General Chemistry

CHEM 101
(3+1+0)

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Introduction

What is Chemistry?

- It's easy to say ***chemistry*** is important because everything is made from chemicals.
  - ***Chemistry*** is the study of matter and the changes it undergoes.
  - ***Chemistry*** sometimes is called the "***central science***" because it connects other sciences to each others, such as biology, physics, geology and environmental science.

Chemistry is the science that deals with the materials of the universe and the changes these materials undergo

Why is Chemistry Important?

1) Chemistry helps you to understand the world around you.

Why does leaves change color in the fall? Why are plants green? What is in soap and how does it clean?

2) A command of chemistry can help keep you safe!

You'll know which household chemicals are dangerous to keep together or mix and which can be used safely.

3) Chemistry teaches useful skills.

Because it is a science, learning chemistry means learning how to be objective and how to reason and solve problems.

4) Chemistry opens up career options.

There are many careers in chemistry, but even if you're looking for a job in another field, the analytical skills you gained in chemistry are helpful. Chemistry applies to the food industry, retail sales, transportation, art, homemaking... really any type of work you can name.

5) Chemistry is fun!

They can glow in the dark, change colors, produces bubbles and change states.

Chapter 1

Chemistry: The Study of Change

Matter & Measurements

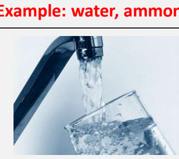
Matter

- ***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.

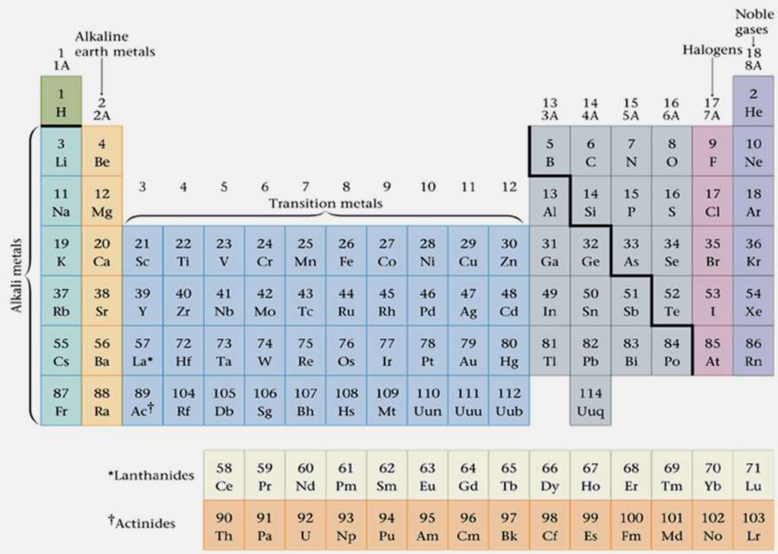
Classifications of Matter

Matter

- **Mixtures**: is a combination of two or more substances in which the substances retain their distinct identities. Example: Air, soft drink, and pizza
  - **Homogeneous mixtures**: The composition is the same throughout. Example: sugar in water, soft drink, milk
  - **Heterogeneous mixtures**: The composition is not uniform throughout. Example: fruit salad, iron filings in sand
- Mixtures → Separation by physical methods → Pure substances
- **Pure substances**: is a form of matter that has a definite composition and distinct properties. Example: water , sugar and gold
  - **Compounds**: is a substance composed of atoms of two or more elements chemically united in fixed proportions. Example: water, ammonia
  - **Elements**: is a substance that cannot be separated into simpler substances by chemical means. Example: aluminum
- Compounds → Separation by chemical methods → Elements

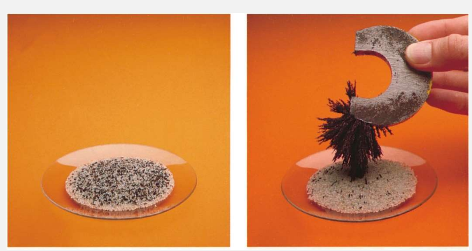


Classifications of Matter



Classifications of Matter

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



magnet

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

The Three States of Matter

States of matter

Gas

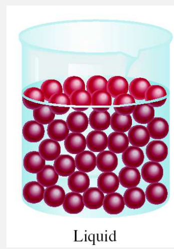
- Molecules are separated by distance.
- Have no definite shape.
- Have no definite volume

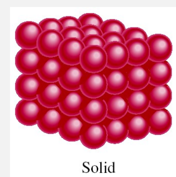
Liquid

- Molecules are close together but not held so rigidly in position and can move past on another.
- Have no definite shape.
- Have a definite volume

