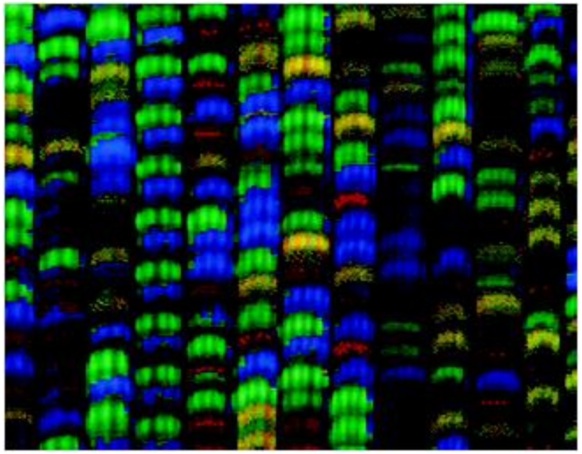
Chapter 1

Chemistry and measurements

1.1 Modern Chemistry: A Brief Glimpse

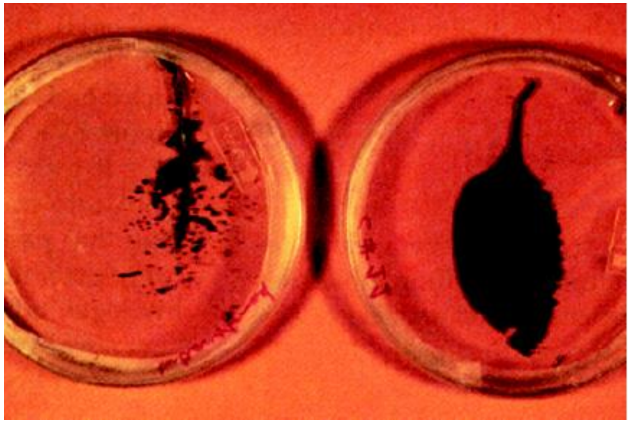
- Health and Medicine
  - Sanitation systems
  - Surgery with anesthesia
  - Vaccines and antibiotics
  - Gene therapy



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

Modern Chemistry: A Brief Glimpse

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



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

1.2 Experiment and Explanation

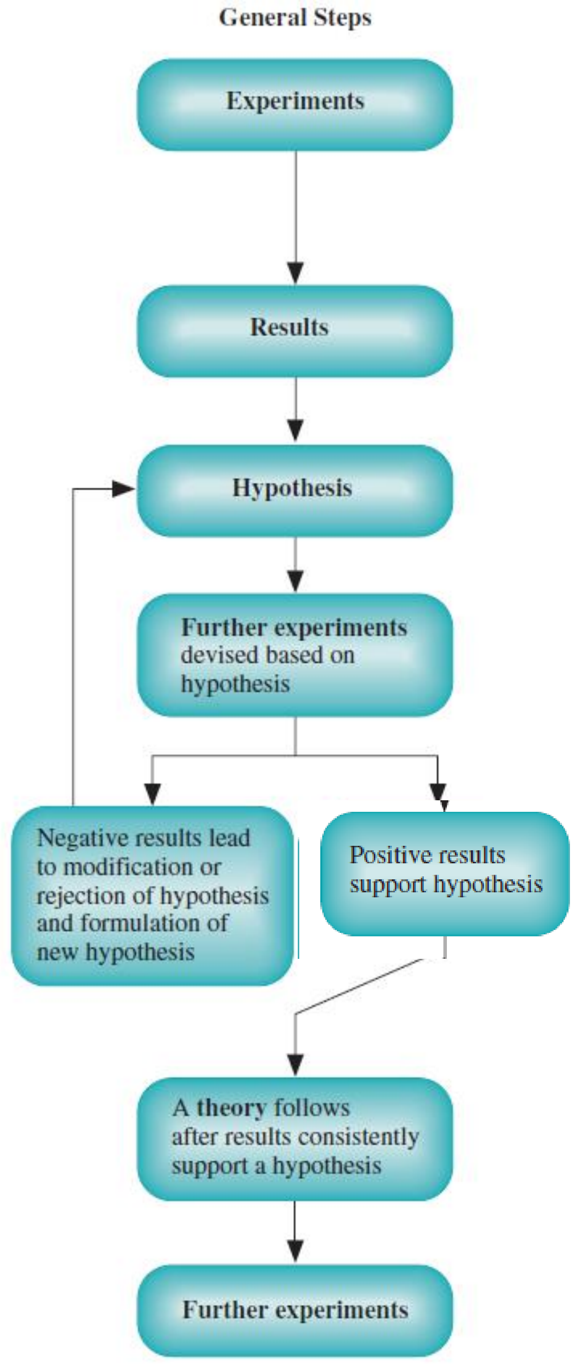
- An **experiment** is *an observation of natural phenomena carried out in a controlled manner so that the results can be duplicated and rational conclusions obtained*.

- A **law** is *a concise statement or mathematical equation about a fundamental relationship or regularity of nature.*

- A **hypothesis** is *a tentative explanation of some regularity of nature.*

- A **theory** is *a tested explanation of basic natural phenomena.*
  Example: molecular theory of gases.

Note: We cannot prove a theory absolutely.
It is always possible that further experiments will show the theory to be limited or that someone will develop a better theory


General Steps
Experiments
Results
Hypothesis
Further experiments devised based on hypothesis
Negative results lead to modification or rejection of hypothesis and formulation of new hypothesis
Positive results support hypothesis
A theory follows after results consistently support a hypothesis
Further experiments

1.3 Law of Conservation of Mass

Antoine Lavoisier (1743–1794), a French chemist, was one of the first to insist on the use of the balance in chemical research. By weighing substances before and after chemical change, he demonstrated the **law of conservation of mass,** which states that "***the total mass remains constant during a chemical change (chemical reaction***)."

Example 1.1 Using the Law of Conservation of Mass

You heat 2.53 grams of metallic mercury in air, which produces 2.73 grams of a red-orange residue. Assume that the chemical change is the reaction of the metal with oxygen in air.

$$\text{Mercury} + \text{oxygen} \longrightarrow \text{red-orange residue}$$

What is the mass of oxygen that reacts? When you strongly heat the red-orange residue, it decomposes to give back the mercury and release the oxygen, which you collect. What is the mass of oxygen you collect?

