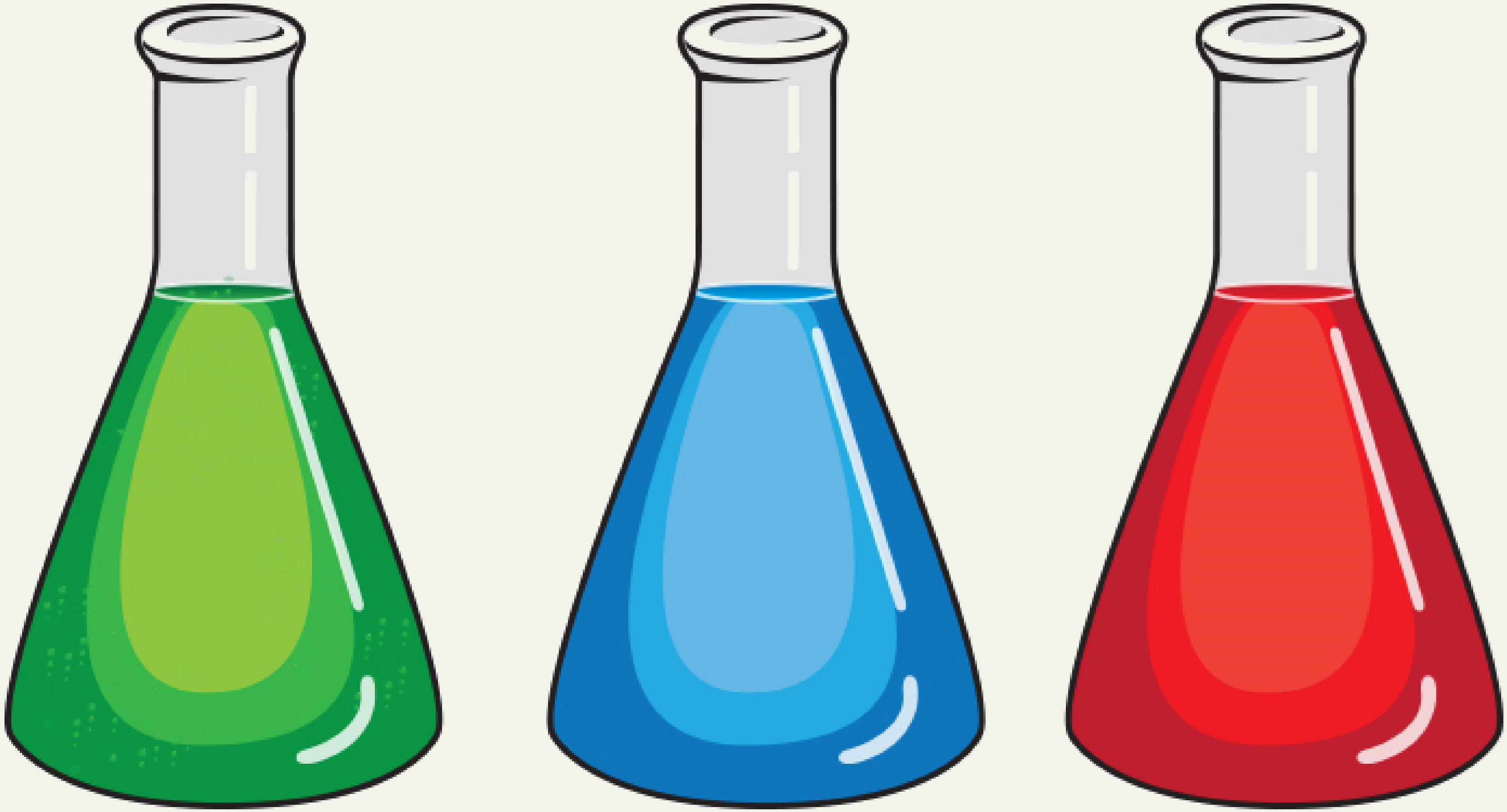
GENERAL CHEMISTRY



The Essential Principles

CHAPTER 1

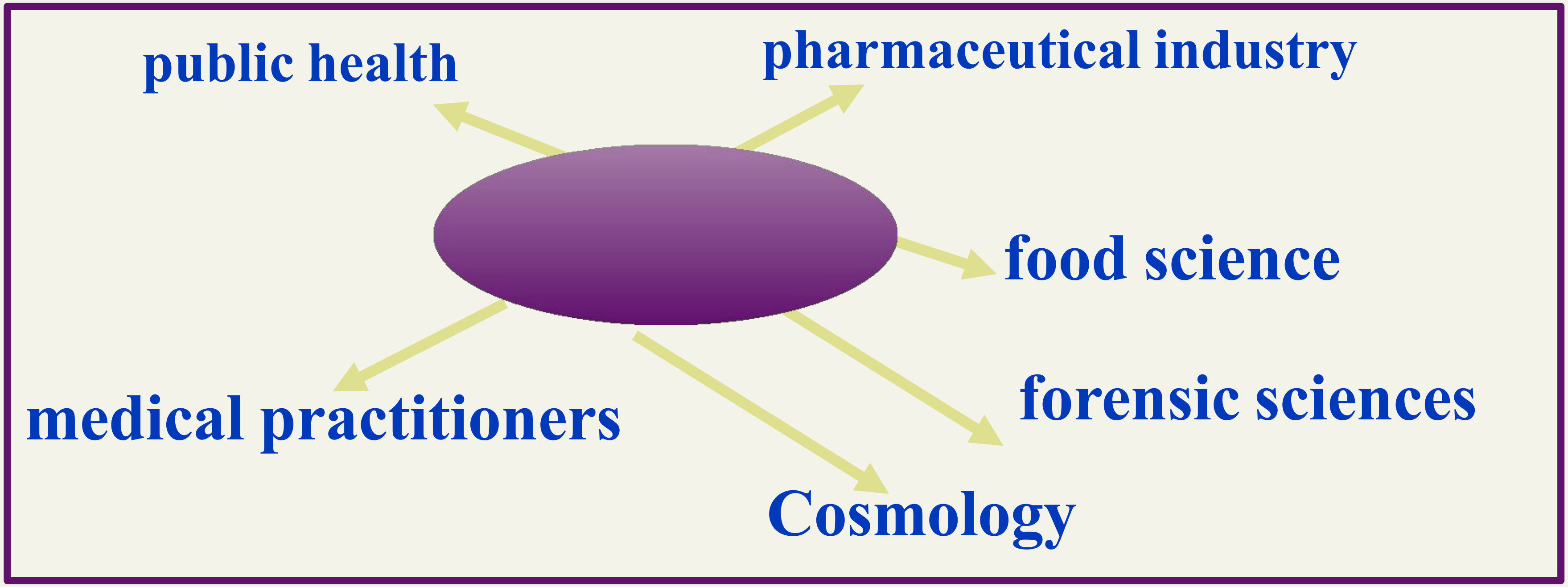
THE Matter: Atoms, Ions and Molecules

Introduction

Chemistry is the science that deals with the study of matter, its properties, composition, and behavior. Chemistry is one of the most important branches of natural sciences, typically related to mining, dyes, medicine, drugs, and some industries such as leather tanning and glass manufacturing. Sometimes chemistry is called the science of matter, i.e. the study of matter and the changes that occur inside it, in particular its properties, structure, composition, behavior, and interactions that occur in that matter. In general, chemistry is associated with other natural sciences such as astronomy, geology, physics and biology, which greatly affects our daily life.

