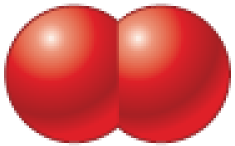
Configuration of Atoms - Example

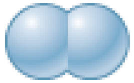
- Manner in which atoms are organized helps determine properties of a substance
  - Water is composed of hydrogen and oxygen atoms



Configuration of Atoms - Example (Continued 1)

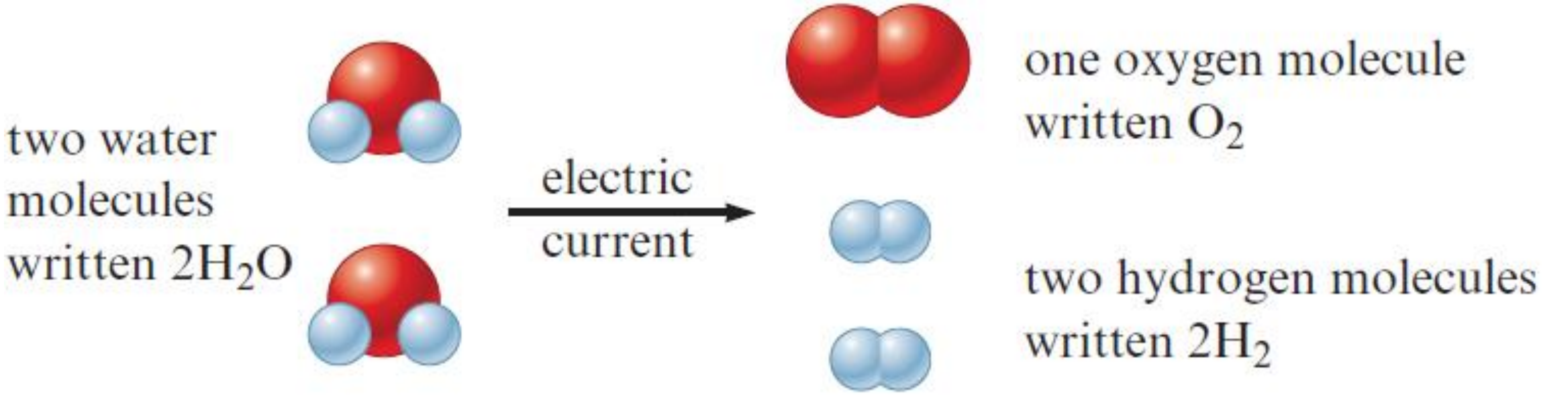
- When current is passed through water, it decomposes to hydrogen and oxygen
  - Both chemical elements exist naturally as diatomic molecules

oxygen molecule  written O_2

hydrogen molecule  written H_2

Configuration of Atoms - Example (Continued 2)

- Decomposition of water into hydrogen and oxygen can be represented as follows:



Fundamental Concepts of Chemistry

- Matter is composed of various types of atoms
- By reorganizing the way the atoms are attached to each other, one substance can be changed to another

Science

- Process for understanding nature and its changes
- Framework for gaining and organizing knowledge
- Procedure for processing and understanding certain types of information
- **Scientific method**: Process that lies at the center of scientific inquiry
  - Methods vary depending on the nature of a scientific problem

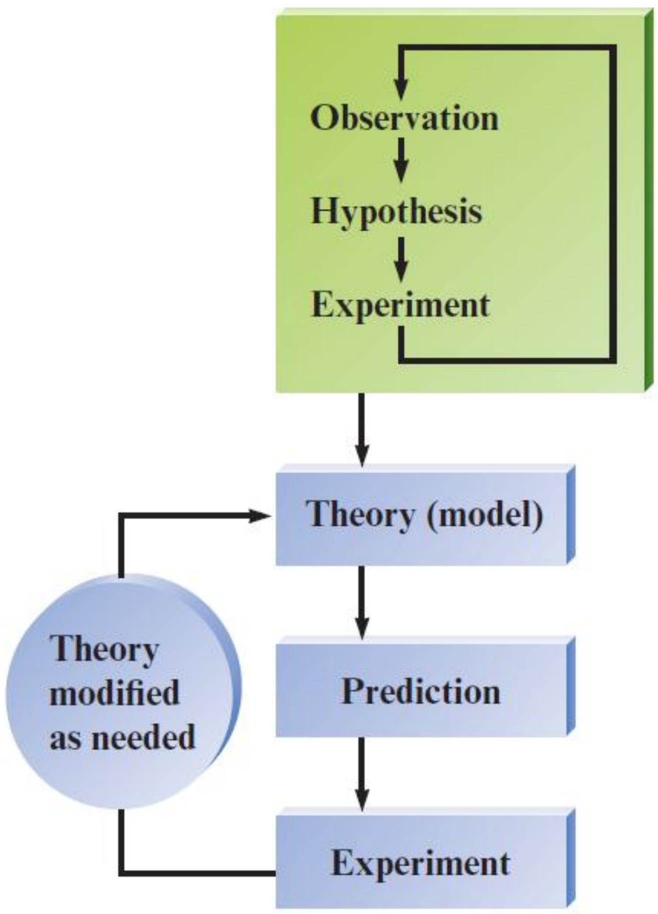
Steps in the Scientific Method

- Make observations
  - Qualitative observations do not involve numbers
  - Quantitative observations (**measurements**) involve numbers and units
- Formulate a hypothesis
  - **Hypothesis**: Possible explanation for an observation

Steps in the Scientific Method (Continued)

- Perform experiments to test the hypothesis
  - Gather new information to test the validity of the hypothesis
  - Help produce new observations
    - Brings the process back to the beginning

Figure 1.3 - Fundamental Steps of the Scientific Method



Scientific Models

- **Theory** (**model**): Set of tested hypotheses that gives an overall explanation of a natural phenomenon
  - Interpretation of why nature behaves in a certain manner
  - Changes with availability of more information
  - Requires constant revision and refining if one hopes to understand nature completely

Scientific Models (Continued 1)

- Observations
  - Events that are witnessed and can be recorded
- **Natural law**: A summary of repeatable observed (measurable) behavior
  - Example - Law of conservation of mass
    - **Law of conservation of mass**: Total mass of materials is unaffected by a chemical change in those materials

Scientific Models (Continued 2)

- Law versus theory
  - Law summarizes what happens
  - Theory attempts to explain why something happens
- Drawbacks
  - Focusing on a theory may limit the ability to see alternative explanations
  - Scientists are humans, and humans have prejudices

Critical Thinking

- What if everyone in the government used the scientific method to analyze and solve society's problems, and politics were never involved in the solutions?
  - How would this be different from the present situation, and would it be better or worse?

Nature of Measurement

- Measurement consists of a number and a scale (unit)
  - Both elements are needed for a measurement to be meaningful
- Systems of measurement
  - English system (used in the United States)
  - Metric system
    - **SI system**: Based on the metric system and units are derived from the metric system

Table 1.1 - Fundamental SI Units

Physical Quantity	Name of Unit	Abbreviation
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Electric current	ampere	A
Amount of substance	mole	mol
Luminous intensity	candela	cd

Table 1.2 - Prefixes Used in the SI System

Prefix	Symbol	Meaning	Exponential Notation
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
—	—	1	10^{0}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

Table 1.3 - Some Examples of Commonly Used Units

Length	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
Mass	A nickel has a mass of about 5 g. A 120-lb person has a mass of about 55 kg.
Volume	A 12-oz can of soda has a volume of about 360 mL.

