Chemistry:
Matter and Measurements

Chapter 1

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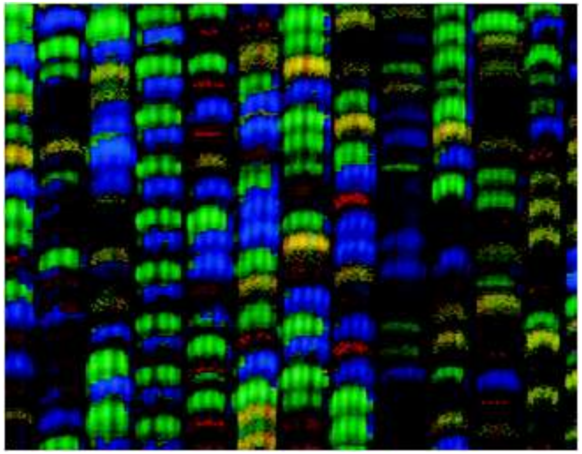
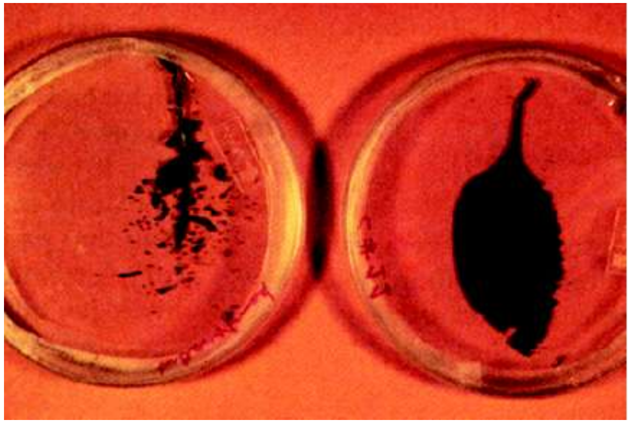
Chemistry

- Health and Medicine
  - Sanitation systems
  - Surgery with anesthesia
  - Vaccines and antibiotics
  - Gene therapy



- Energy and the Environment
  - Fossil fuels
  - Solar energy
  - Nuclear energy

- Materials and Technology
    - Polymers, ceramics, liquid crystals
    - Room-temperature superconductors?
    - Molecular computing?

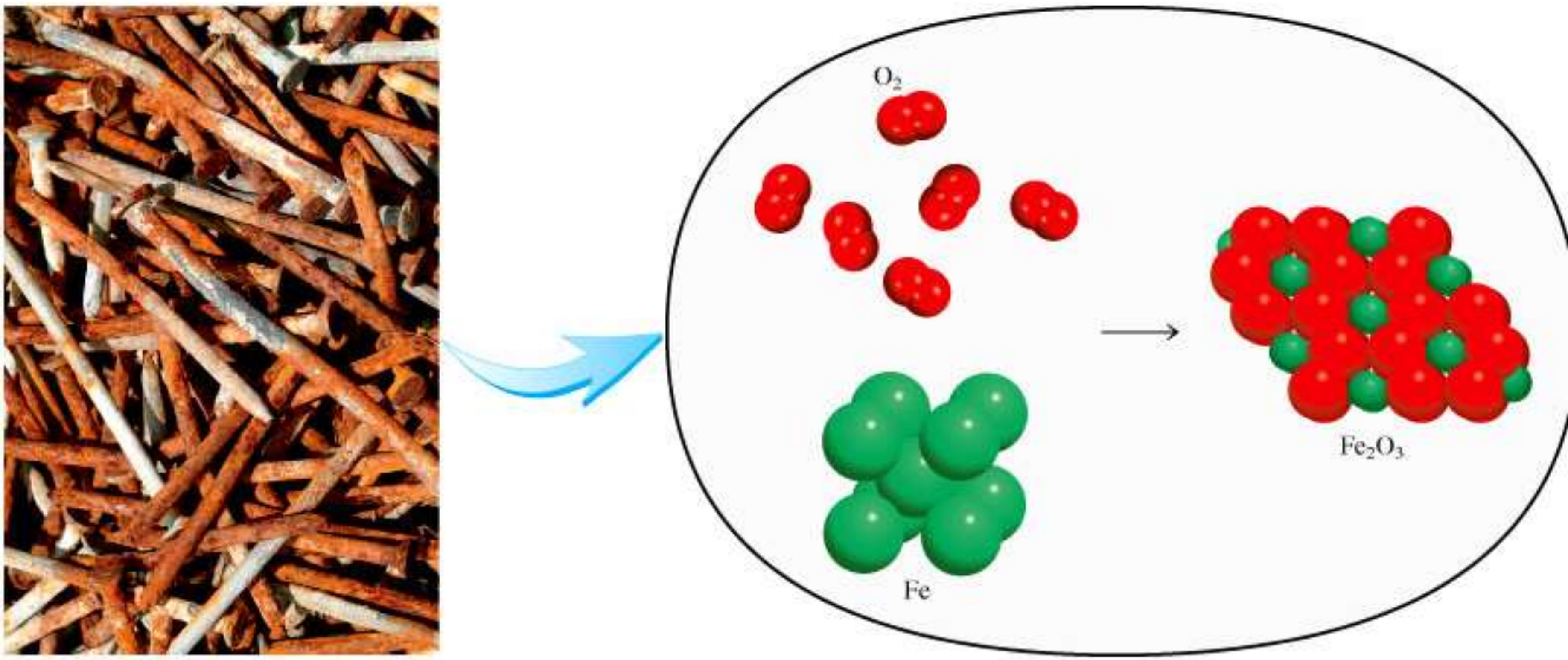


- Food and Agriculture
    - Genetically modified crops
    - “Natural” pesticides
    - Specialized fertilizers