The Importance of Chemistry

Role of Chemistry



Chemistry is the study of matter

- Study the chemical and physical properties of matter.
- Study the chemical and physical changes that matter undergoes.
- Study energy changes (gains or loses) as matter changes.

Matter

- **Matter** - **anything that has mass and occupies space**
- **Properties** - **characteristics of matter scientists can use to categorize different types of matter, such as state, density, reactivity,…etc.**
- **Ways to Categorize matter:**
    1. By State
    2. By Composition

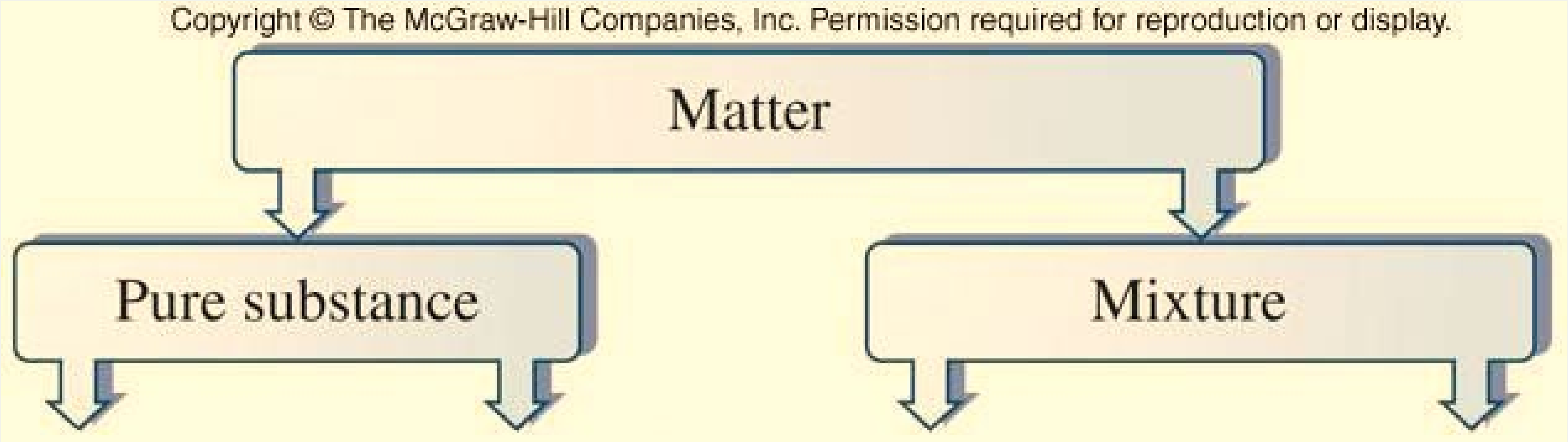
Three **States** of Matter

1. **Gas** - particles widely separated, no definite shape or volume

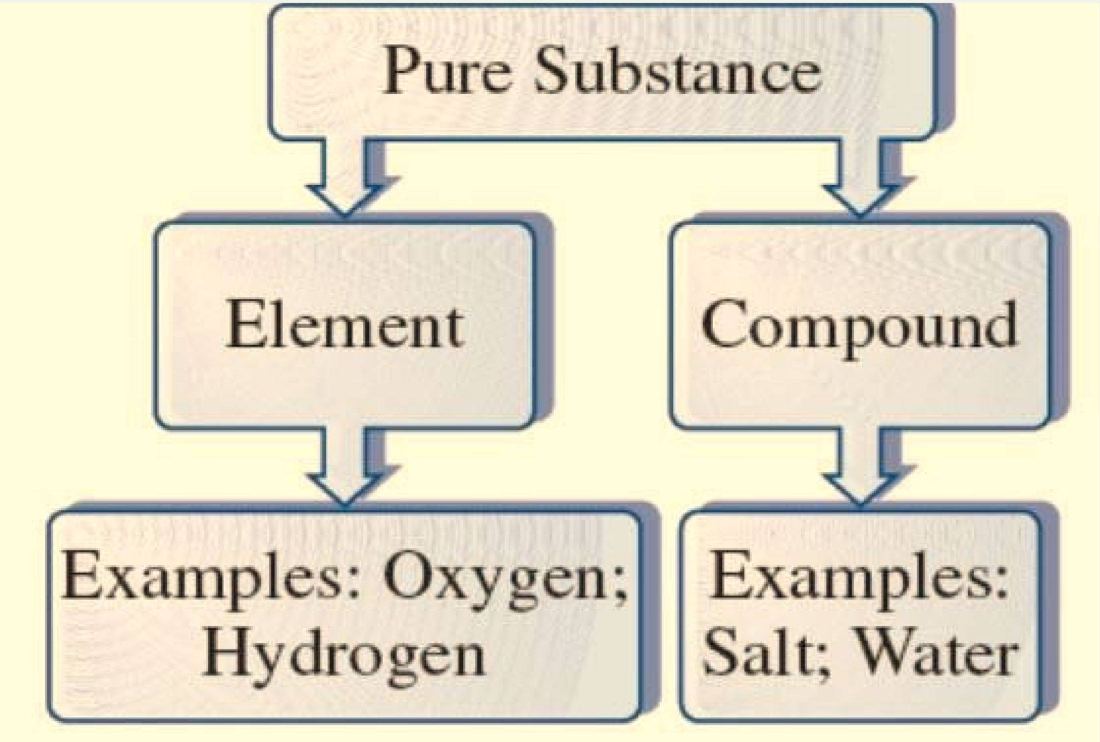
2. **Liquid** - particles closer together, definite volume but no definite shape

