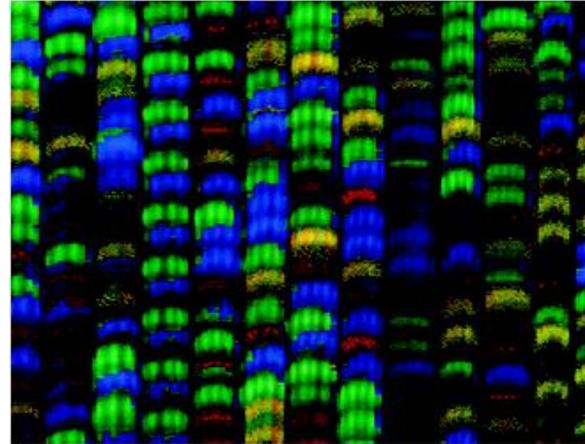


Chapter 1

Chemistry and measurements

1.1 Modern Chemistry: A Brief Glimpse

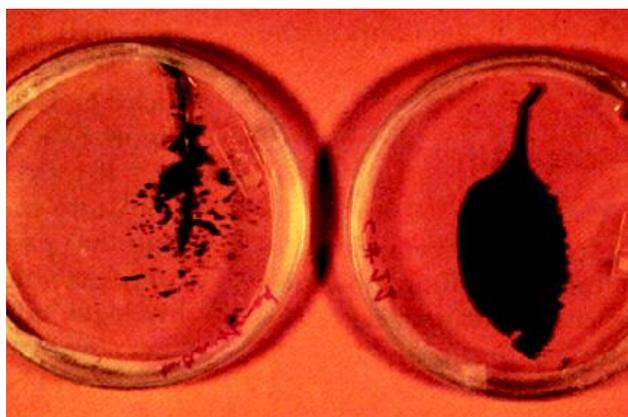
- Health and Medicine
 - Sanitation systems
 - Surgery with anesthesia
 - Vaccines and antibiotics
 - Gene therapy



