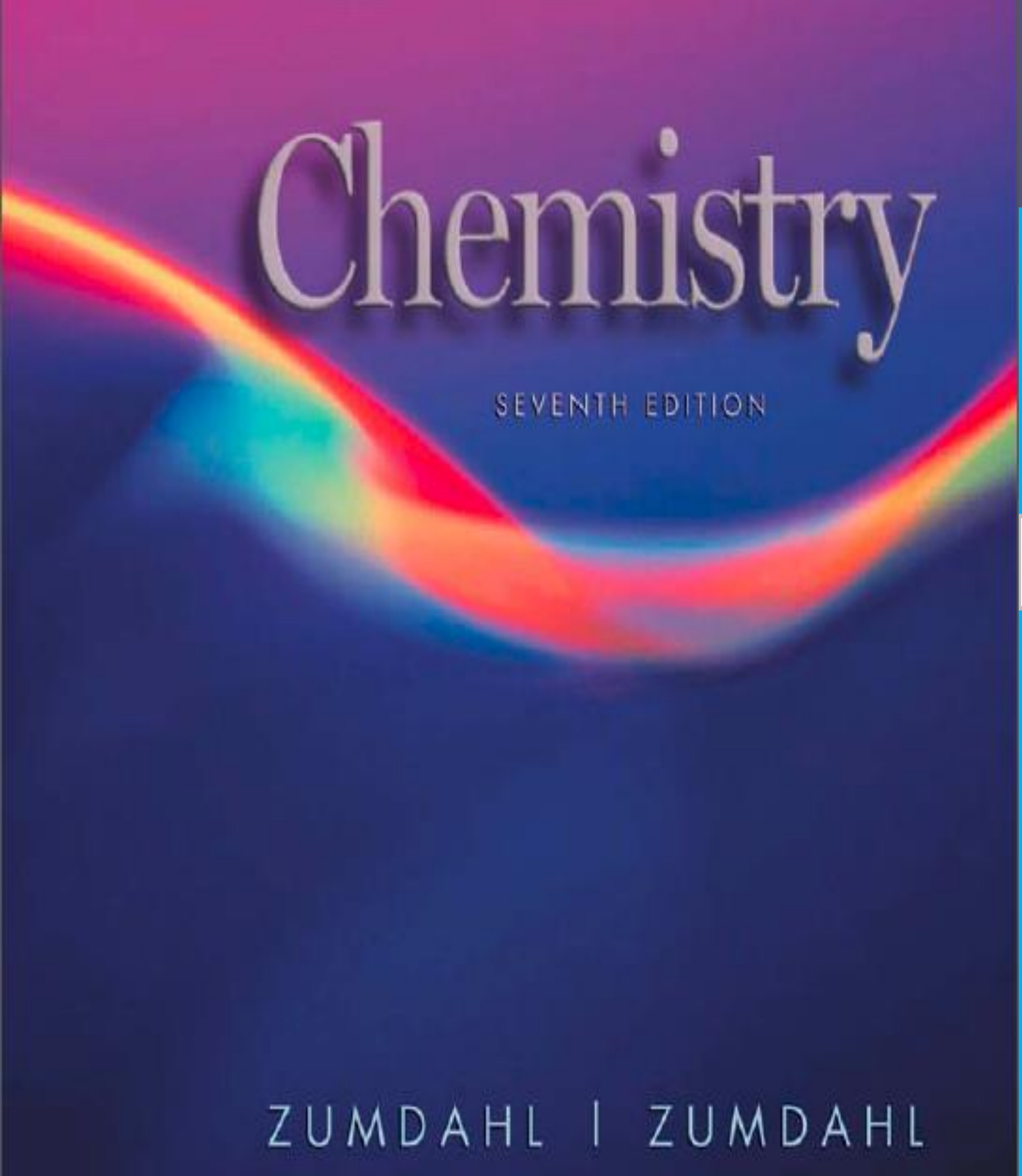


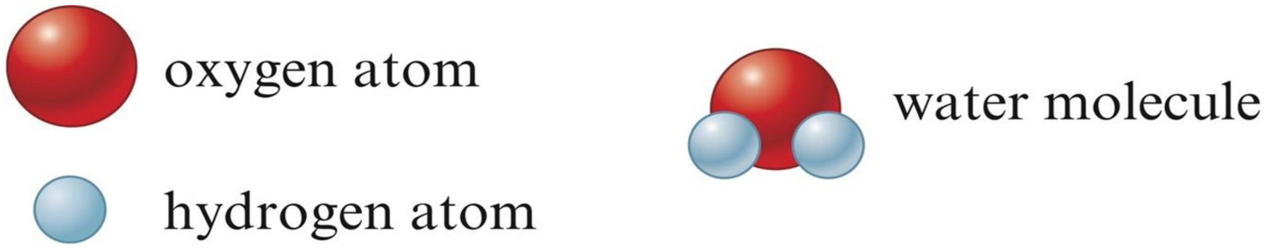
CHAPTER ONE

CHEMICAL FOUNDATION

- A main challenge of chemistry is to understand the connection between the macroscopic world that we experience and the microscopic world of atoms and molecules.
- You must learn to think on the atomic level.

Atoms vs. Molecules

- Matter is composed of tiny particles called atoms.
- Atom: smallest part of an element that is still that element.
- Molecule: Two or more atoms joined and acting as a unit.



Fundamental Steps of the Scientific Method

- Making observations (collecting data)
- Making a prediction (formulating a hypothesis)
- Doing experiments to test the prediction (testing the hypothesis).

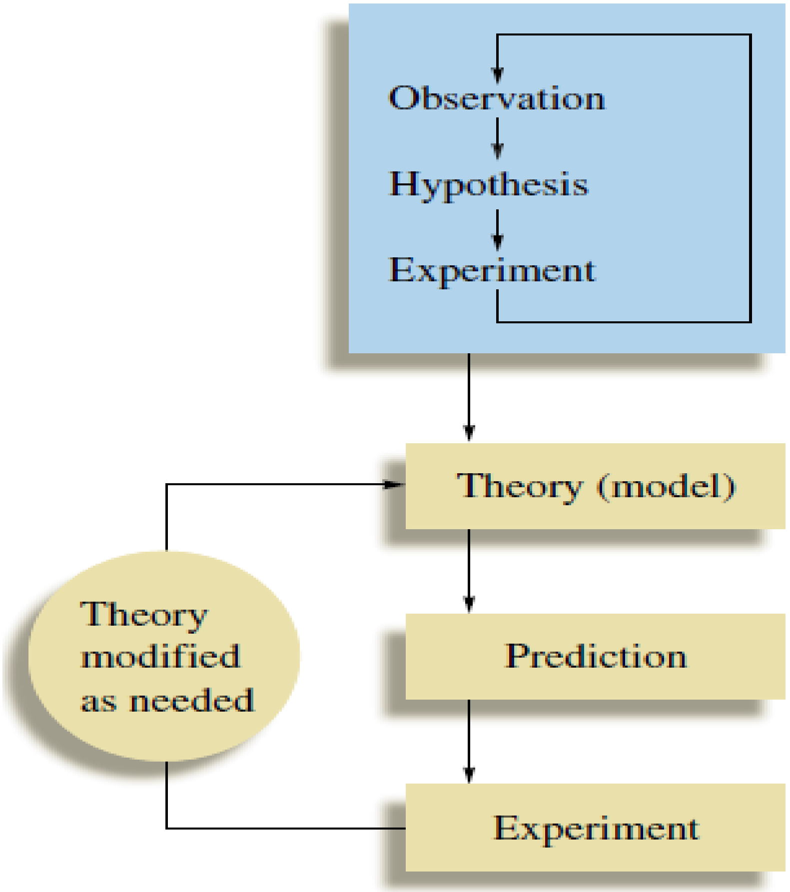


FIGURE 1.4
The fundamental steps of the scientific method.