Solution From the law of conservation of mass,

$$\text{Mass of mercury} + \text{mass of oxygen} = \text{mass of red-orange residue}$$

Substituting, you obtain

$$2.53 \text{ grams} + \text{mass of oxygen} = 2.73 \text{ grams}$$

or

$$\text{Mass of oxygen} = (2.73 - 2.53) \text{ grams} = \mathbf{0.20\ grams}$$

The mass of oxygen collected when the red-orange residue decomposes equals the mass of oxygen that originally reacted **(0.20 grams).**

mass – measure of the quantity of matter

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

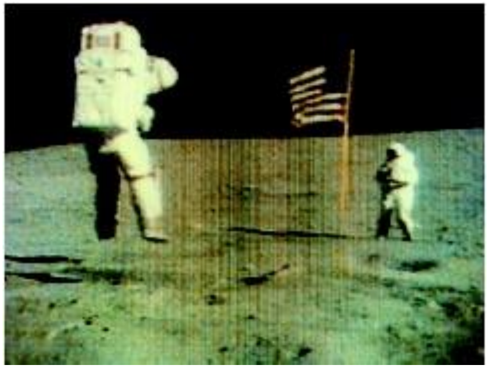
1 kg = 1000 g = 1×10^3 g

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

1.4 Matter: Physical State and Chemical Constitution

There are two principal ways of classifying matter:

(1) by its physical state as a solid, liquid, or gas

(2) by its chemical constitution as an element, compound, or mixture.

(1) Solids, Liquids, and Gases:

- **solid** *the form of matter characterized by **rigidity**;* a solid is **relatively incompressible** and has fixed shape and volume.

- **liquid** *the form of matter that is a **relatively incompressible** fluid;* a liquid has a fixed volume but no fixed shape.

- **gas** *the form of matter that is an **easily compressible** fluid;* a given quantity of gas will fit into a container of almost any size and shape.

States of Matter

Solids:

- Fixed shape and volume
- Particles are close together
- Have restricted motion

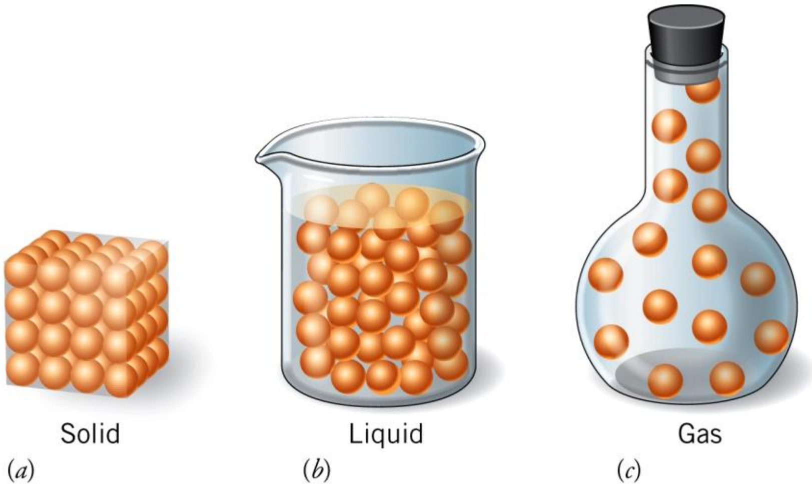
Liquids:

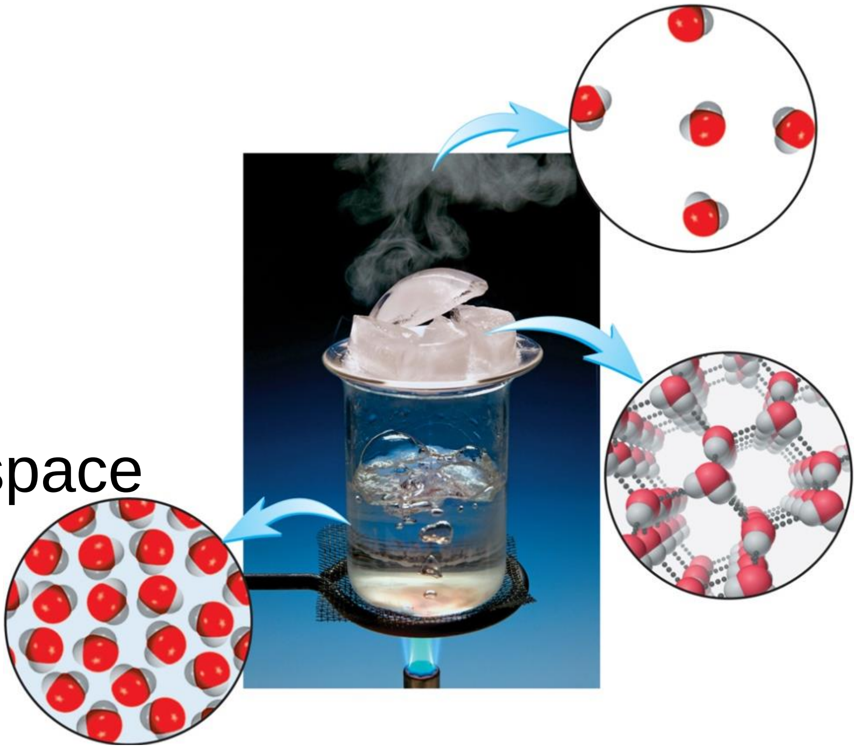
- Fixed volume, but take container shape
- Particles are close together
- Are able to flow

Gases:

- Expand to fill entire container
- Particles separated by lots of space

e.g., Ice, water, steam





(2) Elements, Compounds, and Mixtures

- A **physical change** is *a change in the form of matter but not in its chemical identity.*
  Examples: Ice melting, salt or sugar dissolved in water.
  (Physical property: Melting, boiling, electrical conductivity)
- A **chemical change,** or **chemical reaction,** is *a change in which one or more kinds of matter are transformed into a new kind of matter or several new kinds of matter.*
  Examples: rust formation, burning butane gas in oxygen
  (Chemical property: Describes how a substance undergoes a chemical reaction)

	Chemical	Physical
Magnesium burns when heated		
Magnesium metal tarnishes in air		
Magnesium metal melts at 922 K		
Orange juice lightens when water is added		

- A **substance:** is *a kind of matter that cannot be separated into other kinds of matter by any physical process.*

- **A mixture:** is *a kind of matter that can be separated by physical means into two or more substances*.

  -heterogeneous mixture: *a mixture that consists of physically distinct parts, each with different properties*