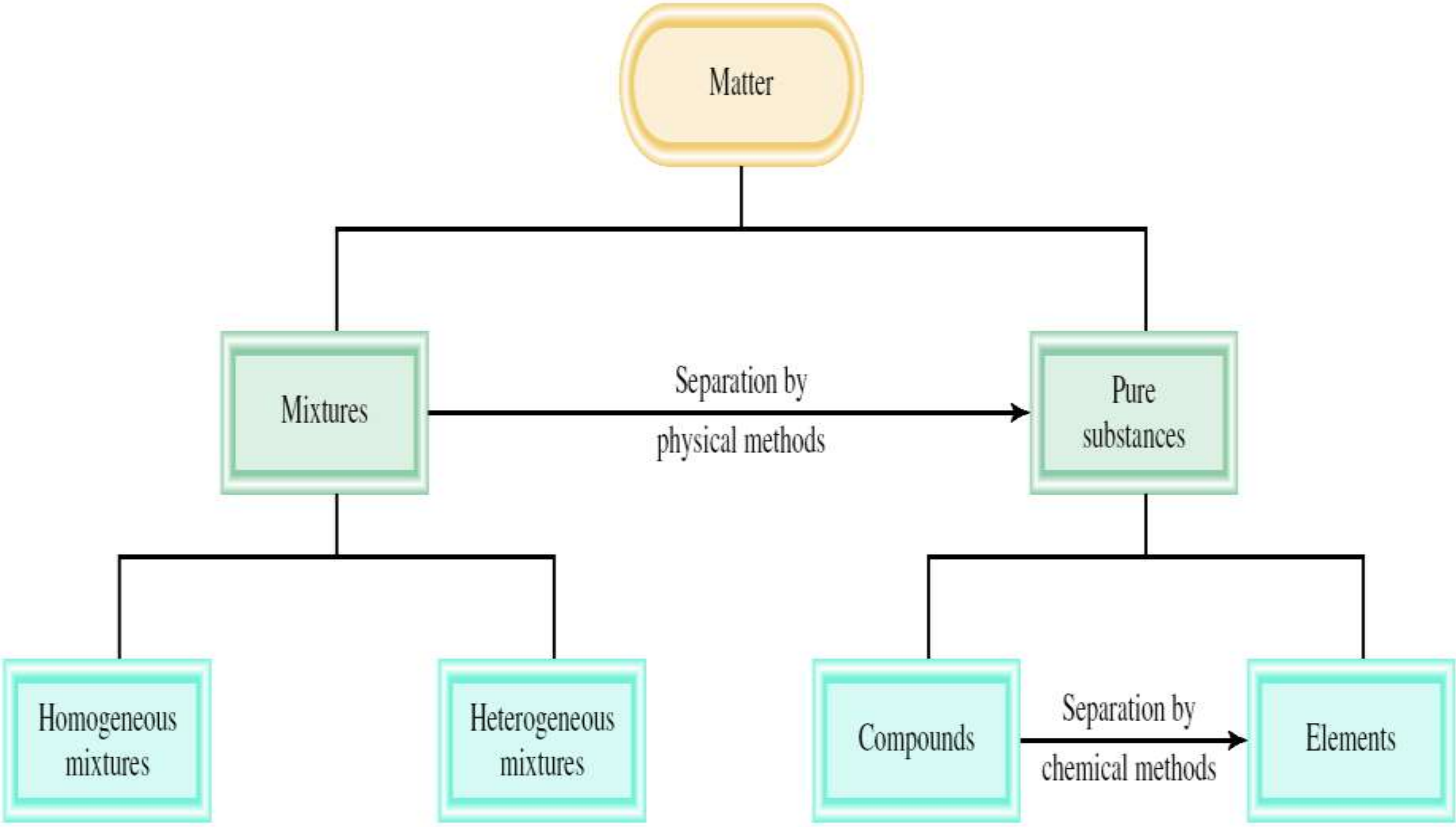
The Study of Chemistry

Macroscopic

Microscopic



Classifications of Matter



Matter

Any thing that occupies space and has mass

Has a definite composition and distinct properties

A combination of two or more substances in which the substances retain their distinct identities

pure substance

Separation by physical methods

mixture

Separation by chemical methods

element

compound

heterogeneous

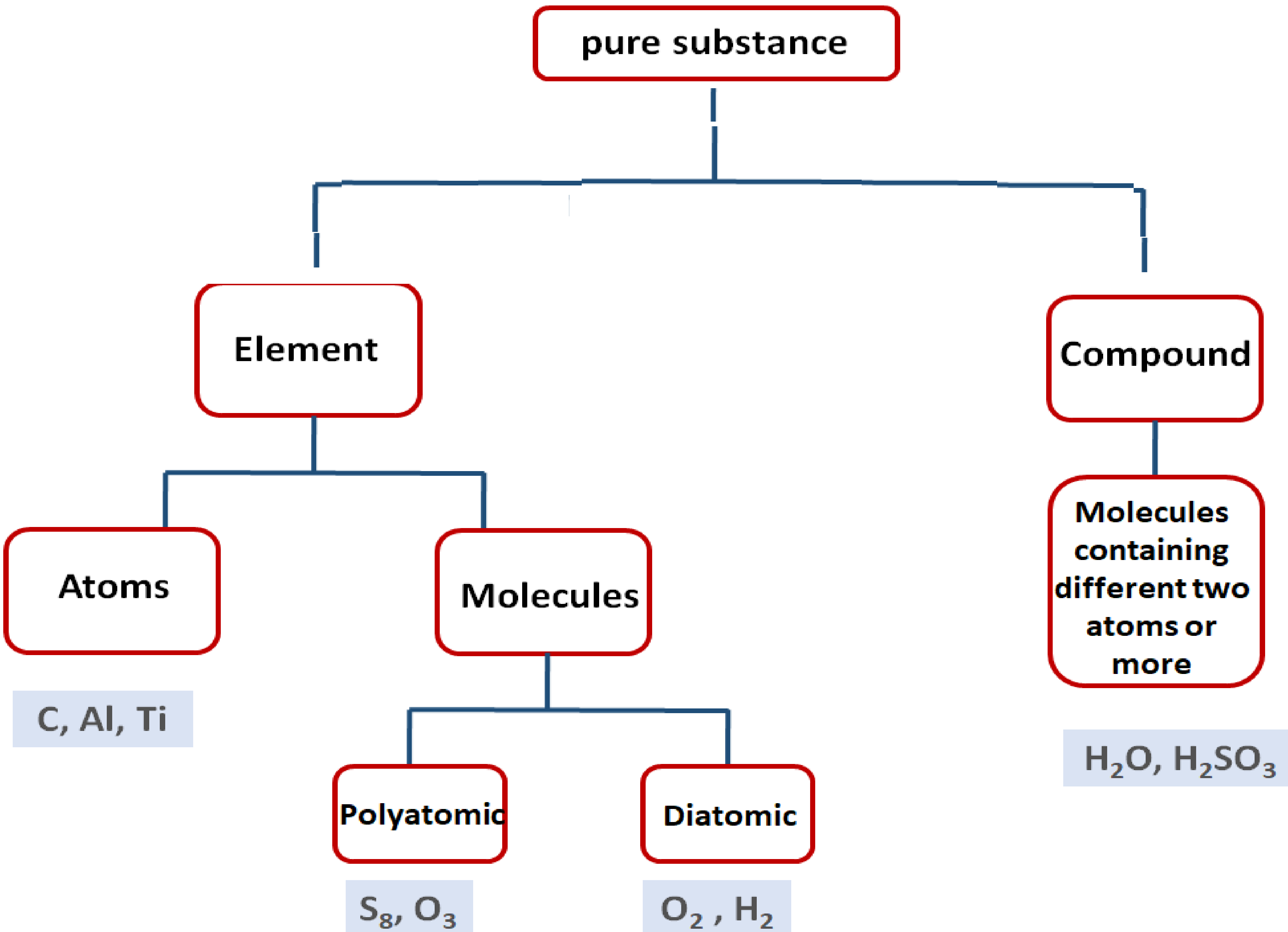
homogeneous

cannot be separated into simpler substances by chemical means

composed of two different elements or more chemically united in fixed

The composition is not uniform

The composition is the same throughout


pure substance
Element
Compound
Atoms
Molecules
Molecules containing different two atoms or more
C, Al, Ti
H_2O, H_2SO_3
Polyatomic
Diatomic
S_8, O_3
O_2 , H_2

Matter

Chemistry is the study of matter and the changes that matter undergoes.

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

Why should YOU study chemistry?

Required for your major

Chemistry impacts our daily lives!

Food and clothing

Health care

Environment

Life itself

Chemistry is the study of matter and the changes it undergoes

Matter is anything that occupies space and has mass.

A ***substance*** is a form of matter that has a definite composition and distinct properties.

liquid nitrogen

gold ingots

silicon crystals

CLASSIFICATIONS OF MATTER

All matter can be classified as either a pure substance or a mixture.