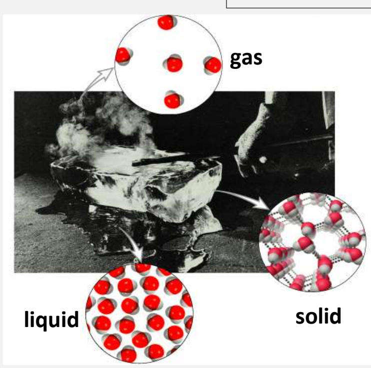
Solid

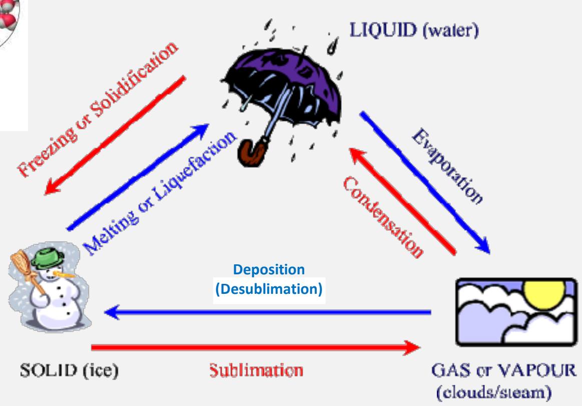
- Molecules are held close together in orderly fashion with little freedom of motion.
- Have a definite shape.
- Have a definite volume





The Three States of Matter





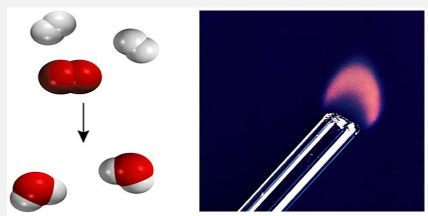
Changes of Matter: Physical and Chemical Changes

- Color, melting point, and boiling point are **physical properties**.
- 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.

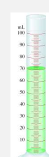
hydrogen burns in air to form water



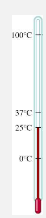
Properties of Matter

Extensive and Intensive Properties

- An ***extensive property*** of a material depends upon how much matter is being considered.
  - mass
  - length
  - volume



- An ***intensive property*** of a material does not depend upon how much matter is being considered.
  - density
  - temperature
  - color



Measurements

International System of Units (*SI*)

- A measured quantity is usually written as a number with an appropriate unit.
- The same is true in chemistry; units are essential to stating measurements correctly.
- For many years, scientists recorded measurements in ***metric units***, which are related decimally, that is, by powers of 10.
- **Metric system** called the ***International System of Units*** (abbreviated *SI*).
- Measurements that we will utilize frequently in our study of chemistry include time, mass, volume, density, and temperature.

International System of Units (*SI*)

TABLE 1.2 SI Base 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

International System of Units (*SI*)

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) $= 1 \times 10^{12}$ m
giga-	G	1,000,000,000, or 10^{9}	1 gigameter (Gm) $= 1 \times 10^{9}$ m
mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) $= 1 \times 10^{6}$ m
kilo-	k	1,000, or 10^{3}	1 kilometer (km) $= 1 \times 10^{3}$ 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) $= 1 \times 10^{-6}$ m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) $= 1 \times 10^{-9}$ m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) $= 1 \times 10^{-12}$ m

International System of Units (*SI*)

Mass and Weight

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

mass – measure of the quantity of matter

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

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

weight – 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

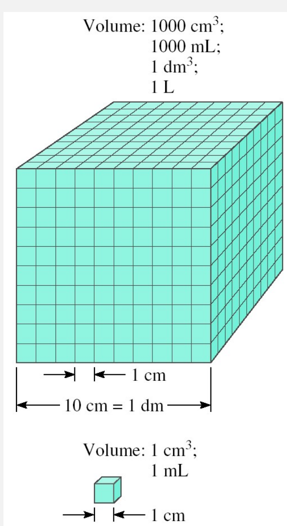
1 kg on earth

0.1 kg on moon

International System of Units (*SI*)

Volume

Volume – SI derived unit for volume is cubic meter (m^3)



1 cm^3 = (1 x 10^{-2} m)^3 = 1 x 10^{-6} m^3

1 dm^3 = (1 x 10^{-1} m)^3 = 1 x 10^{-3} m^3

1 L = 1000 mL = 1000 cm^3 = 1 dm^3

$\boxed{1\ mL = 1\ cm^3}$

International System of Units (*SI*)

Density

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

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

$$1\ g/L = 0.001\ g/mL$$

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

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?

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

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

International System of Units (*SI*)

Density

Density usually decreases with temperature.

