Chapter 1: Essential Ideas

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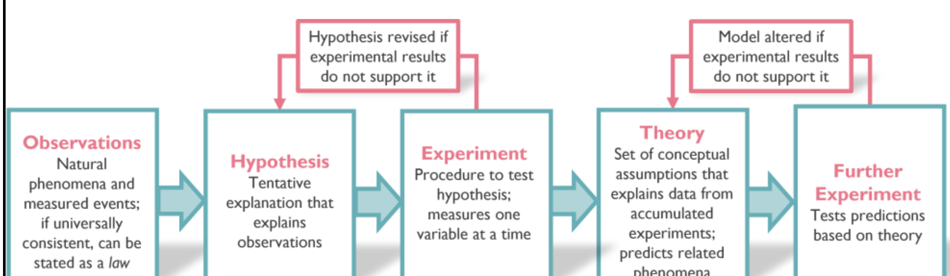
The Study of Chemistry

Chemistry is the study of matter and the changes that matter undergoes.
Matter is anything that has mass and occupies space.



Scientists follow a set of guidelines known as the ***scientific method***:

- Make observations
- Formulate a hypothesis
- Experiment to test hypothesis
  - may evolve into a ***theory***
- Further experimentation to test predictions based on theory
  - theory could become a ***law*** (if others see same thing)

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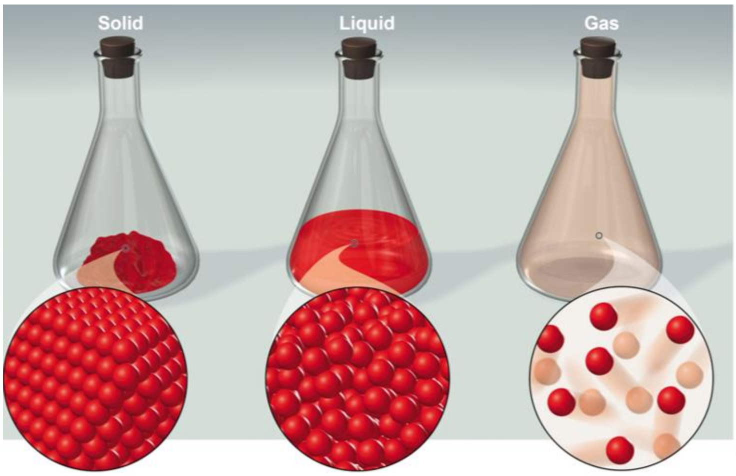
Classification of Matter

Solid, liquid and gas are the three states of matter commonly found on earth. A substance can convert from one state to another without changing the identity of the substance.



Solids Rigid and do NOT conform to the shape of their container.

Liquids conform to the shape of their container.

Gases assume both the shape and volume of their container.

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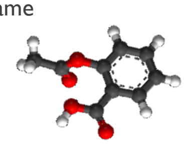
Atoms

Atom: smallest indivisible unit for an element
Molecule: consists of two or more atoms joined by chemical bonds.
Law of Conservation of Mass: No detectable change in mass during an ordinary chemical reaction.

Law of Definite proportions: A chemical compound always contains the same elements in the same proportions by mass
Example: Aspirin will always have 9 Carbon, 8 Hydrogen and 4 Oxygen Atoms ($C_9H_8O_4$)



A **mixture** can be separated by physical means into its components without changing the identities of the components.

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Classification of Matter

Chemists classify matter as either a **substance** or a **mixture** of substances.

Pure Substance is a form of matter that has definite composition and distinct properties.

- Examples: salt (sodium chloride), iron, water, mercury, carbon dioxide, and oxygen

Pure substances may be divided into two classes: **elements** and **compounds**.

- Pure substances that cannot be broken down into simpler substances by chemical changes are called **elements**.
- Pure substances that are comprised of two or more elements are called **compounds**. Compounds may be broken down by chemical changes to yield either elements or other compounds, or both.

Mixture is a physical combination of two or more substances.