Pure Substance:

All components have the same fixed properties and composition

Element

Compound

Examples of Pure Substances:

pure water (not tap water)

carbon

sodium chloride (but not table salt)

NaCl	
Salt water	
Iron	
sugar	
air	
helium	
water	
salad	

compound **element** **homogeneous mixture**
heterogeneous

NaCl	compound
Salt water	homogeneous mixture
Iron	element
sugar	compound
air	homogeneous mixture
helium	element
water	compound
salad	heterogeneous mixture

Mixture:

A combination of two or more substances in which each substance retains its own chemical identity and properties

Examples of mixtures:

Salt water

Air

Orange juice

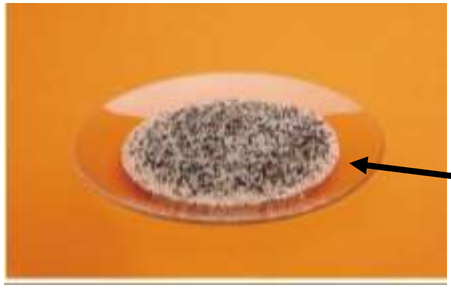
All mixtures can be separated into their individual components using various techniques.

A ***mixture*** is a combination of two or more substances in which the substances retain their distinct identities.

1. ***Homogenous mixture*** – composition of the mixture is the same throughout.

soft drink, milk, solder

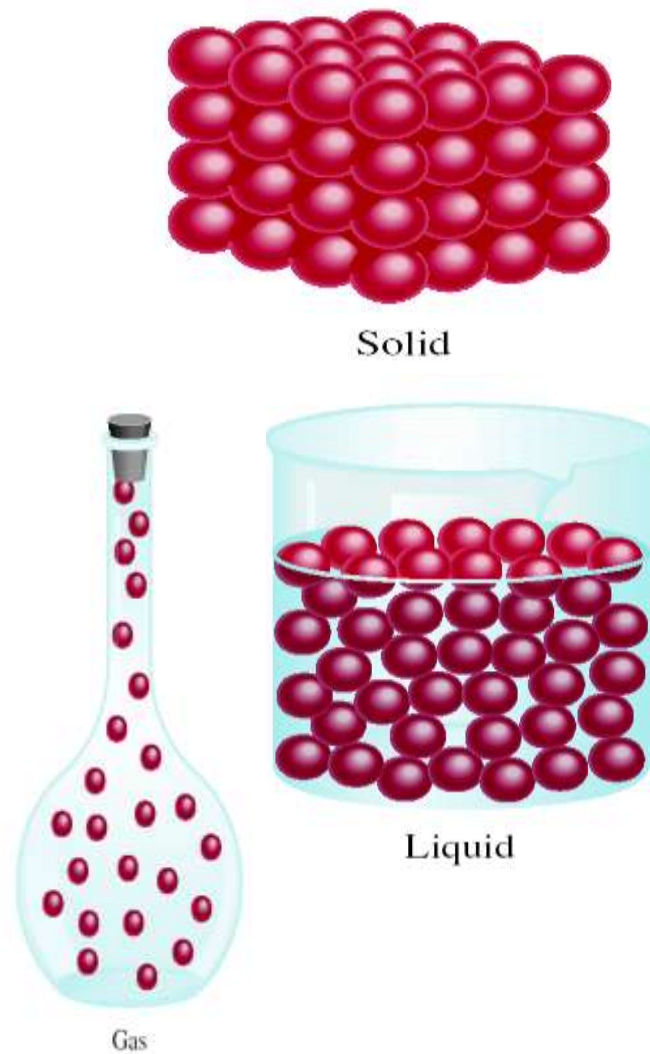
2. ***Heterogeneous mixture*** – composition is not uniform throughout.



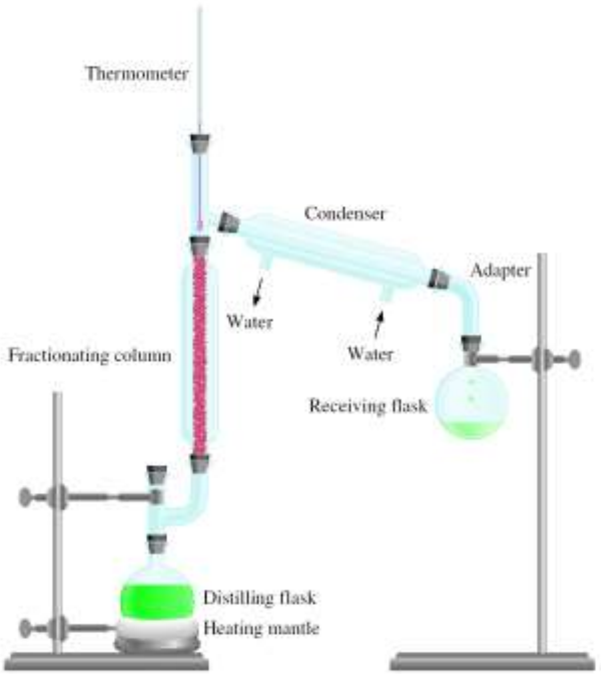
cement, iron filings in sand

Classifications of Matter

- The three physical states of matter:
  - Solid
    - definite shape and volume
  - Liquid
    - definite volume but takes shape of container
  - Gas
    - takes volume and shape of container



Physical means can be used to separate a mixture into its pure components.



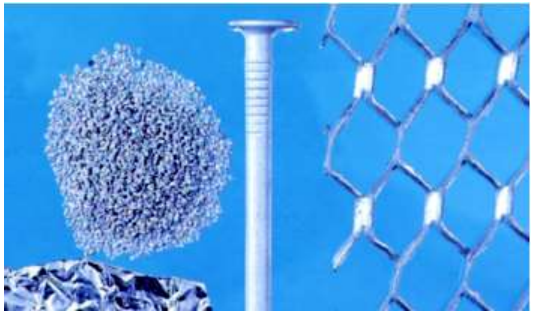
distillation

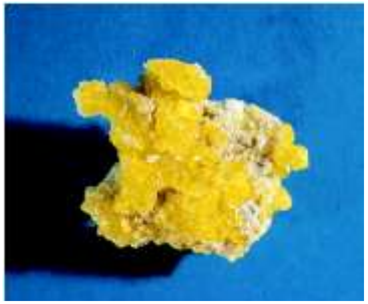
magnet

An ***element*** is a substance that cannot be separated into simpler substances by ***chemical means***.

- 118 elements have been identified
  - 82 elements occur naturally on Earth

    gold, aluminum, lead, oxygen, carbon, sulfur





  - 36 elements have been created by scientists

    technetium, americium, seaborgium

TABLE 1.1 Some Common Elements and Their Symbols

Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bismuth	Bi	Iodine	I	Potassium	K
Bromine	Br	Iron	Fe	Silicon	Si
Calcium	Ca	Lead	Pb	Silver	Ag
Carbon	C	Magnesium	Mg	Sodium	Na
Chlorine	Cl	Manganese	Mn	Sulfur	S
Chromium	Cr	Mercury	Hg	Tin	Sn
Cobalt	Co	Nickel	Ni	Tungsten	W
Copper	Cu	Nitrogen	N	Zinc	Zn

Elements (2)

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Bromine	Br	Iron	Fe	Silicon	Si
Calcium	Ca	Lead	Pb	Silver	Ag
Carbon	C	Magnesium	Mg	Sodium	Na
Chlorine	Cl	Manganese	Mn	Sulfur	S
Chromium	Cr	Mercury	Hg	Tin	Sn
Cobalt	Co	Nickel	Ni	Tungsten	W
Copper	Cu	Nitrogen	N	Zinc	Zn

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A ***compound*** is a substance composed of atoms of two or more elements chemically united in fixed proportions.