Figure 1.5 - Units of Volume

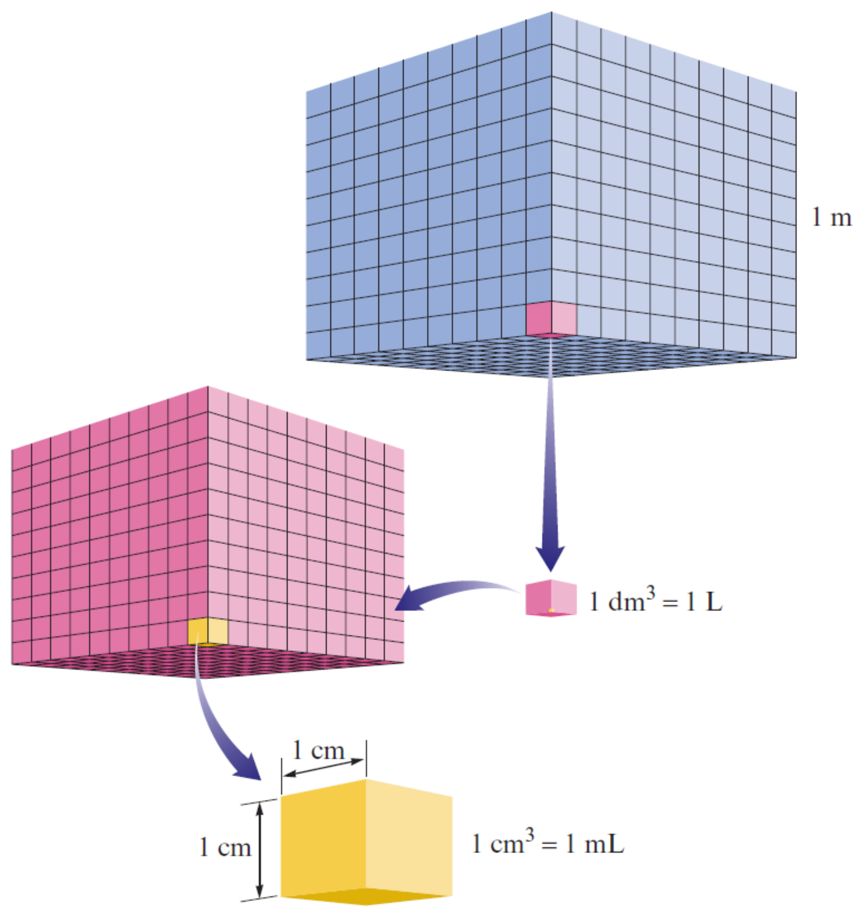
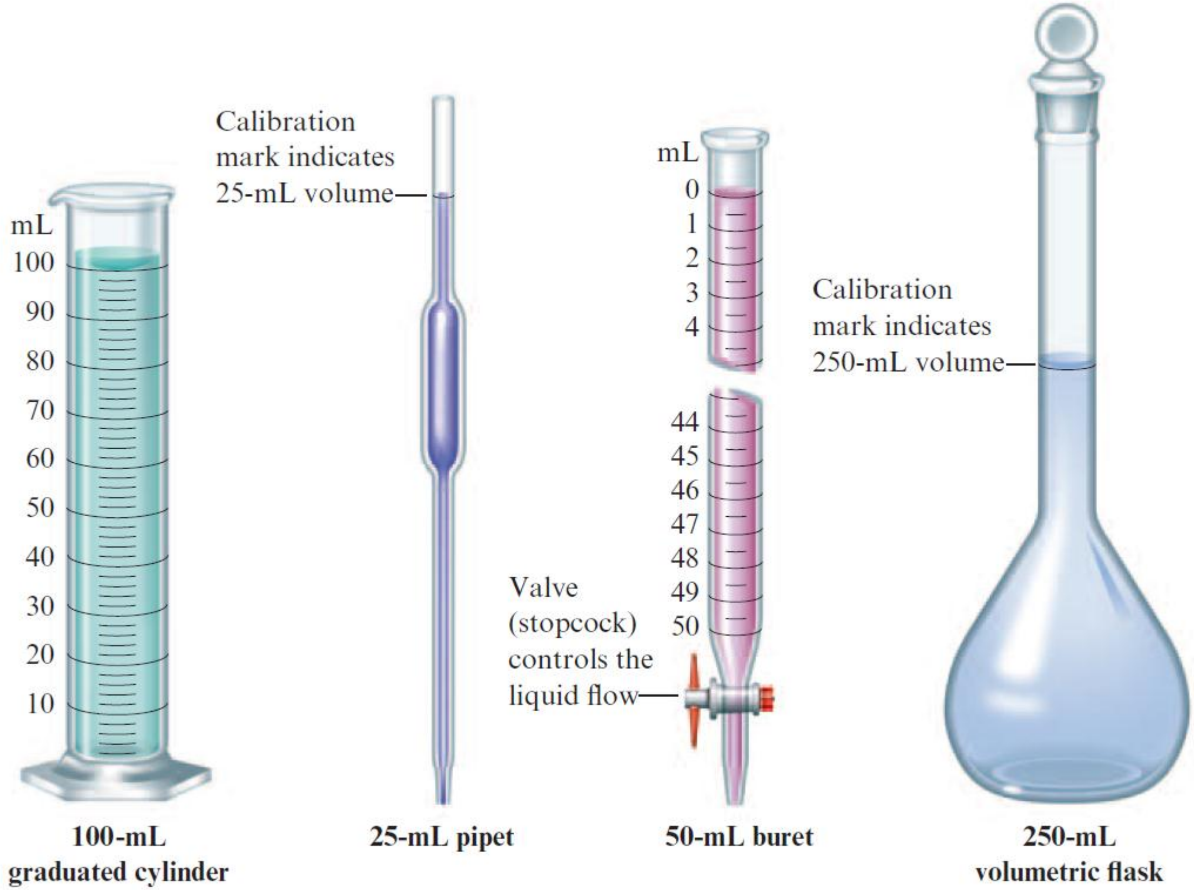


Figure 1.6 - Common Types of Laboratory Equipment Used to Measure Liquid Volume



Mass and Weight

- **Mass**: Measure of the resistance of an object to a change in its state of motion
  - Measured by the force essential to provide an object a certain acceleration
- **Weight**: Force exerted by gravity on an object
  - Varies with the strength of the gravitational field

