Chapter 1
Chemistry: The Central Science

CHEM 101

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What IS Chemistry?

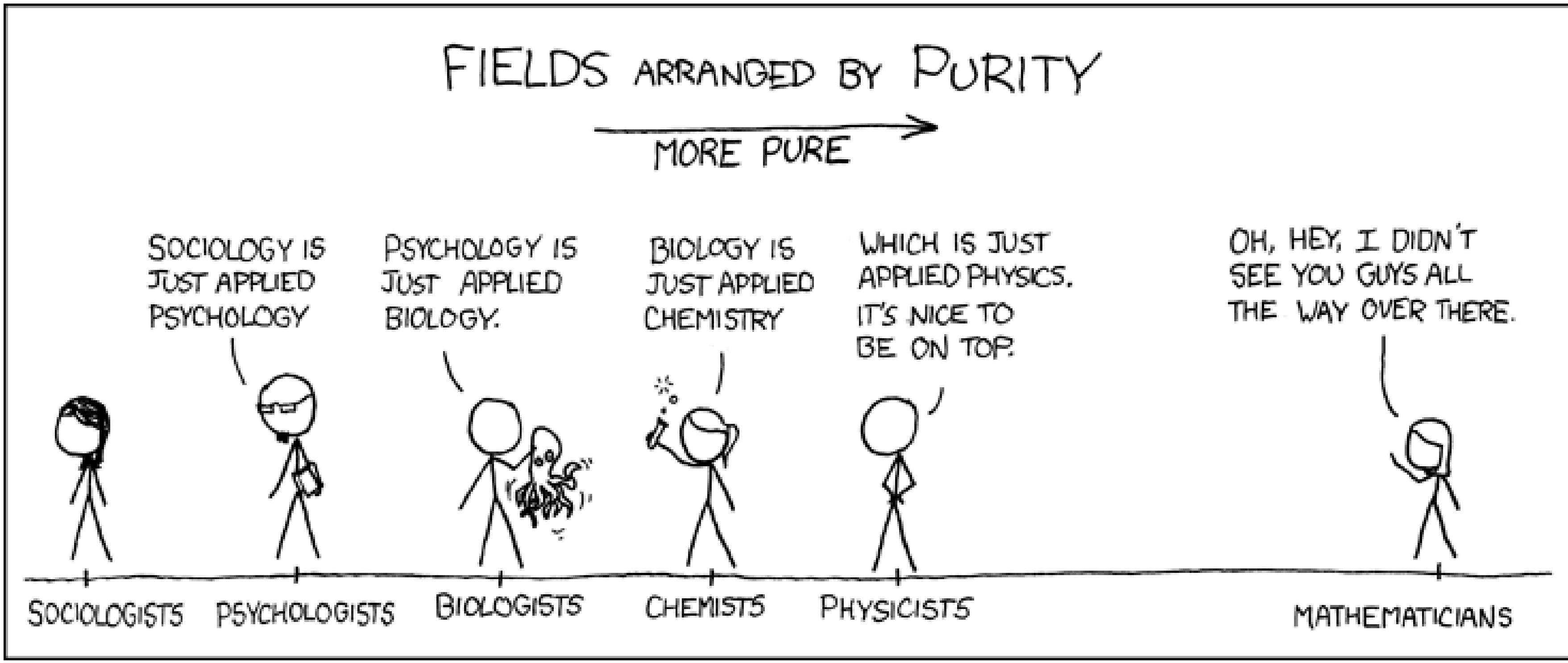
- Text: “The study of matter and the changes that matter undergoes”
- Focus: how matter interacts at the atomic/molecular level

Why "The Central Science"?