  Example: Sand and iron filings

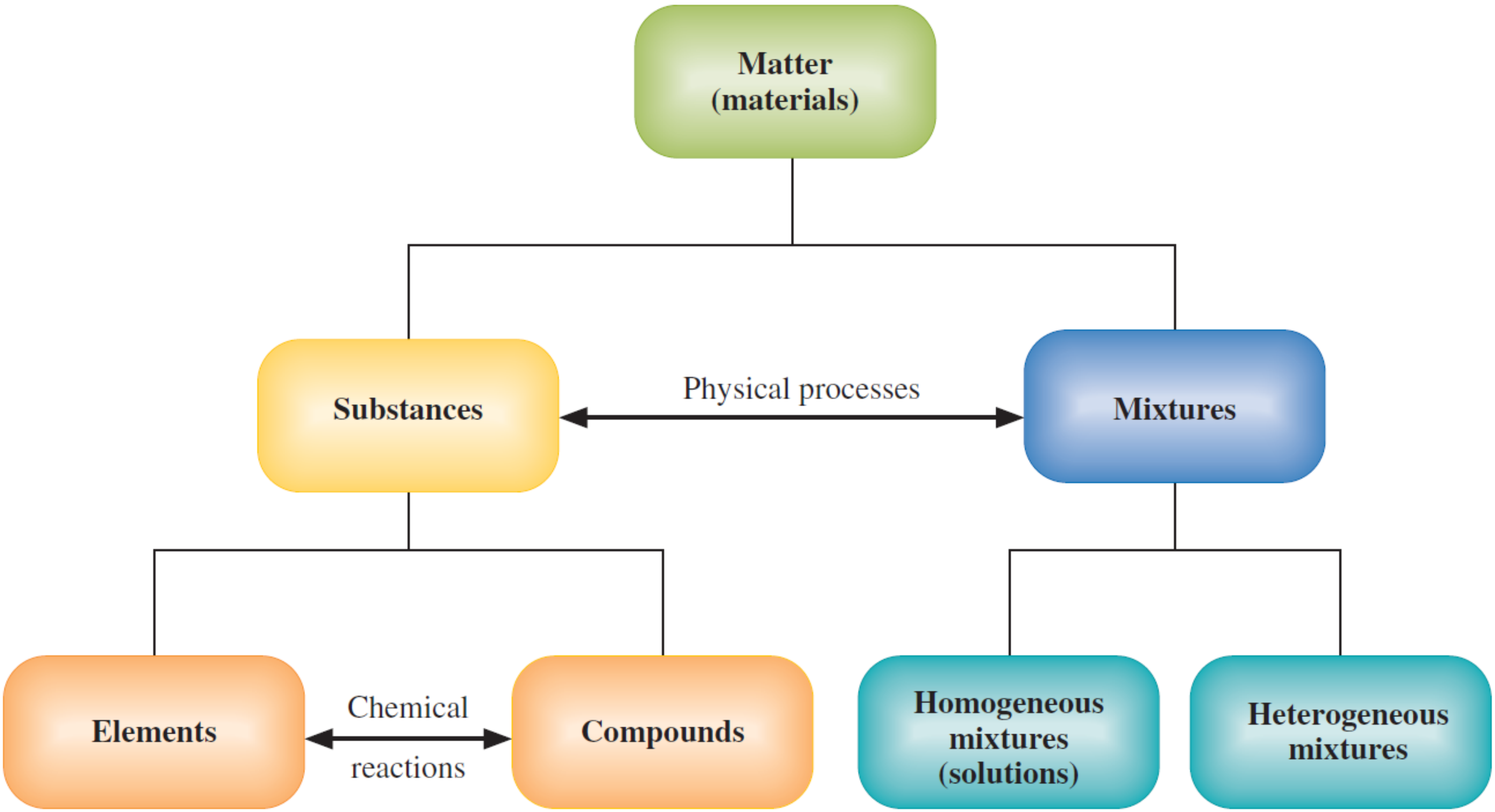
  -**homogeneous mixture** (also known as a **solution**): is *a mixture that is uniform in its properties throughout given samples.* Examples: NaCl solution, Soft drink, Air, Solder

- **an Element:** *A substance that cannot be decomposed by any chemical reaction into simpler substances*

  Fe, Au,Na etc…

- **A Compound:** is *a substance composed of two or more elements chemically combined.*

  H_2O, NaCl, CO_2


Matter
(materials)
Substances
Physical processes
Mixtures
Elements
Chemical
reactions
Compounds
Homogeneous
mixtures
(solutions)
Heterogeneous
mixtures

	Chicken Noodle Soup	Ice (H_2O)	Liquid Dish Soap	Table Salt (NaCl)
Pure substance				
Element				
Compound				
Heterogeneous Mixture				
Homogeneous Mixture				

- A **phase** is *one of several different homogeneous materials present in the portion of matter under study.*

Examples:

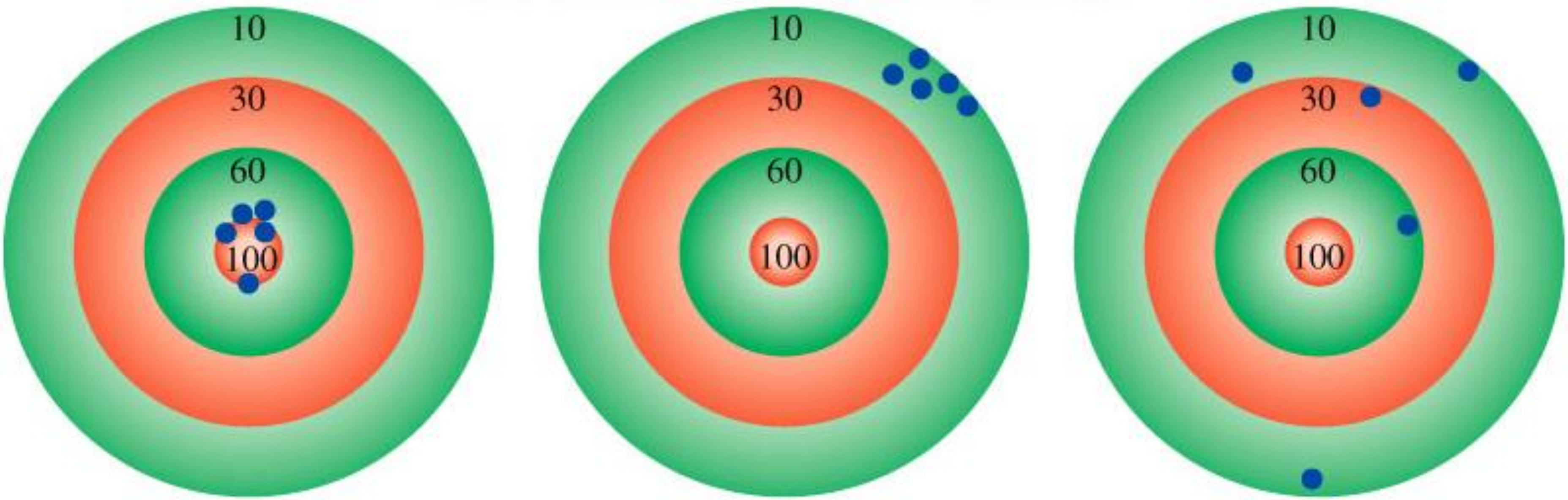
- Ice floating in a solution of sodium chloride in water also consists of two phases, ice and the liquid solution.

- A heterogeneous mixture of talk powder and sugar.