Nature of Measurement

Measurement

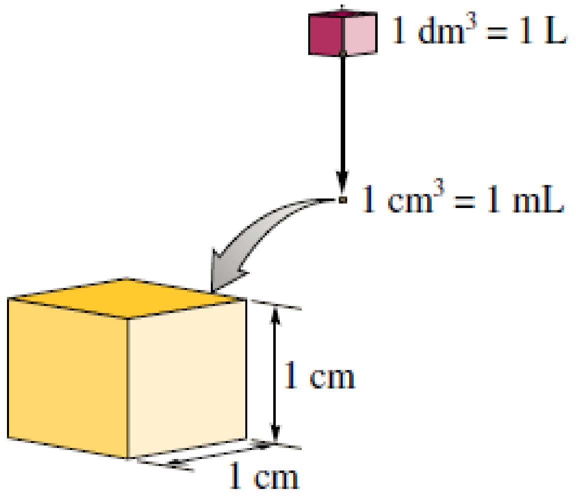
- Quantitative observation consisting of two parts.
  - number
  - scale (unit)
- Examples
  - 20 grams
  - 6.63×10^{-34} joule·second

TABLE 1.1 The 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

- Most scientists in all countries have for many years used the metric system.
- In 1960, an international agreement set up a system of units called the *International System, or the SI system*

- One physical quantity that is very important in chemistry is <u>*volume*</u>, which is not a fundamental SI unit but is derived from length.
- 1 liter = $(1dm)^3$ = $(10cm)^3$ =1000 cm^3
- 1 liter =1000 cm^3 = 1000mL



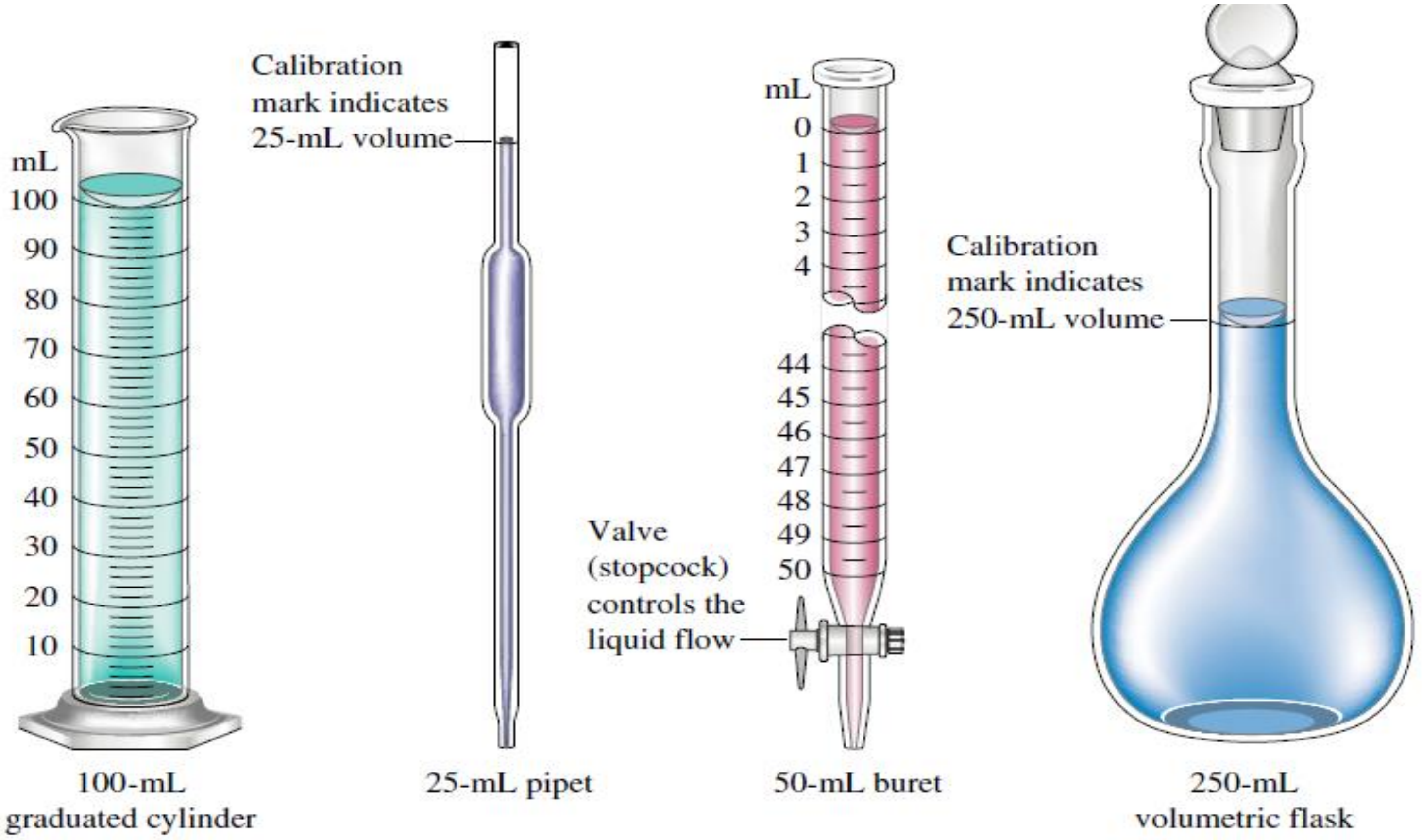


FIGURE 1.7
Common types of laboratory equipment used to measure liquid volume.

Mass ≠ Weight

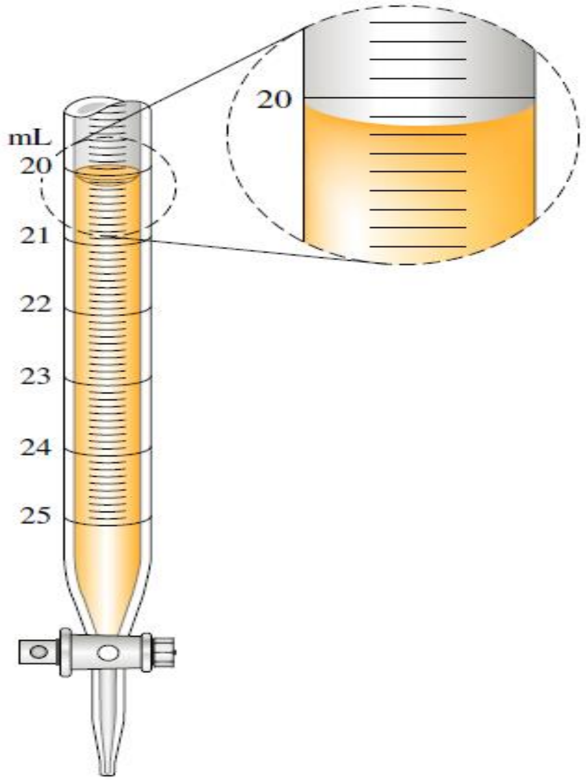
- Mass is a measure of the resistance of an object to a change in its state of motion. Mass does not vary.