Compounds can only be separated into their pure components (elements) by ***chemical*** means.

lithium fluoride

quartz

dry ice – carbon dioxide

Types of Changes

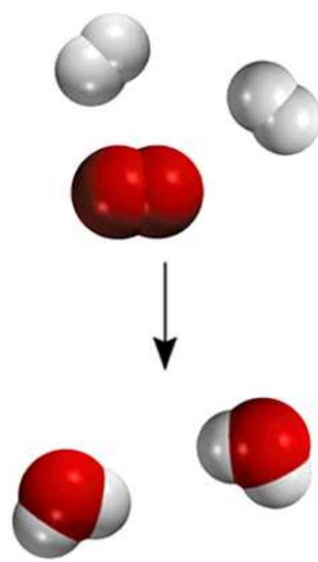
A ***physical change*** does not alter the composition or identity of a substance.

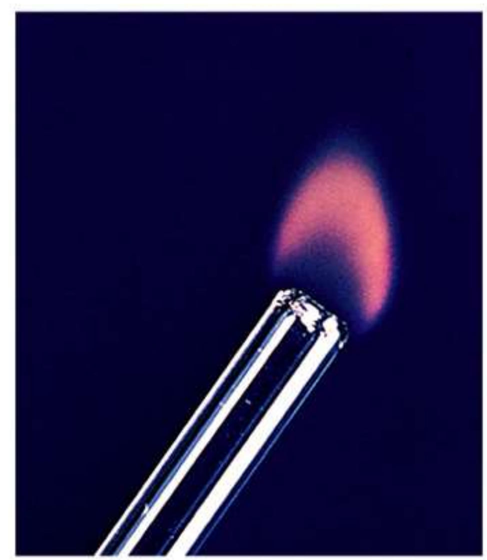
ice melting

sugar dissolving in water

A ***chemical change*** alters the composition or identity of the substance(s) involved.

hydrogen burns in air to form water





The Scientific Method

The ***scientific method*** is a systematic approach to research.

A ***hypothesis*** is a tentative explanation for a set of observations.

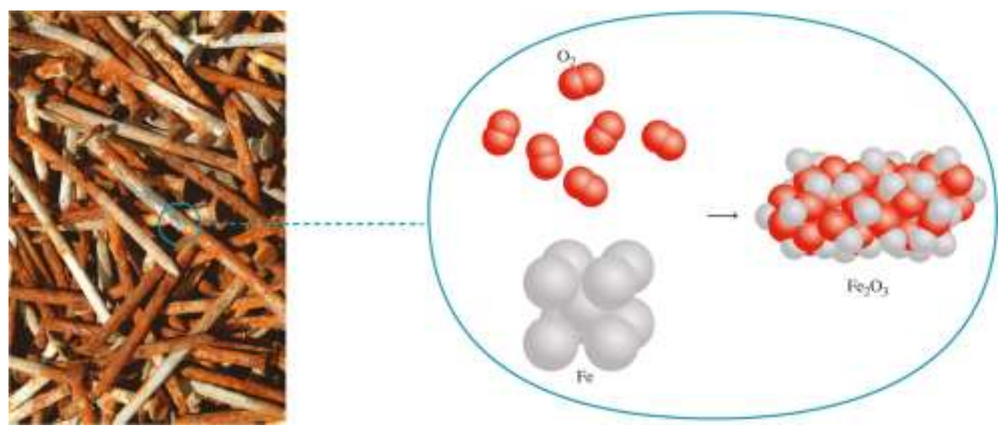
The Scientific Method 2

A ***law*** is a concise statement of a relationship between phenomena that is always the same under the same conditions.

$$\text{Force} = \text{mass} \times \text{acceleration}$$

A ***theory*** is a unifying principle that explains a body of facts and/or those laws that are based on them.

Atomic Theory



Matter - anything that occupies space and has ***mass***.

mass –is a measure of the amount of matter in an object

SI unit of mass is the ***kilogram*** (kg)

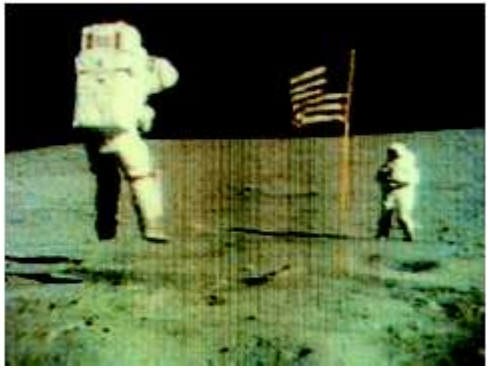
1 kg = 1000 g = 1 x 10^3 g

weight is the force that gravity exerts on an object

weight = *c* x mass

on earth, *c* = 1.0

on moon, *c* ~ 0.1



A 1 kg bar will weigh

1 kg on earth

0.1 kg on moon

Measurement

SI Units

Base Quantity	Name of unit	Symbol
Length	meter	m
Mass	Kilogram	Kg
Time	Second	s
Electrical current	Ampere	A
Temperature	Kelvin	K
Amount of substance	Mole	mol
Luminous intensity	candela	cd
Speed	meter per second	m/s
Pressure	Pascal	Pa

TABLE 1.3 Prefixes Used with SI Units

Prefix	Symbol	Meaning	Example
tera-	T	1,000,000,000,000, or 10^{12}	1 terameter (Tm) = 1×10^{12} m
giga-	G	1,000,000,000, or 10^{9}	1 gigameter (Gm) = 1×10^{9} m
mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) = 1×10^{6} m
kilo-	k	1,000, or 10^{3}	1 kilometer (km) = 1×10^{3} m
deci-	d	1/10, or 10^{-1}	1 decimeter (dm) = 0.1 m
centi-	c	1/100, or 10^{-2}	1 centimeter (cm) = 0.01 m
milli-	m	1/1,000, or 10^{-3}	1 millimeter (mm) = 0.001 m
micro-	μ	1/1,000,000, or 10^{-6}	1 micrometer (μm) = 1×10^{-6} m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) = 1×10^{-9} m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) = 1×10^{-12} m

Measurement in chemistry

The diameter of an atom is approximately 1×10^{-7} mm. What is this diameter when expressed in nanometers?

(a) 1×10^{-18} nm
(b) 1×10^{-15} nm
(c) 1×10^{-9} nm
(d) $\boxed{1 \times 10^{-1} \text{ nm}}$

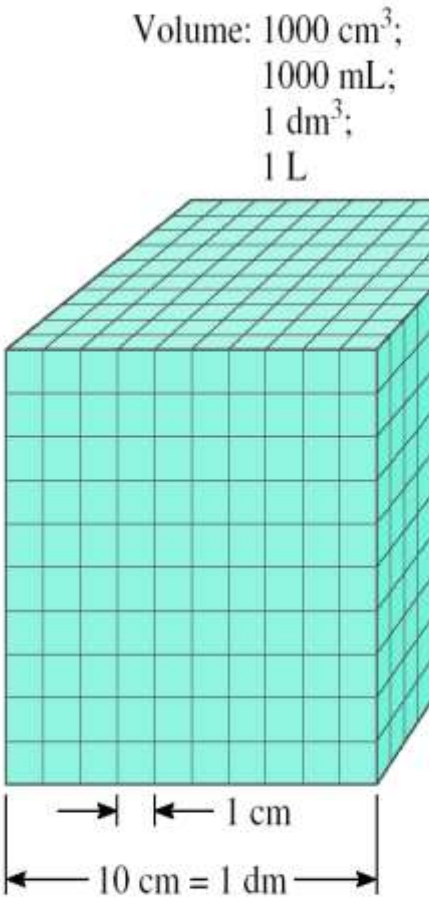
Which of these quantities represents the largest mass?