1.5 Measurement and Significant Figures

Accuracy – how close a measurement is to the *true* value

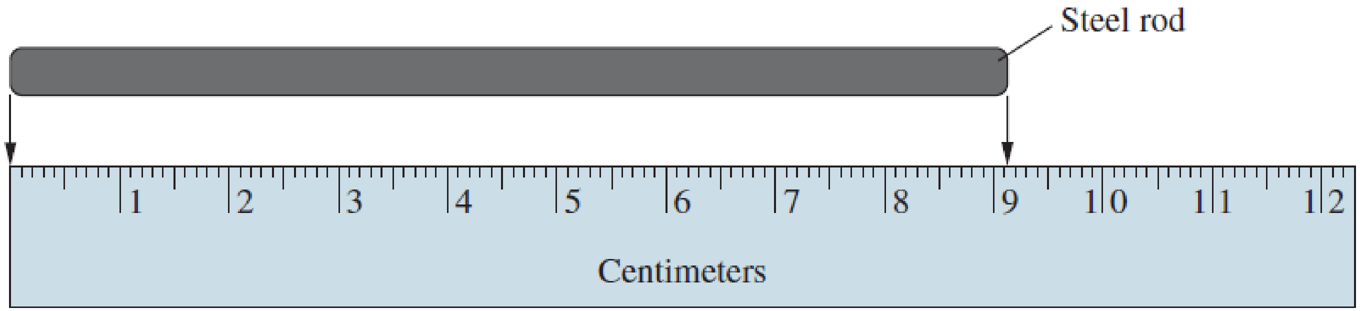
Precision – how close a set of measurements are to each other



accurate
&
precise

precise
but
not accurate

not accurate
&
not precise



The length of the rod is just over 9.1 cm. On successive measurements, we estimate the length by eye at 9.12, 9.11, and 9.13 cm. We record the length as between 9.11 cm and 9.13 cm.

- ✓ To indicate the precision of a measured number, we often use the concept of significant figures.
- ✓ **Significant figures** are *those digits in a measured number (or in the result of a calculation with measured numbers) that* ***include all certain digits plus a final digit having some uncertainty****.*

- ✓ You could report the result as the average, 9.12 cm.
- ✓ The first two digits (9.1) are certain; the next digit (2) is **estimated**, so it has some uncertainty.
- ✓ It would be incorrect to write 9.120 cm for the length of the rod. This would say that the last digit (0) has some uncertainty but that the other digits (9.12) are certain, which is not true.

Number of Significant Figures

9.12 cm → 3 significant figures

9.123 cm → 4 significant figures

- To count the number of significant figures in a given measured quantity, you observe the following rules:

1. All digits are significant except zeros at the beginning of the number and possibly terminal zeros (one or more zeros at the end of a number). Thus, 9.12 cm, 0.912 cm, and 0.00912 cm all contain three significant figures.
2. Terminal zeros ending at the right of the decimal point are significant. Each of the following has three significant figures: 9.00 cm, 9.10 cm, 90.0 cm.
3. Terminal zeros in a number without an explicit decimal point may or may not be significant. If someone gives a measurement as 900 cm, you do not know whether one, two, or three significant figures are intended. If the person writes 900. cm (note the decimal point), the zeros are significant. More generally, you can remove any uncertainty in such cases by expressing the measurement in scientific notation.

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 non zero
Single digit number

n is a positive or
negative integer

Rules for Significant Figures

1. All non-zero numbers are significant.

   e.g., 3.456 has (**4** sig. figs)

2. Zeros between non-zero numbers are significant.

   e.g., 20089 has (**5** sig. figs)

   Can be written as or 2.0089×10^4 (**5** sig. figs)

3. Trailing zeros always count as significant **<u>if number has decimal point</u>**

   e.g., 500. or 5.00×10^2 has **3** sig. figs

Rules for Significant Figures

4. Final zeros on number without decimal point are **NOT significant**

   e.g., 104956000

   (Unknown)

5. Final zeros to right of decimal point are significant

   e.g., 3.00 has **3** sig. figs.

6. Leading zeros, to left of first nonzero digit, are never counted as significant

   e.g., 0.00012 or 1.2×10^{-4} has **2** sig. figs.

How many significant figures does each of the following numbers have?

		Scientific Notation	# of Sig. Figs.
1.	413.97	4.1397×10^2	**5**
2.	0.0006	6×10^{-4}	**1**
3.	5.120063	5.120063	**7**
4.	161000		**Unknown**
5.	3600.	3.600×10^3	**4**

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

N is a single Non-zero digit

n is a positive or negative integer

Q) How many significant figures are in 19.0000?

Q) How many significant figures are in 0.0005650850?

Could be rewritten as 5.650850×10^{-4}

- **Rounding**

1. If this digit is 5 or greater, add 1 to the last digit to be retained and drop all digits farther to the right. Thus, rounding 1.2151 to three significant figures gives 1.22.
2. If this digit is less than 5, simply drop it and all digits farther to the right. Rounding 1.2143 to three significant figures gives 1.21.

Q)Round each of the following to three significant figures. Use scientific notation where needed.

1. 37.459 **37.5 or 3.75×10^1**
2. 5431978 **5.43×10^6**
3. 132.7789003 **133 or 1.33×10^2**
4. 0.00087564 **8.76×10^{-4}**

Q) Round 0.00564458 to four significant figures and express the answer using scientific notation.

A. 5.64×10^{-2}
B. 5.000×10^{-3}
C. 5.645×10^{-4}
D. 0.56446
E. 5.645×10^{-3}

Significant Figures in Calculations

Multiplication and Division

- Number of significant figures in answer = number of significant figures in **least precise** measurement

e.g., $10.54 \times 31.4 \times 16.987$ **$= 5621.9 = 5.62\times10^{3}$**

4 sig. figs. × 3 sig. figs. × 5 sig. figs. = 3 sig. figs.

e.g., $5.896 \div 0.008$ **$= 737 = 7\times10^{2}$**

4 sig. figs. ÷ 1 sig. fig. = 1 sig. fig.

Give the value of the following calculation to the correct number of significant figures.

$$\left(\frac{635.4 \times 0.0045}{2.3589}\right)$$

A. 1.21213

B. 1.212

C. 1.212132774

D. 1.2

E. 1

Significant Figures in Calculations

Addition and Subtraction

- Answer has same number of decimal places as quantity with **fewest number** of decimal places.

e.g.,	12.9753	4 decimal places
	+319.5	1 decimal place
	+ 4.398	3 decimal places
	336.9	1 decimal place

e.g.,	397	0 decimal places
	– 273.15	2 decimal places
	124	0 decimal place

Q) For each calculation, give the answer to the **correct number of significant figures**.

1. 10.0 g + 1.03 g + 0.243 g = **11.3 g** or **1.13×10^{1} g**

2. 19.556 °C – 19.552 °C = **0.004 °C** or **4×10^{-3} °C**

3. 327.5 m × 4.52 m = **1480.3 = 1.48×10^{3} m^2**

4. 15.985 g ÷ 24.12 mL = **0.6627 g/mL** or **6.627×10^{-1} g/mL**

Q) When the expression,

$$412.272 + 0.00031 - 1.00797 + 0.000024 + 12.8$$

is evaluated, the result should be expressed as:

A. 424.06

B. 424.064364

C. 424.1

D. 424.064

E. 424

Q) For the following calculations, give the answer to the correct number of **significant figures**.