- Weight is the force that gravity exerts on an object. Weight varies with the strength of the gravitational field.

- A digit that must be estimated in a measurement is called uncertain.
- A measurement always has some degree of uncertainty. It is dependent on the precision of the measuring device.
- Record the certain digits and the first uncertain digit (the estimated number). These numbers are called the *significant figures of a measurement*.

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: 20.15
  - Uncertain digit: 20.15



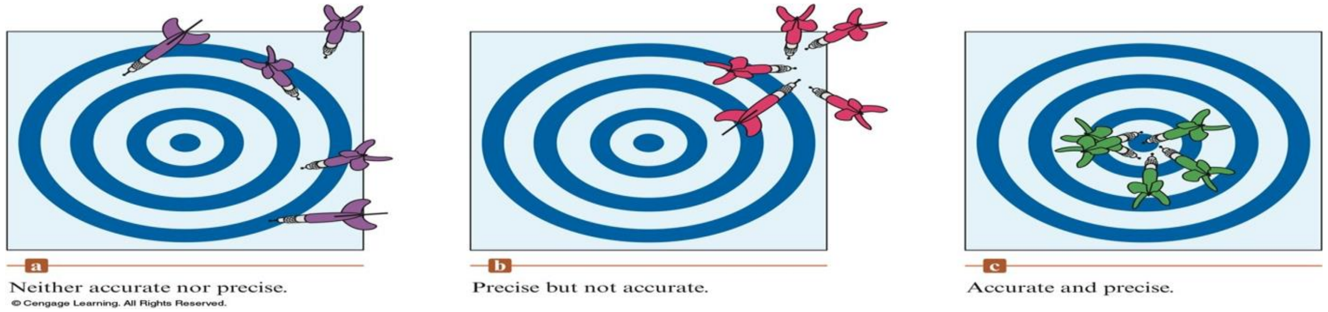
Precision and Accuracy

Accuracy

- Agreement of a particular value with the true value.

Precision

- Degree of agreement among several measurements of the same quantity.



(a) (b) (c)

- The results of several dart throws show the difference between precise and accurate.
- (a) Neither accurate nor precise (large random errors).
- (b) Precise but not accurate (small random errors, large systematic error).
- (c) Bull's-eye! Both precise and accurate (small random errors, no systematic error).

- **Random error (indeterminate error)**
  - Means that a measurement has an equal probability of being high or low.
- **Systematic error(determinate error)**
  - The error occurs in the same direction each time; it is either always high or always low

Sample Exercise 1.1

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?

Solution

- 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.

Rules for Counting Significant Figures

1. Nonzero integers always count as significant figures.
   - 3456 has 4 sig figs (significant figures).
2. There are three classes of zeros.

   A. Leading zeros are zeros that precede all the nonzero digits. These do not count as significant figures.
   - 0.048 has 2 sig figs.

   B. Captive zeros are zeros between nonzero digits. These always count as significant figures.
   - 16.07 has 4 sig figs.

   C. Trailing zeros are zeros at the right end of the number. They are significant only if the number contains a decimal point.
   - 9.300 has 4 sig figs.
   - 150 has 2 sig figs.

3. Exact numbers are numbers that were not obtained using measuring devices but were determined by counting

- 10 experiments, 3 apples, 8 molecules.
- 1 inch = 2.54 cm, exactly.

Exact numbers have an infinite number of significant figures.

Sample Exercise 1.2

- 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.

Solution

a. The number contains three significant figures.

b. The number contains five significant figures

c. This number has four significant figures. Both zeros are significant.

Significant Figures in Mathematical Operations

1. For multiplication or division, the number of significant figures in the result is the same as the number in the least precise measurement used in the calculation.

$$1.342 \times \underline{5.5} = 7.381 \rightarrow \underline{7.4}$$

2. For addition or subtraction, the result has the same number of decimal places as the least precise measurement used in the calculation.

$$\begin{array}{rr} & 23.445 \\ + & 7.83 \\ \hline & 31.275 \end{array} \xrightarrow{\text{Corrected}} 31.28$$

Section 1.7
Dimensional Analysis

❑ Rules for Rounding

- In a series of calculations, carry the extra digits through to the final result, *then round.*

2. If the digit to be removed

a. is less than 5, the preceding digit stays the same. For example, 1.33 rounds to 1.3.

b. is equal to or greater than 5, the preceding digit is increased by 1. For example,1.36 rounds to 1.4.

Dimensional Analysis

❑ It is often necessary to convert a given result from one system of units to another. The best way to do this is by a method called the unit factor method, or more commonly dimensional analysis.

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

Sample Exercise 1.3

- A pencil is 7.00 in long. What is its length in centimeters?

Solution

In this case we want to convert from inches to centimeters.

1 in =2.54 cm

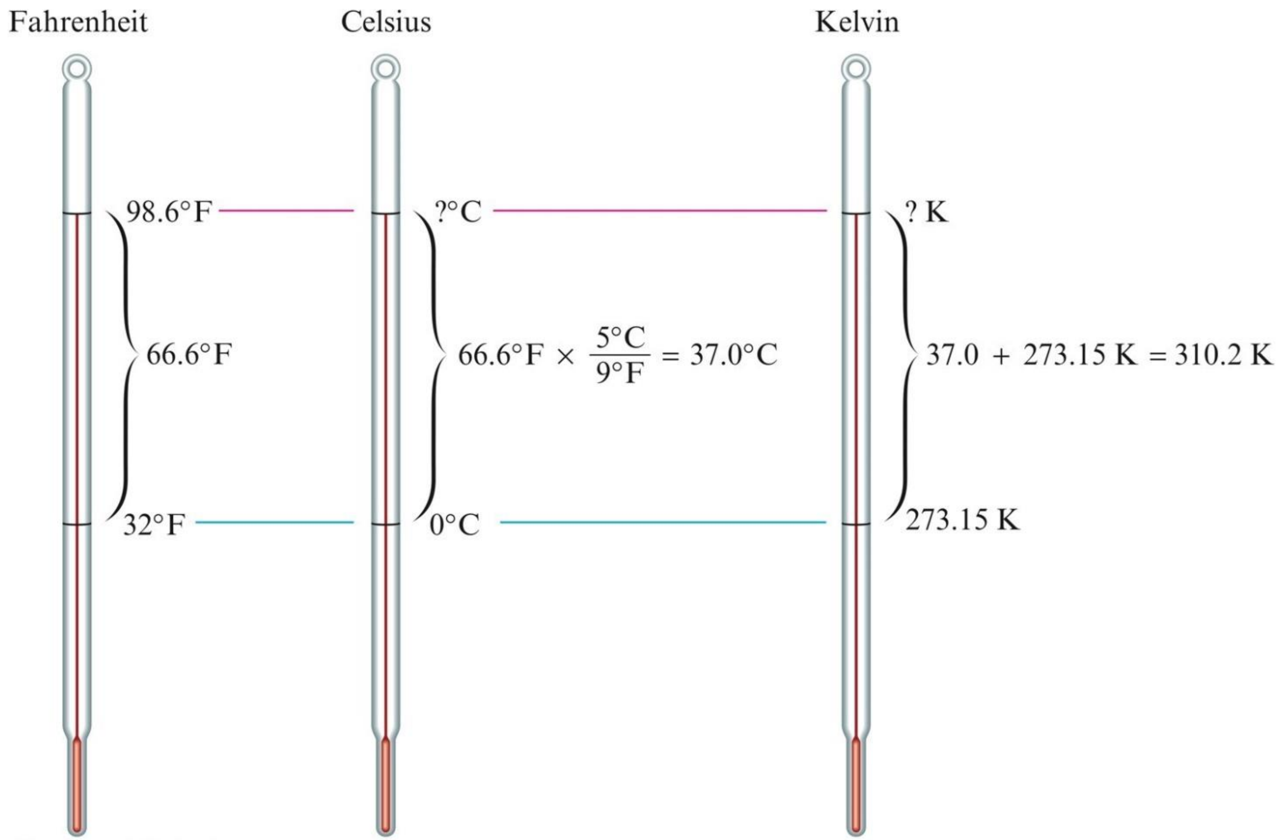
$$7.00\ \cancel{\text{in}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}} = (7.00)(2.54)\ \text{cm} = 17.8\ \text{cm}$$