Critical Thinking

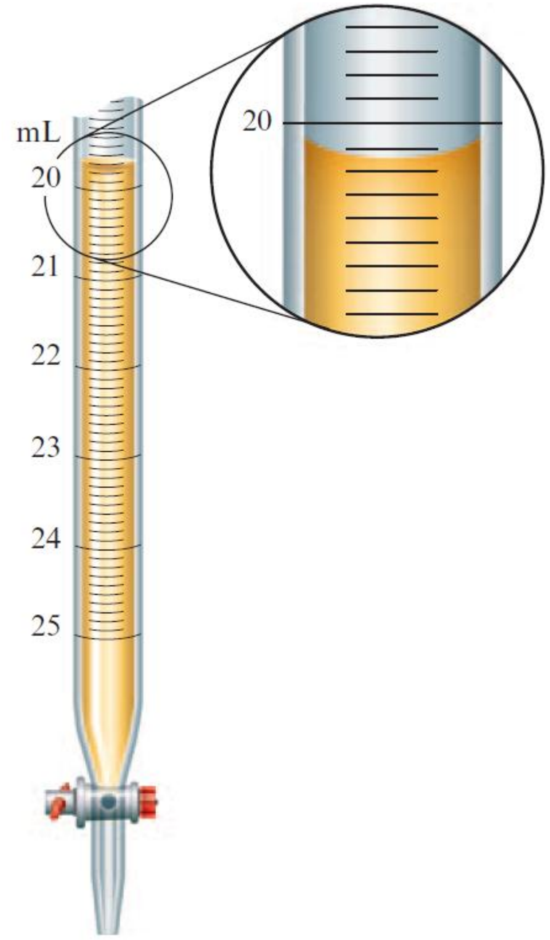
- What if you were not allowed to use units for one day?
  - How would this affect your life for that day?

Certain and Uncertain Digits

- Certain digits
  - Numbers that remain the same regardless of who makes a measurement
- Uncertain digits
  - Digits that must be estimated and therefore vary
- While reporting a measurement, record all certain digits plus the first uncertain digit

Measurement of Volume Using a Buret

- The volume is read at the bottom of the liquid curve (meniscus)
- Meniscus of the liquid occurs at about 20.15 mL
  - Certain digits - **<u>20</u>**.15
  - Uncertain digit - 20.<u>1</u>5



Degree of Uncertainty

- A measurement always has some degree of **uncertainty**
  - Dependent on the precision of the measuring device
- **Significant figures**: Numbers in which the certain digits and the first uncertain digit is recorded
  - Indicate the uncertainty in a measurement, which is always assumed to be ±1 unless indicated otherwise

Example 1.1 - Uncertainty in Measurement

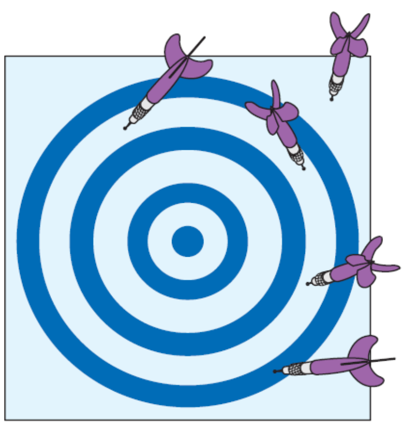
- In analyzing a sample of polluted water, a chemist measured out a 25.00-mL water sample with a pipet
  - At another point in the analysis, the chemist used a graduated cylinder to measure 25 mL of a solution
  - What is the difference between the measurements 25.00 mL and 25 mL?

Example 1.1 - Solution

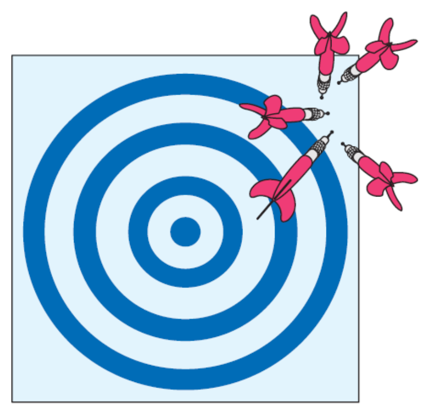
- Even though the two volume measurements appear to be equal, they convey different information
  - The quantity 25 mL means that the volume is between 24 mL and 26 mL, whereas the quantity 25.00 mL means that the volume is between 24.99 mL and 25.01 mL
  - The pipet measures volume with much greater precision than does the graduated cylinder

Precision and Accuracy

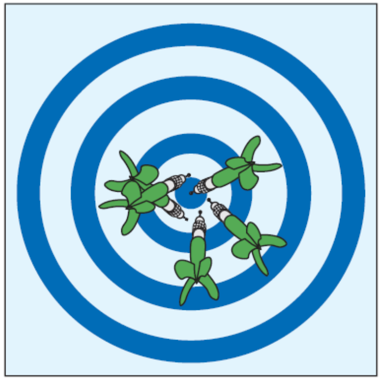
- **Accuracy**: Agreement of a particular value with the true value
- **Precision**: Degree of agreement among several measurements of the same quantity



Neither accurate nor precise



Precise but not accurate

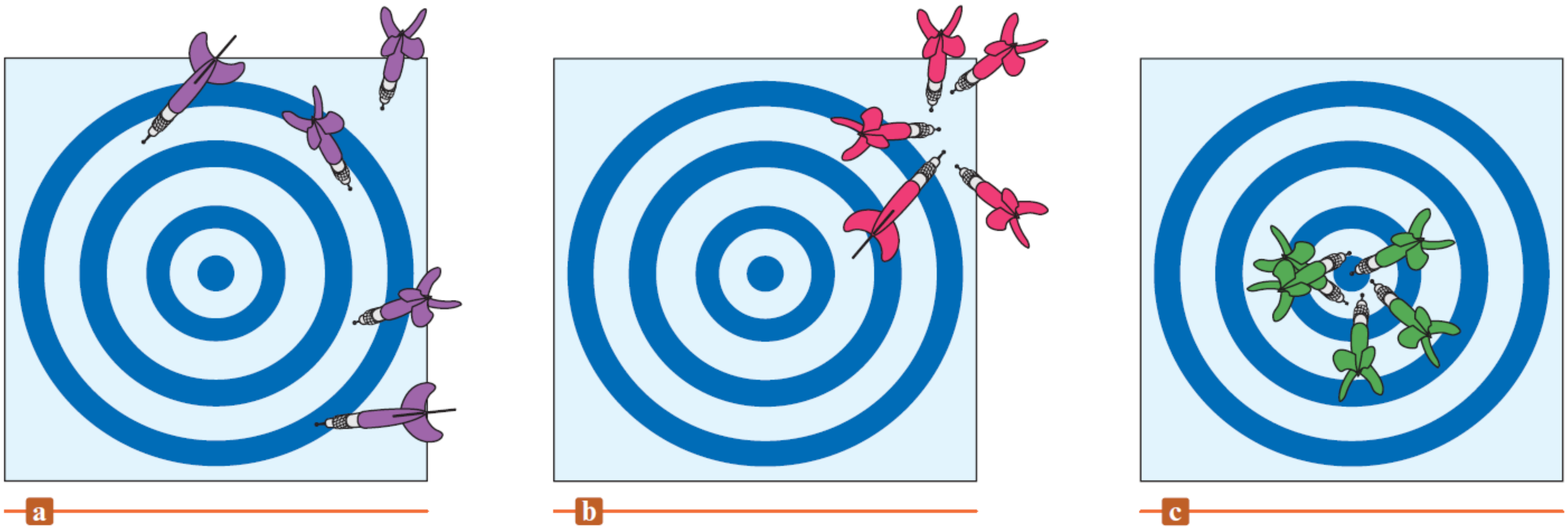


Accurate and precise

Types of Errors

- **Random error** (intermediate error)
  - Measurement has an equal probability of being low or high
  - Occurs while estimating the value of the last digit of a measurement
- **Systematic error** (determinate error)
  - Occurs in the same direction each time
  - Either always high or always low

