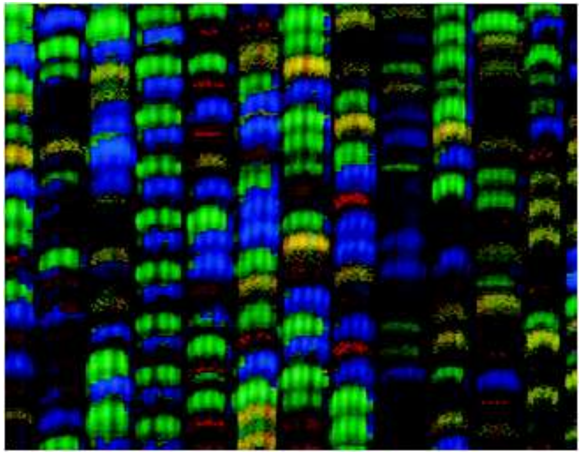
اللهم صلّ وسلّم على نبينا محمد وعلى آله وصحبه أجمعين

Chapter 1

Chemistry and measurements

1.1 Modern Chemistry: A Brief Glimpse

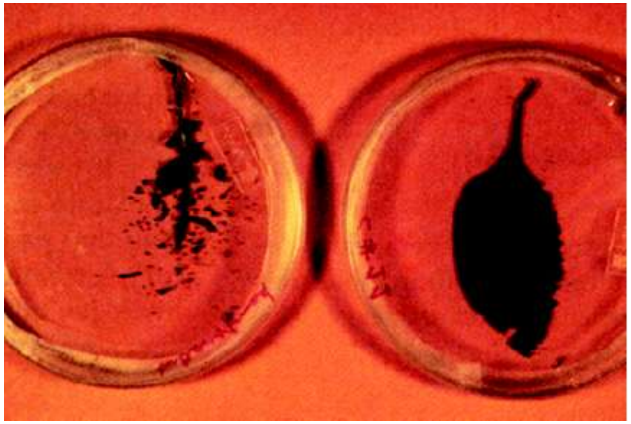
- Health and Medicine
  - Sanitation systems
  - Surgery with anesthesia
  - Vaccines and antibiotics
  - Gene therapy