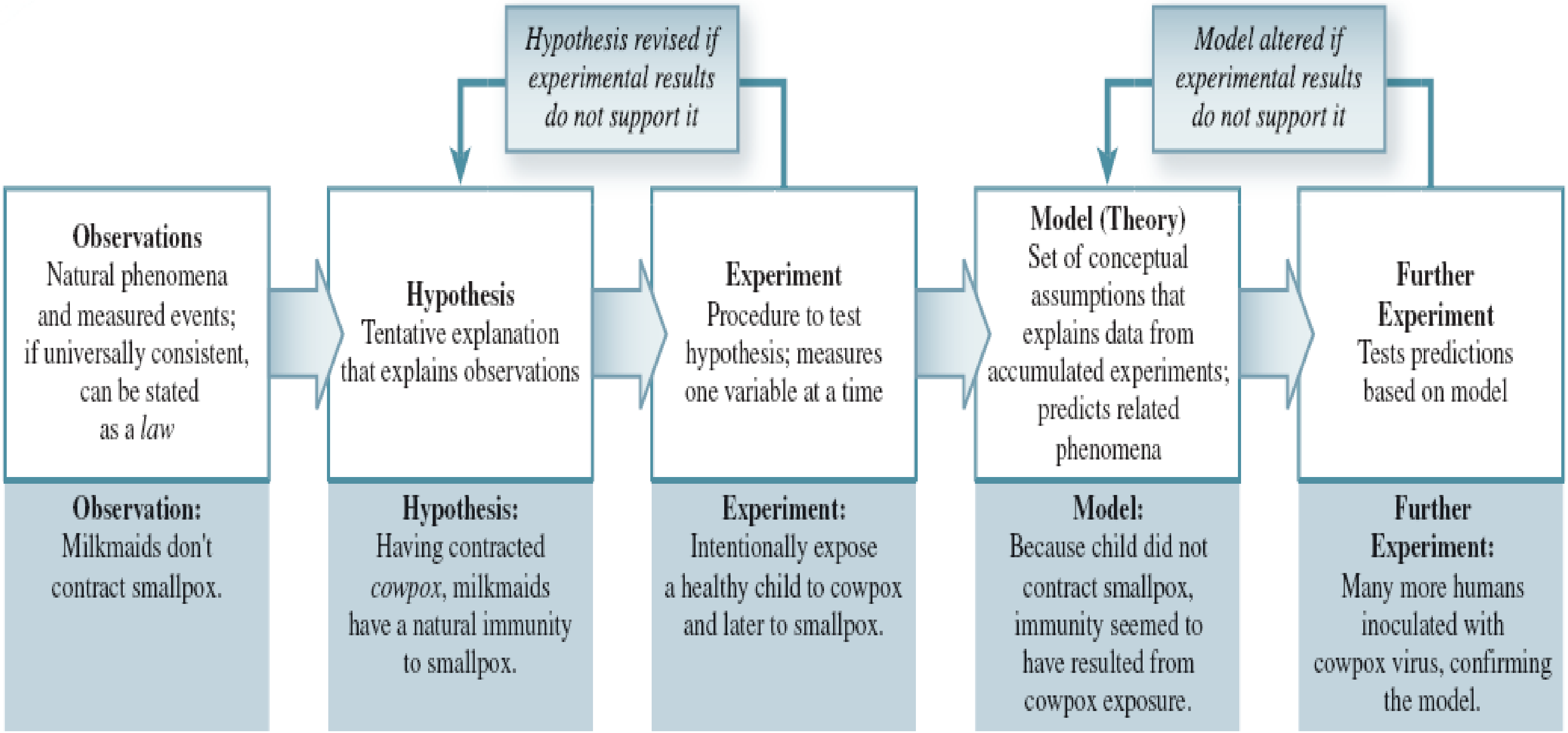
Other scientific disciplines require an understanding of how matter behaves. Some examples:

Physics: electronic devices require an understanding of materials called semiconductors

Biology→Biochemistry→Organic Chemistry (the study of carbon-containing compounds)


FIELDS ARRANGED BY PURITY
MORE PURE
SOCIOLOGY IS JUST APPLIED PSYCHOLOGY
PSYCHOLOGY IS JUST APPLIED BIOLOGY.
BIOLOGY IS JUST APPLIED CHEMISTRY
WHICH IS JUST APPLIED PHYSICS. IT'S NICE TO BE ON TOP.
OH, HEY, I DIDN'T SEE YOU GUYS ALL THE WAY OVER THERE.
SOCIOLOGISTS
PSYCHOLOGISTS
BIOLOGISTS
CHEMISTS
PHYSICISTS
MATHEMATICIANS