- 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

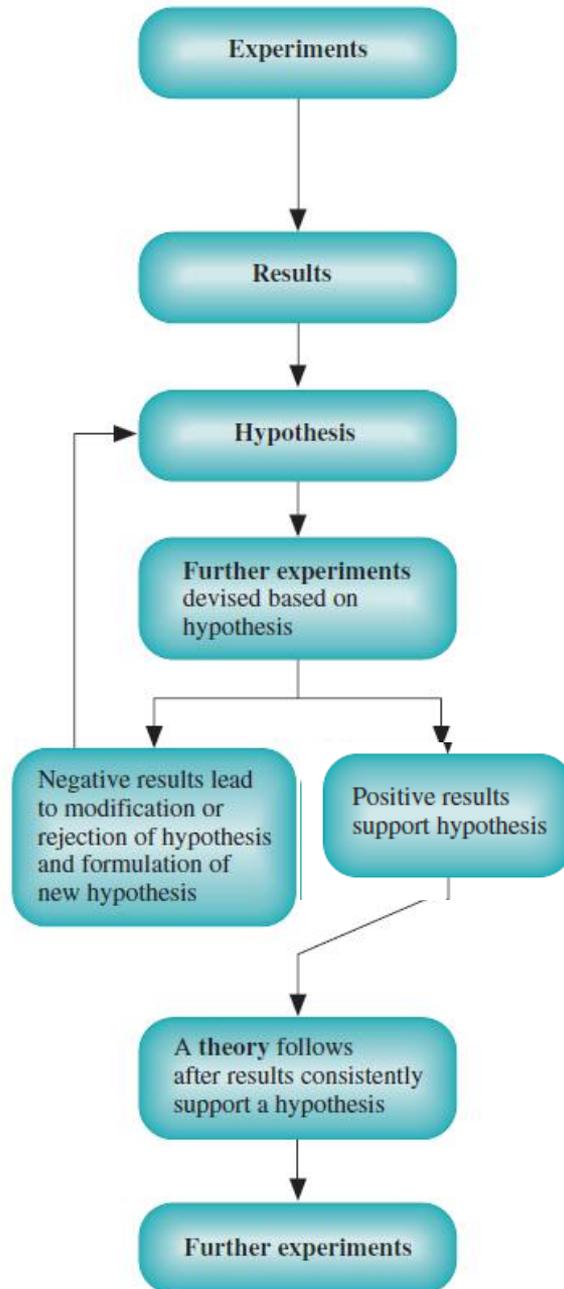
1.2 Experiment and Explanation

- 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



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.



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

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

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)

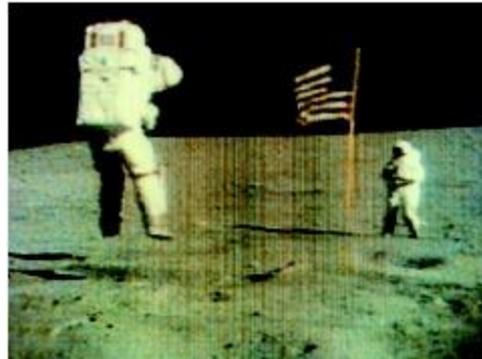
$$1 \text{ kg} = 1000 \text{ g} = 1 \times 10^3 \text{ g}$$

weight – force that gravity exerts on an object

weight = $c \times$ mass

on earth, $c = 1.0$

on moon, $c \sim 0.1$



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

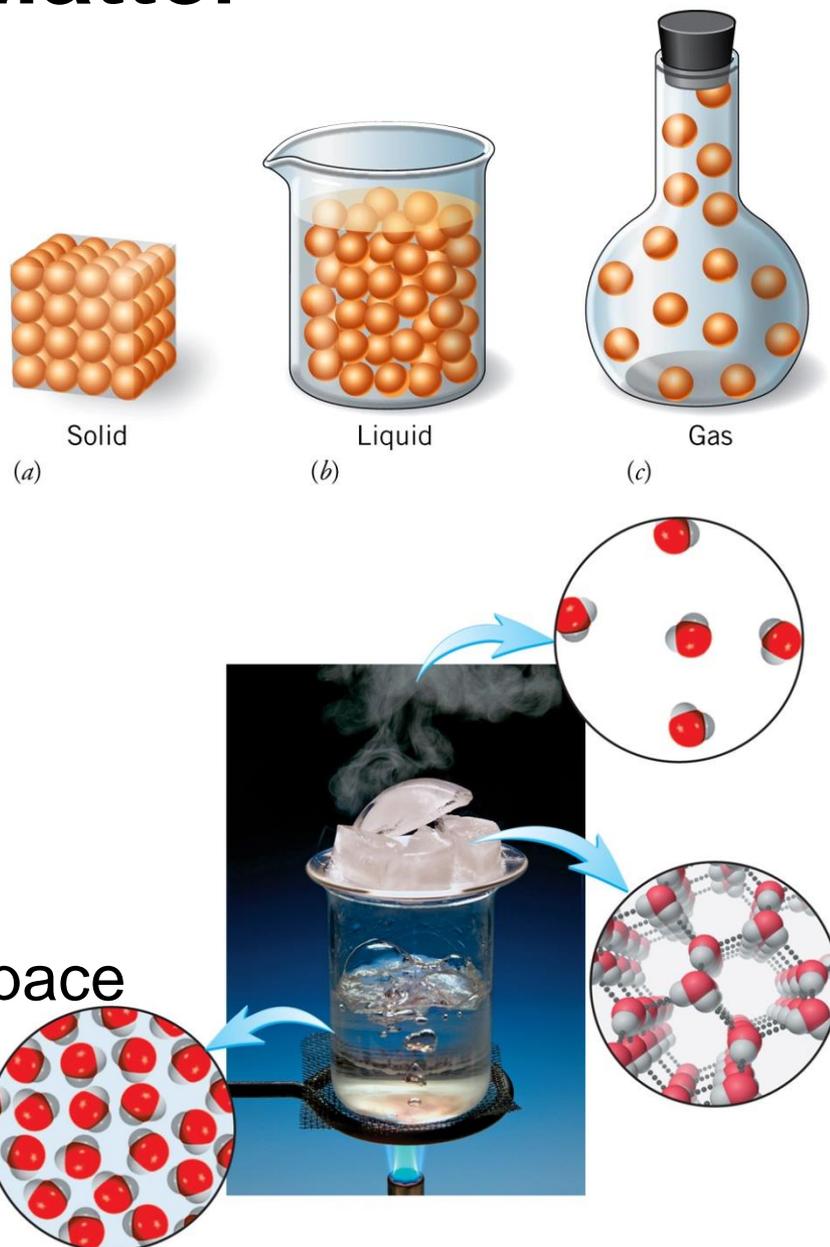
Liquids:

- 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)

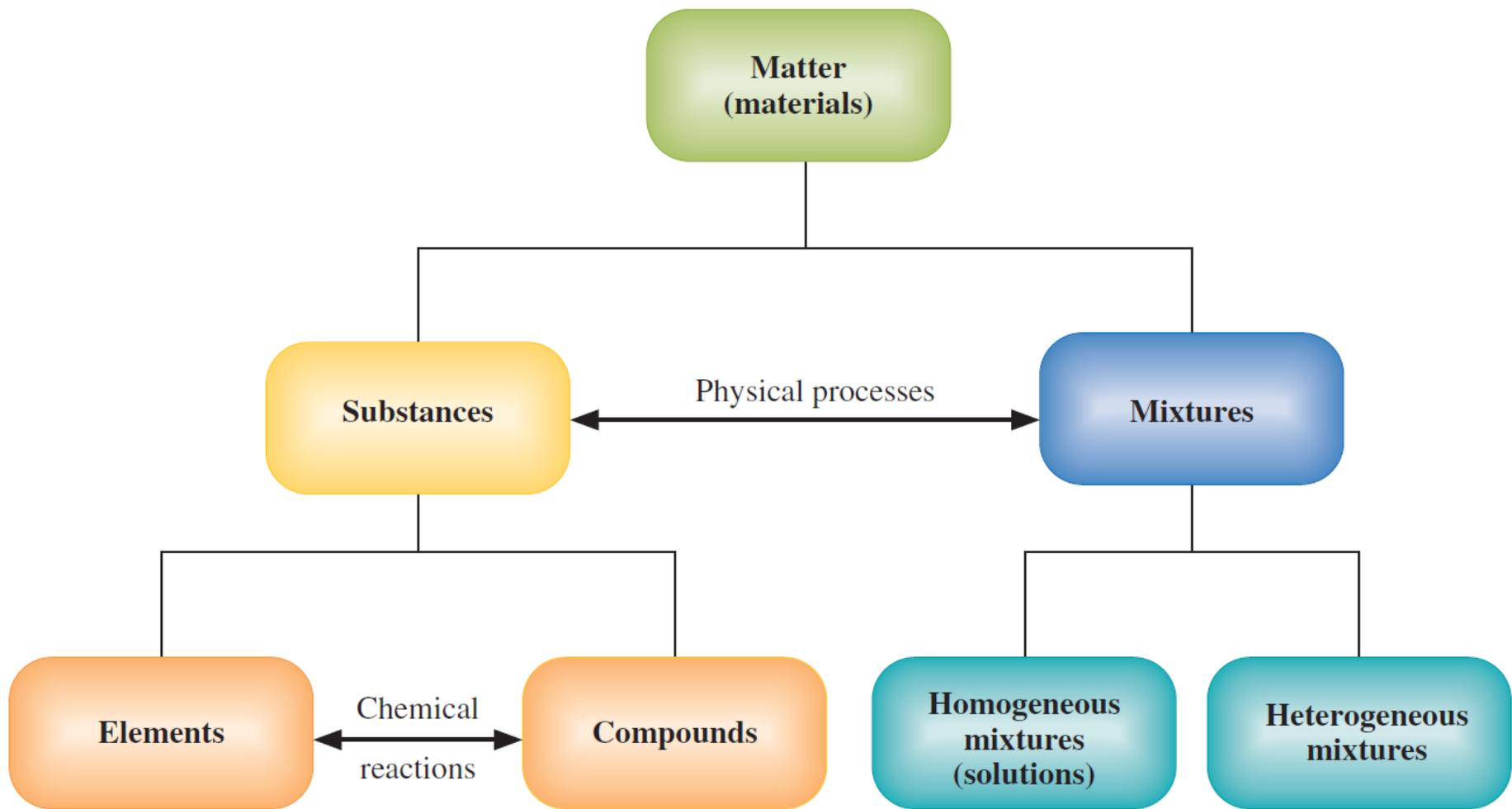
	Chemical	Physical
Magnesium burns when heated		
Magnesium metal tarnishes in air		
Magnesium metal melts at 922 K		
Orange juice lightens when water is added		