3. **Solid** - particles are very close together, define shape and definite volume. Particles are either random (amorphous) or ordered (crystalline)

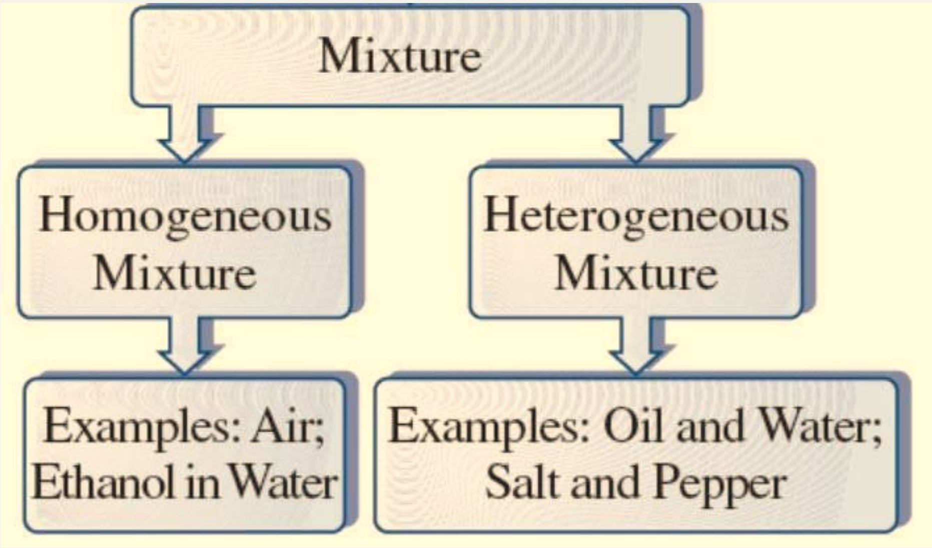
Composition of Matter



- **Pure substance** - a substance that has only one component with specific chemical and physical properties.
- **Mixture** - a combination of two or more pure substances in which each substance retains its own identity, not undergoing a chemical reaction.

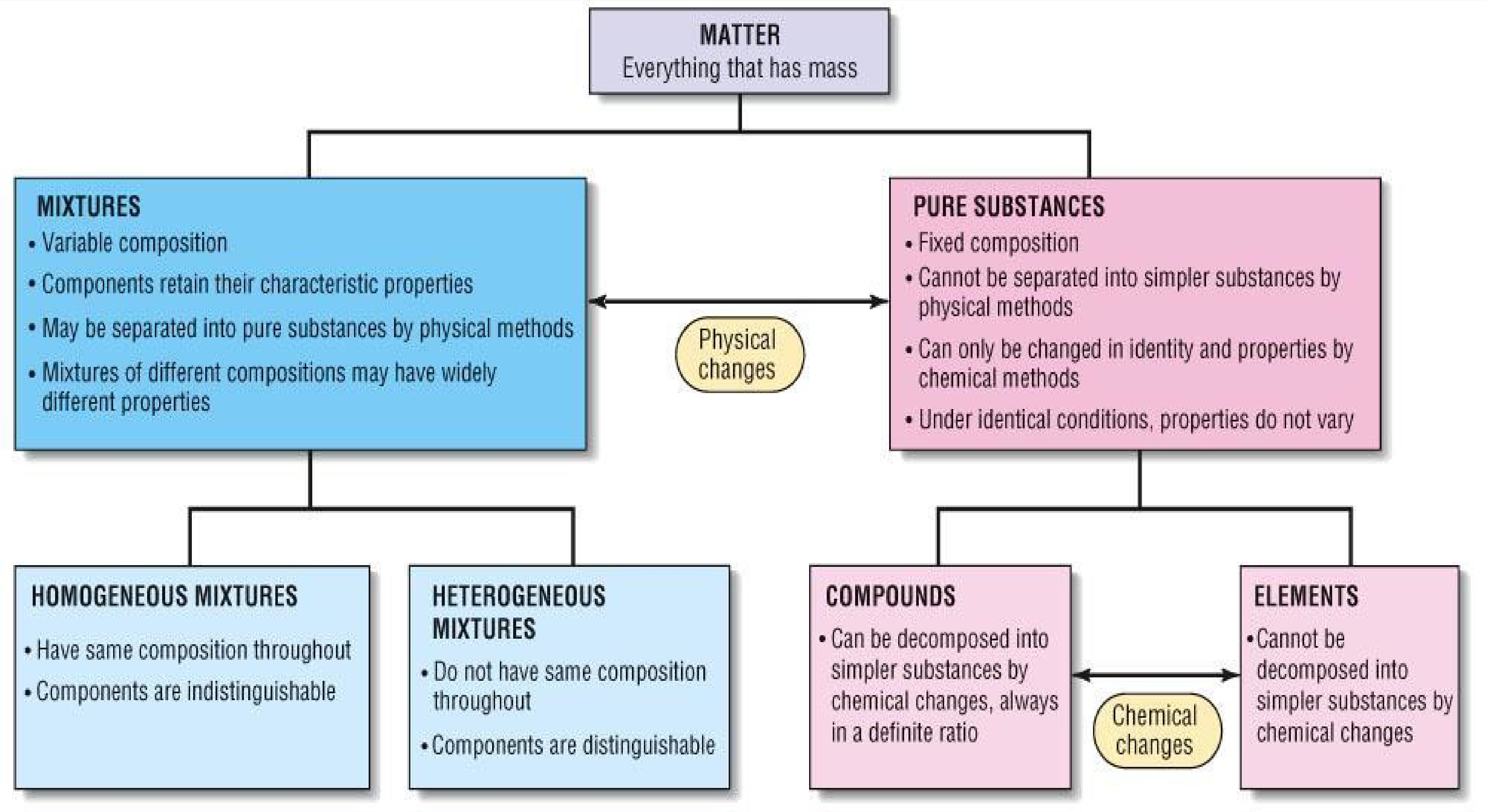


- **Element** - a pure substance that cannot be changed into a simpler form of matter by any chemical reaction
- **Compound** - a pure substance resulting from the combination of two or more elements in a definite, reproducible way, in a fixed ratio