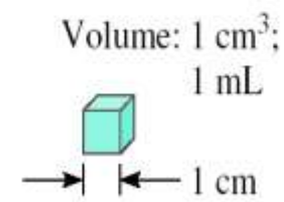
Put all of them in the same unit

(a) 2.0×10^{2} mg $= 2.0 \times 10^{-1}$ g $=$ 0.2 g
(b) 0.0010 kg $= 0.001 \times 10^{3}$ g $=$ 1 g
(c) 1.0×10^{5} μg $= 1 \times 10^{5} \times 10^{-6} = 1 \times 10^{-1}$ g $=$ 0.1 g
(d) $\boxed{2.0 \times 10^{2} \text{ cg}}$ $= 2 \times 10^{2} \times 10^{-2}$ g $=$ 2 g

Volume

- The most commonly used metric units for volume are the liter (L) and the milliliter (mL)
- ***Volume*** – SI derived unit for volume is cubic meter (m^3)
  - A liter is a cube 1 dm long on each side.
  - A milliliter is a cube 1 cm(cm^3) long on each side





$$1\ cm^3 = (1 \times 10^{-2}\ m)^3 = 1 \times 10^{-6}\ m^3$$
$$1\ dm^3 = (1 \times 10^{-1}\ m)^3 = 1 \times 10^{-3}\ m^3$$
$$1\ L = 1000\ ml = 1000\ cm^3 = 1\ dm^3$$
$$1\ ml = 1\ cm^3$$

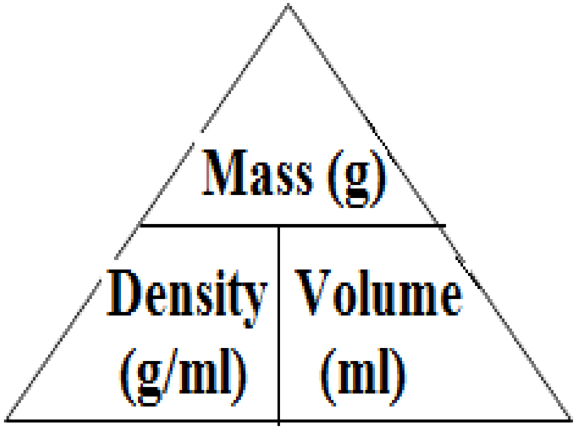
Density

Density is a physical property of a substance.

$$Density = \frac{mass}{Volume}$$

$$d = \frac{m}{V}$$

Mass = density x Vol



Density can be used to predict the behavior materials that are immiscible don’t dissolve in each other

Given 2 immiscible liquids, the lower density liquid will “float” on top of the higher density liquid.

Density can be used to find the mass of a liquid with a known volume.

Density – SI derived unit for density is kg/m^3

$1\ g/cm^3 = 1\ g/mL = 1000\ kg/m^3$

g/cm^3 for solids

g/ml for liquids

g/L for gases

$$\text{density} = \frac{\text{mass}}{\text{volume}} \qquad d = \frac{m}{V}$$

A piece of platinum metal with a density of $21.5\ g/cm^3$ has a volume of $4.49\ cm^3$. What is its mass?

$$d = \frac{m}{V} \qquad m = d \times V = 21.5\ g/\cancel{cm^3} \times 4.49\ \cancel{cm^3} = 96.5\ g$$

A student determines that a piece of an unknown material has a mass of 5.854 g and a volume of 7.57 cm^3. What is the density of the material?

$d = m/V \quad = 5.854\ (g) / 7.75\ (cm^3) = 0.773\ g/cm^3$

The density of mercury, the only metal that is a liquid at room temperature, is 13.6 g/mL.
Calculate the mass of 5.50 mL of the liquid.

Solution

$$d = \frac{m}{V}$$

m = V x d = 13.6 x 5.5 = 74.8 gm

Gold is a precious metal that is chemically unreactive.It is used mainly in jewelry, dentistry, and electronic devices. A piece of gold ingot with a mass of 301 g has a volume of 15.6 cm^3 . Calculate the density of gold.

$$d = \frac{m}{V} = \frac{301}{15.6} = 19.3\ g/cm^3$$

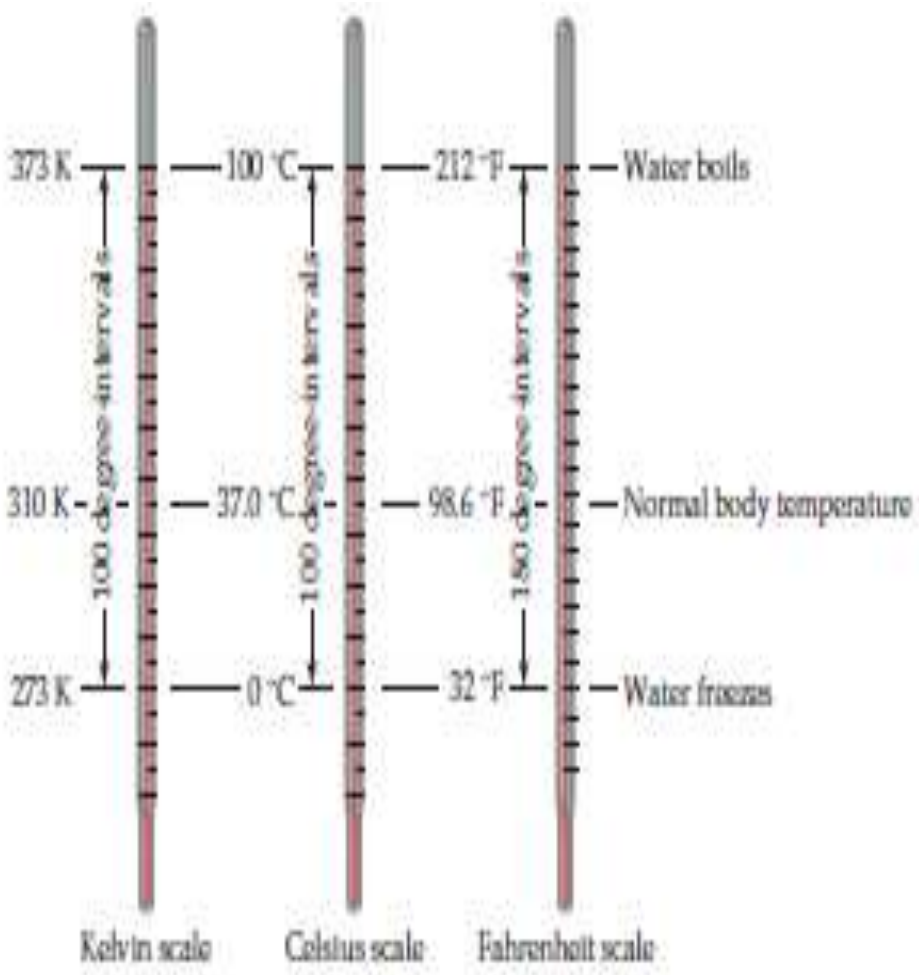
TABLE 1.4

Densities of Some Substances at 25°C

Substance	Density (g/cm^3)
Air*	0.001
Ethanol	0.79
Water	1.00
Mercury	13.6
Table salt	2.2
Iron	7.9
Gold	19.3
Osmium†	22.6

*Measured at 1 atmosphere.
†Osmium (Os) is the densest element known.