- 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.*

  ↳ Reasonable / sensible / wise

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

  ↳ Basic

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

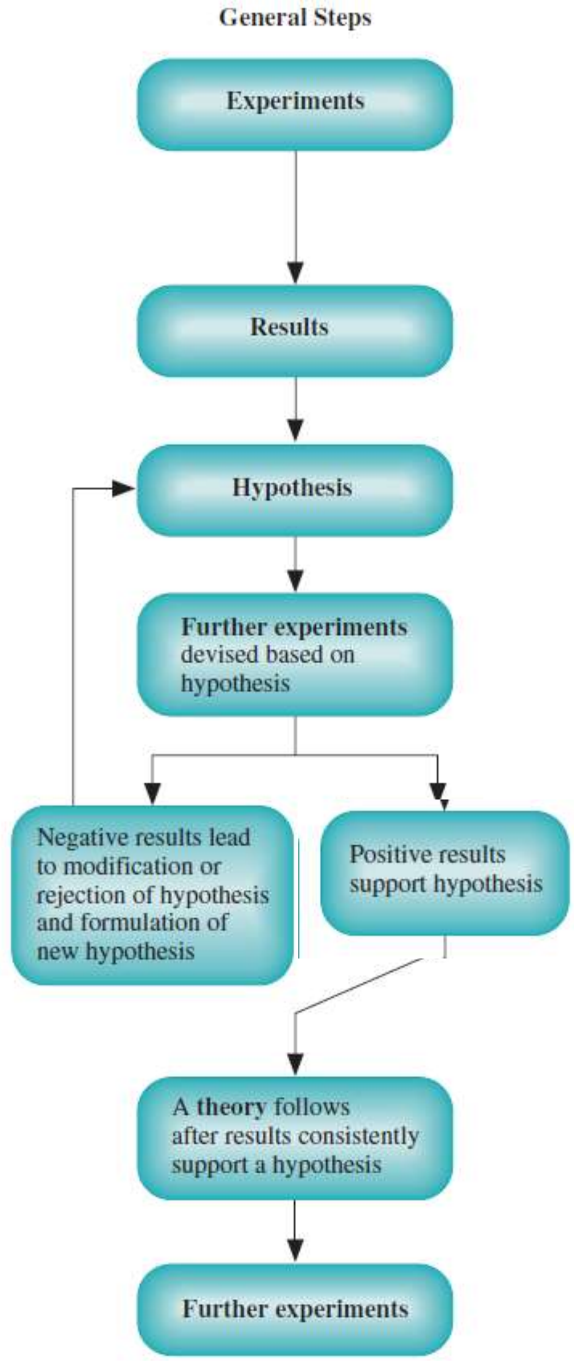
  ↳ مبدئي / أوّلي

- 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**).

Exercise 1.1 You place 1.85 grams of wood in a vessel with 9.45 grams of air and seal the vessel. Then you heat the vessel strongly so that the wood burns. In burning, the wood yields ash and gases. After the experiment, you weigh the ash and find that its mass is 0.28 gram. What is the mass of the gases in the vessel at the end of the experiment?

See Problems 1.37, 1.38, 1.39, and 1.40.

wood = 1.85 grams + Air = 9.45 → Vessel

Vessel —Δ heat→ Wood Burns (while burn) → ash + gases

ash → 0.28g ; gases → ??

** Vessel = 1.85 + 9.45 ⇒ 11.3

→ 11.3 = ash + gases

11.3 = 0.28 + gases ⇒ 11.02 gram

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.

Substance

(1) Solids, Liquids, and Gases:

Resist Motion

- **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

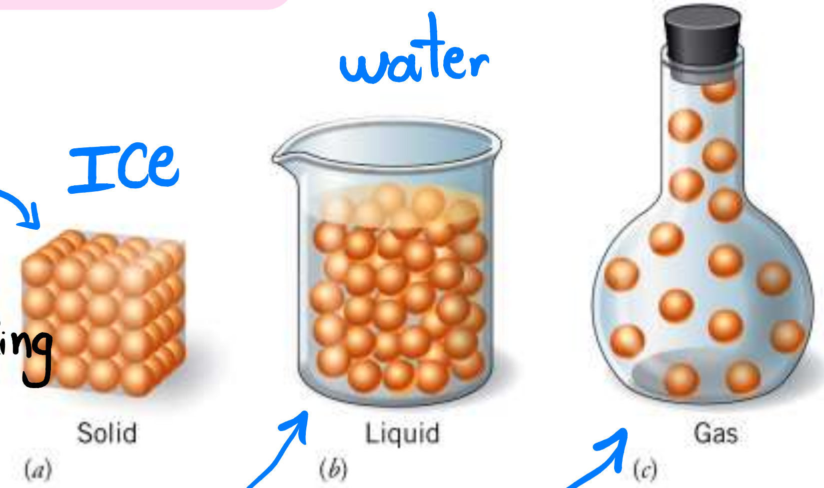
→ Rigid

↓ Close packing

Ice

water

Stream



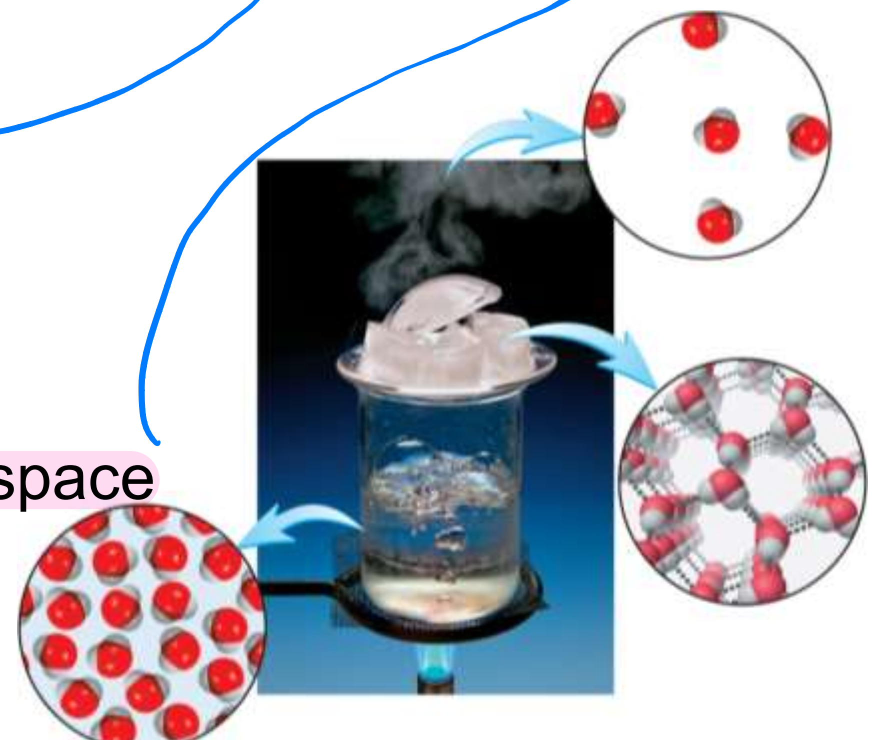
Liquids:

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

Gases: much more less packing

- 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. Solubility
  (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.* ⇒ (Form and break bonds)
  Examples: rust formation, burning butane gas in oxygen
  (Chemical property: Describes how a substance undergoes a chemical reaction) Burn, compositions, rusting / change color

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

Exercise 1.2 Potassium is a soft, silvery-colored metal that melts at 64°C. It reacts vigorously with water, with oxygen, and with chlorine. Identify all of the physical properties given in this description. Identify all of the chemical properties given.

See Problems 1.47, 1.48, 1.49, and 1.50.

Potassium (Soft / Silver-colored / metal)

* melt 64°C

→ React with: Water - Oxygen - Chlorine

Physical properties → Soft - Silver Colored - melt.

Chemical properties → React (water - Oxygen - Chlorine)

metal.

- **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*

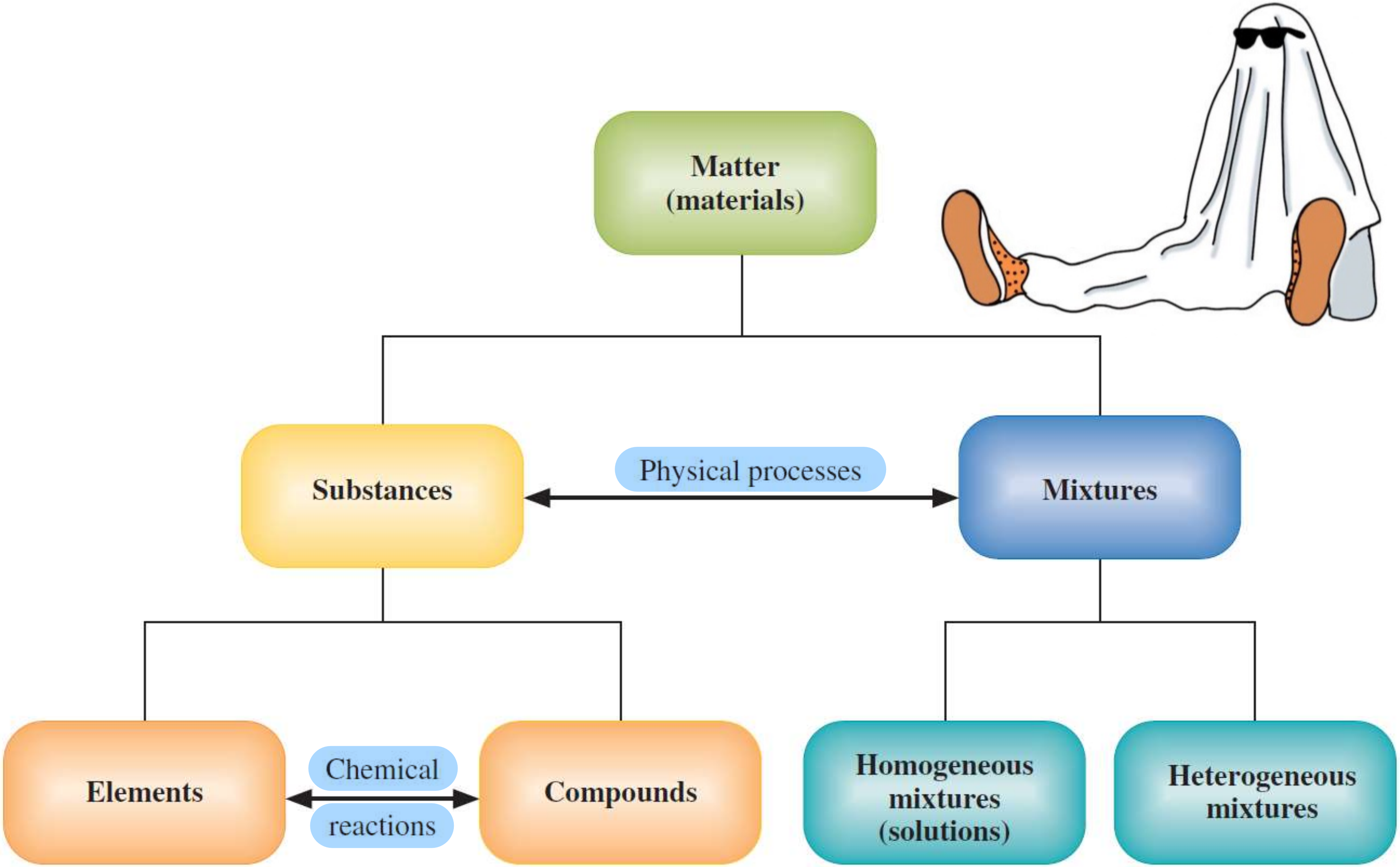
  Fe, Au,Na etc... Can be decomposed during nuclear energy #