- **Mixture** - a combination of two or more pure substances in which each substance retains its own identity
- ***Homogeneous*** - uniform composition, particles well mixed, thoroughly intermingled on the molecular or atomic level
- ***Heterogeneous*** – nonuniform composition, random placement

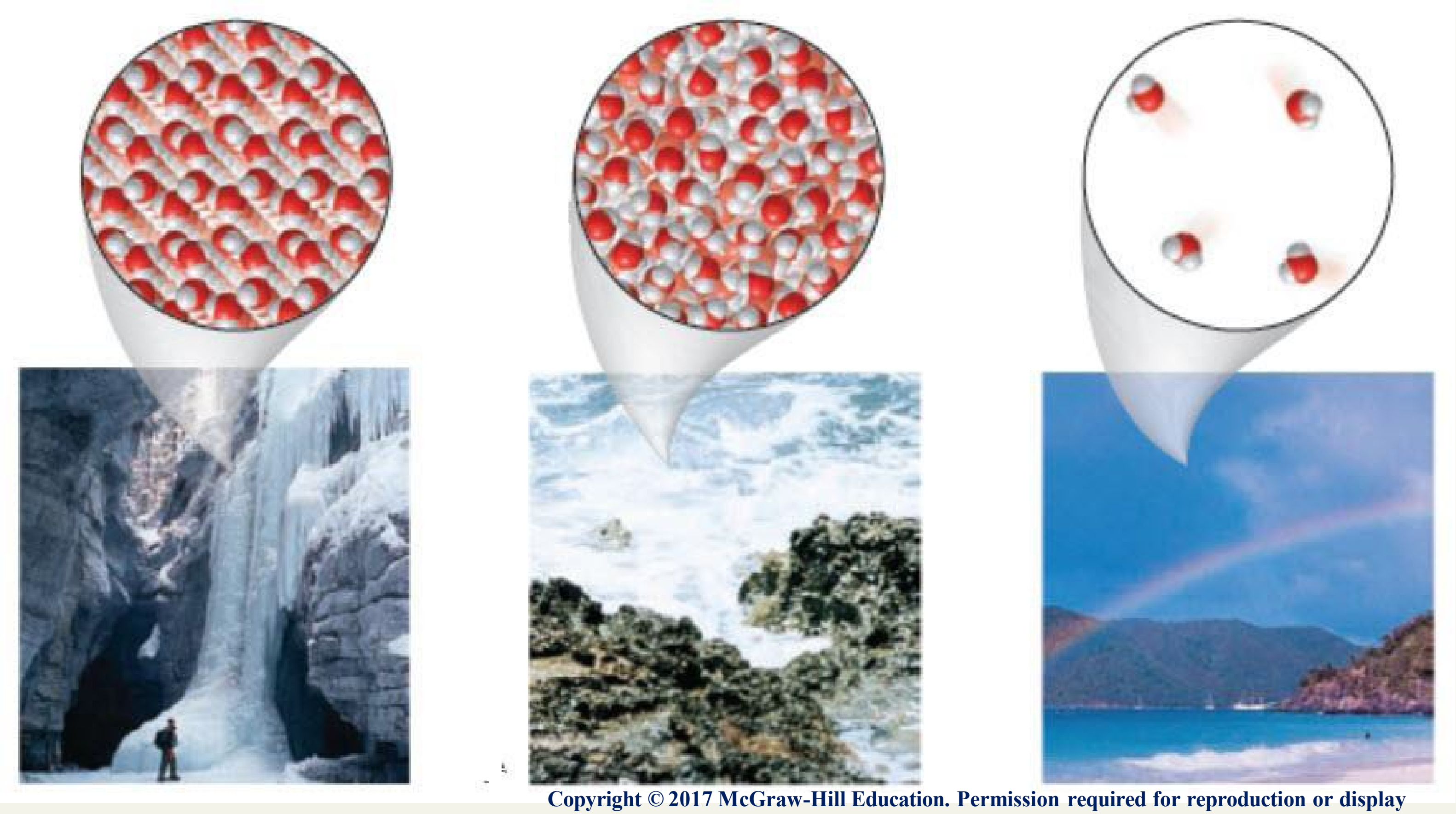
Summary: Classification of Matter



Physical Property vs. Physical Change

- **Physical property** - is observed without changing the composition or identity of a substance
- **Physical change** - produces a recognizable difference in the appearance of a substance without causing any change in its composition or identity
  - conversion from one physical state to another. Example: melting an ice cube

Physical Properties and Physical Change



(a) Solid (b) Liquid (c) Gas

Separation by Physical Properties

Magnetic iron is separated from other nonmagnetic substances, such as sand. This property is used as a large-scale process in the recycling industry.

Chemical Property and Chemical Reaction

- **Chemical property** - results in a change in composition and can be observed only through a chemical reaction
- **Chemical reaction (chemical change) -** a chemical substance is converted in to one or more different substances by rearranging, removing, replacing, or adding atoms

$\boxed{\text{hydrogen + oxygen ❼ water}}$

reactants products

Classification of Properties

Classify the following as either a chemical or physical property:

a. Color
b. Hardness
c. Flammability
d. Odor
e. Taste

Classification of Changes

Classify the following as either a chemical or physical change:

a. Boiling water becomes steam

b. Butter turns rancid

c. Burning of wood

d. Snow melting

e. Decay of leaves in winter

Composition of the Atom

- **Atom** - the basic structural unit of an element
  - The smallest unit of an element that retains the chemical properties of that element
- Atoms consist of three **primary** particles
- protons
- neutrons
- electrons

Electrons, Protons and Neutrons

- **Nucleus** - small, dense, positively charged region in the center of the atom containing:
  - protons - positively charged particles
  - neutrons - uncharged particles

Electrons: negatively charged particles located outside of the nucleus of an atom

- **Protons and electrons** have charges that are equal in magnitude but opposite in sign
  - **A *neutral atom*** (no electrical charge) has the *same number of protons and electrons*

Nomenclature of Chemical Compounds

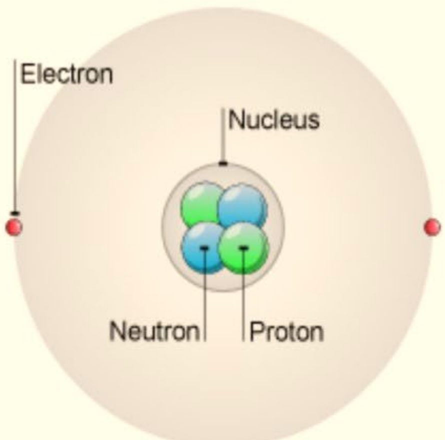
Examples of some anions and cations for the corresponding elements and symbols.