1.2 The Scientific Method



1.3 Classification of Matter

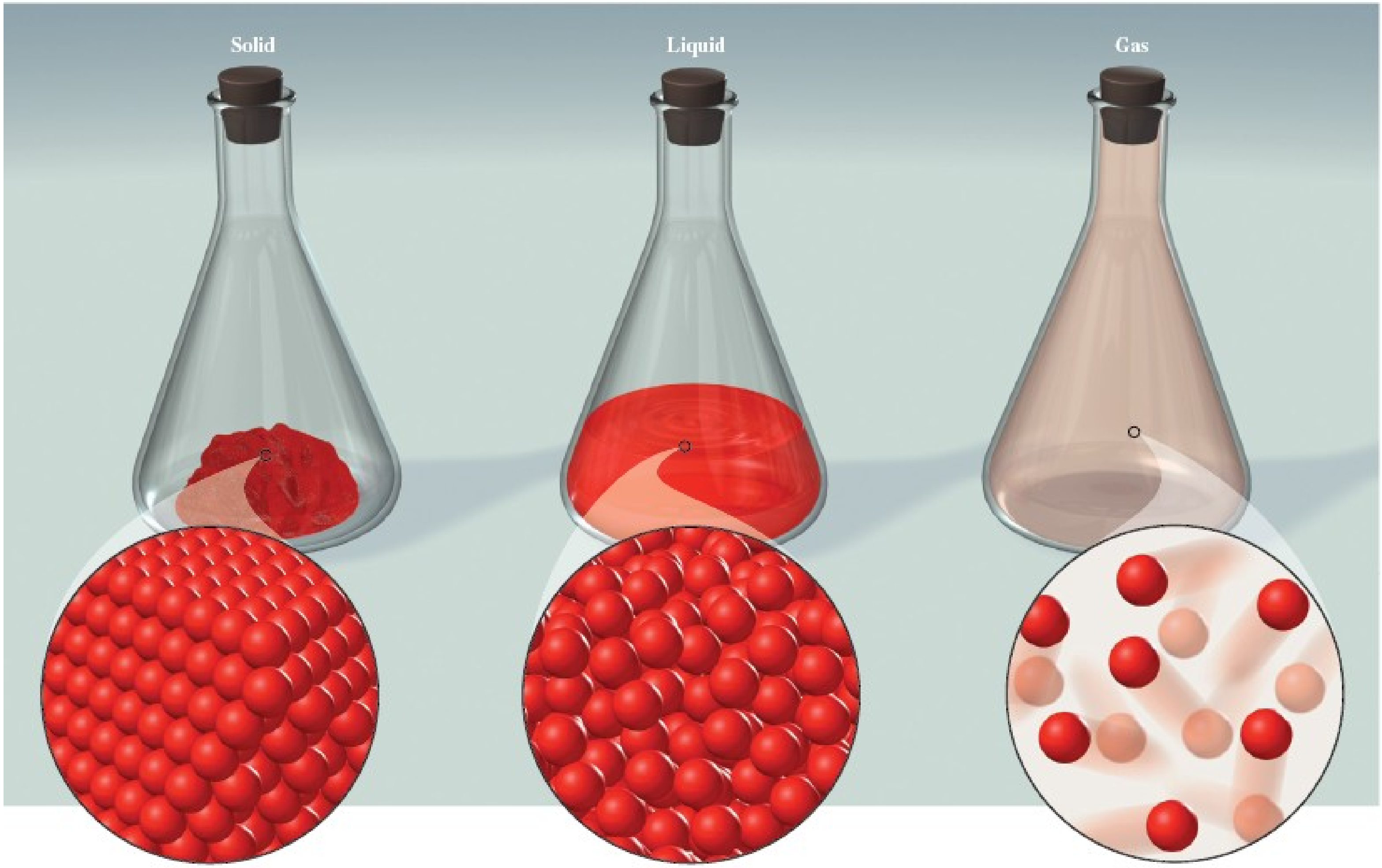
- Matter is either classified as a *substance* or a *mixture* of substances.
- ***Substance***
  - Can be either an ***element*** or a ***compound***
  - Has a definite (constant) composition and distinct properties
  - Examples: sodium chloride, water, oxygen

- States of Matter
  - Solid
    - particles close together in orderly fashion
    - little freedom of motion
    - a solid sample does not conform to the shape of its container
  - Liquid
    - particles close together but not held rigidly in position
    - particles are free to move past one another
    - a liquid sample conforms to the shape of the part of the container it fills

– Gas
  - particles randomly spread apart
  - particles have complete freedom of movement
  - a gas sample assumes both shape and *volume* of container.

– States of matter can be inter-converted without changing chemical composition
  - solid $\rightarrow$ liquid $\rightarrow$ gas (add heat)
  - gas $\rightarrow$ liquid $\rightarrow$ solid (remove heat)

States of Matter



Physical vs. Chemical Change

- In chemistry, we're chiefly concerned with how atoms bond with each other.