Sample Exercise 1.4

- 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?

Three Systems for Measuring Temperature

- Fahrenheit
- Celsius
- Kelvin



Converting Between Scales

$$T_K = T_C + 273.15 \qquad T_C = T_K - 273.15$$

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

- OR

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

- For example, to convert 300.00 K to the Celsius scale, we do the following calculation

$$300.00 - 273.15 = 26.85^0\text{C}$$

Section 1.8

Temperature

Sample Exercise 1.5

- Normal body temperature is 98.6 ^0F. Convert this temperature to the Celsius and Kelvin scales

SOLUTION

- $T_C = (T_F - 32)\frac{5^0C}{9^0F}$ $T_C = (98.60F - 32)\frac{5^0C}{9^0F}$

  $\therefore T_C = 37.0^0C$

- $T_K = T_C + 273.15$ $T_K = 37.0^0C + 273.15 = 310.2K$

Sample Exercise 1.6

- 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?

Solution:

We will first convert 77 K to the Celsius scale:

$$T_C = T_K - 273.15 = 77 - 273.15 = -196°C$$

To convert to the Fahrenheit scale, we will use Equation (1.3):

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

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

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

$$T_F = -281°F - 40 = -321°F$$

- Density
  - Mass of substance per unit volume of the substance.
  - Common units are g/cm^3 or g/mL.

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

Sample Exercise 1.7

A certain mineral has a mass of 17.8 g and a volume of 2.35 cm^3. What is the density of this mineral?

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{17.8\text{ g}}{2.35\text{ cm}^3} = 7.57\text{ g/cm}^3$$

Sample Exercise 1.8

A chemist, trying to identify the main component of a compact disc cleaning fluid, finds that 25.00 cm^3 of the substance has a mass of 19.625 g at 20° C. The following are the names and densities of the compounds that might be the main component:

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

SOLUTION

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{19.625\text{ g}}{25.00\text{ cm}^3} = 0.7850\text{ g/cm}^3$$

- This density corresponds exactly to that of isopropyl alcohol, which is therefore the most likely main component of the cleaner

Section 1.10
Classification of Matter

Matter

- Anything occupying space and having mass.
- Matter exists in three states.
  - Solid
  - Liquid
  - Gas

Solid

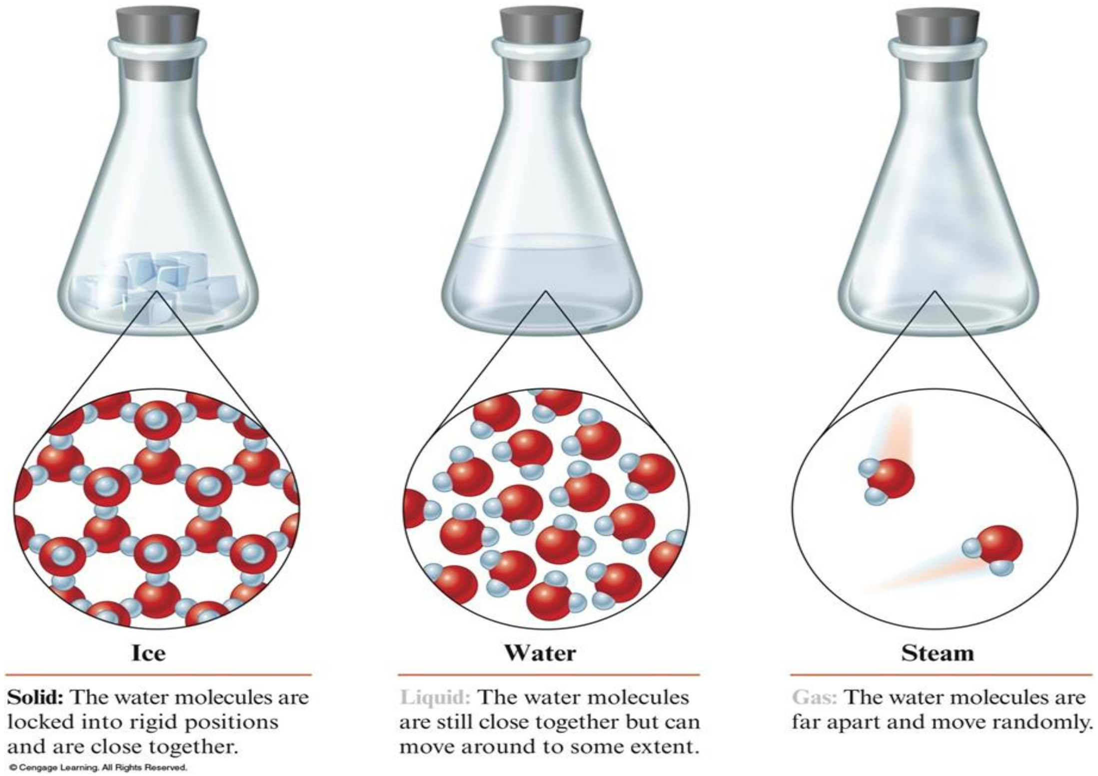
- Rigid
- Has fixed volume and shape

Liquid

- Has definite volume but no specific shape.
- Assumes shape of container

Gas

- Has no fixed volume or shape.
- Takes on the shape and volume of its container



The Three States of Water

Mixtures

- Have variable composition.