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.

International System of Units (*SI*)

EXAMPLE 1.1

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.

Solution We are given the mass and volume and asked to calculate the density. Therefore, from Equation (1.1), we write

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

International System of Units (*SI*)

EXAMPLE 1.2

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.

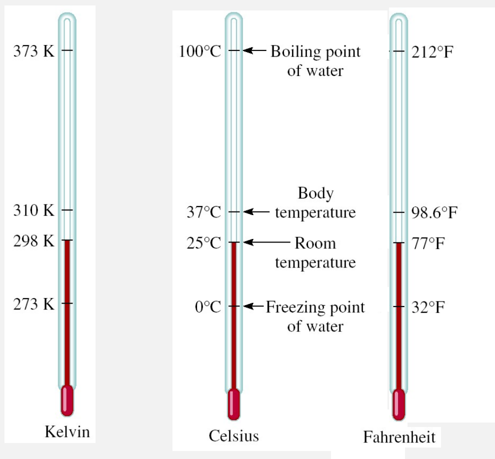
Solution We are given the density and volume of a liquid and asked to calculate the mass of the liquid. We rearrange Equation (1.1) to give

$$m = d \times V$$
$$= 13.6\frac{\text{g}}{\cancel{\text{mL}}} \times 5.50\ \cancel{\text{mL}}$$
$$= \boxed{74.8\ \text{g}}$$

Practice Exercise The density of sulfuric acid in a certain car battery is 1.41 g/mL. Calculate the mass of 242 mL of the liquid.

International System of Units (*SI*)

Temperature Scales



K = ^{0}C + 273.15

273 K = 0 ^{0}C
373 K = 100 ^{0}C

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

32 ^{0}F = 0 ^{0}C
212 ^{0}F = 100 ^{0}C

The ***kelvin (K) is the SI base unit of temperature*****: it is the** ***absolute*** **temperature scale.**

International System of Units (SI)

Temperature Scales

Convert 172.9 ^{0}F to degrees Celsius.

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

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

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

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

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

Handling Numbers

Scientific Notation

The number of atoms in 12 g of carbon:
602,200,000,000,000,000,000,000

6.022×10^{23}

The mass of a single carbon atom in grams:
0.0000000000000000000000199

1.99×10^{-23}

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

N is a number between 1 and 10

n is a positive or negative integer

Any number expressed in this way is said to be written in scientific notation.

Handling Numbers

Scientific Notation

568.762

← move decimal ***left***

n > 0

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

0.00000772

→ move decimal ***right***

n < 0

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

Addition or Subtraction

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

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

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

4.70×10^{4}

Handling Numbers

Scientific Notation

Multiplication

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

$$(4.0 \times 10^{-5}) \times (7.0 \times 10^{3}) =$$
$$(4.0 \times 7.0) \times (10^{-5+3}) =$$
$$28 \times 10^{-2} =$$
$$2.8 \times 10^{-1}$$

Division

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

$$8.5 \times 10^{4} \div 5.0 \times 10^{9} =$$
$$(8.5 \div 5.0) \times 10^{4-9} =$$
$$1.7 \times 10^{-5}$$

Handling Numbers

Significant Figures

Significant figures - are *the meaningful digits in a measured or calculated quantity*.

Guidelines for Using Significant Figures

Handling Numbers

Significant Figures

- 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

Handling Numbers

Significant Figures

For numbers that do not contain decimal points, the trailing zeros (that is, zeros after the last nonzero digit) may or may not be significant.

Thus, 400 cm may have

one significant figure (the digit 4),

two significant figures (40),

or three significant figures (400).

Which is correct? …. we can express the number 400 as

4×10^2 for one significant figure,

4.0×10^2 for two significant figures,

or 4.00×10^2 for three significant figures.

Handling Numbers

Significant Figures

How many significant figures are in each of the following measurements?

24 mL	2 significant figures
3001 g	4 significant figures
0.0320 m^3	3 significant figures
6.4 x 10^4 molecules	2 significant figures
560 kg	2 significant figures

Handling Numbers

Significant Figures

Addition or Subtraction

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

$$\begin{array}{r} 89.332 \\ +1.1 \\ \hline 90.432 \end{array}$$

+1.1 ← one significant figure after decimal point

90.432 ← round off to 90.4

$$\begin{array}{r} 3.70 \\ -2.9133 \\ \hline 0.7867 \end{array}$$

3.70 ← two significant figures after decimal point

0.7867 ← round off to 0.79

Handling Numbers

Significant Figures

Multiplication or Division

The number of significant figures in the result is set by the original number that has the ***smallest*** number of significant figures