- A chemical change involves making or breaking chemical bonds to create new substances.
- A physical change alters a substance without changing its chemical identity.

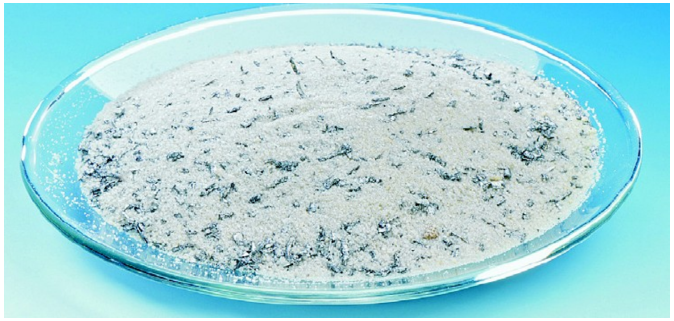
  -e.g. crushing, melting, boiling.

Substances

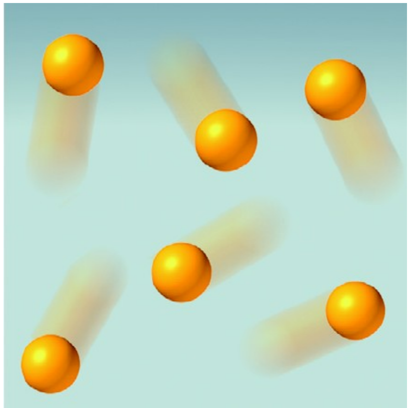
- ***Element***: cannot be separated into simpler substances by chemical means.
  - Examples: iron, mercury, oxygen, and hydrogen
- ***Compounds***: two or more elements chemically combined in definite ratios
  - Cannot be separated by physical means
  - Examples: salt, water and carbon dioxide

Mixtures

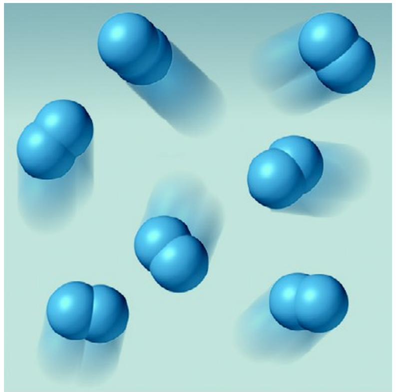
- ***Mixture:*** physical combination of two or more substances
  - Substances retain distinct identities
  - No universal constant composition
  - Can be separated by physical means
    - Examples: sugar/iron; sugar/water



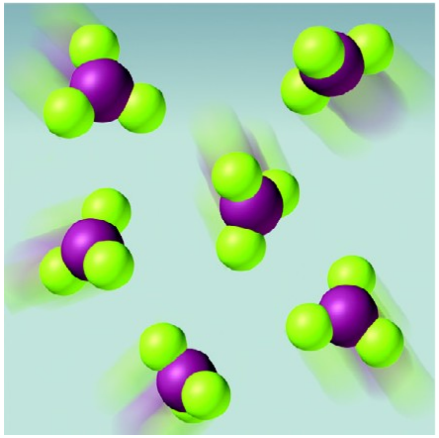
Molecular Comparison of Substances and Mixtures