Homogeneous Mixture

- Having visibly indistinguishable parts; solution.
  - Solid solution as brass (copper and zinc)
  - Liquid solution as sugar dissolved in water
  - Gases solution as air

Heterogeneous Mixture

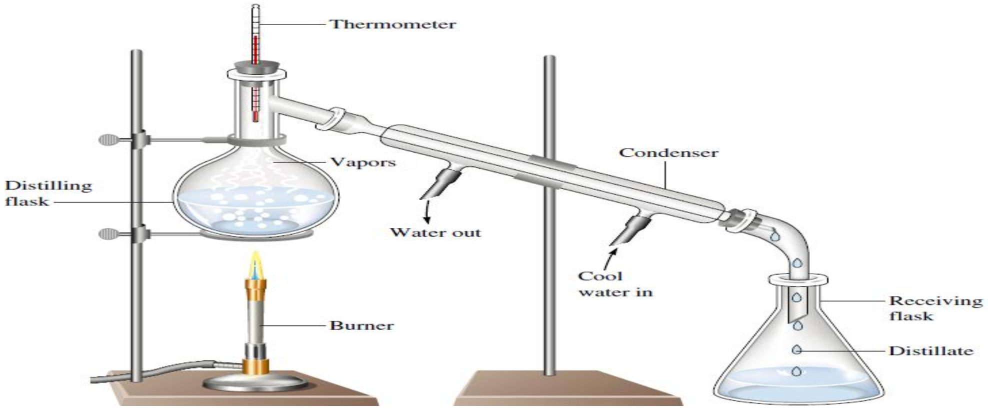
- Having visibly distinguishable parts.
  - For example Sand in water and iced tea with ice cubes

A pure substance

- is one with constant composition for example water

Physical Change

- **Change in the form of a substance, not in its chemical composition**.
  - Example: boiling or freezing water
- **Can be used to separate a mixture into pure compounds**, but it will not break compounds into elements.
  - Distillation
  - Filtration
  - Chromatography



Chemical Change

- A given substance becomes a new substance or substances with different properties and different composition.
  - Example: Bunsen burner, and iron rust

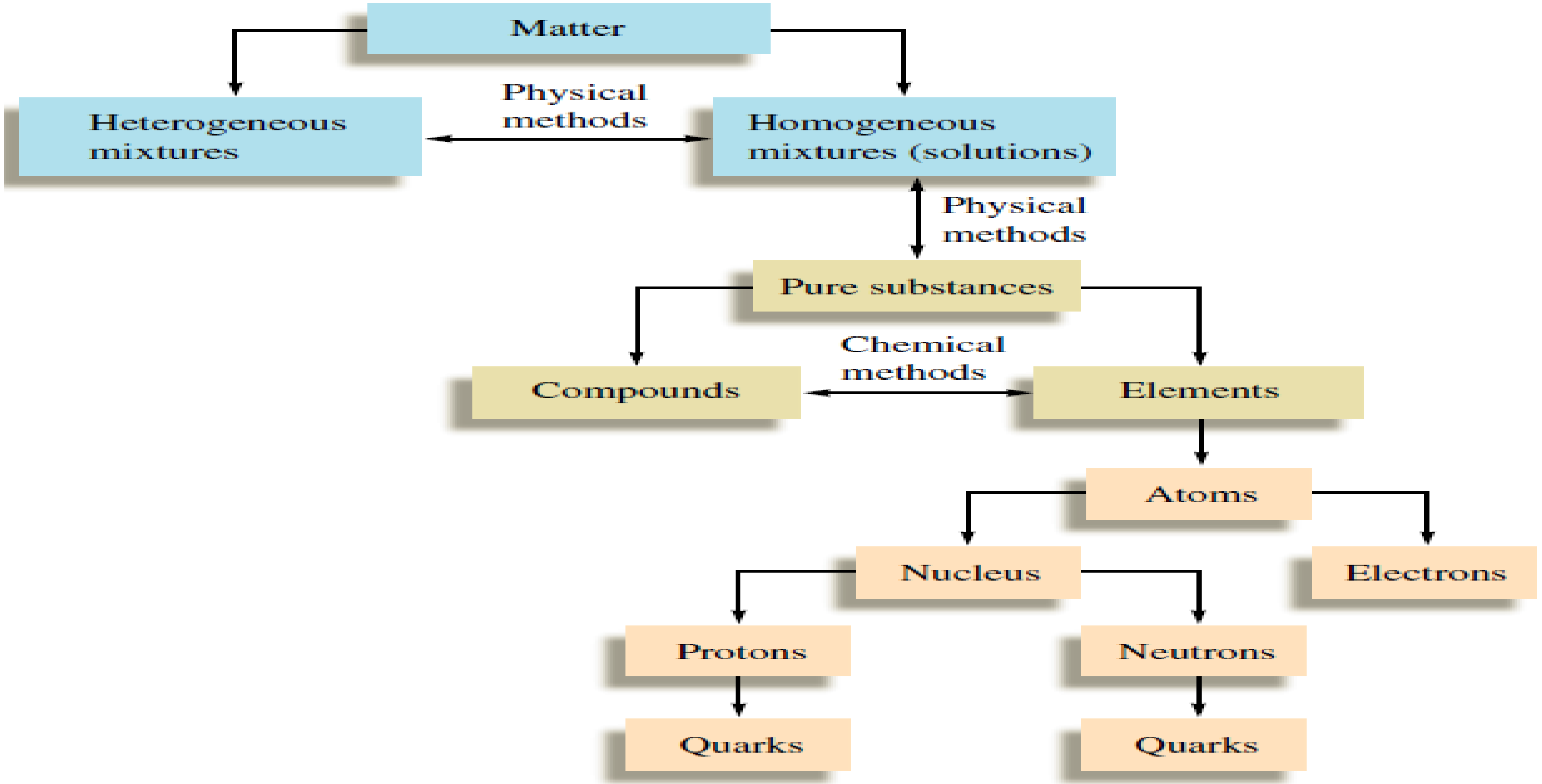
compound

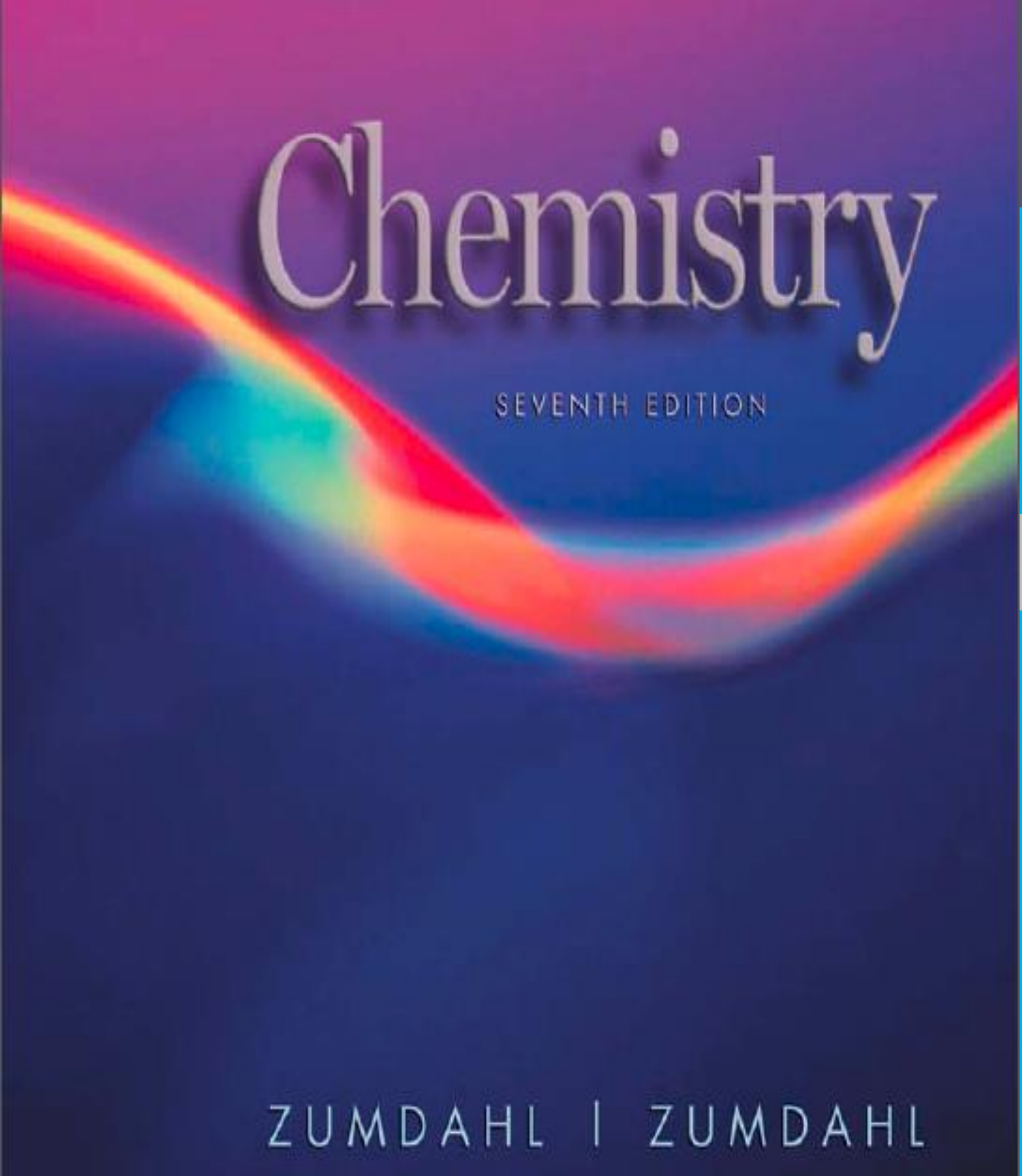
- is a substance with *constant composition* that can be broken down into elements by chemical processes
- example of a chemical process is the electrolysis of water