  H_2 ,O_2

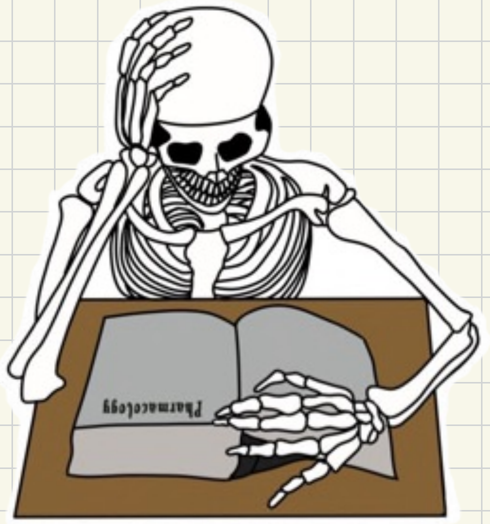
- **A Compound:** is *a substance composed of two or more elements* ***chemically*** *combined.*

  H_2O, NaCl, CO_2 → Solid

  NaCl Solution → homogenous mixture #


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

* 2 or more stustances physically combined ___ mixture
* Element + element chemically combined ___ compound
* Element + element physically combined ___ mixture
* Compound chemical process ___ element + element
* Compound + compound physical process ___ mixture



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

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

→ water + Oil

Examples:

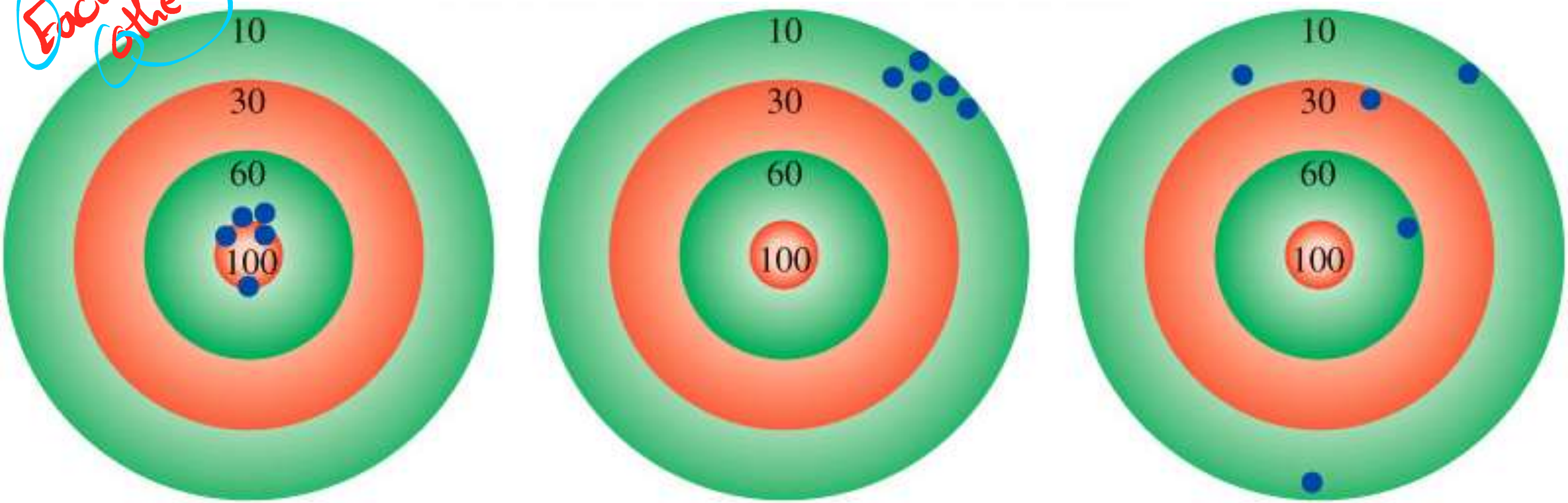
- Ice floating in a solution of sodium chloride in water also consists of two phases, ice (phase) and the liquid solution (phase).

- A heterogeneous mixture of talk powder and sugar.