1. $$\frac{(71.359\ \text{m} - 71.357\ \text{m})}{(3.2\ \text{s} \times 3.67\ \text{s})} = \frac{(0.00\underline{2}\ \text{m})}{(1\underline{1}.744\ \text{s}^2)}$$

= (0.002/12)=(1.666 × 10^{-4}) =2 × 10^{-4} m/s^2

2. $$\frac{(13.674\ \text{cm} \times 4.35\ \text{cm} \times 0.35\ \text{cm})}{(856\ \text{s} + 1531.1\ \text{s})}$$

$$= \frac{(2\underline{0}.818665\ \text{cm}^3)}{(238\underline{7}.1\ \text{s})}$$

= (21/2387)= 0.0088 cm^3/s

Or 8.8 x 10^{-3}

Exact Numbers

(1)Numbers that come from definitions

12 in. = 1 ft

60 s = 1 min

(2)Numbers that come from direct count

– Number of people in small room

- Have no uncertainty
- <u>Assume they have infinite number of significant figures</u>
- The number of significant figures in a calculation result **depends only on** the numbers of significant figures in **quantities having uncertainties**

Q) If you have 9 coins in a jar. What is the total mass of the 9 coins when each coin has a mass of 3.0 grams ?

$$3.0 \text{ g} \times 9 = 27 \text{ g}$$

The number 9 is exact and does not determine the number of significant figures

Q) How many feet are there in 36.00 inches? Express the answer with the correct number of significant figures: (1 ft.=12 in.)

$$36.00 \text{ in} \times \left(\frac{\text{ft.}}{12 \text{ in.}} \right) =$$

A. 3 ft.

B. 3.0 ft

C. 3.00 ft.

D. 3.000 ft.

E. 3.00000 ft.

Q) For the following calculation, give the answer to the correct number of significant figures.

$$\frac{(14.5\ \text{cm} \times 12.334\ \text{cm})}{(2.223\ \text{cm} - 1.04\ \text{cm})}$$

A. 179 cm^2

B. 1.18 cm

C. 151.2 cm

D. 151 cm

E. 178.843 cm^2

$$\frac{(17\underline{8}.843\ \text{cm}^2)}{(1.1\underline{8}3\ \text{cm})}$$

$$= 179 / 1.18$$
$$= 151.694$$
$$=152$$

$$=151.177$$

3 sig.fig.

$$=151$$

Note: Do not round intermediate answers !

(3.0 x 2.001) + 0.05

2 sig.fig x 4 sig.fig. = 2 sig.fig. = 6.0

= $6.\underline{0}03 + 0.05$

6.0 + 0.05 = 1 d.p. + 2 d.p = 1 d.p

= 6.053

= 6.0 (no rounding)

= 6.1 (rounding) Final answer

Perform the following calculations and round the answers to the correct number of significant figures (units of measurement have been omitted).

a. $\dfrac{2.568 \times 5.8}{4.186}$

b. $5.41 - 0.398$

c. $3.38 - 3.01$

d. $4.18 - 58.16 \times (3.38 - 3.01)$

a = 3.6
b = 5.01
c = 0.37
d = - 17

1.6 SI Units

International System of units (*metric system*)

French *le Système International d'Unités*

TABLE 1.1	**SI Base Units**	
Quantity	Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	kelvin	K
Amount of substance	mole	mol
Electric current	ampere	A
Luminous intensity	candela	cd

TABLE 1.2	**Selected SI Prefixes**	
Prefix	Multiple	Symbol
mega	10^{6}	M
kilo	10^{3}	k
deci	10^{-1}	d
centi	10^{-2}	c
milli	10^{-3}	m
micro	10^{-6}	μ*
nano	10^{-9}	n
pico	10^{-12}	p

*Greek letter mu, pronounced "mew."

In this chapter, we will discuss four base quantities: length, mass, time, and temperature.

(Q) The SI unit of length is:

A. millimeter

B. meter

C. yard

D. centimeter

E. foot

Examples:

$2.54 \text{ cm} = 2.54 \times 10^{-2} \text{ m}$

$1 \text{ mL} = 10^{-3} \text{ L}$

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

$1 \text{ ng} = 10^{-9} \text{ g}$

$1{,}130{,}000 \text{ m} = 1.13 \times 10^{6} \text{ m} =$ **1.13 Mm**

TABLE 1.5 SI Prefixes—Their Meanings and Values[a]

Prefix	Meaning	Symbol	Prefix Value[b] (numerical)	Prefix Value[b] (power of ten)
exa		E		10^{18}
peta		P		10^{15}
tera		T		10^{12}
giga	billions of	G	1000000000	10^{9}
mega	millions of	M	1000000	10^{6}
kilo	thousands of	k	1000	10^{3}
hecto		h		10^{2}
deka		da		10^{1}
deci	tenths of	d	0.1	10^{-1}
centi	hundredths of	c	0.01	10^{-2}
milli	thousandths of	m	0.001	10^{-3}
micro	millionths of	μ	0.000001	10^{-6}
nano	billionths of	n	0.000000001	10^{-9}
pico	trillionths of	p	0.000000000001	10^{-12}
femto		f		10^{-15}
atto		a		10^{-18}

[a]Prefixes in red type are used most often.

[b]Numbers in these columns can be interchanged with the corresponding prefix.

TABLE 1.3 Some Non-SI Metric Units Commonly Used in Chemistry

Measurement	Unit	Abbreviation	Value in SI Units
Length	angstrom	Å	1 Å = 0.1 nm = 10^{-10} m
Mass	atomic mass unit	u (amu)	1 u = 1.66054×10^{-27} kg (rounded to six digits)
	metric ton	t	1 t = 10^3 kg
Time	minute	min.	1 min. = 60 s
	hour	h	1 h = 60 min. = 3600 s
Temperature	degree Celsius	°C	$T_K = t_{°C} + 273.15$
Volume	liter	L	1 L = 1000 cm^3

TABLE 1.4 Some Useful Conversions

Measurement	English Unit	English/SI Equality[a]
Length	inch	1 in. = 2.54 cm
	yard	1 yd = 0.9144 m
	mile	1 mi = 1.609 km
Mass	pound	1 lb = 453.6 g
	ounce (mass)	1 oz = 28.35 g
Volume	gallon	1 gal = 3.785 L
	quart	1 qt = 946.4 mL

Laboratory Measurements

- **Four common**
  1. Length
  2. Volume
  3. Mass
  4. Temperature

Laboratory Measurements

1. **Length**

– SI Unit is meter (m)

– Meter too large for most laboratory measurements

– Commonly use

- Centimeter (cm)

  $1\ cm = 10^{-2}\ m = 0.01\ m$

- Millimeter (mm)