Elements

- are substances that cannot be decomposed into simpler substances by chemical or physical means.

The Organization of Matter





CHAPTER TWO

ATOMS,IONS,and MOLECULES

Early History of Chemistry

- Greeks were the first to attempt to explain why chemical changes occur.
- Alchemy dominated for 2000 years.
  - Several elements discovered.
  - Mineral acids prepared.
- Robert Boyle was the first "chemist".
  - Performed quantitative experiments.
  - Developed first experimental definition of an element.

Three Important Laws

- Law of conservation of mass (Lavoisier):
  - Mass is neither created nor destroyed in a chemical reaction.

- Law of definite proportion (Proust):
  - A given compound always contains exactly the same proportion of elements by mass.

- Law of multiple proportions (Dalton):
  - When two elements form a series of compounds, the ratios of the masses of the second element that combine with 1 gram of the first element can always be reduced to small whole numbers.

Illustrating the Law of Multiple Proportions

Sample Exercise 2.1

The following data were collected for several compounds of nitrogen and oxygen

	Mass of Nitrogen That Combines with 1 g of Oxygen
Compound A	1.750 g
Compound B	0.8750 g
Compound C	0.4375 g

Show how these data illustrate the law of multiple proportions

Solution

For the law of multiple proportions to hold, the ratios of the masses of nitrogen combining with 1 gram of oxygen in each pair of compounds should be small whole numbers.We therefore compute the ratios as follows:

$$\frac{A}{B} = \frac{1.750}{0.875} = \frac{2}{1}$$

$$\frac{B}{C} = \frac{0.875}{0.4375} = \frac{2}{1}$$

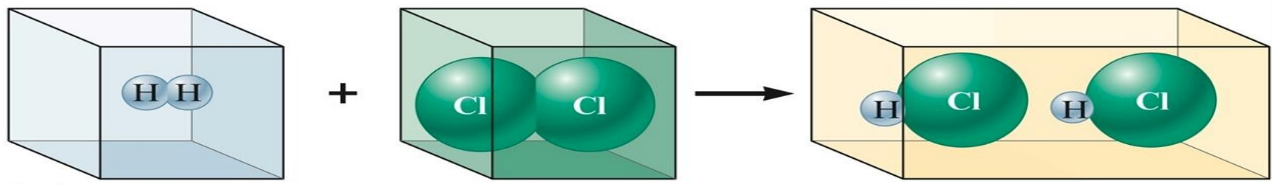
$$\frac{A}{C} = \frac{1.750}{0.4375} = \frac{4}{1}$$

Dalton's Atomic Theory (1808)

- *Each element is made up of tiny particles called atoms.*
- *The atoms of a given element are identical; the atoms of different elements are different in some fundamental way or ways.*
- *Chemical compounds are formed when atoms of different elements combine with each other. A given compound always has the same relative numbers and types of atoms.*
- *Chemical reactions involve reorganization of the atoms—changes in the way they are bound together. The atoms themselves are not changed in a chemical reaction.*

Gay-Lussac and Avogadro (1809—1811)

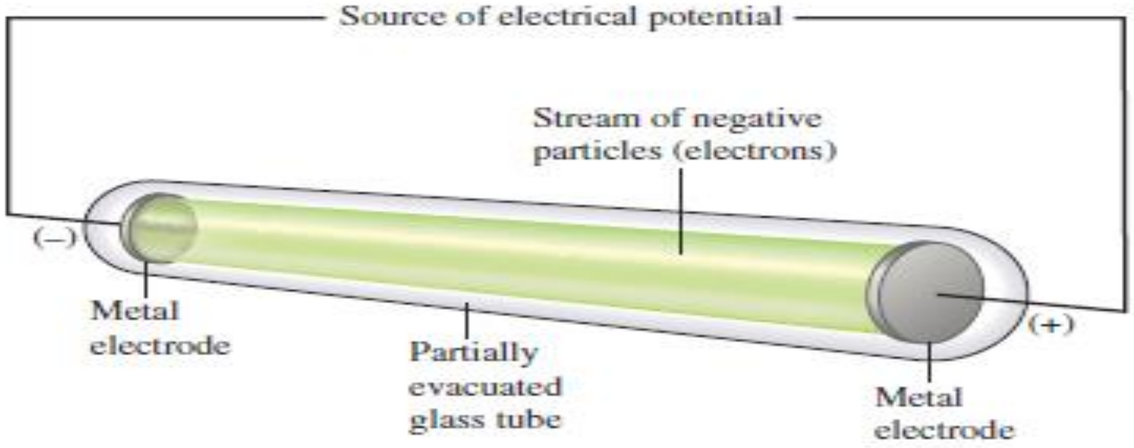
- **Gay—Lussac**
  - Measured (under same conditions of T and P) the volumes of gases that reacted with each other.
- **Representing Gay—Lussac's Results**



- **Avogadro's Hypothesis**
  - At the same T and P, equal volumes of different gases contain the same number of particles.
    - Volume of a gas is determined by the number, not the size, of molecules.