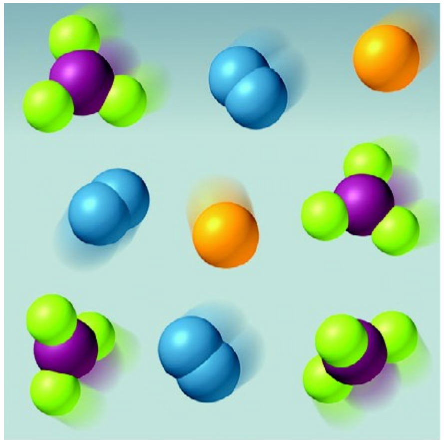
Atoms of an element



Molecules of an element



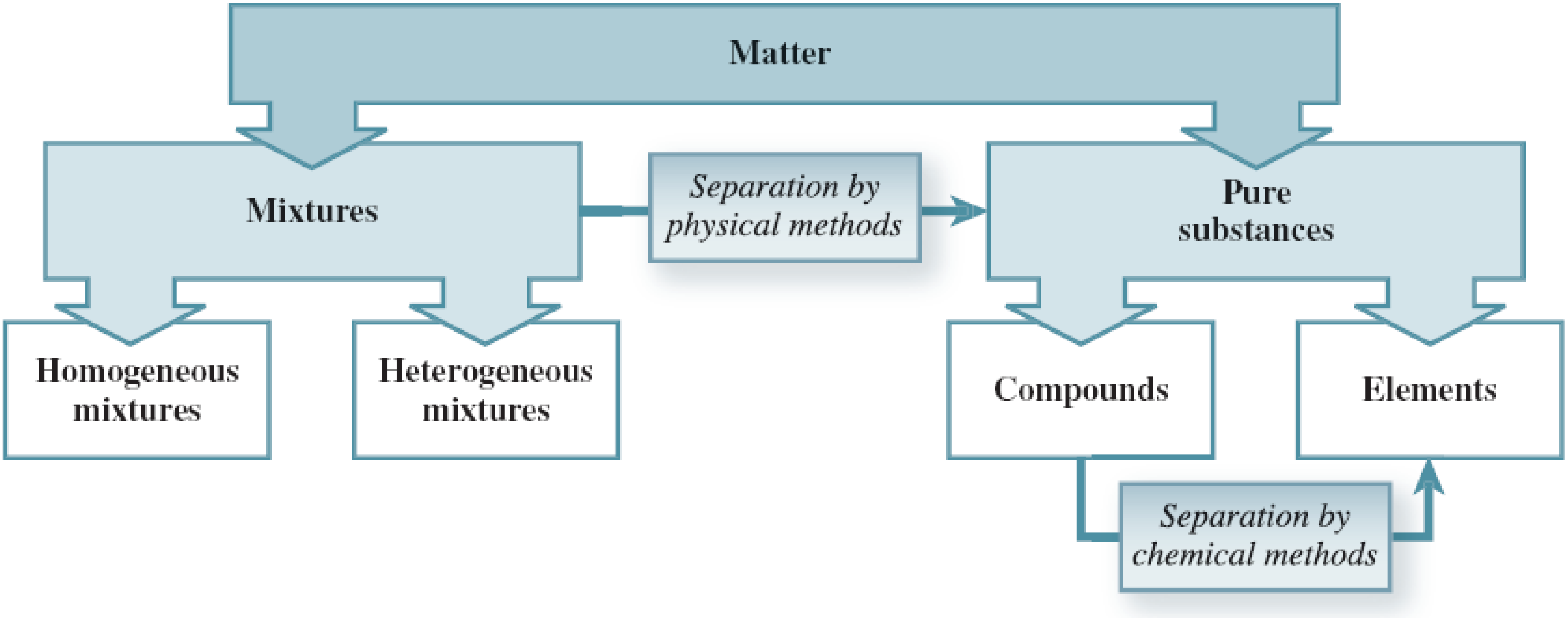
Molecules of a compound



Mixture of two elements
and a compound

- Types of Mixtures
  - ***Homogeneous***: composition of the mixture is *uniform* throughout
    - Example: sugar dissolved in water
  - ***Heterogeneous***: composition is *not* uniform throughout
    - Example: sugar mixed with iron filings

Classification of Matter



Classify the following

Aluminum foil

Baking soda

Milk

Air

Copper wire

Aluminum foil:

substance, element

Baking soda:

substance, compound

Milk:

mixture, homogeneous

Air:

mixture, homogeneous

Copper wire:

substance, element

1.3 Scientific Measurement

- Used to measure quantitative properties of matter
- SI base units

TABLE 1.2 Base SI Units

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

SI Prefixes

TABLE 1.3 Prefixes Used with SI Units

Prefix	Symbol	Meaning	Example
Tera-	T	1×10^{12} (1,000,000,000,000)	1 teragram (Tg) = 1×10^{12} g
Giga-	G	1×10^{9} (1,000,000,000)	1 gigawatt (GW) = 1×10^{9} W
Mega-	M	1×10^{6} (1,000,000)	1 megahertz (MHz) = 1×10^{6} Hz
Kilo-	k	1×10^{3} (1,000)	1 kilometer (km) = 1×10^{3} m
Deci-	d	1×10^{-1} (0.1)	1 deciliter (dL) = 1×10^{-1} L
Centi-	c	1×10^{-2} (0.01)	1 centimeter (cm) = 1×10^{-2} m
Milli-	m	1×10^{-3} (0.001)	1 millimeter (mm) = 1×10^{-3} m
Micro-	μ	1×10^{-6} (0.000001)	1 microliter (μL) = 1×10^{-6} L
Nano-	n	1×10^{-9} (0.000000001)	1 nanosecond (ns) = 1×10^{-9} s
Pico-	p	1×10^{-12} (0.000000000001)	1 picogram (pg) = 1×10^{-12} g