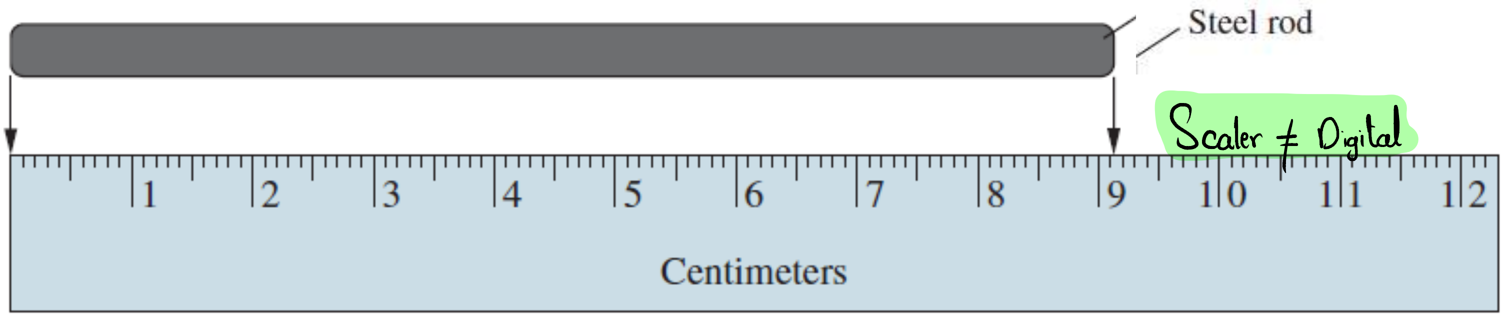
1.5 Measurement and Significant Figures

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

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

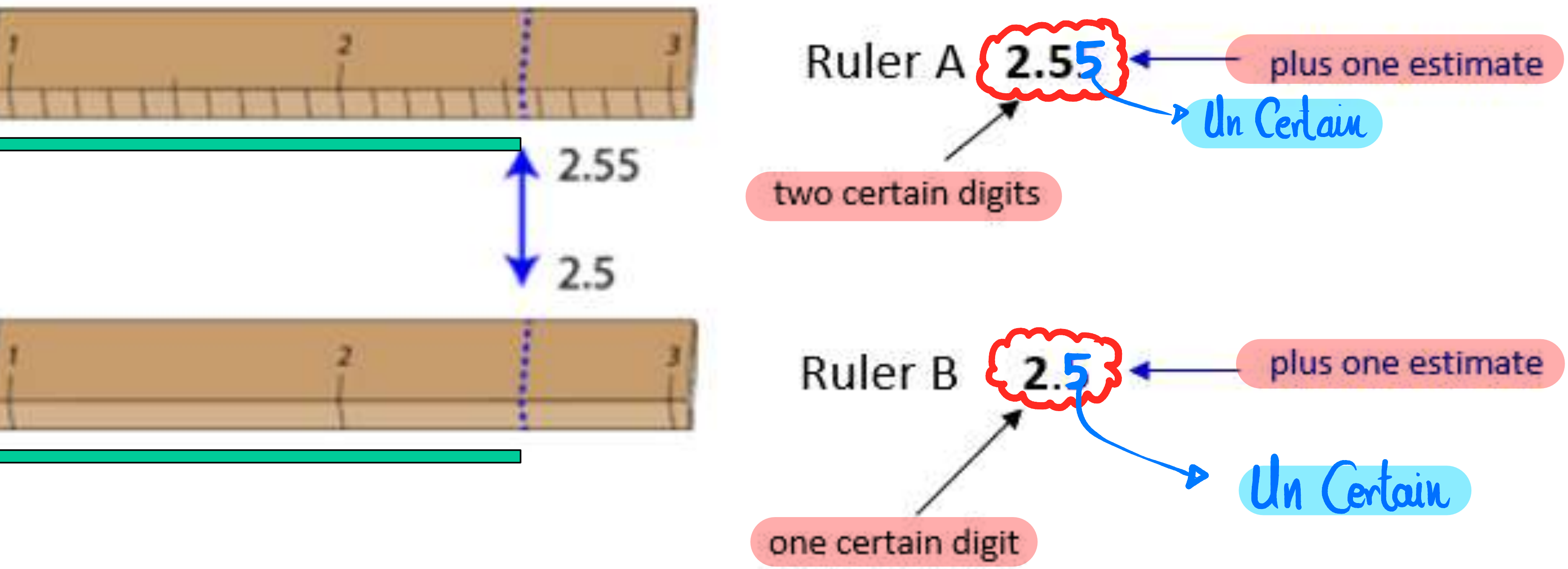


accurate & precise	precise but not accurate	not accurate & not precise



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
Un Certain
two certain digits
Ruler B
2.5
plus one estimate
Un Certain
one certain digit

Number of Significant Figures

9.12 cm → 3 significant figures

9.123 cm → 4 significant figures → Accuracy كلما زاد العدد بعد الفاصلة تزداد الـ

- 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.

3 Sf

e.x: 0.010213 → 5 S.f

To count the number of significant figures, follow the following rules:

JU_MED

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Rules:

1. All digits from (1-9) are significant, no matter where the digit is located

2. الصفر (0)

a) Zeros at the beginning of a number: 0.2, 0.02

Zeros at the beginning of a number are always not significant

0.096 2 sig. fig.

b) Terminal Zeros: 900, 900.0

Number with decimal point:

All terminal zeros are significant

900.0 4 sig. fig.
900. 3 sig. fig.
900.00 5 sig. fig.

Number with no decimal point: 900, 720

We don't know how many significant figures are in the number

لا يمكن الجزم كم significant figure في

900 → maybe 1, 2 or 3 significant figures, you can't decide

823 → 3 significant figures

c) Zeros between non-zero numbers

All these zeros are significant

0.1012 : 4 sig. fig. (not significant → first 0; significant → 0 between 1 and 2)

*If the number is written in scientific notation, then the ten raised to the power n (10^n) is not included in the calculation of the number of significant figures

9×10^2 1 sig. fig. 9.00×10^2 3 sig. fig.
9.0×10^2 2 sig. fig.