Types of Errors - Example



Large random errors

Small random errors and a large systematic error

Small random errors and no systematic error

Example 1.2 - Precision and Accuracy

- To check the accuracy of a graduated cylinder, a student filled the cylinder to the 25-mL mark using water delivered from a buret and then read the volume delivered

Example 1.2 - Precision and Accuracy (Continued)

- Following are the results of five trials:

Trial	Volume Shown by Graduated Cylinder	Volume Shown by the Buret
1	25 mL	26.54 mL
2	25 mL	26.51 mL
3	25 mL	26.60 mL
4	25 mL	26.49 mL
5	25 mL	26.57 mL
Average	25 mL	26.54 mL

- Is the graduated cylinder accurate?

Example 1.2 - Solution

- The results of the trials show good precision (for a graduated cylinder)
  - The student has good technique
  - Note that the average value measured using the buret is significantly different from 25 mL
    - Thus, this graduated cylinder is not very accurate
    - It produces a systematic error (in this case, the indicated result is low for each measurement)

Rules for Counting Significant Figures

- Nonzero integers - Always count as significant figures
  - Example - 3456 has 4 sig figs (significant figures)

Rules for Counting Significant Figures (Continued 1)

- Zeros - Classes
  - Leading zeros - Zeros that precede all the nonzero digits
    - Do not count as significant figures
    - Example - 0.048 has 2 sig figs
  - Captive zeros - Zeros between nonzero digits
    - Always count as significant figures
    - Example - 16.07 has 4 sig figs

Rules for Counting Significant Figures (Continued 2)

- Trailing zeros - Zeros at the right end of the number
  - Significant only if the number contains a decimal point
  - Examples - 9.300 has 4 sig figs, and 150 has 2 sig figs

Rules for Counting Significant Figures (Continued 3)

- Exact numbers
  - Determined by counting and not by using a measuring device
  - Assumed to have an infinite number of significant figures
  - Examples
    - 1 inch = 2.54 cm, exactly
    - 9 pencils (obtained by counting)

Exponential Notation

- Example
  - 300 can be written as 3.00×10^2, and it contains three significant figures
- Advantages
  - Number of significant figures can be easily indicated
  - Fewer zeros are required to write a very large or very small number

Interactive Example 1.3 - Significant Figures

- Give the number of significant figures for each of the following results:
  a. A student's extraction procedure on tea yields 0.0105 g of caffeine
  b. A chemist records a mass of 0.050080 g in an analysis
  c. In an experiment a span of time is determined to be 8.050×10^{-3} s

Interactive Example 1.3 - Solution

a. The number contains three significant figures
   - The zeros to the left of the 1 are leading zeros and are not significant, but the remaining zero (a captive zero) is significant

b. The number contains five significant figures
   - The leading zeros (to the left of the 5) are not significant

Interactive Example 1.3 - Solution (Continued)

- The captive zeros between the 5 and the 8 are significant, and the trailing zero to the right of the 8 is significant because the number contains a decimal point

c. This number has four significant figures

- Both zeros are significant

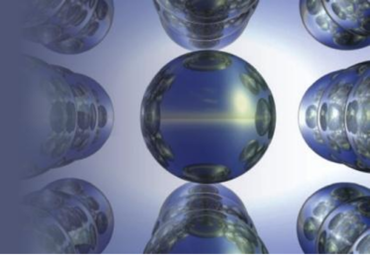
Rules for Significant Figures in Mathematical Operations

- Multiplication or division
  - Number of significant figures in the result is the same as the number in the least precise measurement used in the calculation
  - Example

$$4.56 \times 1.4 = 6.38 \xrightarrow{\text{corrected}} 6.4$$

Limiting term has two significant figures

Two significant figures



Rules for Significant Figures in Mathematical Operations (Continued)

- Addition or subtraction
  - Result has the same number of decimal places as the least precise measurement used in the calculation
  - Example

$$\begin{array}{r} 12.11 \\ 18.0 \\ 1.013 \\ \hline 31.123 \end{array} \xrightarrow{\text{corrected}} 31.1$$

18.0 ← Limiting term has one decimal place

31.1 ↑ One decimal place

Rules for Rounding

- In a series of calculations, round off only after carrying the extra digits through to the final result
- If the digit to be removed is:
  - Less than 5 - Preceding digit stays the same
    - Example - 1.33 rounds to 1.3
  - Greater than or equal to 5 - Preceding digit is increased by 1
    - Example - 1.36 to 1.4
- Do not round sequentially

Interactive Example 1.4 - Significant Figures in Mathematical Operations

- Carry out the following mathematical operations, and give each result with the correct number of significant figures
  a. $1.05 \times 10^{-3} \div 6.135$
  b. $21 - 13.8$

Interactive Example 1.4 - Significant Figures in Mathematical Operations (Continued)

c. As part of a lab assignment to determine the value of the gas constant (*R*), a student measured the pressure (*P*), volume (*V*), and temperature (*T*) for a sample of gas, where

$$R = \frac{PV}{T}$$

- The following values were obtained:
  - P = 2.560
  - T = 275.15
  - V = 8.8
- Calculate *R* to the correct number of significant figures

Interactive Example 1.4 - Solution

a. The result is 1.71×10^{-4}, which has three significant figures because the term with the least precision (1.05×10^{-3}) has three significant figures
b. The result is 7 with no decimal point because the number with the least number of decimal places (21) has none

Interactive Example 1.4 - Solution (Continued 1)

c.

$$R = \frac{PV}{T} = \frac{(2.560)(8.8)}{275.15}$$

- The correct procedure for obtaining the final result can be represented as follows:

$$\frac{(2.560)(8.8)}{275.15} = \frac{22.528}{275.15} = 0.0818753$$

$$= 0.082 = 8.2 \times 10^{-2} = R$$

Interactive Example 1.4 - Solution (Continued 2)

- The final result must be rounded to two significant figures because 8.8 (the least precise measurement) has two significant figures
- To show the effects of rounding at intermediate steps, carry out the calculation as follows:

$$\frac{(2.560)(8.8)}{275.15} = \frac{22.528}{275.15} = \frac{\overset{\substack{\text{Rounded to two}\\ \text{significant figures}\\ \downarrow}}{23}}{275.15}$$

Interactive Example 1.4 - Solution (Continued 3)

- Now we proceed with the next calculation

$$\frac{23}{275.15} = 0.0835908$$

- Rounded to two significant figures, this result is

$$0.084 = 8.4 \times 10^{-2}$$

- Note that intermediate rounding gives a significantly different result than that obtained by rounding only at the end

Interactive Example 1.4 - Solution (Continued 4)

- Again, we must reemphasize that in your calculations you should round only at the end
- Rounding is carried out at intermediate steps in this text (to always show the correct number of significant figures)
  - The final answer given in the text may differ slightly from the one you obtain (rounding only at the end)

Learning Chemistry

- Involves solving various types of problems
- Questions to ask while approaching a problem
  - Where am I going?
  - What do I know?
  - How do I get there?