Element	Cation	Element	Anions
Li	Lithium Li^{1+} Group 1	C	**Carbide C^{4-} Group 4**
Na	Sodium Na^{+} Group 1	Si	**Silicide Si^{4-} Group 4**
K	Potasium K^{+} Group 1	N	**Nitride N^{3-} Group 5**
Ba	Barium Ba^{2+} Group 2	P	**Phosphide P^{3-} Group 5**
Mg	Magnesium Mg^{2+} Group 2	O	**Oxide O^{2-} Group 6**
Be	Beryllium Be^{2+} Group 2	S	**Sulfide S^{2-} Group 6**
Ca	Calcium Ca^{2+} Group 2	F	**Fluoride F^{-} Group 7**
Al	Aluminium Al^{3+}	Cl	**Chloride Cl^{-} Group 7**

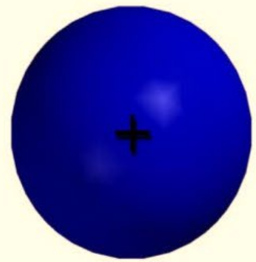
Names and Symbols of Most Polyatomic Compounds

Name	Formula	Name	Formula
Ammonium	NH_4^+	Hydroxide	**OH^-**
Chlorate	ClO_3^-	Chromate	**CrO_4^{2-}**
Nitrate	NO_3^-	Dichromate	**$Cr_2O_7^{2-}$**
Carbonate	CO_3^{2-}	Phosphite	**PO_3^{3-}**
Nitrite	NO_2^-	Phosphate	**PO_4^{3-}**
Hydrogen Carbonate	HCO_3^-	Thiocyanate	**SCN^-**
Sulfate	SO_4^{2-}	Cyanide	**CN^{1-}**

Composition of the Atom

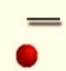


Atom



proton □ positive charge
mass = 1.673 x 10^{-27} kg

neutron □ no charge
mass = 1.675 x 10^{-27} kg



electron □ negative charge
mass = 9.109 x 10^{-31} kg

Selected Properties of the Three Basic Subatomic Particles

Mass(amu) **Charge** **Name**

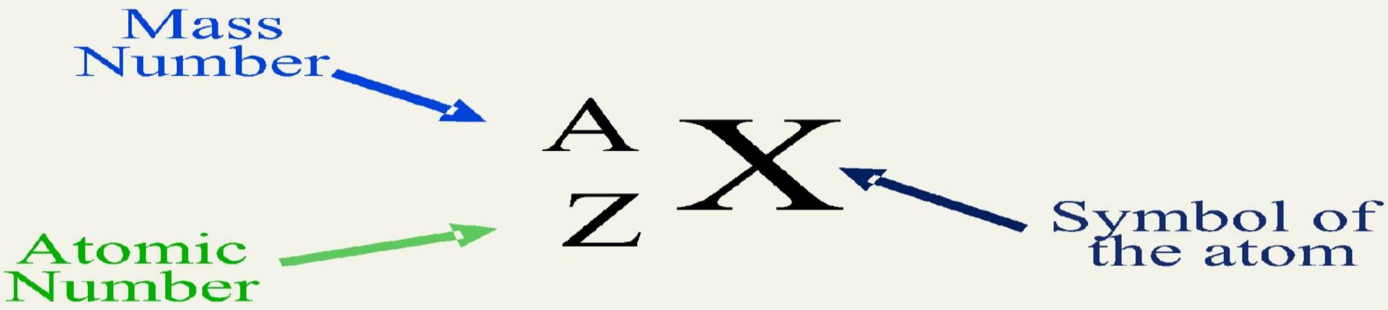
Mass (kg) **Electrons (e^-)** -1 5.486×10^{-4}

9.109×10^{-31} **Protons (p^+)** $+1$ 1.007

1.673×10^{-27} **Neutrons (n)** 0 1.009

1.675×10^{-27}

Symbolic Representation of an Element



- **Atomic number (Z)**

  - the number of protons in the atom

- **Mass number** (A)

  - sum of the number of protons and neutrons

Determining the Composition of an Atom

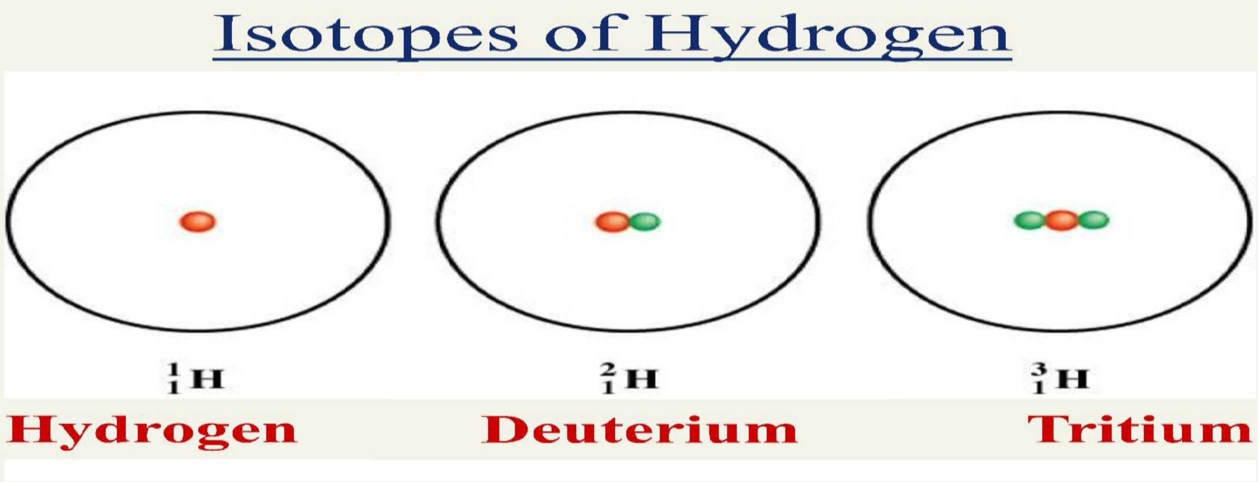
Calculate the number of protons, neutrons and electrons in each of the following:

$$^{11}_{5}B$$

$$^{55}_{26}Fe$$

Isotopes

- **Isotopes** - atoms of the same element having different masses
  – contain same number of protons
  – contain *different* numbers of *neutrons*



Isotopes

- **Isotopes of the same element have *identical chemical properties, however,* some isotopes have different nuclear reactivity such as radioactivity.**

In the periodic table, the atomic mass is the weighted average of the masses of all the isotopes that make up the element.