A Comparison of Temperature Scales



$$K = {}^0C + 273.15$$

$273\ K = 0\ {}^0C$

$373\ K = 100\ {}^0C$

$${}^0F = \left(\frac{9}{5} \times {}^0C\right) + 32$$

$32\ {}^0F = 0\ {}^0C$

$212\ {}^0F = 100\ {}^0C$

Convert 172.9 ^{0}F to degrees Celsius.

$$^0F = \frac{9}{5} x\ ^0C + 32$$

$$^0F - 32 = \frac{9}{5}\ x\ ^0C$$

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

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

$$^0C = \frac{5}{9} x\ (172.9 - 32) = 78.3$$

(a) Solder is an alloy made of tin and lead that is used in electronic circuits. A certain solder has a melting point of 224°C. What is its melting point in degrees Fahrenheit? (b) Helium has the lowest boiling point of all the elements at -452°F. Convert this temperature to degrees Celsius. (c) Mercury, the only metal that exists as a liquid at room temperature, melts at -38.9°C. Convert its melting point to kelvins.

Solution

$$^{0}F = \frac{9}{5} x\ ^{0}C + 32 \quad =\left(\frac{9}{5}\ x\ \ 224\ ^{0}C\right) +32 = \ 435\ \ ^{0}F$$

$$^{0}C = \frac{5}{9}\ x\ (^{0}F - 32) \quad = \frac{5}{9}\ x\ (\ -452^{0}F - 32) = -269\ ^{0}C$$

$$K = {}^{0}C + 273.15 \quad = \quad -38.9 + 273.15 = 234.3\ K$$

Units of Energy

- The SI unit of energy is the joule (J). $1\ \text{J} = 1\ \dfrac{\text{kg m}^2}{\text{s}^2}$

shows that a mass of 2 kg moving at a speed of 1m/s possesses a kinetic energy of 1 J:

$E_k = \frac{1}{2}\,mv^2 = \frac{1}{2}\,(2\ \text{kg})(1\ \text{m/s})^2 = 1\ \text{kg-m}^2/\text{s}^2 = 1\ \text{J}$

Kilojoule (kJ)   1 kJ = 1000 J

- An older, non-SI unit is still in widespread use: the calorie (cal). Originally defined as the amount of energy needed to raise the temperature of 1 g of water from 14.5^{o}C to 15.5^{o}C

1 cal = 4.184 J

Kilocalories (kcal)   1 kcal = 1000 cal

Chemistry In Action

On 9/23/99, \$125,000,000 Mars Climate Orbiter entered Mar's atmosphere 100 km (62 miles) lower than planned and was destroyed by heat.

force = mass x acceleration

$= 0.4536 \text{ kg} \times 9.81 \text{ m/s}^2$

$= 4.4\,5 \text{ kg m/s}^2$

$= 4.45 \text{ N}$

1 lb $\cancel{=}$ 1 N

1 lb = 4.45 N

"This is going to be the cautionary tale that will be embedded into introduction to the metric system in elementary school, high school, and college science courses till the end of time."

Dimensional Analysis Method of Solving Problems

How many mL are in 1.63 L?

Conversion Unit 1 L = 1000 mL

$$1.63\ \cancel{L} \times \frac{1000\ mL}{1\cancel{L}} = 1630\ mL$$

$$\cancel{1.63\ L \times \frac{1L}{1000\ mL} = 0.001630\ \frac{L^2}{mL}}$$

The speed of sound in air is about 343 m/s. What is this speed in miles per hour?

conversion units

meters to miles

seconds to hours

1 mi = 1609 m 1 min = 60 s 1 hour = 60 min

$$343\ \frac{\cancel{m}}{\cancel{s}} \times \frac{1\ \text{mi}}{1609\ \cancel{m}} \times \frac{60\ \cancel{s}}{1\ \cancel{min}} \times \frac{60\ \cancel{min}}{1\ \text{hour}} = 767\ \frac{\text{mi}}{\text{hour}}$$

A person’s average daily intake of glucose (a form of sugar) is 0.0833 pound (lb). What is this mass in milligrams (mg)? (1 lb = 453.6 g.)

Solution

The sequence of conversions is

$$\text{pounds} \rightarrow \text{grams} \rightarrow \text{milligrams}$$

$$\frac{453.6\ \text{g}}{1\ \text{lb}} \text{ and } \frac{1\ \text{mg}}{1\times10^{-3}\ \text{g}}$$

$$?\ \text{mg} = 0.0833\ \cancel{\text{lb}} \times \frac{453.6\ \cancel{\text{g}}}{1\ \cancel{\text{lb}}} \times \frac{1\,\text{mg}}{1\times10^{-3}\ \cancel{\text{g}}} = 3.78\times10^{4}\ \text{mg}$$

An average adult has 5.2 L of blood. What is the volume of blood in m^3?

$$\frac{1\times10^{-2}\ \text{m}}{1\,\text{cm}}\times\frac{1\times10^{-2}\ \text{m}}{1\,\text{cm}}\times\frac{1\times10^{-2}\ \text{m}}{1\,\text{cm}}=\left(\frac{1\times10^{-2}\ \text{m}}{1\,\text{cm}}\right)^3$$

$$1\,\text{cm}^3=1\times10^{-6}\ \text{m}^3.$$

$$?\ \text{m}^3=5.2\ \cancel{\text{L}}\times\frac{1000\ \text{cm}^3}{1\ \cancel{\text{L}}}\times\left(\frac{1\times10^{-2}\ \text{m}}{1\,\text{cm}}\right)^3=5.2\times10^{-3}\ \text{m}^3$$

Scientific Notation 1

- The number of atoms in 12 g of carbon:

602,200,000,000,000,000,000,000

$$6.022 \times 10^{23}$$

- The mass of a single carbon atom in grams:

0.0000000000000000000000199

$$1.99 \times 10^{-23}$$

$$N \times 10^{n}$$

N is a number between 1 and 10

n is a positive or negative integer

Scientific Notation 2

568.762

$\leftarrow$ move decimal left

$n > 0$

$$568.762 = 5.68762 \times 10^{2}$$

0.00000772

$\rightarrow$ move decimal right

$n < 0$

$$0.00000772 = 7.72 \times 10^{-6}$$

- <u>Addition or Subtraction</u>

1. Write each quantity with the same exponent ***n***
2. Combine N_1 and N_2
3. The exponent, ***n***, remains the same

$$4.31 \times 10^{4} + 3.9 \times 10^{3} =$$

$$4.31 \times 10^{4} + 0.39 \times 10^{4} =$$

$$4.70 \times 10^{4}$$

Scientific Notation 3

- <u>Multiplication</u>

1. Multiply N_1 and N_2
2. Add exponents n_1 and n_2

$$\left(4.0\times10^{-5}\right)\times\left(7.0\times10^{3}\right)=$$
$$\left(4.0\times7.0\right)\times\left(10^{-5+3}\right)=$$
$$28\times10^{-2}=$$
$$2.8\times10^{-1}$$

