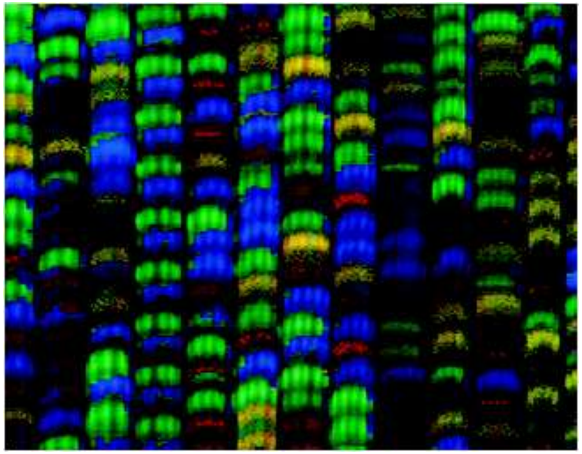
Chapter 1

Chemistry and measurements

1.1 Modern Chemistry: A Brief Glimpse

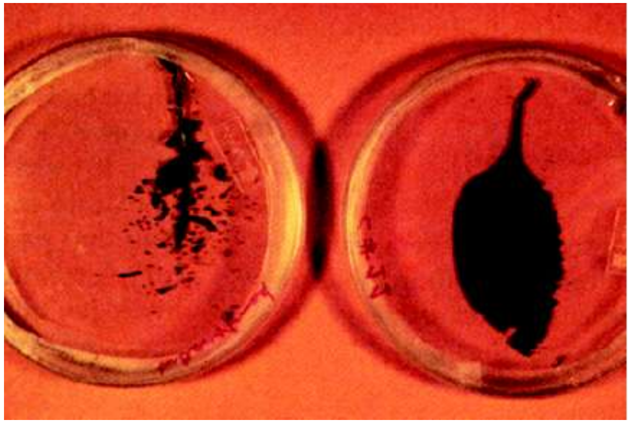
- Health and Medicine
  - Sanitation systems
  - Surgery with anesthesia
  - Vaccines and antibiotics
  - Gene therapy



- 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

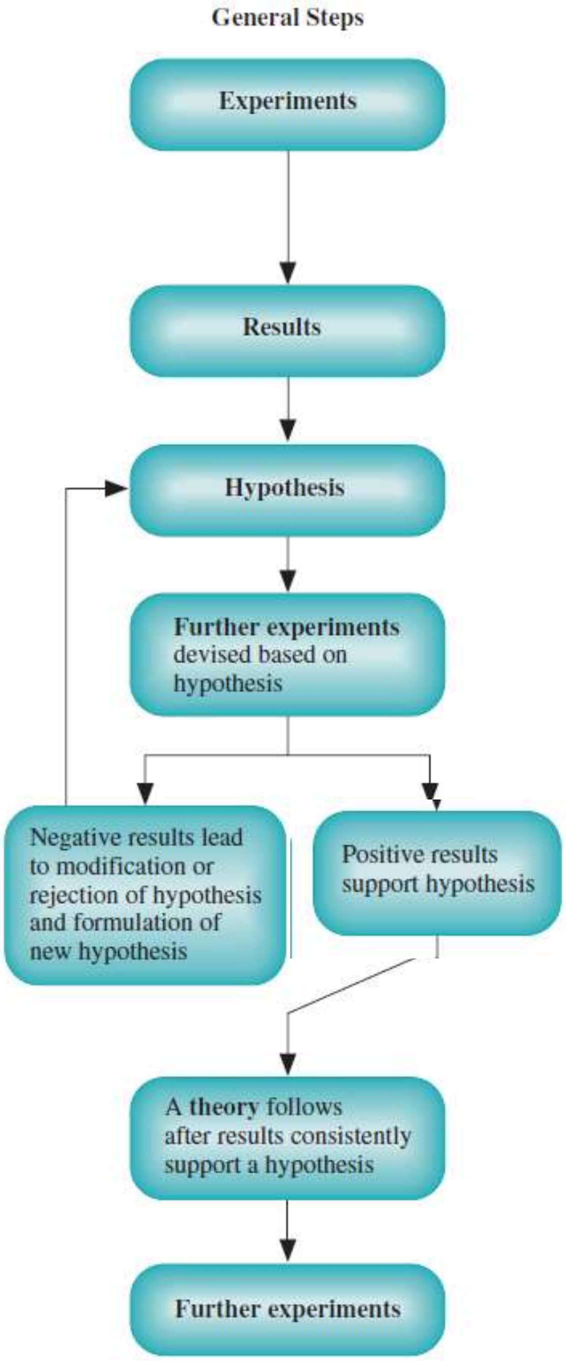
- An **experiment** is *an observation of natural phenomena carried out in a controlled manner so that the results can be duplicated and rational conclusions obtained*.