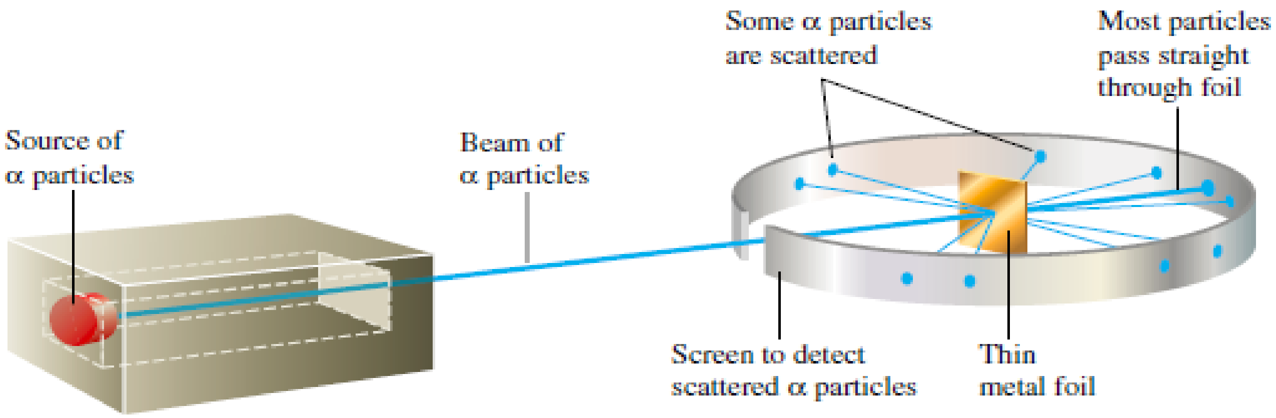
J. J. Thomson (1898—1903)

- Postulated the existence of negatively charged particles, that we now call electrons, using cathode-ray tubes.
- Determined the charge-to-mass ratio of an electron. $\frac{e}{m}$ = -1.76 x 10^8 C/g
- The atom must also contain positive particles that balance exactly the negative charge carried by electrons.



Ernest Rutherford (1911)

- Explained the nuclear atom.
- The atom has a dense center of positive charge called the nucleus.
- Electrons travel around the nucleus at a large distance relative to the nucleus.



Rutherford's experiment on α-particle bombardment of metal foil.

- The atom contains:
  - *Electrons* – found outside the nucleus; negatively charged.
  - *Protons* – found in the nucleus; positive charge equal in magnitude to the electron's negative charge.
  - *Neutrons* – found in the nucleus; no charge; virtually same mass as a proton.

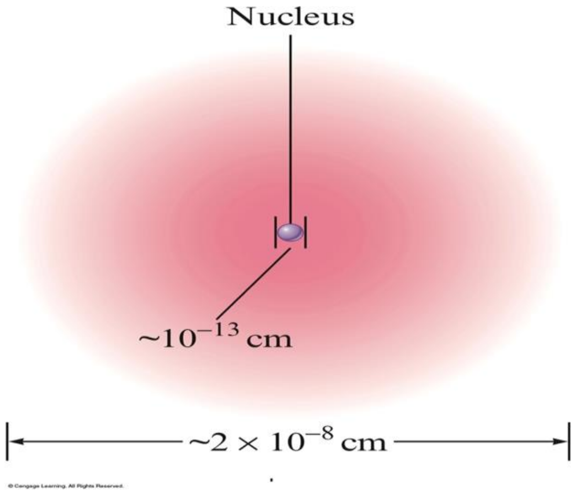
TABLE 2.1 The Mass and Charge of the Electron, Proton, and Neutron

Particle	Mass	Charge*
Electron	9.11×10^{-31} kg	1−
Proton	1.67×10^{-27} kg	1+
Neutron	1.67×10^{-27} kg	None

*The magnitude of the charge of the electron and the proton is 1.60×10^{-19} C.

Section 2.5
The Modern View of Atomic Structure: An Introduction

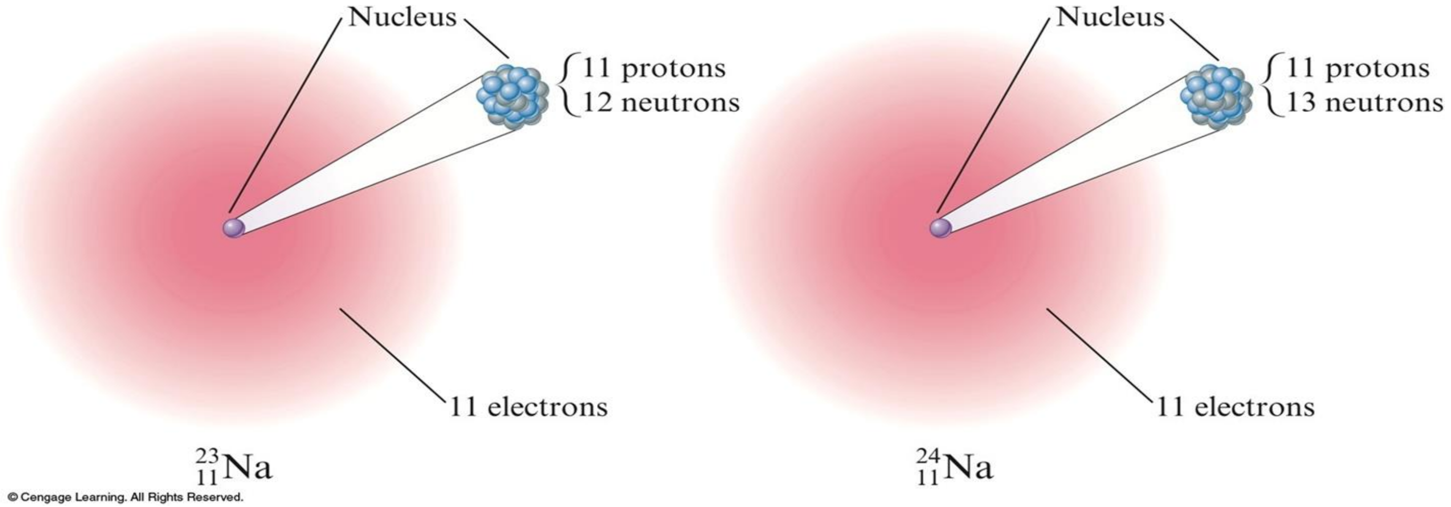
- **The nucleus is:**
  - Small compared with the overall size of the atom.
  - Extremely dense; accounts for almost all of the atom's mass.
- **Electrons**
  - reside outside the nucleus in the relatively large remaining atomic volume.
  - Electrons—negative charge, small mass (11840 of proton)
- the atoms of different elements, which have different numbers of protons and electrons, show different chemical behavior.



Isotopes

- Atoms with the same number of protons but different numbers of neutrons.
- Show almost identical chemical properties; chemistry of atom is due to its electrons.
- In nature most elements contain mixtures of isotopes.