- A **homogeneous mixture** is uniform throughout. Also called a **solution**.
  - (Examples: seawater, apple juice)
- A **heterogeneous mixture** is not uniform throughout.
  - (Examples: trail mix, chicken noodle soup)

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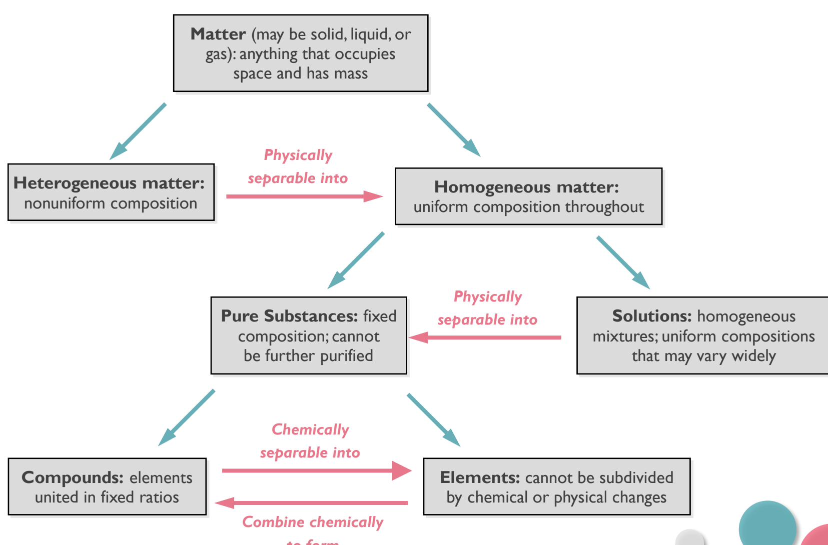
Types of Matter



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The Properties of Matter

There are two general types of properties of matter:

Quantitative properties are measured and expressed with a number.
(Quantity = An Amount)
Qualitative properties do not require measurement and are usually based on observation.

Examples:

Extracts from Pacific-yew bark kill cancer cells.

Compound "13a" is twenty times more effective than paclitaxel in killing ovarian cancer cells.

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The Properties of Matter

A ***physical property*** is one that can be observed and measured without changing the identity of the substance.
Examples: color, melting point, boiling point, density

A ***physical change*** is one in which the state of matter changes, but the identity of the matter does not change.
Examples: changes of state (melting, freezing, condensation)

Examples of ***physical processes***, *mixtures are separated , but the identities of the matter does not change*

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The Properties of Matter

A **chemical property** is one a substance exhibits as it interacts with another substance.
Examples: flammability, corrosiveness
A **chemical change** is one that results in a change of composition; the original substances no longer exist.
Examples: digestion, combustion, oxidation

An **extensive property** depends on the amount of matter.
Examples: mass, volume
An **intensive property** does not depend on the amount of matter.
Examples: temperature, density

Properties that can be measured are called *quantitative* properties.
A measured quantity must always include a unit.

The **English system** has units such as the foot, gallon, pound, etc.
The **metric system** includes units such as the meter, liter, kilogram, etc.

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SI Base Units

The revised metric system is called the **International System of Units** (abbreviated **SI Units**) and was designed for universal use by scientists.

SI Base 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

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SI Base Units

The magnitude of a unit may be tailored to an application using prefixes.

Prefix	Symbol	Meaning	Example (base unit grams)
Tera–	T	1×10^{12}	$1\ Tg = 1 \times 10^{12}\ g$
Giga–	G	1×10^{9}	$1\ Gg = 1 \times 10^{9}\ g$
Mega–	M	1×10^{6}	$1\ Mg = 1 \times 10^{6}\ g$
Kilo–	k	1×10^{3}	$1\ kg = 1 \times 10^{3}\ g$
Hecto–	h	1×10^{2}	$1\ hg = 1 \times 10^{2}\ g$
Deka–	da	1×10^{1}	$1\ dag = 1 \times 10^{1}\ g$
Unit (g, L, etc.)		1	**1 g = 1 g**
Deci–	d	1×10^{-1}	$1\ dg = 1 \times 10^{-1}\ g$
Centi–	c	1×10^{-2}	$1\ cg = 1 \times 10^{-2}\ g$
Milli–	m	1×10^{-3}	$1\ mg = 1 \times 10^{-3}\ g$
Micro–	μ	1×10^{-6}	$1\ \mu g = 1 \times 10^{-6}\ g$
Nano–	n	1×10^{-9}	$1\ ng = 1 \times 10^{-9}\ g$
Pico–	p	1×10^{-12}	$1\ pg = 1 \times 10^{-12}\ g$

The great mighty king hector died unexpectedly drinking chocolate milk many nights past.