- 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

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

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

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

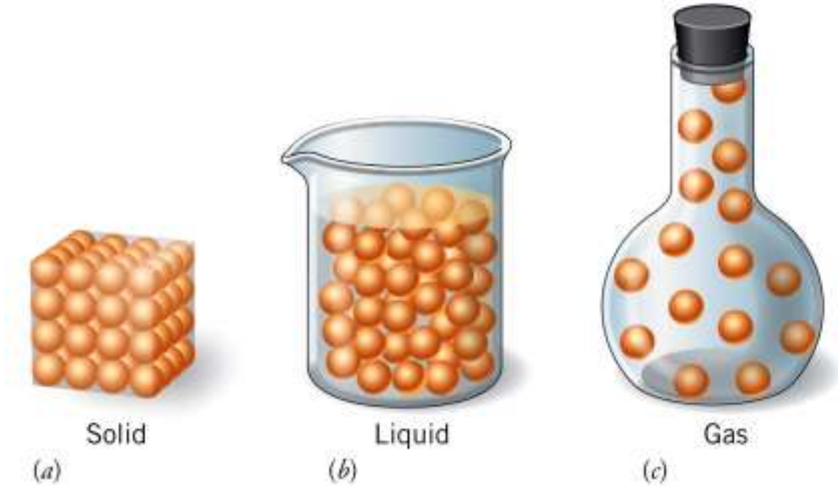
Liquids:

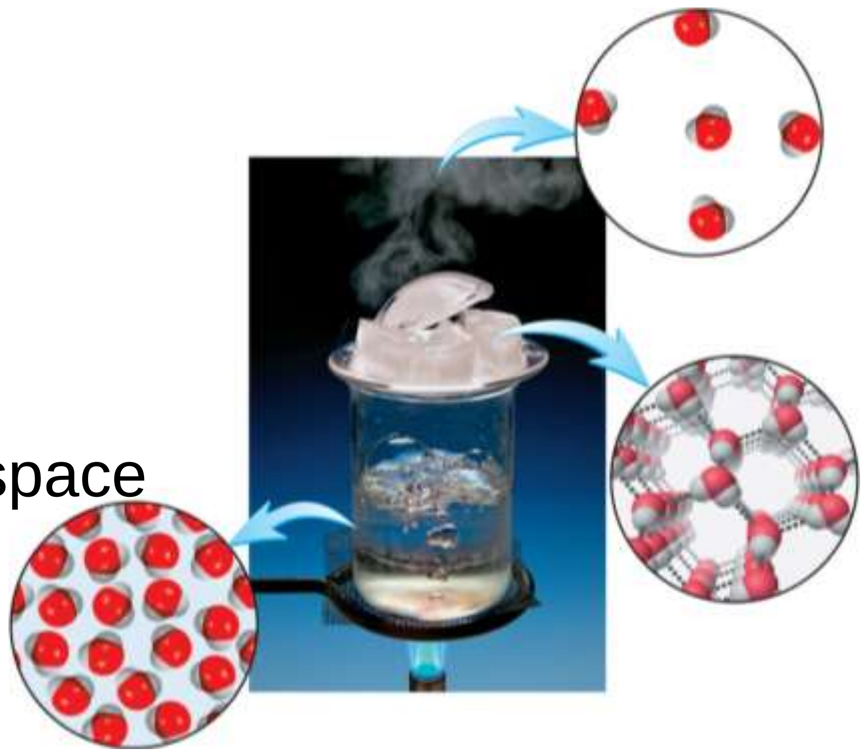
- Fixed volume, but take container shape
- Particles are close together
- Are able to flow

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* ***<u>cannot be</u>*** *separated into other kinds of matter by any <u>physical process</u>.*