Dimensional Analysis (Unit Factor Method)

- Method for converting a given result from one system of units to another

English–metric equivalents →

Length	1 m = 1.094 yd
	2.54 cm = 1 in
Mass	1 kg = 2.205 lb
	453.6 g = 1 lb
Volume	1 L = 1.06 qt
	1 ft^3 = 28.32 L

Problem-Solving Strategy - Converting from One Unit to Another

- Use the equivalence statement that relates the two units
- Derive the appropriate unit factor by looking at the direction of the required change to cancel the unwanted units
- Multiply the quantity to be converted by the unit factor to give the quantity with the desired units

Interactive Example 1.6 - Unit Conversions II

- You want to order a bicycle with a 25.5-in frame, but the sizes in the catalog are given only in centimeters
  - What size should you order?

Interactive Example 1.6 - Solution

- Where are we going?
  - To convert from inches to centimeters
- What do we know?
  - The size needed is 25.5 in
- How do we get there?
  - Since we want to convert from inches to centimeters, we need the equivalence statement 2.54 cm = 1 in

Interactive Example 1.6 - Solution (Continued)

- The correct unit factor in this case is

$$\frac{2.54\ \text{cm}}{1\ \text{in}}$$

$$25.5\ \cancel{\text{in}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}} = 64.8\ \text{cm}$$

Interactive Example 1.7 - Unit Conversions III

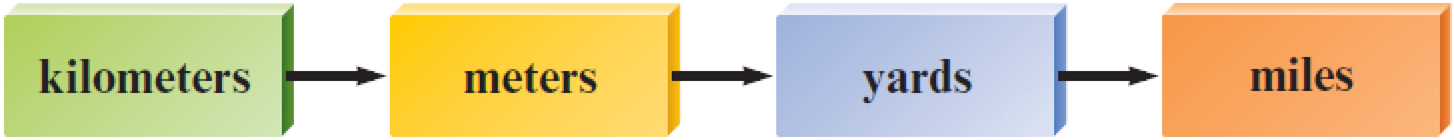
- A student has entered a 10.0-km run
  - How long is the run in miles?

Interactive Example 1.7 - Solution

- Where are we going?
  - To convert from kilometers to miles
- What do we know?
  - The run is 10.00 km long
- How do we get there?
  - This conversion can be accomplished in several different ways

Interactive Example 1.7 - Solution (Continued 1)

- Since we have the equivalence statement 1 m = 1.094 yd, we will proceed by a path that uses this fact
- Before we start any calculations, let us consider our strategy
  - We have kilometers, which we want to change to miles
  - We can do this by the following route:



Interactive Example 1.7 - Solution (Continued 2)

- To proceed in this way, we need the following equivalence statements:

  1 km = 1000 m

  1 m = 1.094 yd

  1760 yd = 1 mi

- To make sure the process is clear, we will proceed step by step

Interactive Example 1.7 - Solution (Continued 3)

- Kilometers to meters

$$10.0\ \cancel{\text{km}} \times \frac{1000\ \text{m}}{1\ \cancel{\text{km}}} = 1.00 \times 10^4\ \text{m}$$

- Meters to yards

$$1.00 \times 10^4\ \cancel{\text{m}} \times \frac{1.094\ \text{yd}}{1\ \cancel{\text{m}}} = 1.094 \times 10^4\ \text{yd}$$

  - Note that we should have only three significant figures in the result

Interactive Example 1.7 - Solution (Continued 4)

- However, since this is an intermediate result, we will carry the extra digit
- Remember, round off only the final result

- Yards to miles

$$1.094 \times 10^4 \ \cancel{\text{yd}} \times \frac{1 \text{ mi}}{1760 \ \cancel{\text{yd}}} = 6.216 \text{ mi}$$

- Note in this case that 1 mi equals exactly 1760 yd by designation

Interactive Example 1.7 - Solution (Continued 5)

- Thus 1760 is an exact number
- Since the distance was originally given as 10.0 km, the result can have only three significant figures and should be rounded to 6.22 mi

10.0 km = 6.22 mi

- Alternatively, we can combine the steps:

$$10.0\ \cancel{\text{km}} \times \frac{1000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1.094\ \cancel{\text{yd}}}{1\ \cancel{\text{m}}} \times \frac{1\ \text{mi}}{1760\ \cancel{\text{yd}}} = 6.22\ \text{mi}$$

Interactive Example 1.9 - Unit Conversions V

- A Japanese car is advertised as having a gas mileage of 15 km/L
  - Convert this rating to miles per gallon

Interactive Example 1.9 - Solution

- Where are we going?
  - To convert gas mileage from 15 kilometers per liter to miles per gallon
- What do we know?
  - The gas mileage is 15 km/L

Interactive Example 1.9 - Solution (Continued)

- How do we get there?
  - We use the following unit factors to make the required conversion:

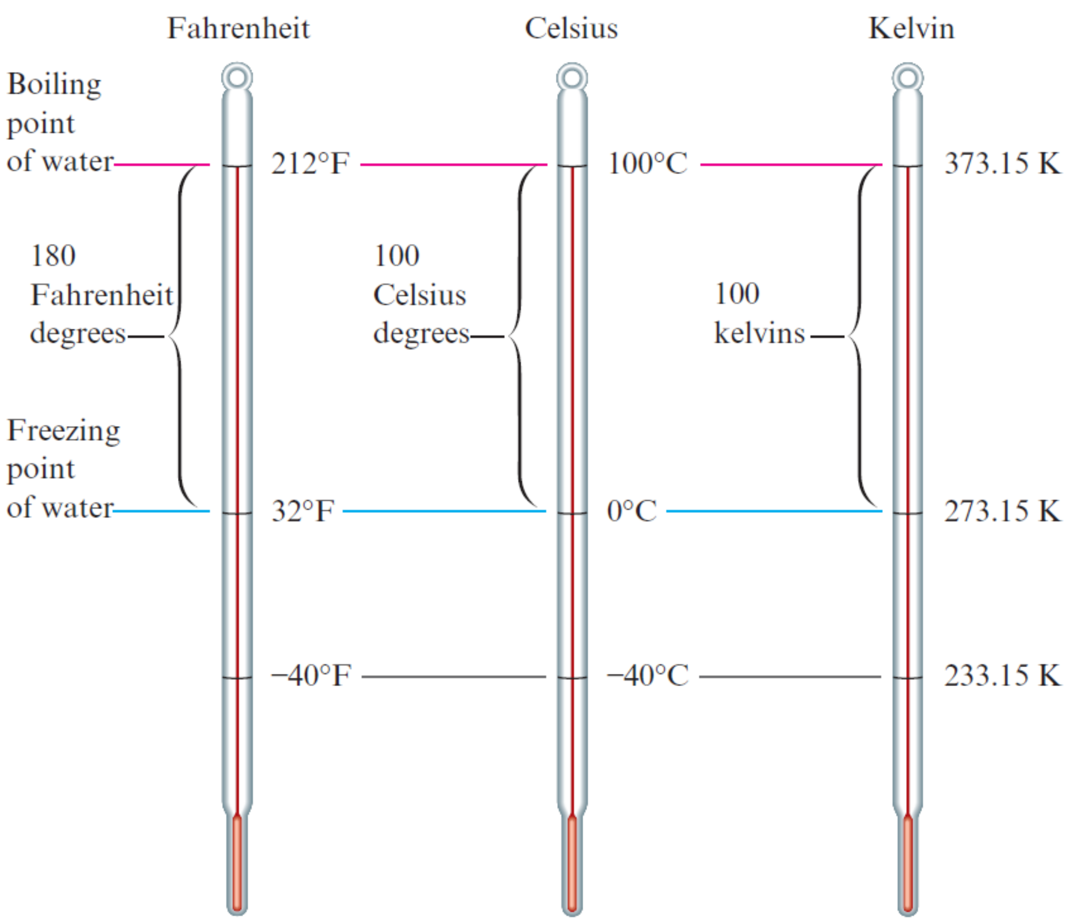
Result obtained by rounding only at the end of the calculation