4.51 x 3.6666 = 16.536366 = 16.5

3 sig figs — round to 3 sig figs

6.8 ÷ 112.04 = 0.0606926 = 0.061

2 sig figs — round to 2 sig figs

Handling Numbers

Significant Figures

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?

$$\frac{6.64 + 6.68 + 6.70}{3} = 6.67333 = 6.67 = \cancel{7}$$

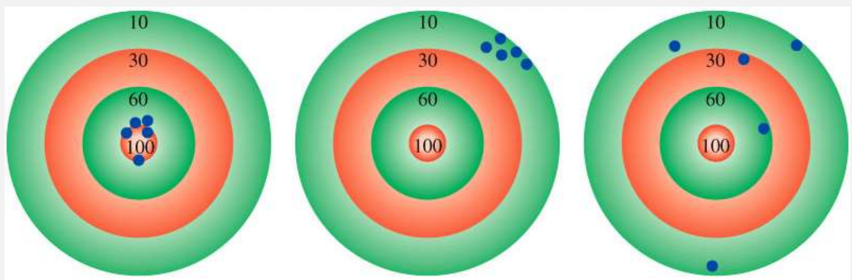
Because 3 is an ***exact number***

Handling Numbers

Significant Figures

Accuracy and Precision

Accuracy – how close a measurement is to the *true* value
Precision – how close a set of measurements are to each other



accurate & precise	precise but not accurate	not accurate & not precise

Handling Numbers

Significant Figures

Accuracy and Precision

Three students are asked to determine the mass of a piece of copper wire. (*The true mass of the wire is 2.000 g*)
The results of two successive weighing's by each student are

	Student A	Student B	Student C
	1.964 g	1.972 g	2.000 g
	1.978 g	1.968 g	2.002 g
Average value	1.971 g	1.970 g	2.001 g
	not accurate & not precise	precise but not accurate	accurate & precise

Dimensional Analysis in Solving Problems

- ***Dimensional analysis*** **-** The procedure we use to convert between units in solving chemistry problems.
- ***Dimensional analysis*** is based on the relationship between different units that express the same physical quantity.

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.

given quantity x conversion factor = desired quantity

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

Dimensional Analysis in Solving Problems

How many mL are in 1.63 L?

Conversion Unit 1 L = 1000 mL

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

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

Dimensional Analysis in Solving Problems

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

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

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 mg = 1×10^{-3} 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:

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

Check As an estimate, we note that 1 lb is roughly 500 g and that 1 g = 1000 mg. Therefore, 1 lb is roughly 5×10^5 mg. Rounding off 0.0833 lb to 0.1 lb, we get 5×10^4 mg, which is close to the preceding quantity.

Practice Exercise A roll of aluminum foil has a mass of 1.07 kg. What is its mass in pounds?

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 L = 1000 cm^3 and 1 cm = 1×10^{-2} 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 cm^3 = 1×10^{-6} m^3. Now we can write

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

Check From the preceding conversion factors you can show that 1 L = 1×10^{-3} m^3. Therefore, 5 L of blood would be equal to 5×10^{-3} m^3, which is close to the answer.

Practice Exercise The volume of a room is 1.08×10^8 dm^3. What is the volume in 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 77 K) is 0.808 g/cm^3. Convert the density to units of kg/m^3.

Strategy The problem can be stated as

$$? \text{ kg/m}^3 = 0.808 \text{ g/cm}^3$$

Two separate conversions are required for this problem: g $\longrightarrow$ kg and $cm^3 \longrightarrow m^3$. Recall that 1 kg = 1000 g and 1 cm = 1×10^{-2} m.

Solution In Example 1.7 we saw that 1 cm^3 = 1×10^{-6} m^3. The conversion factors are

$$\frac{1 \text{ kg}}{1000 \text{ g}} \quad \text{and} \quad \frac{1 \text{ cm}^3}{1 \times 10^{-6} \text{ m}^3}$$

Finally,

$$? \text{ kg/m}^3 = \frac{0.808 \cancel{\text{g}}}{1 \cancel{\text{cm}^3}} \times \frac{1 \text{ kg}}{1000 \cancel{\text{g}}} \times \frac{1 \cancel{\text{cm}^3}}{1 \times 10^{-6} \text{ m}^3} = \boxed{808 \text{ kg/m}^3}$$

Check Because 1 m^3 = 1×10^6 cm^3, we would expect much more mass in 1 m^3 than in 1 cm^3. Therefore, the answer is reasonable.

Practice Exercise The density of the lightest metal, lithium (Li), is 5.34×10^2 kg/m^3. Convert the density to g/cm^3.