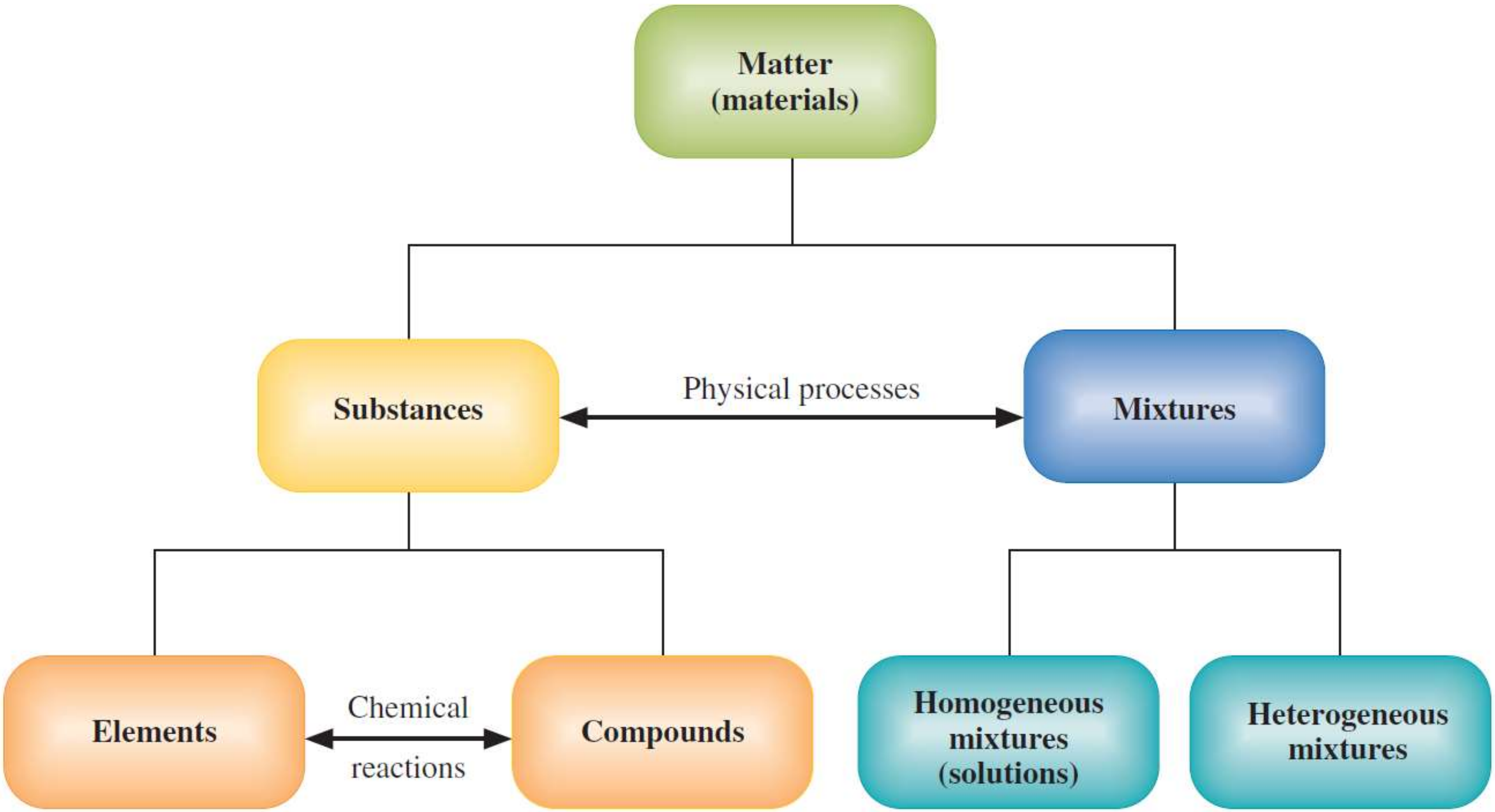
- **A mixture:** is *a kind of matter that* ***<u>can be</u>*** *separated by physical means into two or more substances*.