$$\frac{15\ \cancel{\text{km}}}{\cancel{\text{L}}} \times \frac{1000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1.094\ \cancel{\text{yd}}}{1\ \cancel{\text{m}}} \times \frac{1\ \text{mi}}{1760\ \cancel{\text{yd}}} \times \frac{1\ \cancel{\text{L}}}{1.06\ \cancel{\text{qt}}} \times \frac{4\ \cancel{\text{qt}}}{1\ \text{gal}} = 35\ \text{mi/gal}$$

Systems for Measuring Temperature

- Celsius scale and Kelvin scale
  - Used in physical sciences
- Fahrenheit scale
  - Used in engineering sciences

Figure 1.9 - The Three Major Temperature Scales



Converting between Scales

- Kelvin and Celsius scales
  - Differ in their zero points
  - Conversion requires an adjustment for the different zero points

$$\text{Temperature (Kelvin)} = \text{temperature (Celsius)} + 273.15$$

$$\text{Temperature (Celsius)} = \text{temperature (Kelvin)} - 273.15$$

Converting between Scales (Continued 1)

- Fahrenheit and Celsius scales
  - Degree sizes and the zero points are different

212 – 32 = 180 Fahrenheit degrees = 100 – 0 = 100 Celsius degrees

$$\text{Unit factor is: } \frac{180°\text{F}}{100°\text{C}} \text{ or } \frac{9°\text{F}}{5°\text{C}}$$

Converting between Scales (Continued 2)

- Consider the different zero points

$$\left(T_F - 32°F\right)\frac{5°C}{9°F} = T_C$$

T_F - Temperature on the Fahrenheit scale

T_C - Temperature on the Celsius scale

$$T_F = T_C \times \frac{9°F}{5°C} + 32°F$$

Example 1.12 - Temperature Conversions II

- One interesting feature of the Celsius and Fahrenheit scales is that -40° C and -40° F represent the same temperature
  - Verify that this is true

Example 1.12 - Solution

- Where are we going?
  - To show that –40° C = –40° F
- What do we know?
  - The relationship between the Celsius and Fahrenheit scales
- How do we get there?
  - The difference between 32° F and –40° F is 72° F
  - The difference between 0° C and –40° C is 40° C

Example 1.12 - Solution (Continued)

- The ratio of these is

$$\frac{72°\text{F}}{40°\text{C}} = \frac{8 \times 9°\text{F}}{8 \times 5°\text{C}} = \frac{9°\text{F}}{8 \times 5°\text{C}}$$

- Thus –40° C is equivalent to –40° F

Relationship between the Fahrenheit and Celsius Scales

$$\frac{\text{Number of Fahrenheit degrees}}{\text{Number of Celsius degrees}} = \frac{T_F - (-40)}{T_C - (-40)} = \frac{9°F}{5°C}$$

$$\frac{T_F + 40}{T_C + 40} = \frac{9°F}{5°C}$$

- T_F and T_C represent the same temperature, but not the same number

Interactive Example 1.13 - Temperature Conversions III

- Liquid nitrogen, which is often used as a coolant for low-temperature experiments, has a boiling point of 77 K
  - What is this temperature on the Fahrenheit scale?

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Interactive Example 1.13 - Solution

- Where are we going?
  - To convert 77 K to the Fahrenheit scale
- What do we know?
  - The relationship between the Kelvin and Fahrenheit scales
- How do we get there?
  - We will first convert 77 K to the Celsius scale

$$T_C = T_K - 273.15 = 77 - 273.15 = -196^\circ \text{ C}$$

Interactive Example 1.13 - Solution (Continued)

- Now convert to the Fahrenheit scale

$$\frac{T_F + 40}{T_C + 40} = \frac{9°F}{5°C} \qquad \frac{T_F + 40}{-196°C + 40} = \frac{T_F + 40}{-156°C} = \frac{9°F}{5°C}$$

$$T_F + 40 = \frac{9°F}{5°C}\left(-156°C\right) = -281°F$$

$$T_F = -281°F - 40 = -321°F$$

Exercise

- Convert the following Celsius temperatures to Kelvin and to Fahrenheit degrees
  a. Temperature of someone with a fever, 39.2° C

     312.4 K; 102.6° F

  b. Cold wintery day, –25° C

     248 K; –13° F

Exercise (Continued)

- Convert the following Celsius temperatures to Kelvin and to Fahrenheit degrees

c. Lowest possible temperature, –273° C

0 K; –459° F

d. Melting-point temperature of sodium chloride, 801° C

1074 K; 1470° F

Density - An Introduction

- Property of matter that is used as an identification tag for substances
- Mass of substance per unit volume of the substance

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

Interactive Example 1.14 - Determining Density

- A chemist, trying to identify an unknown liquid, finds that 25.00 cm^3 of the substance has a mass of 19.625 g at 20° C

Interactive Example 1.14 - Determining Density (Continued)

- The following are the names and densities of the compounds that might be the liquid:

Compound	Density in g/cm^3 at 20°C
Chloroform	1.492
Diethyl ether	0.714
Ethanol	0.789
Isopropyl alcohol	0.785
Toluene	0.867

- Which of these compounds is the most likely to be the unknown liquid?

Interactive Example 1.14 - Solution

- Where are we going?
  - To calculate the density of the unknown liquid
- What do we know?
  - The mass of a given volume of the liquid
- How do we get there?
  - To identify the unknown substance, we must determine its density

Interactive Example 1.14 - Solution (Continued 1)

- Density can be determined by using its definition

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

$$= \frac{19.625\ \text{g}}{25.00\ \text{cm}^3}$$

$$= 0.7850\ \text{g/cm}^3$$

- This density corresponds exactly to that of isopropyl alcohol, which therefore most likely is the unknown liquid

Interactive Example 1.14 - Solution (Continued 2)

- However, note that the density of ethanol is also very close
- To be sure that the compound is isopropyl alcohol, we should run several more density experiments
  - In the modern laboratory, many other types of tests could be done to distinguish between these two liquids

Table 1.5 - Densities of Various Common Substances* at 20° C

Substance	Physical State	Density (g/cm^3)
Oxygen	Gas	0.00133
Hydrogen	Gas	0.000084
Ethanol	Liquid	0.789
Benzene	Liquid	0.880
Water	Liquid	0.9982
Magnesium	Solid	1.74
Salt (sodium chloride)	Solid	2.16
Aluminum	Solid	2.70
Iron	Solid	7.87
Copper	Solid	8.96
Silver	Solid	10.5
Lead	Solid	11.34
Mercury	Liquid	13.6
Gold	Solid	19.32

*At 1 atmosphere pressure.