For Example:-

Chlorine consists of chlorine-35 and chlorine-37 in a 3:1 ratio

What is the mass given? **35.5?!**

$$\mathbf{(35x3/4 + 37\ x1/4) = 35.5}$$

Atomic Mass Determination

- Neon has three naturally occurring isotopes

	Mass	Abundance
^{20}Ne	20.0 amu	90.48%
^{21}Ne	21.0 amu	0.27%
^{22}Ne	22.0 amu	9.25%

- What is the atomic mass of Neon?

$$= 20.0 \times \frac{90.48}{100} + 21.0 \times \frac{0.27}{100} + 22.0 \times \frac{9.25}{100} = 20.2 \text{ amu}$$

Units of Measurement

A measurement is useless without its units

SI Units (International system of units).

– Fundamental units (basic units)

- There are 7 fundamental units in the **SI** system

Quantity	Unit	Symbol
✦ Length	Meter	m
✦ Mass	kilogram	kg
✦ Time	second	s
✦ Temperature	Kelvin	K
✦ Amt. substance	mole	mol
✦ Current	ampere	A
✦ Luminous intensity	candela	cd

-Derived units : All other units are derived from the above basic units

Metric System Prefixes

- Basic units are the units of a quantity without any metric prefix (Memorize)

Prefix	Abbreviation	Meaning	Decimal Equivalent	Equality with major metric units (g, m, or L are represented by x in each)
mega	M	10^{6}	1,000,000.	$1\ \mathrm{M}x = 10^{6}x$
kilo	k	10^{3}	1000.	$1\ \mathrm{k}x = 10^{3}x$
deka	da	10^{1}	10.	$1\ \mathrm{da}x = 10^{1}x$
deci	d	10^{-1}	0.1	$1\ \mathrm{d}x = 10^{-1}x$
centi	c	10^{-2}	0.01	$1\ \mathrm{c}x = 10^{-2}x$
milli	m	10^{-3}	0.001	$1\ \mathrm{m}x = 10^{-3}x$
micro	μ	10^{-6}	0.000001	$1\ \mu x = 10^{-6}x$
nano	n	10^{-9}	0.000000001	$1\ \mathrm{n}x = 10^{-9}x$

Units of Measurement expressed in Scientific Notation

- Used to express very large or very small numbers easily and with the correct number of value and accuracy.
- Represents a number as a power of ten
- Example:

$$4{,}300 = 4.3 \times 1{,}000 = 4.3 \times 10^3$$

Scientific Notation Rules

- **To convert a number <u>greater</u> than 1 to scientific notation, the original decimal point is moved *x* places to the <u>left</u>, and the resulting number is multiplied by 10^x**
- The exponent ***x*** is a *positive* number equal to the number of places the decimal point moved

$$6200 = 6.20 \times 10^3$$

Scientific Notation Rules

- **To convert a number <u>less</u> than 1 to scientific notation, the original decimal point is moved *x* places to the <u>right</u>, and the resulting number is multiplied by 10^{-x}**

- The exponent x is a *negative* number equal to the number of places the decimal point moved

$$0.0062 = 6.2 \times 10^{-3}$$

Scientific Notation Example

- When a number is exceedingly large or small, scientific notation *must be used* to input the number into a calculator:

0.000000000000000000000006692 g

must be entered into calculator as:

6.692×10^{-24}

Represent the following numbers in scientific notation:

1. 0.00018
2. 3004
3. 305
4. 0.00304

Unit Conversions

- **Factor-Label Method** (Dimensional Analysis)
  - Uses **Conversion Factors** to**:**
    - Convert from one unit to another within the same system (Examples: mg in g, μm in km,….etc)
    - Convert units from one system to another (Examples: ft in m, m in mile, gal in L,….etc)

Unit Conversion - Example

- To convert from one unit to another you must know the conversion factor, which is the relationship between the two units
  - **The Relationship:**

$$1\ \text{m} = 1000\ \text{mm}$$

  - **The Conversion Factor:**

$$\frac{1\ \text{m}}{1000\ \text{mm}} \quad \text{or} \quad \frac{1000\ \text{mm}}{1\ \text{m}}$$

Using Conversion Factors

Convert 12 inch to cm

- **The Relationship (English system):**

  1 inch = 2.54 cm (GIVEN)

- **The Conversion Factor:**

  $$\frac{1 \text{ inch}}{2.54 \text{ cm}} \quad \text{or} \quad \frac{2.54 \text{ cm}}{1 \text{ inch}}$$

Data Given: 12 inch

Use Conversion Factor with inch in denominator

12 inch x 2.54 cm/ 1 inch = 30.5 cm

Multistep Conversion - Example

Convert 0.0047 kilograms to milligrams

- **The Relationships (metric system):**

$$1\ \text{kg} = 10^3\ \text{g} \quad \text{and} \quad 10^3\ \text{mg} = 1\ \text{g}$$

- **The Conversion Factors:**

$$\frac{1\ \text{kg}}{10^3\ \text{g}} \text{ or } \frac{10^3\ \text{g}}{1\ \text{kg}} \quad \text{and} \quad \frac{1\ \text{g}}{10^3\ \text{mg}} \text{ or } \frac{10^3\ \text{mg}}{1\ \text{g}}$$

Data Given: 0.0047 kg

$$0.0047\ \text{kg} \times \frac{10^3\ \text{g}}{1\ \text{kg}} \times \frac{1\ \text{mg}}{10^{-3}\ \text{g}} = 4.7 \times 10^3\ \text{mg}$$