  -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* ***<u>cannot be</u>*** *decomposed by any* ***<u>chemical reaction</u>*** *into simpler substances*
  Fe, Au,Na etc…

- **A Compound:** is *a substance composed of two or more elements* ***<u>chemically</u>*** *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)
substance				
Element				
Compound				
Heterogeneous Mixture				
Homogeneous Mixture				

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

Examples:

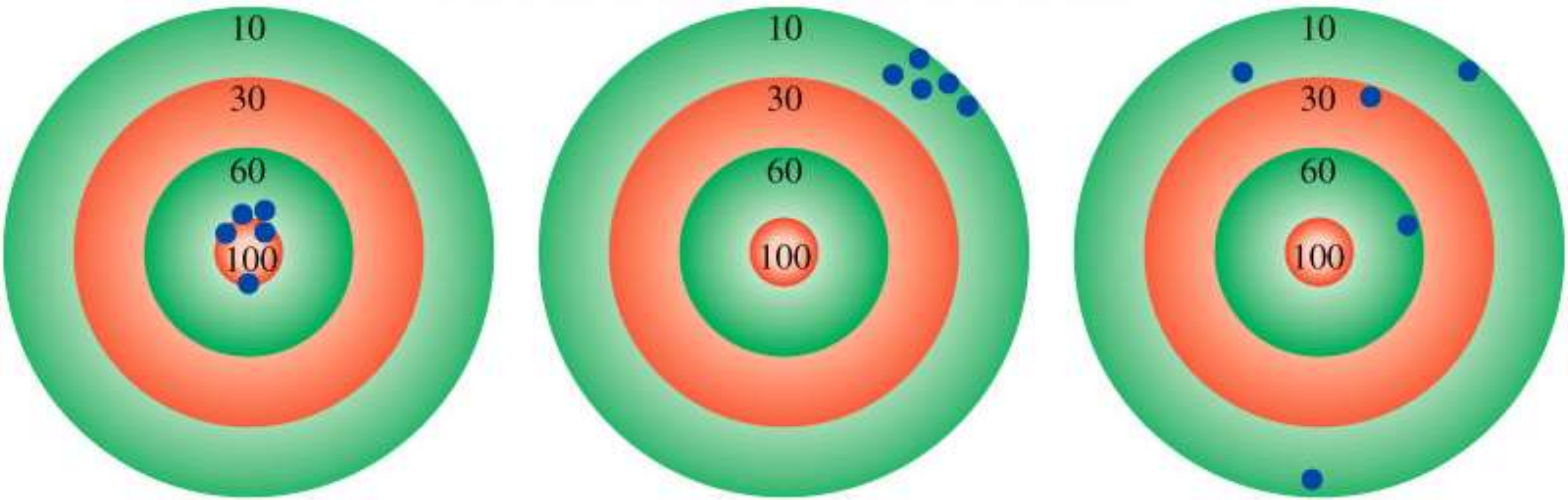
- Ice floating in a solution of sodium chloride in water also consists of two phases, ice and the liquid solution.

- A heterogeneous mixture of talk powder and sugar.

1.5 Measurement and Significant Figures

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

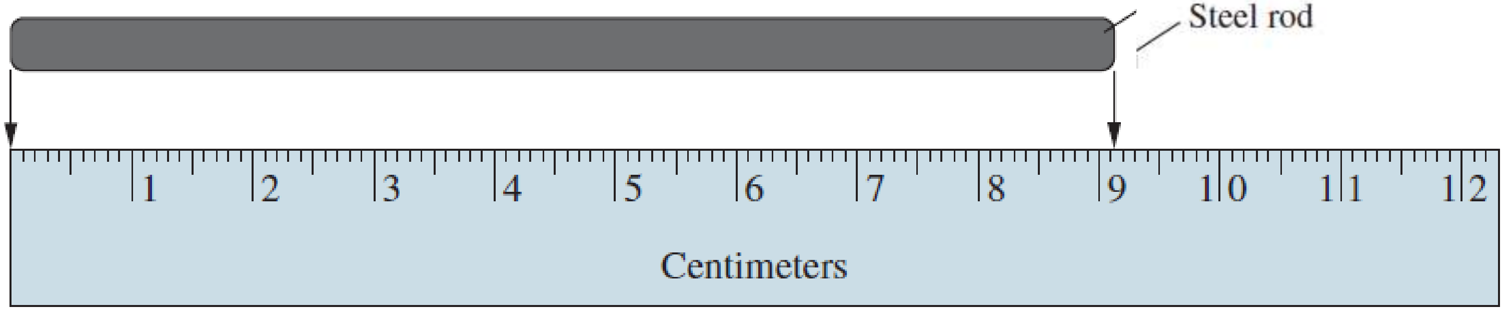
Precision – how close a set of measurements are to each other