  $1\ mm = 10^{-3}\ m = 0.001\ m$

2. Volume

- Dimensions of $(\text{length})^3$
- SI unit for Volume = m^3
- Most laboratory measurements use V in liters (L)

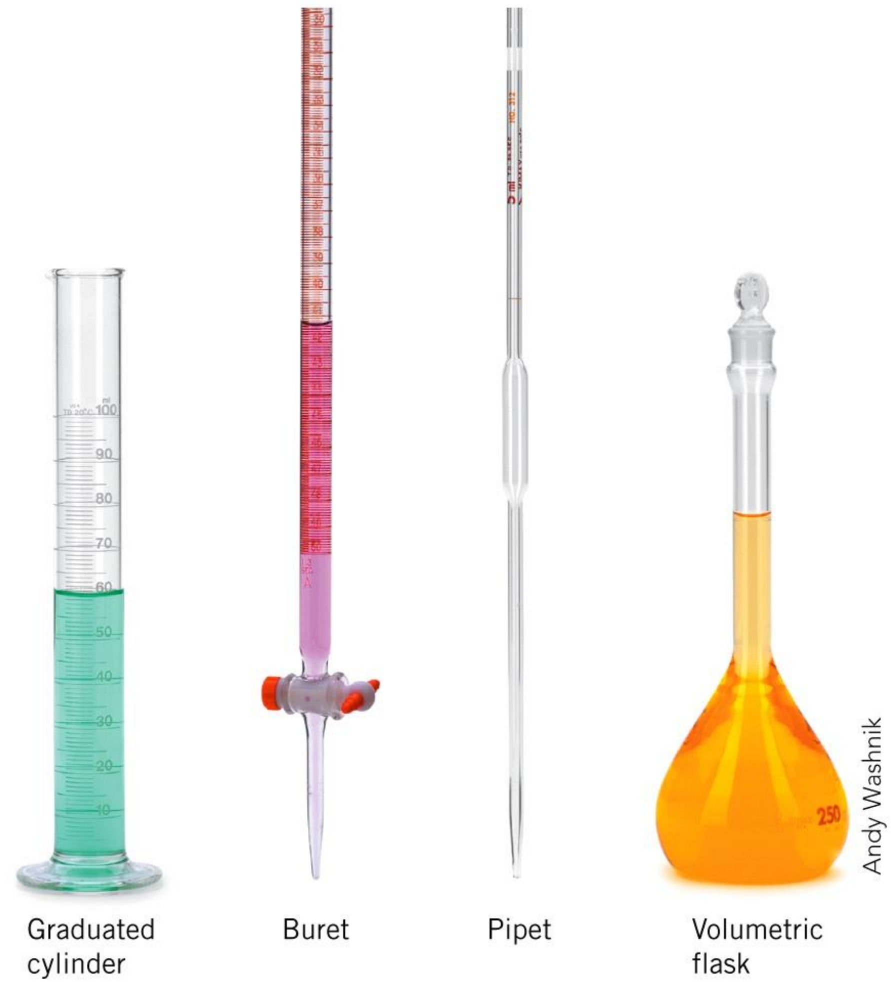
  $1\ L = 1\ dm^3$

  Chemistry glassware marked in L or mL

  $1\ L = 1000\ mL$

- What is a mL?

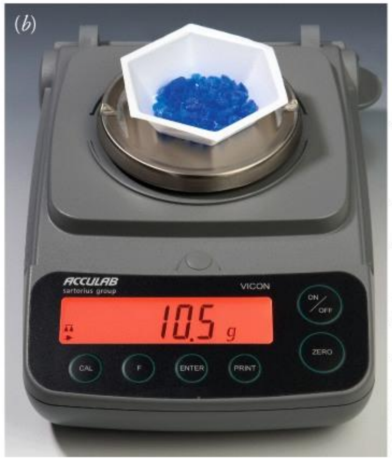
  $1\ mL = 1\ cm^3$



3. Mass

- SI unit is kilogram (kg)
  - Frequently use grams (g) in laboratory as more realistic size
- 1 kg = 1000 g   1 g = 0.001 kg = $\frac{1}{1000}$ g
- Mass is measured by comparing weight of sample with weights of known standard masses
- Instrument used = balance

Michael Watson



© 2005 Richard Megna/Fundamental Photographs

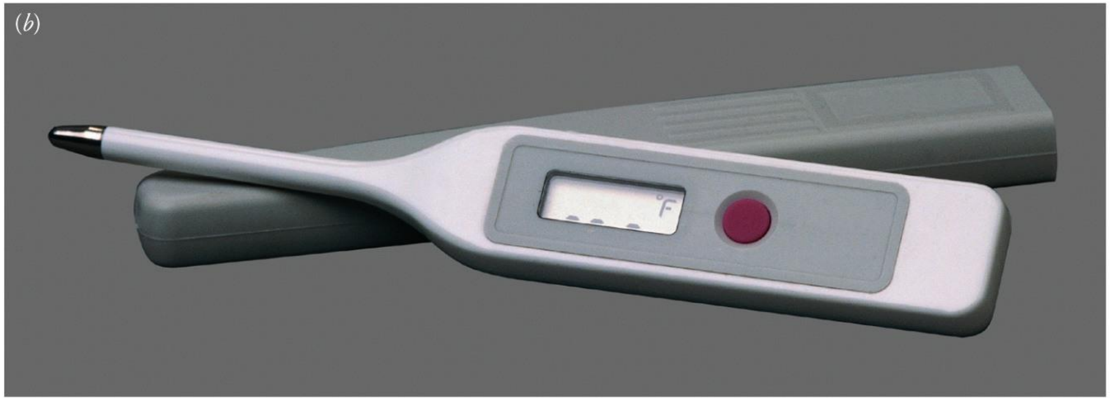
Charles D. Winters/Photo Researchers

4. Temperature

- Measured with thermometer
- Three common scales

A. Fahrenheit scale

- Common in US
- Water freezes at 32 °F and boils at 212 °F
- 180 degree units between melting and boiling points of water

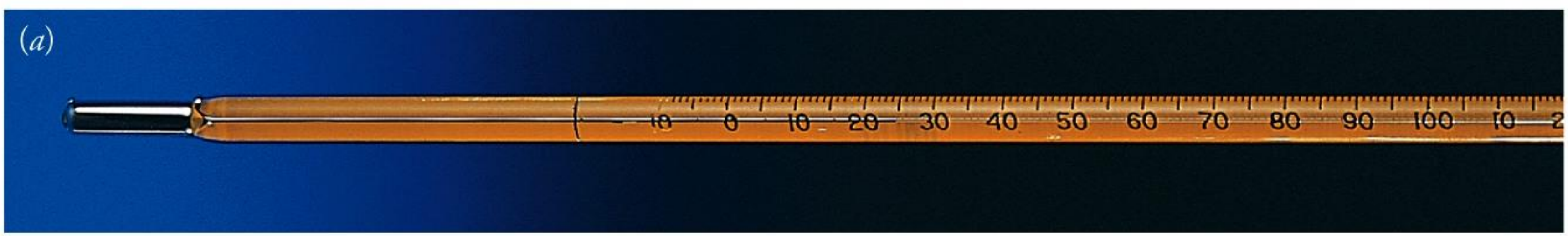


Corbis Images

4. Temperature

B. Celsius scale

- Most common for use in science
- Water freezes at 0 °C
- Water boils at 100 °C
- 100 degree units between melting and boiling points of water