Practice Unit Conversions

1. Convert 5.5 inches to millimeters

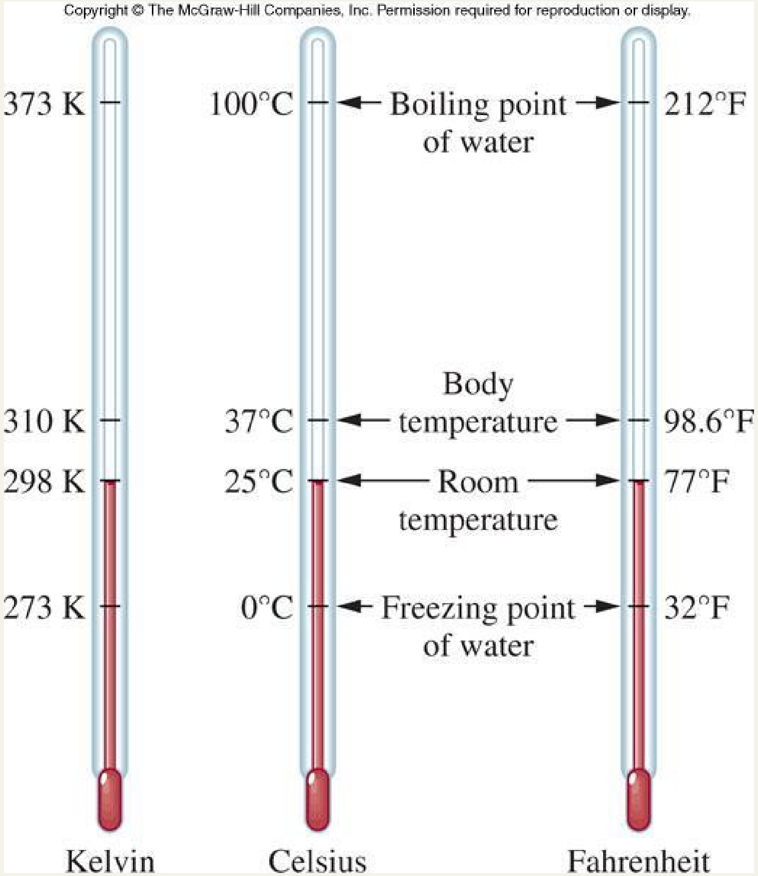
Answer 1.4×10^{2} mm

2. Convert 50.0 mg to kg

Answer 5.00×10^{-5} kg

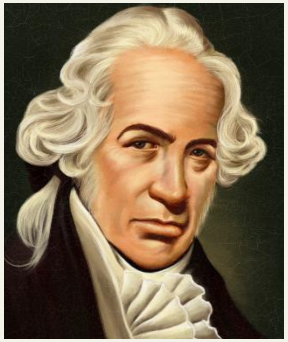
Additional Experimental Quantities

- **Temperature** - the degree of “hotness” of an object



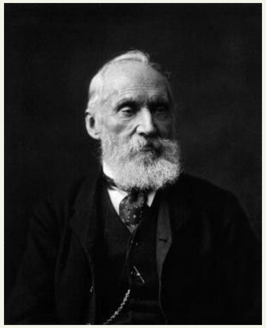
Temperature

	MP water	BP water
• Fahrenheit	32 ^{o}F	212 ^{o}F
• Celsius	0.0 ^{o}C	100 ^{o}C
• Kelvin	273 K	373 K



Daniel Gabriel **Fahrenheit**
1686-1736
Germany

Anders **Celsius**
1701-1744
Sweden



Lord **Kelvin**
1824-1907
England

Kelvin Temperature Scale

- The **Kelvin (K)** scale is another temperature scale
- It is of particular importance because it is directly related to molecular motion
- As molecular speed increases, the Kelvin temperature proportionately increases

$$T_K = T_{^oC} + 273$$

Conversions Between Fahrenheit and Celsius

$$T_{^oF} = 1.8 \times T_{^oC} + 32$$

$$T_{^oC} = \frac{T_{^oF} - 32}{1.8}$$

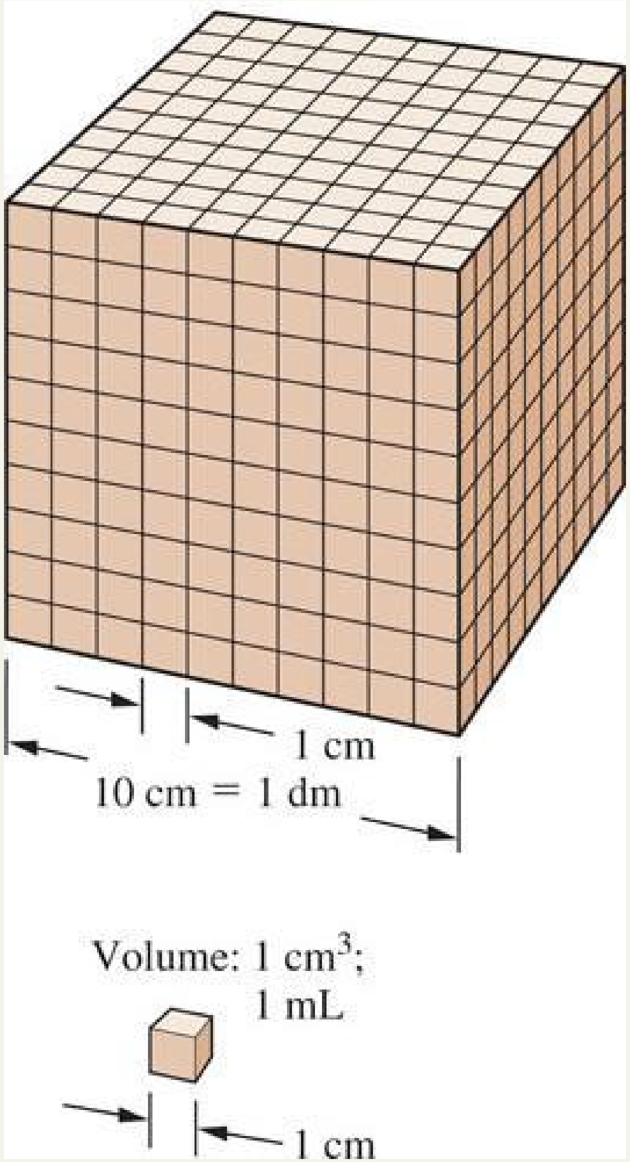
1. Convert 75^{o}C to ^{o}F
2. Convert -10^{o}F to ^{o}C