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Using the conversion chart

1 T(base unit) = 1×10^{12} (base unit)
1 G(base unit) = 1×10^{9} (base unit)
1 M(base unit) = 1×10^{6} (base unit)
1 k(base unit) = 1×10^{3} (base unit)
1 h(base unit) = 1×10^{2} (base unit)
1 da(base unit) = 1×10^{1} (base unit)
1 (base unit) = 1 (base unit)
1 d(base unit) = 1×10^{-1} (base unit)
1 c(base unit) = 1×10^{-2} (base unit)
1 m(base unit) = 1×10^{-3} (base unit)
1 μ(base unit) = 1×10^{-6} (base unit)
1 n(base unit) = 1×10^{-9} (base unit)
1 p(base unit) = 1×10^{-12} (base unit)

How do we use this chart? So let's say our base units are liters? How could we find our conversion factor?

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Using the chart

1 T(unit) = 1 x 10^{12} (unit)
1 G(unit) = 1 x 10^{9} (unit)
1 M(unit) = 1 x 10^{6} (unit)
1 k(unit) = 1 x 10^{3} (unit)
1 h(unit) = 1 x 10^{2} (unit)
1 da(unit) = 1 x 10^{1} (unit)
1 (base unit) = 1 (base unit)
1 d(unit) = 1 x 10^{-1} (unit)
1 c(unit) = 1 x 10^{-2} (unit)
1 m(unit) = 1 x 10^{-3} (unit)
1 µ(unit) = 1 x 10^{-6} (unit)
1 n(unit) = 1 x 10^{-9} (unit)
1 p(unit) = 1 x 10^{-12} (unit)

Replace the word base unit with the symbol for liter (L) (gives an equality statement). →

1 TL = 1 x 10^{12} L
1 GL = 1 x 10^{9} L
1 ML = 1 x 10^{6} L
1 kL = 1 x 10^{3} L
1 hL = 1 x 10^{2} L
1 daL = 1 x 10^{1} L
1 L = 1 L
1 dL = 1 x 10^{-1} L
1 cL = 1 x 10^{-2} L
1 mL = 1 x 10^{-3} L
1 µL = 1 x 10^{-6} L
1 nL = 1 x 10^{-9} L
1 pL = 1 x 10^{-12} L

This can be done with all standard units of measurement (grams, meter, second, kelvin, etc.). You need to know these conversions!

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Conversion Factors

A **conversion factor** is a fraction in which the same quantity is expressed one way in the numerator and another way in the denominator. Created from an equality statement (shows the relationship between two units. When we multiply by a conversion factor, we are only multiplying by 1, so nothing is changed!

Conversion Factors

Equality Statement

1 hour = 60 minutes

$$\frac{1 \text{ hour}}{60 \text{ minutes}}$$

$$\frac{60 \text{ minutes}}{1 \text{ hour}}$$

Divide one side of the equality statement by the other side of the equality statement and you have one of the two conversion factors….They mean the same thing!! Both are saying there is 1 hour in 60 minutes, meaning either fraction is correct because both came from the equality statement.

We are going to use this concept of conversion factors to convert (or move) between SI base units…

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How many kg are in 3.6 g?

1 T(base unit) = 1 x 10^{12} (base unit)
1 G(base unit) = 1 x 10^{9} (base unit)
1 M(base unit) = 1 x 10^{6} (base unit)
1 k(base unit) = 1 x 10^{3} (base unit)
1 h(base unit) = 1 x 10^{2} (base unit)
1 da(base unit) = 1 x 10^{1} (base unit)
1 (base unit) = 1 (base unit)
1 d(base unit) = 1 x 10^{-1} (base unit)
1 c(base unit) = 1 x 10^{-2} (base unit)
1 m(base unit) = 1 x 10^{-3} (base unit)
1 μ(base unit) = 1 x 10^{-6} (base unit)
1 n(base unit) = 1 x 10^{-9} (base unit)
1 p(base unit) = 1 x 10^{-12} (base unit)

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

Equality Statement….We can make conversion factors. Think of conversion factors as the roads to the next unit….