- **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.*
H₂O, NaCl, CO₂



	Chicken Noodle Soup	Ice (H₂O)	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.*

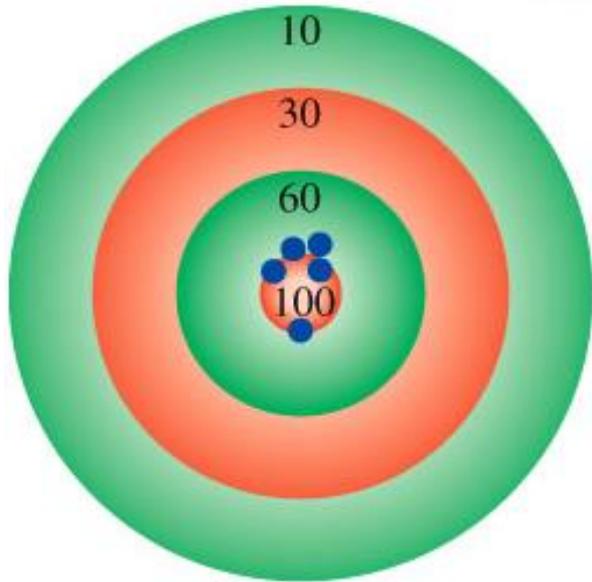
Examples:

- 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.