1. Ans. 167 ^{o}F 2. Ans. -23^{o}C

Density

- **Density**
  - the ratio of mass to volume
  - use to characterize a substance as each substance has a unique density
  - Units for density include:
    - g/mL
    - g/cm^3
    - g/cc

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



Volume =

Length x width x height

Volume =

1 m x 1 m x 1 m =

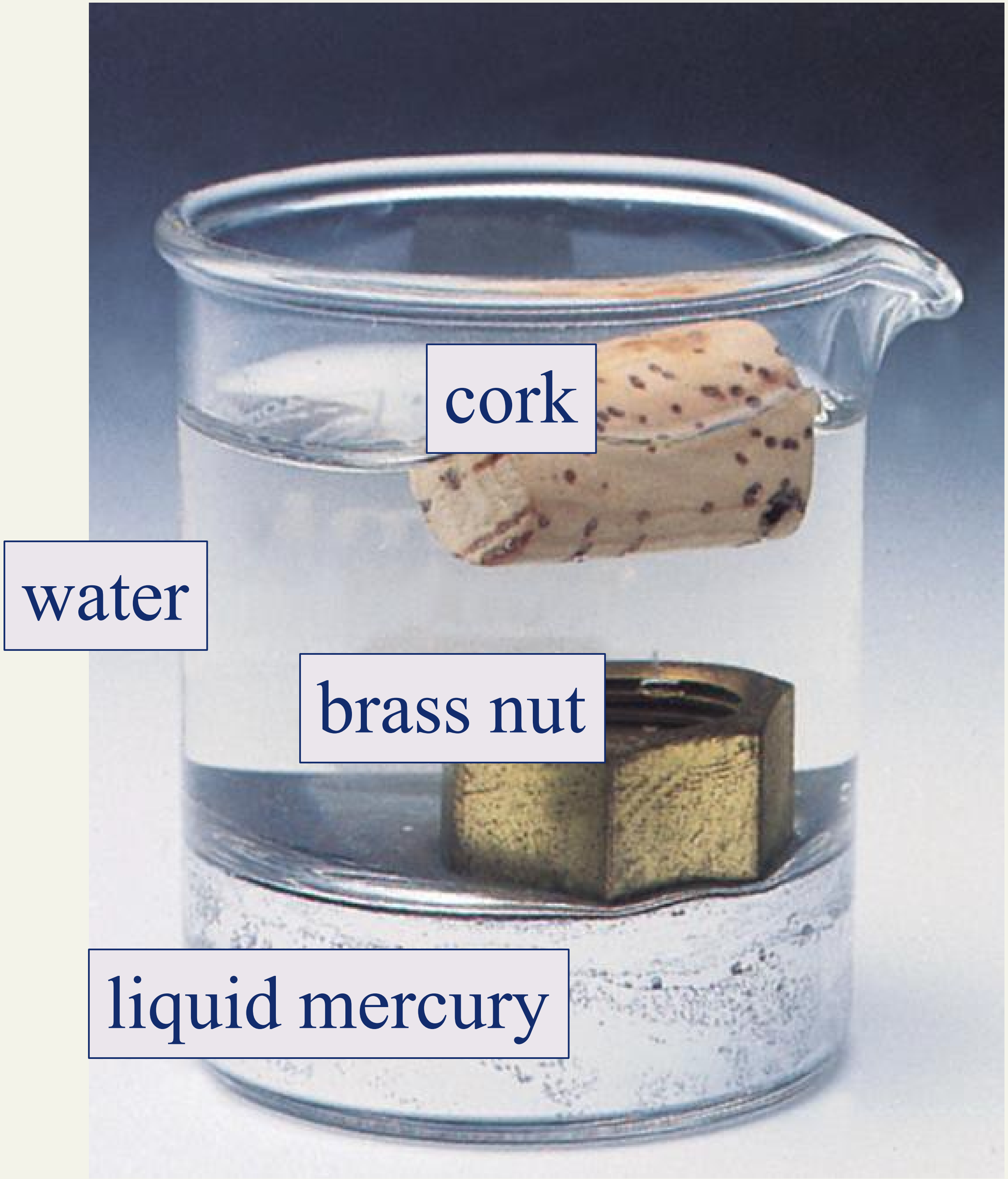
$1\ m^3$

$1\ m^3 = 1000\ L$

$1\ dm^3 = 1\ L$

$1\ cm^3 = 1\ mL$

Density Examples



Calculating Density

A 2.00 cm^3 sample of aluminum is found to weigh 5.40 g. Calculate the density in g/cm^3 and g/mL.

Use the expression:

- Density (d) = m/V

Substitute information given into the expression

$$d = \frac{5.40\ g}{2.00\ cm^3} = 2.70\ g/cm^3$$

- Since 1 cm^3 = 1 mL,

$$= 2.70\ g/mL$$

Use Density in Calculation

Calculate the volume, in mL, of a liquid that has a density of 1.20 g/mL and a mass of 5.00 g.

- Density can be written as a Conversion Factor

$$\frac{1.20\ \text{g}}{1\ \text{mL}} \quad \text{or} \quad \frac{1\ \text{mL}}{1.20\ \text{g}}$$

- Multiply the Data Given (g) by the Conversion Factor with the unit g in the denominator

$$5.00\ \cancel{\text{g}} \times \frac{1\ \text{mL}}{1.20\ \cancel{\text{g}}} = 4.17\ \text{mL}$$

A particular solution of salt water contains 20 grams of salt and 200 grams of water. Find out the density of the salt water? (note = density of water is 1 g/ml)

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

$$d = \frac{m_{salt} + m_{water}}{V_{solution}}$$

Since the density of water is 1 g/ml, then the volume of water and hence the volume of the solution is 200 ml

$$d = \frac{20 + 200}{200}$$

$$d = \frac{220\ g}{200\ ml} = 1.1.\ g/ml$$

Density Calculations

- Air has a density of 0.0013 g/mL. What is the mass of 6.0-L sample of air?

(Ans. 7.8 g)

- Calculate the mass in grams of 10.0 mL if mercury (Hg) if the density of Hg is 13.6 g/mL.

(Ans. 136 g)