- *Mass*: measure of the *amount of matter*
  - (*weight* refers to gravitational pull)
- Temperature:
  - Celsius
    - Represented by °C
    - Based on freezing point of water as 0°C and boiling point of water as 100°C
  - Kelvin
    - Represented by K (no degree sign)
    - The *absolute* scale
    - Units of Celsius and Kelvin are equal in magnitude
  - Fahrenheit (the English system) (°F)

Equations for Temperature Conversions

$$^{o}C = (^{o}F - 32) \times \frac{5}{9}$$

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

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

Temperature Conversions

A clock on a local bank reported a temperature reading of 28^{o}C. What is this temperature on the Kelvin scale?

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

$$K = 28\ {}^{o}C + 273.15 = 301K$$

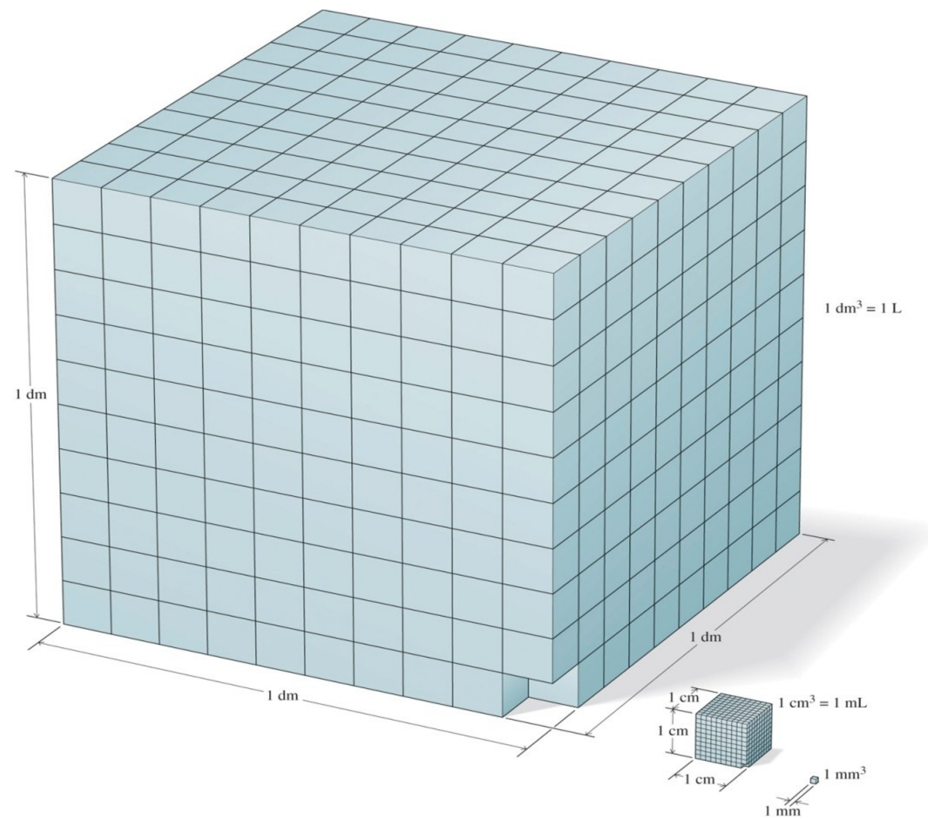
Practice

Convert the temperature reading on the local bank (28°C) into the corresponding Fahrenheit temperature.

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

$$^{o}\,\mathrm{F} = \frac{9}{5} \times 28\ ^{o}\mathrm{C} + 32 = 82\ ^{o}\mathrm{F}$$

- Volume: meter cubed (m^3)
  - Derived unit
  - The unit liter (L) is more commonly used in the laboratory setting. It is equal to a decimeter cubed (dm^3).



- Density: Ratio of mass to volume
  - Formula:

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

  - d = density (g/mL)
  - m = mass (g)
  - V = volume (mL or cm^3)

(*gas densities are usually expressed in g/L)

Practice

The density of a piece of copper wire is 8.96 g/cm^3. Calculate the volume in cm^3 of a piece of copper with a mass of 4.28 g.