*انتبه: نتحدث من الأصفار على اليمين وليس عن جميع الأرقام

Scientific Notation

The number of atoms in 12 g of carbon:

you can't know How many S.F Because there is no (decimal)

(1 Mole) = Avogadro Number = 602200000000000000000000

Scientific Notation: → 6.022×10^{23}

The mass of a single carbon atom in grams:

Zeros at the beginning of the number aren't Significant

0.0000000000000000000000199

Scientific Notation: → 1.99×10^{-23}

× → 11×10^{2}

✓ → 1.1×10^{3} (decimal)

$N \times 10^{n}$

N is non zero Single digit number (1-10) 10 is not included (1-9)

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)

   Scientific Notation:

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

   To the Right

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

S.f

1. $5. \times 10^2$ = 1
2. 5.0×10^2 = 2
3. 5.00×10^2 = 3

Rules for Significant Figures

4. Final zeros on number "without decimal point" are
→ **Not significant**

Or We don't know how many significant numbers

Zero's between non- zero's are significant

e.g., 104956000 = Unknown

(more than one answer)

Q) 10056 ?? 5

521369 ?? 6

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		**More than one answer**
5.	3600.	3.600×10^{3}	**4**

Number of s.f should be the same before the scientific notation and after

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

N is a single Non-zero digit

n is a positive or negative integer

أدعلك تَشيلي الأصفار

Q) How many significant figures are in 19.0000? (6)

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** "We only round on the last step"

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 Scientific Notation ← ضروري	**37.5 or 3.75×10^1**
2. 5431978 5.431978×10^6	**5.43×10^6** 543 ✗ wrong answer **5430000 is awrong answer**
3. 132.7789003	**133 or 1.33×10^2**
4. 0.00087564 8.7564×10^{-4}	**8.76×10^{-4}** Scientific notation. More easy

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

We should do the scientific notation first

A. 5.64×10^{-2}

B. 5.000×10^{-3}

C. 5.645×10^{-4}

D. 0.56446

E. 5.645×10^{-3}

*Round to two significant figures:

5.245 = 5.2

*Round to four significant figures:

5.239921 = 5.240

Significant Figures in Calculations

#

Multiplication and Division

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

Scientific Notation يفضل تحويلها أو لأشر لـ #

e.g., $10.54 \times 31.4 \times 16.987$ **= 5621.9 =** $\boxed{\mathbf{5.62 \times 10^3}}$

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

1) ⇒ 5621.94 → 2) 5.62194×10^3 → 3) 5.62×10^3

e.g., $5.896 \div 0.008$ **= 737 =** $\boxed{\mathbf{7 \times 10^2}}$ = 7.37×10^2 ✗

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

⇒ 7 → 1 SF

Scientific Notation يفضل تحويلها أو لأشر لـ

:) Smart question.

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

لما يكون السؤال
(× − ÷) لحاله عادي
أطلع الجواب النهائي على
طول بعين حدي كم
S.F لازم يكون
في

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
336.8733 →	**336.9** / 3.369×10^2	1 decimal place

ما تنسى الـ Rounding

e.g.,	397	0 decimal places
	– 273.15	2 decimal places
123.85 →	**124** / 1.24×10^2	0 decimal place

ممنوع أحط لـ Scientific Notation قبل ما أطلع الجواب
ـه بعد عادي بحس =
وممكن تخلي زي ما هو

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

11.3 = 11.273 → 1.1273×10^1

1.13×10^1

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**

$0.66272803 = 6.6272803 \times 10^{-1}$

6.627×10^{-1}

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 → 4.241×10^2

D. 424.064

E. 424

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

1. $$\frac{(71.359\ \text{m} - 71.357\ \text{m})}{(3.2\ \text{s} \times 3.67\ \text{s})} = \frac{(0.002\ \text{m})}{(11.744\ \text{s}^2)}$$

$$= (0.002/12) = (1.666 \times 10^{-4}) = 2 \times 10^{-4}\ \text{m/s}^2$$

Rounding بنعمل بالخطوة الأخيرة بس

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

$$= \frac{(20.818665\ \text{cm}^3)}{(2387.1\ \text{s})} = (21/2387) = 0.0088\ \text{cm}^3/\text{s}$$