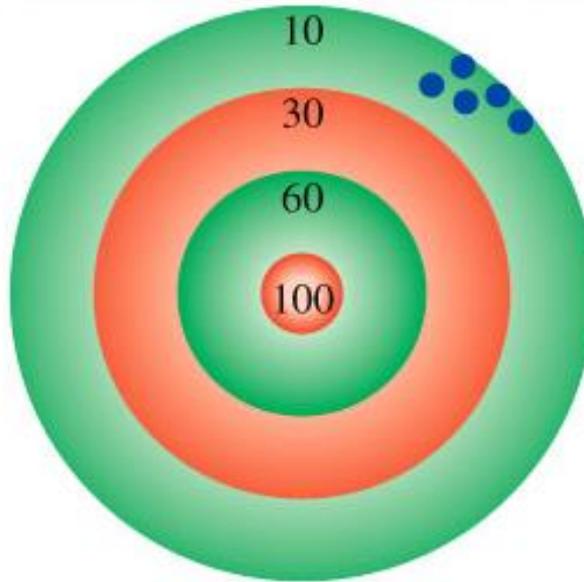
1.5 Measurement and Significant Figures

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

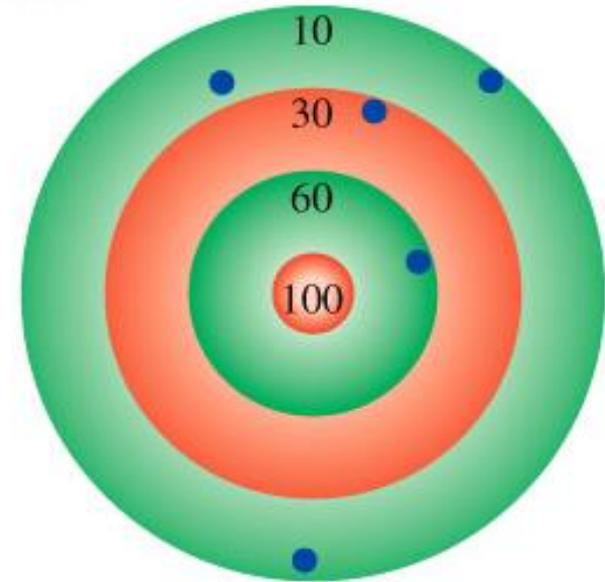
Precision – how close a set of measurements are to each other



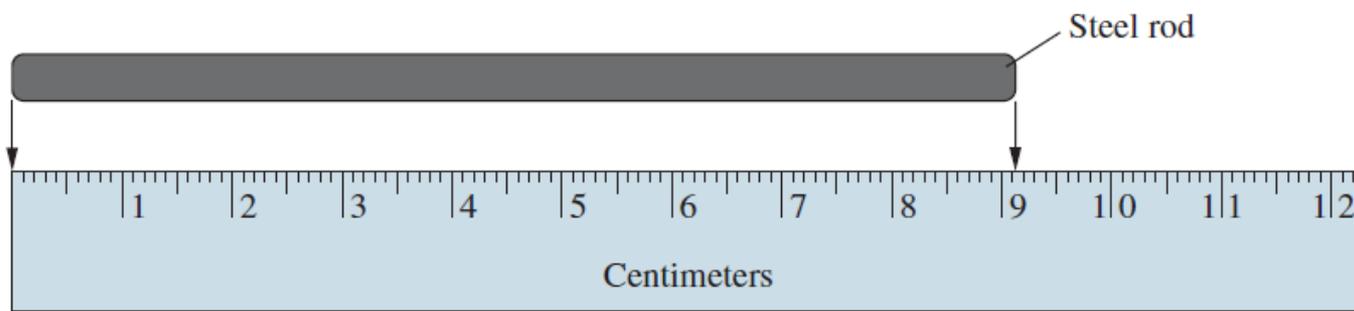
accurate
&
precise



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:

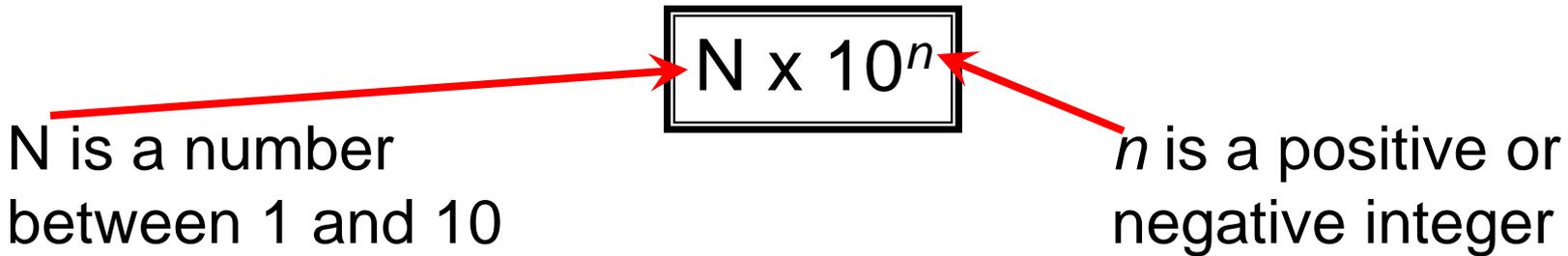
602,200,000,000,000,000,000,000

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

The mass of a single carbon atom in grams:

0.000000000000000000000000199

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


$$N \times 10^n$$

N is a number
between 1 and 10

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., 20,089 has (**5** sig. figs)

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

3. Trailing zeros always count as significant **if number has decimal point**

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., 104,956,000 or 1.04956×10^8
has **6** sig. figs.