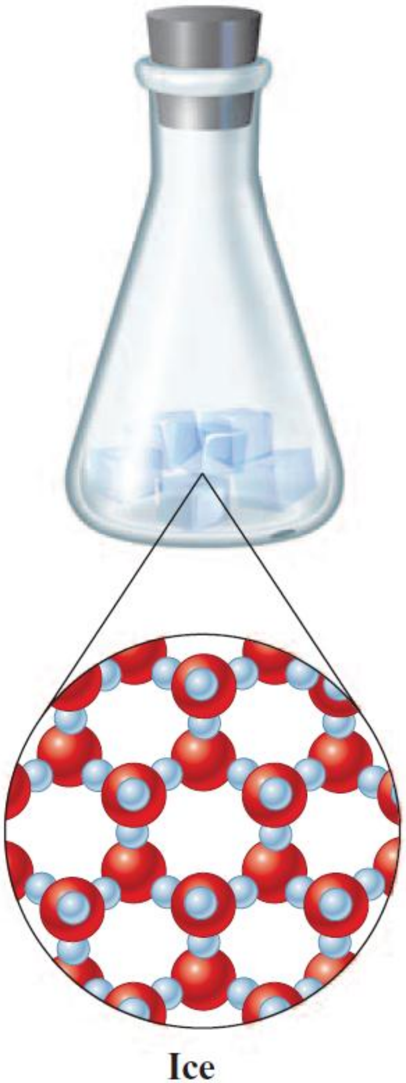
Matter

- Anything that occupies space and has mass
- Is complex and has many levels of organization
- Exists in three **states**
  - Solid
  - Liquid
  - Gas

Properties of a Solid

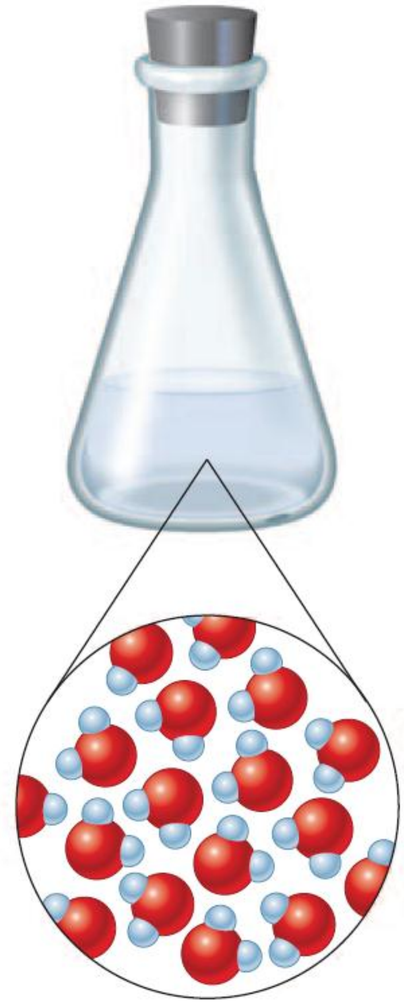
- Rigid
- Fixed volume and shape
- Slightly compressible

Solid: The water molecules are locked into rigid positions and are close together.



Properties of a Liquid

- Definite volume
- No specific shape
  - Assumes the shape of its container
- Slightly compressible

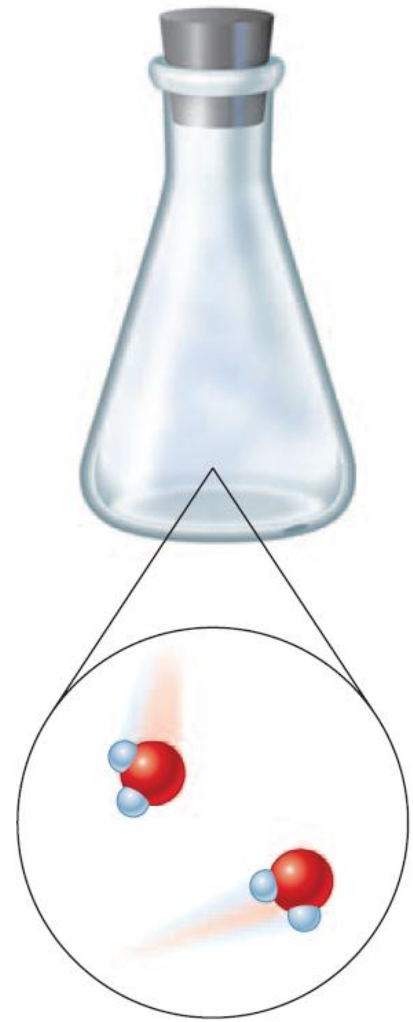


Water

Liquid: The water molecules are still close together but can move around to some extent.

Properties of a Gas

- No fixed volume or shape
  - Takes on the shape and volume of its container
- Highly compressible



Steam

Gas: The water molecules are far apart and move randomly.

Mixtures

- Main characteristic - Variable composition
- Classification
  - **Homogeneous mixture**: Has visibly indistinguishable parts
    - Often called a **solution**
  - **Heterogeneous mixture**: Has visibly distinguishable parts

Mixtures (Continued)

- Can be separated into **pure substances** by physical methods
  - Pure substances have constant composition