Michael Watson

4. Temperature

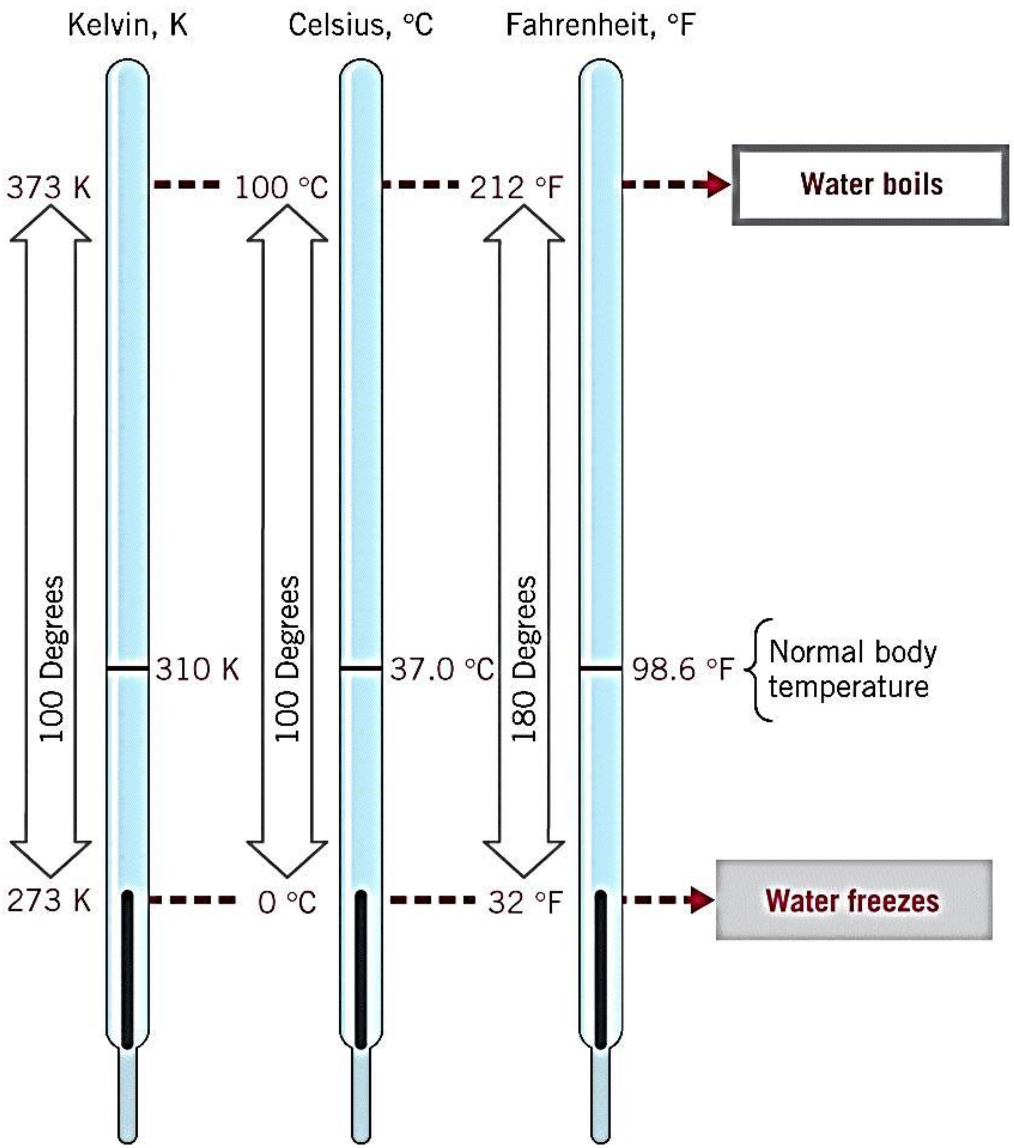
C. Kelvin scale

- SI unit of temperature is **kelvin (K)**
  - **Note:** No degree symbol in front of K
- Water freezes at 273.15 K and boils at 373.15 K
  - 100 degree units between melting and boiling points
- Only difference between Kelvin and Celsius scale is zero point

Absolute Zero

- Zero point on Kelvin scale
- Corresponds to nature's lowest possible temperature

Temperature Conversions



How to convert between°F and °C?

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

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

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

Temperature Conversions

- Common laboratory thermometers are marked in Celsius scale
- How to convert to Kelvin scale

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

$$273.15\ K = 0\ {}^{0}C$$
$$373.15\ K = 100\ {}^{0}C$$

- Amounts to adding 273.15 to Celsius temperature

Example: What is the Kelvin temperature of a solution at 25 °C?

$$T_K = (25\ {}^{\circ}C + 273.15\ {}^{\circ}C)\frac{1\,K}{1\ {}^{\circ}C} \quad \mathbf{= 298\ K}$$

1. Convert 121 °F to the Celsius scale.

$$^{0}F = \frac{9}{5} \times {}^{0}C + 32 \qquad t_C = \left(t_F - 32\ °F\right)\left(\frac{5\ °C}{9\ °F}\right)$$

$$t_C = \left(121\ {}^{o}F - 32\ °F\right)\left(\frac{5\ °C}{9\ °F}\right)$$ **= 49 °C**

2. Convert 121 °F to the Kelvin scale.

- We already have in °C so…

$$T_K = (t_C + 273.15\,°C)\frac{1\,K}{1\,°C} = (49 + 273.15\,°C)\frac{1\,K}{1\,°C}$$

***T*$_K$ = 332 K**

3. Convert 77 K to the Celsius scale.

$$T_K = (t_C + 273.15\ °C)\frac{1\ K}{1\ °C} \qquad t_C = (T_K - 273.15\ K)\frac{1\ °C}{1\ K}$$

$$t_C = (77\ K - 273.15\ K)\frac{1\ °C}{1\ K}\ \mathbf{= -196\ °C}$$

4. Convert 77 K to the Fahrenheit scale.

– We already have in °C so

$$t_F = \left(\frac{9\ °F}{5\ °C}\right)(-196\ °C) + 32\ °F\ \mathbf{= -321\ °F}$$

The melting point of UF_6 is 64.53 °C. What is the melting point of uranium UF_6 on the Fahrenheit scale?