Or 8.8×10^{-3}

* Question :-

1)

$$\frac{(71.359 - 71.357)}{(3.2 \times 3.67)} = \frac{.002}{11.744} = .0001702\ldots \rightarrow 1.702 \times 10^{-4} = 2 \times 10^{-4}$$

1-sf ✓

2-sf, 3-sf

2-sf

2)

$$\frac{(13.674 \times 4.35 \times .35)}{(856 + 1531.1)} = \frac{20.818665}{2387.1} = .00872123\ldots$$

5 × 3 × 2 → 2-sf

2sf

0-dp, 1dp = 0-dp

4sf

2-sf ↓

8.72123×10^{-3}

Scientific Notation

8.7×10^{-3}

Types

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
- Assume they have infinite number of significant figures

كلما زاد ال S.f ← يزداد ال Certainty

- The number of significant figures in a calculation result **depends only on** the numbers of significant figures in **quantities having uncertainties**

Explanation: Exact Number:

1. $\dfrac{6.00\ kg}{2} = \boxed{3}$ → (1) Significant figure

6.00 kg: Counted _ not exact → (3) Significant figure

2 → (1) Significant figure

NoT exact Number = NoT Counted

2. $\dfrac{6.00\ kg}{2} = \boxed{3.00}$ → (3) Significant figure

6.00 kg: Counted _ not exact → (3) Significant figure

2 → (∞) Significant figure

exact Number = Counted

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 ?

Counted ← $3.0\ g \times 9 = 27\ g$ → Un-Certain

$2 \times \infty = 2$

The number 9 is exact and does not determine the number of significant figures → ∞ Significant figure

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

$$36.00\ \text{in} \times \left(\frac{\text{ft.}}{12\ \text{in.}}\right) = 4 \times \infty = 4$$

4 - Significant figure

∞ exact Number

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

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

=151.177

3 sig.fig.

=151

Rounding "intermediate steps" →

= 179 / 1.18

= 151.694

=152

Note: Do not round intermediate answers !

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Example 1.2:

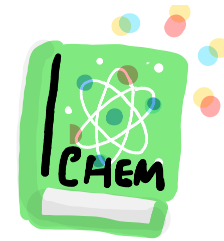
Perform the following calculations and round the answers to the correct number of significant figures (units of measurement have been omitted).

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

Exercise 1.3 Give answers to the following arithmetic setups. Round to the correct number of significant figures.

a. $\frac{5.61 \times 7.891}{9.1}$ b. $8.91 - 6.435$

c. $6.81 - 6.730$ d. $38.91 \times (6.81 - 6.730)$

See Problems 1.61 and 1.62.

A) $\frac{5.61 \times 7.891}{9.1} = \frac{3\text{ Sf} \times 4\text{ S.f}}{2\text{ S.f}} = \frac{3\text{ Sf}}{2\text{ Sf}} = (2)\text{ Sf}$

$\frac{44.26851}{9.1} = 4.86467143 \Rightarrow$ (4.9)

B) $8.91 - 6.435 = 2.475 \Rightarrow$ (2.48)

2-dp − 3-dp = 2-dp

C) $6.81 - 6.730 = 0.08 \Rightarrow$ (0.08)

2dp − 3dp = 2dp

---

d) $38.91 \times (\ C = 0.08\ ) = 3.1128 \Rightarrow$ (3)

4-Sf 1-Sf = 1Sf

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

* radius = 1.58 in
* Density = 7.88 g / cm^3
* Mass = ??

r = in → cm exact

Un Certain

$1.58 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}}$

3-SF × ∞

4.0132 cm → 4.01 cm

→ 3-SF

$D = \frac{M}{V}$

$7.88 = \frac{M}{(\frac{4}{3})(\pi)(4.0132)^3 cm^3}$

$M = 7.88 \times \frac{4}{3} \times \frac{22}{7} \times (64.6356)$

3 3

$509.329 \times \frac{4}{3} \times \frac{22}{7}$

3

509.329 × ---- × ---- =

2134,331 = 2.13×10^3

Done by: Joud Jaber

Thank You