- Isotopes are identified by:
  - Atomic Number (Z) – number of protons
  - Mass Number (A) – number of protons plus number of neutrons



Sample Exercise 2.2

- Write the symbol for the atom that has an atomic number of 9 and a mass number of 19.
- How many electrons and how many neutrons does this atom have?
- ***Solution***

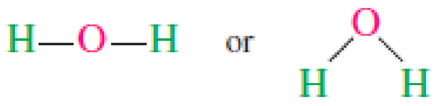
The symbol is F (fluorine) $^{19}_{9}F$

Electrons number = 9

Number of neutrons = 10

Chemical Bonds

- **Chemical bonds**
  - The forces that hold atoms together in compounds
- **Covalent Bonds**
  - Bonds form between atoms by sharing electrons.
  - Resulting collection of atoms is called a molecule.
- **Chemical formula**
  - in which the symbols for the elements are used to indicate the types of atoms present and subscripts are used to indicate the relative numbers of atoms. For example, the formula of carbon dioxide CO_2, and glucose $C_6H_{12}O_6$
- **Structural formula**
  - in which the individual bonds are shown (indicated by lines)
  - Structural formulas may or may not indicate the actual shape of the molecule



Chemical Bonds

- Ionic Bonds
  - Bonds form due to force of attraction between oppositely charged ions.

$$Na^{+} + Cl^{-} \rightarrow NaCl$$

  - *Ion* – atom or group of atoms that has a net positive or negative charge.
  - *Cation* – positive ion; lost electron(s).

$$Na \rightarrow Na^{+} + e^{-}$$

  - *Anion* – negative ion; gained electron(s).

$$Cl + e^{-} \rightarrow Cl^{-}$$

Section 2.7
An Introduction to the Periodic Table

Alkali metals · Alkaline earth metals · Transition metals · Halogens · Noble gases

1 1A	2 2A	3	4	5	6	7	8	9	10	11	12	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La*	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac†	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Uub	113 Uut	114 Uuq	115 Uup			

*Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
†Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

FIGURE 2.21
The periodic table.

The Periodic Table

- Metals vs. Nonmetals
  - Most of the elements in periodic table are metals.
- Physical properties of Metals
  - have efficient conduction of heat and electricity
  - malleability (they can be hammered into thin sheets)
  - ductility (they can be pulled into wires)
  - have a lustrous appearance.
- Chemically
  - metals tend to lose electrons to form positive ions

- Nonmetals
  - appear in the upper-right corner of the table except hydrogen, a nonmetal that resides in the upper left corner.
  - The nonmetals lack the physical properties that characterize the metals.
- Chemically
  - they tend to gain electrons in reactions with metals to form negative ions
- Groups or Families
  - elements in the same vertical columns; have similar chemical properties
- Periods
  - horizontal rows of elements
  - Horizontal row 1 is called the first period (it contains H and He); row 2 is called the second period (elements Li through Ne); and so on.

.

- Group 1A (alkali metals)
  - lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr)
  - Are very active elements
  - Form ions with a 1+ charge when they react with nonmetals
- Group 2A (alkaline earth metals)
  - Beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra)
  - They all form ions with a 2+ charge when they react with nonmetals.
- Group 7A (halogens)
  - Fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At)
  - All form diatomic molecules.
  - Form salts containing ions with a 1- charge when react with metal
- Group 8A (noble gases)
  - helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn)

TABLE 2.2 The Symbols for the Elements That Are Based on the Original Names

Current Name	Original Name	Symbol
Antimony	Stibium	Sb
Copper	Cuprum	Cu
Iron	Ferrum	Fe
Lead	Plumbum	Pb
Mercury	Hydrargyrum	Hg
Potassium	Kalium	K
Silver	Argentum	Ag
Sodium	Natrium	Na
Tin	Stannum	Sn
Tungsten	Wolfram	W

Naming Compounds

- Binary Compounds
  - Composed of two elements
  - Ionic and covalent compounds included
- Binary Ionic Compounds
  - Metal—nonmetal
- Binary Covalent Compounds
  - Nonmetal—nonmetal
- Binary Ionic Compounds (Type I)

1. The cation is always named first and the anion second.
2. A monatomic cation takes its name from the name of the parent element.
3. A monatomic anion is named by taking the root of the element name and *adding –ide.*

Compound	Ions Present	Name
$NaCl$	Na^+, Cl^-	Sodium chloride
KI	K^+, I^-	Potassium iodide
CaS	Ca^{2+}, S^{2-}	Calcium sulfide
Li_3N	Li^+, N^{3-}	Lithium nitride
$CsBr$	Cs^+, Br^-	Cesium bromide
MgO	Mg^{2+}, O^{2-}	Magnesium oxide

Sample Exercise 2.3

Name each binary compound.

a. CsF **b.** $AlCl_3$ **c.** LiH

Solution

a. CsF is cesium fluoride.
b. $AlCl_3$ is aluminum chloride.
c. LiH is lithium hydride.

Notice that, in each case, the cation is named first, and then the anion is named.

Binary Ionic Compounds (Type II)

- Metals in these compounds form more than one type of positive ion.
- Charge on the metal ion must be specified.
- Roman numeral indicates the charge of the metal cation
- Another system for naming these ionic compounds that is seen in the older literature was used for metals that form only two ions
  - *The ion with the higher charge has a name ending in -ic,*
    and the one with the lower charge has a name ending in -ous.
  - For example, Fe^{+3} is called the ferric ion, and Fe^{+2} is called the ferrous ion

TABLE 2.4 Common Type II Cations

Ion	Systematic Name
Fe^{3+}	Iron(III)
Fe^{2+}	Iron(II)
Cu^{2+}	Copper(II)
Cu^{+}	Copper(I)
Co^{3+}	Cobalt(III)
Co^{2+}	Cobalt(II)
Sn^{4+}	Tin(IV)
Sn^{2+}	Tin(II)
Pb^{4+}	Lead(IV)
Pb^{2+}	Lead(II)
Hg^{2+}	Mercury(II)
Hg_2^{2+}*	Mercury(I)
Ag^{+}	Silver†
Zn^{2+}	Zinc†
Cd^{2+}	Cadmium†

*Note that mercury(I) ions always occur bound together to form Hg_2^{2+} ions.
†Although these are transition metals, they form only one type of ion, and a Roman numeral is not used.

Sample Exercise 2.5

give the systematic name of the following

CuCl	Copper(I) chloride
Fe_2O_3	Iron(III)oxide
HgO	mercury(I)oxide

Sample Exercise 2.6

Give the systematic name for each of the following compounds:

a. $CoBr_2$ **b. $CaCl_2$** **c. Al_2O_3**

Solution

1.

Formula	Name	Comments
a. $CoBr_2$	Cobalt(II) bromide	Cobalt is a transition metal; the compound name must have a Roman numeral. The two Br^- ions must be balanced by a Co^{2+} ion.
b. $CaCl_2$	Calcium chloride	Calcium, an alkaline earth metal, forms only the Ca^{2+} ion. A Roman numeral is not necessary.
c. Al_2O_3	Aluminum oxide	Aluminum forms only the Al^{3+} ion. A Roman numeral is not necessary.

Binary Covalent Compounds (Type III)

- Formed between two nonmetals.
  - The first element in the formula is named first, using the full element name.
  - The second element is named as if it were an anion.
  - Prefixes are used to denote the numbers of atoms present.
  - The prefix *mono-* is never used for naming the first element.

Compound	Systematic Name	Common Name
N_2O	Dinitrogen monoxide	Nitrous oxide
NO	Nitrogen monoxide	Nitric oxide
NO_2	Nitrogen dioxide	
N_2O_3	Dinitrogen trioxide	
N_2O_4	Dinitrogen tetroxide	
N_2O_5	Dinitrogen pentoxide	

Table 2.6 | Prefixes Used to Indicate Number in Chemical Names

Prefix	Number Indicated
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Sample Exercise 2.7

Name each of the following compounds:

a. PCl_5 **b. PCl_3** **c. SO_2**

Solution

Formula	Name
a. PCl_5	Phosphorus pentachloride
b. PCl_3	Phosphorus trichloride
c. SO_2	Sulfur dioxide

Flowchart for Naming Binary Compounds