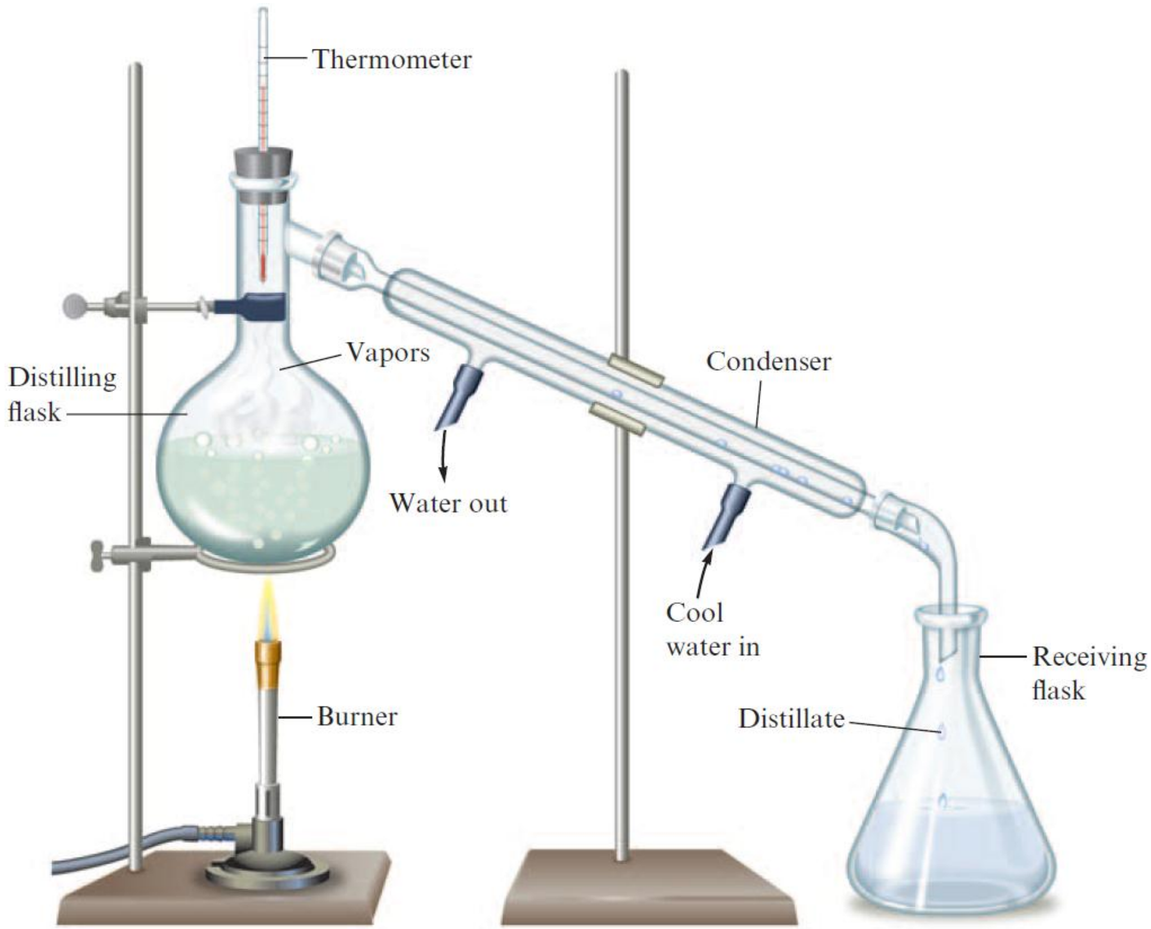
Physical Change

- Change in the form of a substance, not in its chemical composition
- Example - Boiling or freezing of water
- Used to separate a mixture into pure compounds
  - Will not break compounds into elements

Methods for Separating Components in a Mixture

- **Distillation**: Depends on the volatility differences of the components
  - Heat the mixture in a distillation device
    - Most volatile component vaporizes at the lowest temperature
    - Vapor is passed through a condenser, where it goes back to its liquid state
  - Does not give a pure substance in the receiving flask

Figure 1.12 - Simple Laboratory Distillation Apparatus



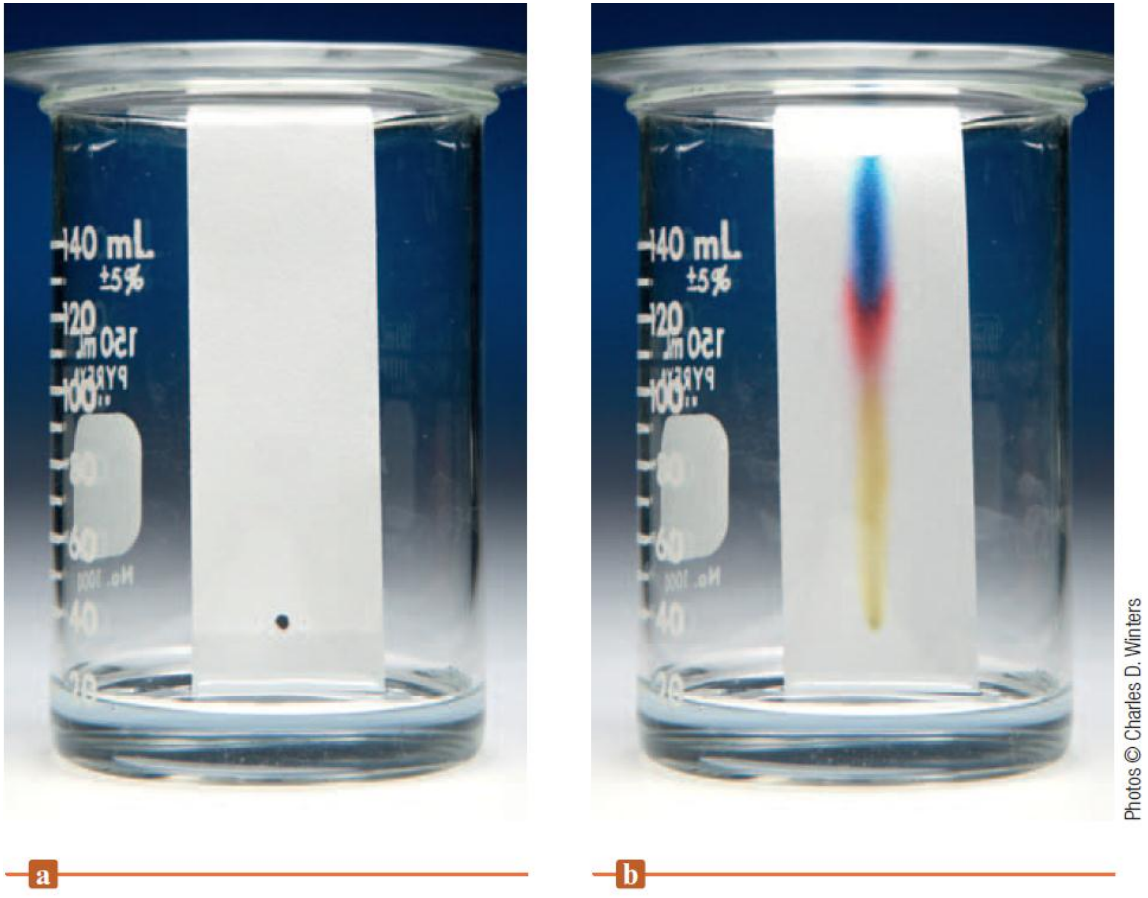
Methods for Separating Components in a Mixture
(Continued)

- **Filtration**: Used when a mixture comprises a solid and a liquid
- **Chromatography**: Series of methods that use a system with two states (phases) of matter
  - Mobile phase - Liquid or gas
  - Stationary phase - Solid

Chromatography

- Separation is facilitated by difference in the component’s affinity for the two phases
  - Component with high affinity for the mobile phase will quickly go through the chromatographic system and vice versa
- **Paper chromatography**: Uses a strip of porous paper for the stationary phase

Figure 1.13 - Paper Chromatography of Ink



Compounds and Elements

- **Compound**: Substance with a constant composition that can be broken down into its elements via chemical processes
- **Element**: Substance that cannot be broken down into simpler substances by physical or chemical means

Chemical Change

- A given substance becomes a new substance or substances with different properties and different composition
  - Only manner by which compounds can be decomposed to elements

Figure 1.14 - The Organization of Matter