<u>Division</u>

1. Divide N_1 and N_2
2. Subtract exponents n_1 and n_2

$$8.5\times10^{4}\div5.0\times10^{9}=$$
$$\left(8.5\div5.0\right)\times10^{4-9}=$$
$$1.7\times10^{-5}$$

Significant Figures 1

- Any digit that is not zero is significant

  1.234 kg **4** significant figures
- Zeros between nonzero digits are significant

  606 m **3** significant figures
- Zeros to the left of the first nonzero digit are **not** significant

  0.08 L **1** significant figure
- If a number is greater than 1, then all zeros to the right of the decimal point are significant

  2.0 mg **2** significant figures
- If a number is less than 1, then only the zeros that are at the end and in the middle of the number are significant

  0.00420 g **3** significant figures

(1) Express 568.762 in scientific notation:

$$568.762 = 5.68762 \times 10^{2}$$

Note that the decimal point is moved to the left by two places and $n = 2$

(2) Express 0.00000772 in scientific notation:

$$0.00000772 = 7.72 \times 10^{-6}$$

Here the decimal point is moved to the right by six places and $n = -6$

EXAMPLE 1.4

Determine the number of significant figures in the following measurements: (a) 478 cm, (b) 6.01 g, (c) 0.825 m, (d) 0.043 kg, (e) 1.310×10^{22} atoms, (f) 7000 mL.

Solution (a) Three, because each digit is a nonzero digit. (b) Three, because zeros between nonzero digits are significant. (c) Three, because zeros to the left of the first nonzero digit do not count as significant figures. (d) Two. Same reason as in (c). (e) Four, because the number is greater than one so all the zeros written to the right of the decimal point count as significant figures. (f) This is an ambiguous case. The number of significant figures may be four (7.000×10^3), three (7.00×10^3), two (7.0×10^3), or one (7×10^3). This example illustrates why scientific notation must be used to show the proper number of significant figures.

Significant Figures 2

- Addition or Subtraction

- The answer cannot have more digits to the right of the decimal point than any of the original numbers.

$$\begin{array}{rl} 89.332 & \\ \underline{+1.1} & \leftarrow \text{one significant figure after decimal point} \\ 90.432 & \leftarrow \text{roundoff to } 90.4 \end{array}$$

$$\begin{array}{rl} 3.70 & \\ \underline{-2.9133} & \leftarrow \text{two significant figure after decimal point} \\ 0.7867 & \leftarrow \text{roundoff to } 0.79 \end{array}$$

Significant Figures 3

- <u>Multiplication or Division</u>
- The number of significant figures in the result is set by the original number that has the ***smallest*** number of significant figures.

$$\underset{\substack{\uparrow \\ \text{3 sig figs}}}{4.51} \times 3.6666 = \underset{\substack{\uparrow \\ \text{round to} \\ \text{3 sig figs}}}{16.536366} = 16.5$$

$$\underset{\substack{\uparrow \\ \text{2 sig figs}}}{6.8} \div 112.04 = \underset{\substack{\uparrow \\ \text{round to} \\ \text{2 sig figs}}}{0.0606926} = 0.061$$

Significant Figures 4

Exact Numbers

Numbers from definitions or numbers of objects are considered to have an infinite number of significant figures.

The average of three measured lengths: 6.64, 6.68 and 6.70?

$$\frac{6.64 + 6.68 + 6.70}{3} = 6.67333 = 6.67$$

Because 3 is an **exact number**

EXAMPLE 1.5

Carry out the following arithmetic operations to the correct number of significant figures: (a) 11,254.1 g + 0.1983 g, (b) 66.59 L − 3.113 L, (c) 8.16 m × 5.1355, (d) 0.0154 kg ÷ 88.3 mL, (e) 2.64×10^3 cm + 3.27×10^2 cm.

Solution In addition and subtraction, the number of decimal places in the answer is determined by the number having the lowest number of decimal places. In multiplication and division, the significant number of the answer is determined by the number having the smallest number of significant figures.

(a)
$$\begin{array}{r} 11{,}254.1\ \text{g} \\ +\quad 0.1983\ \text{g} \\ \hline 11{,}254.2983\ \text{g} \end{array} \longleftarrow \text{round off to } 11{,}254.3\ \text{g}$$

(b)
$$\begin{array}{r} 66.59\ \text{L} \\ -\ 3.113\ \text{L} \\ \hline 63.477\ \text{L} \end{array} \longleftarrow \text{round off to } 63.48\ \text{L}$$

(c) $8.16\ \text{m} \times 5.1355 = 41.90568\ \text{m} \longleftarrow$ round off to 41.9 m

(d) $\dfrac{0.0154\ \text{kg}}{88.3\ \text{mL}} = 0.000174405436\ \text{kg/mL} \longleftarrow$ round off to 0.000174 kg/mL or 1.74×10^{-4} kg/mL

(e) First we change 3.27×10^2 cm to 0.327×10^3 cm and then carry out the addition $(2.64\ \text{cm} + 0.327\ \text{cm}) \times 10^3$. Following the procedure in (a), we find the answer is 2.97×10^3 cm.

Dimensional Analysis Method of Solving Problems

1. Determine which unit conversion factor(s) are needed
2. Carry units through calculation
3. If all units cancel except for the ***desired unit(s)***, then the problem was solved correctly.

$$\text{given quantity} \times \text{conversion factor} = \text{desired quantity}$$

$$\cancel{\text{given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

EXAMPLE 1.6

A person's average daily intake of glucose (a form of sugar) is 0.0833 pound (lb). What is this mass in milligrams (mg)? (1 lb = 453.6 g.)

Strategy The problem can be stated as

$$? \text{ mg} = 0.0833 \text{ lb}$$

The relationship between pounds and grams is given in the problem. This relationship will enable conversion from pounds to grams. A metric conversion is then needed to convert grams to milligrams ($1 \text{ mg} = 1 \times 10^{-3} \text{ g}$). Arrange the appropriate conversion factors so that pounds and grams cancel and the unit milligrams is obtained in your answer.

Solution The sequence of conversions is

$$\text{pounds} \longrightarrow \text{grams} \longrightarrow \text{milligrams}$$

Using the following conversion factors

$$\frac{453.6 \text{ g}}{1 \text{ lb}} \quad \text{and} \quad \frac{1 \text{ mg}}{1 \times 10^{-3} \text{ g}}$$

we obtain the answer in one step:

$$? \text{ mg} = 0.0833 \cancel{\text{lb}} \times \frac{453.6 \cancel{\text{g}}}{1 \cancel{\text{lb}}} \times \frac{1 \text{ mg}}{1 \times 10^{-3} \cancel{\text{g}}} = \boxed{3.78 \times 10^{4} \text{ mg}}$$

EXAMPLE 1.7

An average adult has 5.2 L of blood. What is the volume of blood in m^3?

Strategy The problem can be stated as

$$? \text{ m}^3 = 5.2 \text{ L}$$

How many conversion factors are needed for this problem? Recall that $1 \text{ L} = 1000 \text{ cm}^3$ and $1 \text{ cm} = 1 \times 10^{-2} \text{ m}$.

Solution We need two conversion factors here: one to convert liters to cm^3 and one to convert centimeters to meters:

$$\frac{1000 \text{ cm}^3}{1 \text{ L}} \quad \text{and} \quad \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}$$

Because the second conversion factor deals with length (cm and m) and we want volume here, it must therefore be cubed to give

$$\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \times \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \times \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} = \left(\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}\right)^3$$

This means that $1 \text{ cm}^3 = 1 \times 10^{-6} \text{ m}^3$. Now we can write

$$? \text{ m}^3 = 5.2 \cancel{\text{L}} \times \frac{1000 \text{ cm}^3}{1 \cancel{\text{L}}} \times \left(\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}\right)^3 = \boxed{5.2 \times 10^{-3} \text{ m}^3}$$

Example 1.8

Liquid nitrogen is obtained from liquefied air and is used to prepare frozen goods and in low-temperature research.