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

$$V = \frac{m}{d} = \frac{4.28\ \text{g}}{8.96\ \dfrac{\text{g}}{\text{cm}^3}} = 0.478\ \text{cm}^3$$

1.4 Properties of Matter

- ***Quantitative***: expressed using *numbers*
- ***Qualitative***: expressed using *properties*
- ***Physical properties***: can be observed and measured without changing the substance
  - Examples: color, melting point, states of matter
- ***Physical changes***: the identity of the substance stays the same
  - Examples: changes of state (melting, freezing)

- ***Chemical properties***: must be determined by the chemical *changes* that are observed
  - Examples: flammability, acidity, corrosiveness, reactivity
- ***Chemical changes***: after a chemical change, the original substance no longer exists
  - Examples: combustion, digestion

- ***Extensive property*: depends on amount of matter**
  – Examples: mass, length

- ***Intensive property*: does not depend on amount**
  – Examples: density, temperature, color

1.5 Uncertainty in Measurement

- Exact: numbers with defined values
  - Examples: counting numbers, conversion factors based on definitions
- Inexact: numbers obtained by any method other than counting
  - Examples: measured values in the laboratory

- Significant figures
  - Used to express the uncertainty of inexact numbers obtained by measurement
  - The last digit in a measured value is an uncertain digit - an estimate

- Guidelines for significant figures
  - Any non-zero digit is significant
  - Zeros between non-zero digits are significant
  - Zeros to the left of the first non-zero digit are not significant
  - Zeros to the right of the last non-zero digit are significant if decimal is present
  - Zeros to the right of the last non-zero digit are not significant if decimal is not present

Simplified rule:

If you can’t write the number without the zeros, they are space-fillers and not significant.

e.g. can’t write “eleven hundred” without the zeros…1100 = 2 sigfigs.
But 1100.0 = 5 sigfigs…there has to be a reason for all those zeros.

0.0012 is 2 sigfigs but 0.001200 is 4 sigfigs

Practice

Determine the number of significant figures in each of the following.

345.5 cm

4 significant figures

0.0058 g

2 significant figures

1205 m

4 significant figures

250 mL

2 significant figures

250.00 mL

5 significant figures

- Calculations with measured numbers
  - Addition and subtraction
    - Answer cannot have more digits to the right of the decimal than any of original numbers
    - Example:

$$\begin{array}{rl} 102.50 & \text{two digits after decimal point} \\ \underline{+\quad 0.231} & \text{three digits after decimal point} \\ 102.731 & \text{round to } 102.73 \end{array}$$

- Multiplication and division
  - Final answer contains the smallest number of significant figures
  - Example:

$$1.4 \times 8.011 = 11.2154 \longrightarrow \text{round to } 11$$

(Limited by 1.4 to <u>two</u> significant figures in answer)

- Exact numbers
  – Do not limit answer because exact numbers have an infinite number of significant figures
  – Example:

    A penny minted after 1982 has a mass of 2.5 g. If we have three such pennies, the total mass is

    $$3 \text{ x } 2.5 \text{ g} = 7.5 \text{ g}$$

  – In this case, 3 is an exact number and does not limit the number of significant figures in the result.

- Multiple step calculations
  - It is best to retain at least one extra digit until the end of the calculation to minimize rounding error.
- Rounding rules
  - If the number is less than 5 round “down”.
  - If the number is 5 or greater round “up”.

Practice

105.5 L + 10.65 L = 11

Calculator answer: 116.15 L

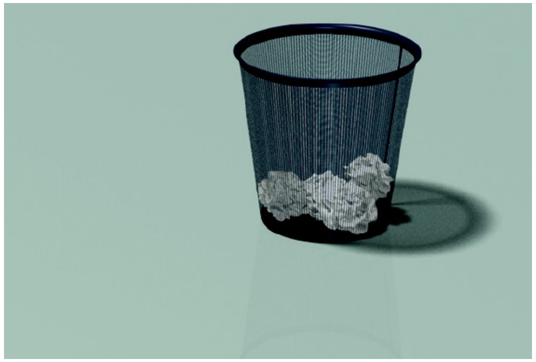
Round to: 116.2 L ⟵ Answer to the tenth position