A. 67.85 °F

B. 96.53 °F

C. 116.2 °F

D. 337.5 °F

E. 148.2 °F

$$t_{\mathrm{F}} = \left(\frac{9\ °\mathrm{F}}{5\ °\mathrm{C}}\right) t_{\mathrm{C}} + 32\ °\mathrm{F}$$

$$t_{\mathrm{F}} = \left(\frac{9\ °\mathrm{F}}{5\ °\mathrm{C}}\right) 64.53\ °\mathrm{C} + 32\ °\mathrm{F}$$

SI Units

- All physical quantities will have units **derived** from these seven SI base units

e.g., Area

- Derived from SI units based on definition of area
- length × width = area
- meter × meter = area

$$m \times m = m^2$$

- SI unit for area = square meters = m^2

Note: Units undergo same kinds of mathematical operations that numbers do

TABLE 1.3 Derived Units

Quantity	Definition of Quantity	SI Unit
Area	Length squared	m^2
Volume	Length cubed	m^3
Density	Mass per unit volume	kg/m^3
Speed	Distance traveled per unit time	m/s
Acceleration	Speed changed per unit time	m/s^2
Force	Mass times acceleration of object	$kg \cdot m/s^2$ (= newton, N)
Pressure	Force per unit area	$kg/(m \cdot s^2)$ (= pascal, Pa)
Energy	Force times distance traveled	$kg \cdot m^2/s^2$ (= joule, J)

- What is the SI derived unit for velocity?

$$\text{Velocity } (v) = \frac{\text{distance}}{\text{time}}$$

$$\text{Velocity units} = \frac{\text{meters}}{\text{seconds}} = \frac{\text{m}}{\text{s}}$$

- What is the SI derived unit for volume of a cube?

$$\text{Volume } (V) = \text{length} \times \text{width} \times \text{height}$$

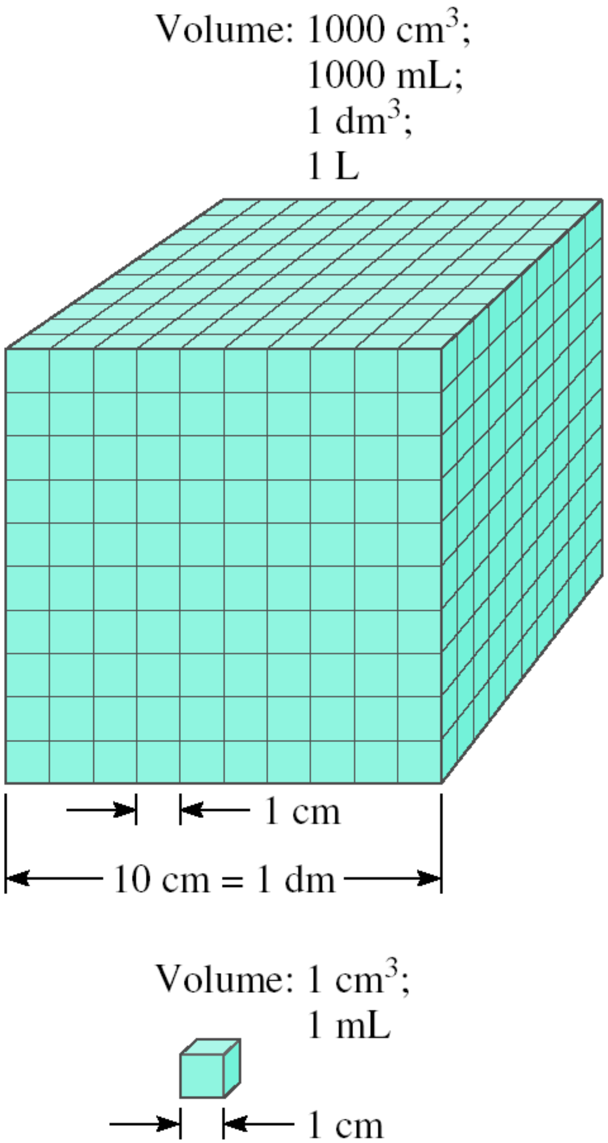
$$V = \text{meter} \times \text{meter} \times \text{meter}$$

$$\boldsymbol{V = \text{m}^3}$$

What is the SI derived unit for acceleration (hint: acceleration = distance/time^2)?

A. mm/min
B. yd/hr^2
C. m/s^2
D. m/s
E. ft^3

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



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

$$1\ dm^3 = (1 \times 10^{-1}\ m)^3 = 1 \times 10^{-3}\ m^3$$

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

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

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.

given quantity x conversion factor = desired quantity

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

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.)

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

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

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

Q) A liquid helium storage tank has a volume of 275 L. What is the volume in m^3?

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

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

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

Example: How to convert a person's height from 68.0 in to cm? if 2.54 cm = 1 in.

Example: Convert 0.097 m to mm.

Example: Convert 3.5 m^3 to cm^3.

Q) Convert speed of light from 3.00×10^8 m/s to mi/hr

(1 mi = 1.609 km)

The Toyota Camry hybrid electric car has a gas mileage rating of 56 miles per gallon. What is this rating expressed in units of kilometers per liter?

1 gal = 3.784 L   1 mile = 1.609 km

A. 1.3×10^2 km L^{-1}

B. 24 km L^{-1}

C. 15 km L^{-1}

D. 3.4×10^2 km L^{-1}

E. 9.2 km L^{-1}

$$56 \frac{\cancel{mi}}{\cancel{gal}} \times \frac{1\,\cancel{gal}}{3.784\,L} \times \frac{1.609\,km}{1\,\cancel{mi}}$$

The volume of a basketball is 433.5 in^3. Convert this to mm^3. (1 in. = 2.54 cm)

A. 1.101×10^{-2} mm^3

B. 7.104×10^{6} mm^3

C. 7.104×10^{4} mm^3

D. 1.101×10^{4} mm^3

E. 1.101×10^{6} mm^3

Example 1.7: The world's oceans contain approximately 1.35×10^{9} km^3 of water. What is this volume in liters?

1km = 1×10^3 m = $1 \times 10^3 \times 10^2$ cm = 1×10^5 cm ($km^3 \rightarrow cm^3$)

($cm^3 \rightarrow$ L) (1 cm^3 = 1mL) (1mL= 1×10^{-3}L)

Density

- Ratio of object's mass to its volume

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

- Units (depends on what units we use for mass and volume.
  - **g/mL** or **g/cm^3**
  - **Or g/L** or **kg/L**

- A student weighs a piece of gold that has a volume of 11.02 cm^3 of gold. She finds the mass to be 212 g. What is the density of gold?

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

$$d = \frac{212 \text{ g}}{11.02 \text{ cm}^3} = \mathbf{19.3\ g/cm^3}$$

Another student has a piece of gold with a volume of 1.00 cm^3. What does it weigh? **19.3 g**

What if it were 2.00 cm^3 in volume? **38.6 g**

(Q) If the density of an object is 2.87 x 10^{-4} lbs/cubic inch, what is its density in g/mL? (1 lb = 454 g, 1 inch = 2.54 cm)