The density of the liquid at its boiling point (−196°C or 77K) is 0.808 g/cm^3. Convert the density to units of kg/m^3.

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liquid nitrogen

Example 1.8

Solution

In Example 1.7 we saw that $1\,\text{cm}^3 = 1\times10^{-6}\ \text{m}^3$.

The conversion factors are

$$\frac{1\,\text{kg}}{1000\ \text{g}} \text{ and } \frac{1\,\text{cm}^3}{1\times10^{-6}\ \text{m}^3}$$

Finally

$$?\ \text{kg/m}^3 = \frac{0.808\ \cancel{\text{g}}}{1\ \cancel{\text{cm}^3}} \times \frac{1\,\text{kg}}{1000\ \cancel{\text{g}}} \times \frac{1\ \cancel{\text{cm}^3}}{1\times10^{-6}\ \text{m}^3} = 808\,\text{kg/m}^3$$

Check Because $1\,\text{m}^3 = 1\times10^{6}\ \text{cm}^3$,

we would expect much more mass in $1\,\text{m}^3$ than $1\,\text{cm}^3$

Therefore, the answer is reasonable.

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Example 1.9

A modern pencil "lead" is actually composed primarily of graphite, a form of carbon.

Estimate the mass of the graphite core in a standard No. 2 pencil before it is sharpened.

The volume of a cylinder *V* is given by

$$V = \pi r^2 l$$

where *r* is the radius and *l* is the length.

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Example 1.9

Solution

Converting the diameter of the lead to units of cm gives

$$2\text{ mm}\times\frac{1\text{ cm}}{10\text{ mm}}=0.2\text{ cm}$$

Which, along with the length of the lead, gives

$$V=\pi\left(\frac{0.2\,\text{cm}}{2}\right)^2\times 18\ \text{cm} \qquad =0.57\text{ cm}^3$$

Rearranging Equation (1.1) gives

$$\text{m}=\text{d}\times\text{v}$$

$$=2.2\,\frac{\text{g}}{\text{cm}^3}\times 0.57\text{ cm}^3 \quad =1\text{ g}$$

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- 1- What is the equivalent temperature for 98.6 Fahrenheit in Kelvin?

(a) 37 **(b) 310** (c) 471.6 (d) -273.1

2 . An unknown sample has a mass of 13.9 g and a volume of 17.4 mL. What is its density (g/mL)?

(a) 0.798 (b) 1.04 (c) 3.16 (d) 4.62

3. Which of the following is equal to 100 mL?

a) 0.001 L b) 0.01 L **c) 0.1 L** d) 0.0001 L

4 - 32° Celsius is ____________ kelvin:

a) 240 b) -240 c) 305 d) -305

5- Absolute zero is:

a) -273 K b) -273°F c) -273°C

6- A solid has a mass of 20 grams and occupied a volume of 5.0 mL. What is its density in g/mL?

a) 4.0 b) 0.25 c) 100 d) 2.5

Ex. What is the Kelvin temperature of a solution at 25 °C?

- $$T_{\mathrm{K}} = (25\ ^{\circ}\mathrm{C} + 273.15\ ^{\circ}\mathrm{C})\frac{1\,\mathrm{K}}{1\ ^{\circ}\mathrm{C}}$$ **= 298 K**

- **Convert 77 K to the Celsius scale.**

$$T_{\mathrm{K}} = (t_{\mathrm{C}} + 273.15\ ^{\circ}\mathrm{C})\frac{1\,\mathrm{K}}{1\ ^{\circ}\mathrm{C}} \qquad t_{\mathrm{C}} = (T_{\mathrm{K}} - 273.15\,\mathrm{K})\frac{1\ ^{\circ}\mathrm{C}}{1\,\mathrm{K}}$$

$$t_{\mathrm{C}} = (77\,\mathrm{K} - 273.15\,\mathrm{K})\frac{1\ ^{\circ}\mathrm{C}}{1\,\mathrm{K}}$$ **= –196 °C**

- The density of an object is the ratio of its mass to its volume. What is the derived SI unit for density?

  a. kg m/s^3 b. kg m/s   **c. kg/m^3**   d. m/s^2

. What is the number needed to complete the following: 1 dm = _ m?

  a. 10   b. 20   c. **1**   **d. 0.1**

- . The SI base units of temperature and mass, respectively, are

  a. degree and gram.   **b. kelvin and kilogram.**

  c. Celsius and milligram.   d. degree and kilogram.

-

- . The SI prefixes giga and micro, indicate respectively:

  a**. 10^9 and 10^{-6}**   b. 10^{-9} and 10^{-6}

  c. 10^6 and 10^{-3} d. 10^3 and 10^{-3}

Units of Energy

Joule (J)

- KE possessed by 2 kg object moving at speed of 1 m/s.

$$1J = \frac{1}{2}(2kg)\left(\frac{1m}{1s}\right)^2 \qquad 1J = \frac{1kg \cdot m^2}{s^2}$$

- If calculated value is greater than 1000 J, use kJ
- 1 kJ = 1000 J

Units of Energy

calorie (cal)

- Energy needed to raise T of 1 g H_2O by 1 °C
  - **1 cal = 4.184 J (exactly)**
  - 1 kcal = 1000 cal
  - 1 kcal = 4.184 kJ

1 nutritional Calorie (Cal)

- (note capital C)
- 1 Cal = 1000 cal = 1 kcal
- 1 kcal = 4.184 kJ

Which is a unit of energy?

Pascal newton joule watt ampere

Question 1

Which of the following is an example of a physical property?

A) combustibility **B)** corrosiveness

C) Explosiveness **D)** density **E)** A and D

Question 2

Which of the following represents the greatest mass?

A) 2.0×10^3 mg = 2g **B)** 10.0 dg = 1g

C) 0.0010 kg = 1g **D)** 1.0×10^6 μg = 1g

E) 3.0×10^{12} pg = 3g

Question 3

Convert 240 K and 468 K to the Celsius scale.

A) 513°C and 741°C **B)** -59°C and 351°C

C) -18.3°C and 108°C **D)** -33°C and 195°C

Question 4

Calculate the volume occupied by 4.50×10^2 g of gold (density = 19.3 g/cm^3).

A) 23.3 cm^3 **B)** 8.69×10^3 cm

C) 19.3 cm^3 **D)** 450 cm^3

Question 5

Convert 240 K and 468 K to the Celsius scale.

A) 513°C and 741°C **B)** -59°C and 351°C

C) -18.3°C and 108°C**D)** -33°C and 195°C

Question 6

Convert 240 K and 468 K to the Celsius scale.

A) 513°C and 741°C **B)** -59°C and 351°C

B) **C)** -18.3°C and 108°C **D)** -33°C and 195°C

Question 7

The melting point of bromine is -7°C. What is this melting point expressed in °F?

A) 45°F **B)** -28°F **C)** -13°F **D)** 19°F

E) None of these is within 3°F of the correct answer.

Question 8

The value of 345 mm is a measure of

A) Temperature **B)** density **C)** volume

D) distance **E)** Mass