1.0267 cm x 2.508 cm x 12.599 cm = 32.44

Calculator answer: 32.4419664 cm^3

Round to: 32.44 cm^3 ⟵ round to the smallest number of significant figures

- Accuracy and precision
  - Two ways to gauge the quality of a set of measured numbers
  - ***Accuracy***: how close a measurement is to the true or accepted value
  - ***Precision***: how closely measurements of the same thing are to one another



both accurate and precise

not accurate but precise

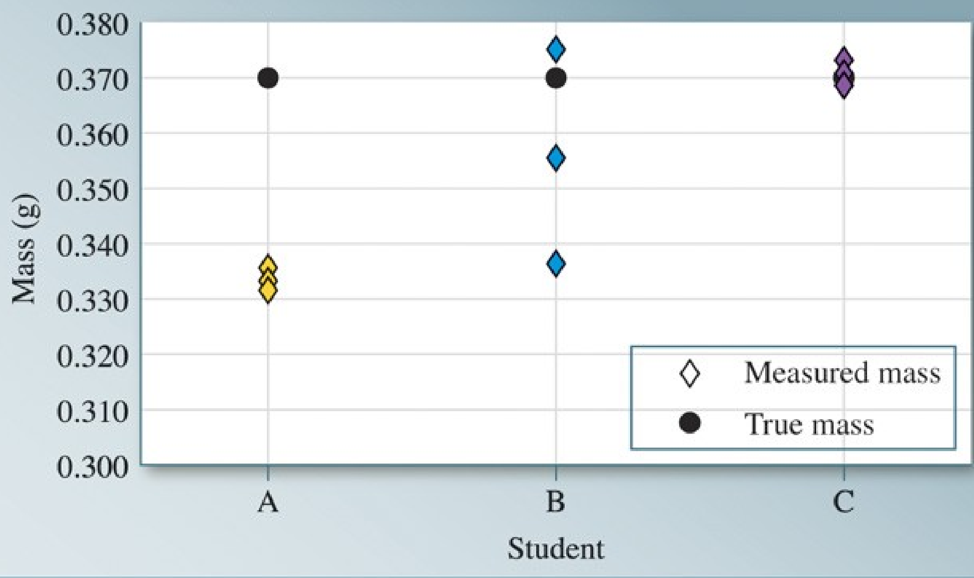
neither accurate nor precise

Describe accuracy and precision for each set

Student A	Student B	Student C
0.335 g	0.357 g	0.369 g
0.331 g	0.375 g	0.373 g
0.333 g	0.338 g	0.371 g
Average:		
0.333 g	0.357 g	0.371 g

- True mass is 0.370 grams

	Student A	Student B	Student C
Measurement 1	0.335 g	0.357 g	0.369 g
Measurement 2	0.331 g	0.375 g	0.373 g
Measurement 3	0.333 g	0.338 g	0.371 g



Student A’s results are precise but not accurate.

Student B’s results are neither precise nor accurate.

Student C’s results are both precise and accurate.

1.6 Using Units and Solving Problems

- ***Conversion factor*:** a fraction in which the same quantity is expressed one way in the numerator and another way in the denominator
  - Example: by definition, 1 inch = 2.54 cm

$$\frac{1\,\text{in}}{2.54\,\text{cm}} \qquad \frac{2.54\,\text{cm}}{1\,\text{in}}$$

- ***Dimensional analysis***: a problem solving method employing conversion factors to change one measure to another often called the "factor-label method"
  - Example: Convert 12.00 inches to meters
    - Conversion factors needed:

    2.54 cm = 1 in and 100 cm = 1 meter

$$12.00\,\cancel{\text{in}} \times \frac{2.54\,\cancel{\text{cm}}}{1\,\cancel{\text{in}}} \times \frac{1\,\text{m}}{100\,\cancel{\text{cm}}} = 0.3048\,\text{m}$$

*Note that neither conversion factor limited the number of significant figures in the result because they both consist of exact numbers.

Notes on Problem Solving

- Read carefully; find information given and what is asked for
- Find appropriate equations, constants, conversion factors
- Check for sign, units and significant figures
- Check for reasonable answer

Practice

The Food and Drug Administration (FDA) recommends that dietary sodium intake be no more than 2400 mg per day. What is this mass in pounds (lb), if 1 lb = 453.6 g?

$$2400\ \cancel{\text{mg}} \times \frac{1\ \cancel{\text{g}}}{1000\ \cancel{\text{mg}}} \times \frac{1\ \text{lb}}{453.6\ \cancel{\text{g}}} = 5.3 \times 10^{-3}\ \text{lb}$$

Key Points

- Scientific method
- Classifying matter
- SI conversions
- Density
- Temperature conversions
- Physical vs chemical properties and changes
- Precision vs accuracy
- Dimensional analysis