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. 161,000	1.61×10^5	3
5. 3600.	3.600×10^3	4

$$N \times 10^n$$

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., $10.54 \times 31.4 \times 16.987 = 5621.9 = 5.62 \times 10^3$

4 sig. figs. \times 3 sig. figs. \times 5 sig. figs. = 3 sig. figs.

e.g., $5.896 \div 0.008 = 737 = 7 \times 10^2$

4 sig. figs. \div 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	<u>3 decimal places</u>
	<hr/>	
	336.9	1 decimal place

e.g.,	397	0 decimal places
	- 273.15	<u>2 decimal places</u>
	<hr/>	
	124	0 decimal place

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

1. $10.0 \text{ g} + 1.03 \text{ g} + 0.243 \text{ g} =$ **11.3 g** or **$1.13 \times 10^1 \text{ g}$**

2. $19.556 \text{ }^\circ\text{C} - 19.552 \text{ }^\circ\text{C} =$ **0.004 $^\circ\text{C}$** or **$4 \times 10^{-3} \text{ }^\circ\text{C}$**

3. $327.5 \text{ m} \times 4.52 \text{ m} =$ **1480.3 = $1.48 \times 10^3 \text{ m}^2$**

4. $15.985 \text{ g} \div 24.12 \text{ mL} =$ **0.6627 g/mL** or **$6.627 \times 10^{-1} \text{ g/mL}$**

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.
$$\frac{(71.359 \text{ m} - 71.357 \text{ m})}{(3.2 \text{ s} \times 3.67 \text{ s})} = \frac{(0.002 \text{ m})}{(11.744 \text{ s}^2)}$$
$$= (0.002/12) = (1.666 \times 10^{-4}) = 2 \times 10^{-4} \text{ m/s}^2$$

2.
$$\frac{(13.674 \text{ cm} \times 4.35 \text{ cm} \times 0.35 \text{ cm})}{(856 \text{ s} + 1531.1 \text{ s})}$$
$$= \frac{(20.818665 \text{ cm}^3)}{(2387.1 \text{ s})} = (21/2387) = 0.0088 \text{ cm}^3/\text{s}$$
$$\text{Or } 8.8 \times 10^{-3}$$

Exact Numbers

(1) Numbers that come from definitions

$$12 \text{ in.} = 1 \text{ ft}$$

$$60 \text{ s} = 1 \text{ min}$$

(2) Numbers that come from direct count

– Number of people in small room

- Have no uncertainty
- Assume they have infinite number of significant figures
- 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.

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

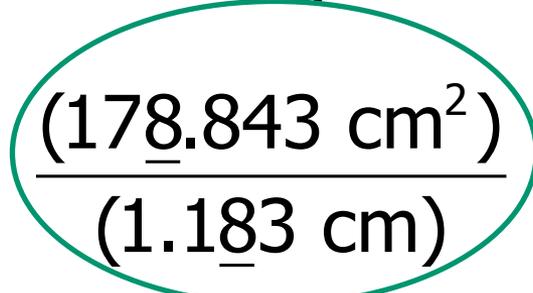
A. 179 cm^2

B. 1.18 cm

C. 151.2 cm

D. 151 cm

E. 178.843 cm^2


$$\frac{(178.843 \text{ cm}^2)}{(1.183 \text{ cm})}$$

$$\begin{aligned} &= 179 / 1.18 \\ &= 151.694 \\ &= 152 \end{aligned}$$

$$= 151.177$$

↓ 3 sig.fig.

$$= 151$$

Note: Do not round intermediate answers !

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

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

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