accurate
&
precise

precise
but
not accurate

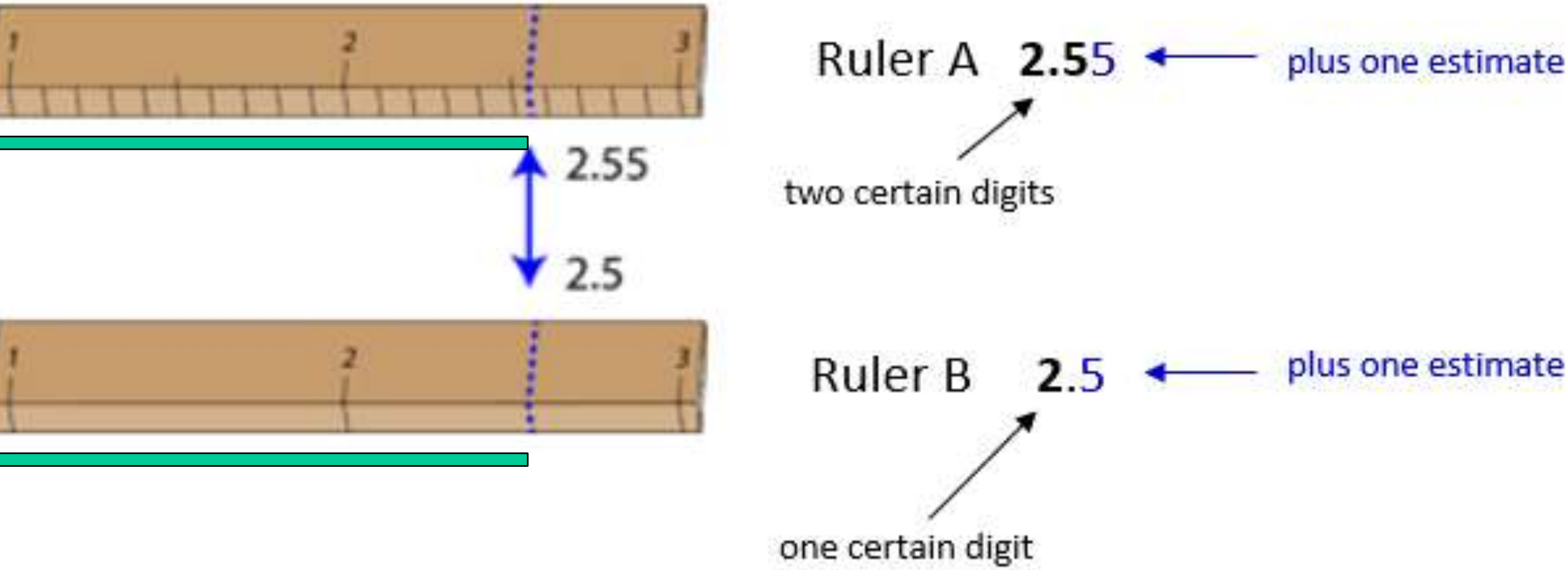
not accurate
&
not precise



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.


1
2
3
2.55
2.5
Ruler A 2.55
plus one estimate
two certain digits
Ruler B 2.5
plus one estimate
one certain digit

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:

602200000000000000000000

6.022×10^{23}

The mass of a single carbon atom in grams:

0.000000000000000000000000199

1.99×10^{-23}

$N \times 10^{n}$

N is non zero
Single digit number
(1-10) 10 is not included

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**

Or We don't know how many significant numbers

e.g., 104956000

(more than one answer)

Q) 10056 ??
521369 ??

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		**More than one answer**
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}

Note the decimal point

→7 sig. figs.

- **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}**

   5430000 is awrong answer

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 = 5.62×10^3**

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

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

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}

1.7029972752043596730245231607629 e^{-4}

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

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

=151.177

3 sig.fig.

=151

Rounding intermediate steps →

= 179 / 1.18

= 151.694

=152

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

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

1.165 A steel sphere has a radius of 1.58 in. If this steel has a density of 7.88 g/cm3, what is the mass of the sphere in grams?

$V = (4/3)\pi r^3$ 1 in = 2.54 cm