What are the two possible conversion factors from the equality statement?

$$\frac{1\ \text{kg}}{10^3\text{g}} \text{ and } \frac{10^3\text{g}}{1\ \text{kg}}$$

Which is the correct one to use? We let the units tell us. Write down what is given in the problem and those units will need to be crossed out….

We were given 3.6 g and asked to convert it to kg. So the first thing we write down is 3.6 g. Then we will use the conversion factor that allows us to cancel the grams so that we are left with kg.

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Conversion Practice

Calculate the number of kg that are found in 1.3 g?

How many Mg are in 43.7 mg?

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Mass

Mass is a measure of the amount of matter in an object or sample.

Because gravity varies from location to location, the weight of an object varies depending on where it is measured. But mass doesn't change.

The SI base unit of mass is the kilogram (kg), but in chemistry the smaller gram (g) is often used.

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

Atomic mass unit (amu) is used to express the masses of atoms and other similar sized objects.

$$1 \text{ amu} = 1.6605378 \times 10^{-24} \text{ g}$$

Temperature

There are two temperature scales used in chemistry:

The Celsius scale (°C): Freezing point (pure water): 0°C; Boiling point (pure water): 100°C

The Kelvin scale (K): The "absolute" scale; Lowest possible temperature: 0 K (absolute zero)

$$\boxed{K = °C + 273.15}$$

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Practice

Normal human body temperature can range over the course of a day from about 36°C in the early morning to about 37°C in the afternoon. Express these two temperatures and the range that they span using the kelvin scale.

36°C + 273.15 = 309.15 K

37°C + 273.15 = 310.15 K

What range do they span?

310.15 K – 309.15 K= 1K

The Fahrenheit scale is common in the United States.

Freezing Point (pure water)	32°F
Boiling Point (pure water)	212°F

There are 180 degrees between freezing and boiling in Fahrenheit (212°F–32°F) but only 100 degrees in Celsius (100°C–0°C).

The size of a degree on the Fahrenheit scale is only $\frac{9}{5}$ of a degree on the Celsius scale.

$$\boxed{\text{Temp in °F} = \left(\frac{9\ °\text{F}}{5\ °\text{C}}\right)(\text{temp in °C}) + 32\ °\text{F}}$$

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Practice

A body temperature above 39°C constitutes a high fever. Convert this temperature to the Fahrenheit scale.

Gold Boils at 5173°F. What is the Boiling Point of Gold in Celsius?

Chlorine melts at –100.95 °C what is that in °F?

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Practice

Perform the following conversions:

41°F to °C

68 °C to Kelvin

426 K to °C

681 K to °F

Normal body temperature is 37°C. What is this in Kelvin?

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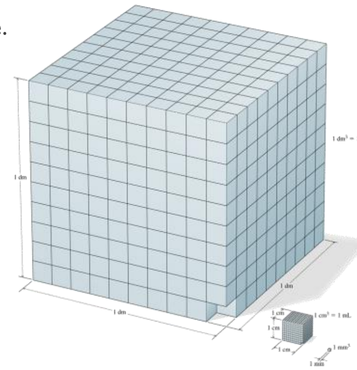
Derived Units: Volume and Density

The **density** of a substance is the ratio of mass to volume.

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

d = density
m = mass
V = volume



There are many units (such as volume) that require units not included in the base SI units.
The derived SI unit for volume is the meter cubed (m^3).
A more practical unit for volume is the **liter** (L).

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

SI-derived unit: kilogram per cubic meter (kg/m^3)

Common units based on state of matter:

$\frac{g}{cm^3}$ (solids)		$g\ cm^{-3}$ (solids)
$\frac{g}{mL}$ (liquids)	**Same As:**	$g\ mL^{-1}$ (liquids)
$\frac{g}{L}$ (gases)		$g\ L^{-1}$ (gases)

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Practice

Density at 20 °C	
Substance	D ($g\ mL^{-1}$)
Magnesium	1.74
Aluminum	2.70
Titanium	4.50
Copper	8.93
Lead	11.34
Mercury	13.55
Gold	19.32

A piece of metal has mass = 215.8 g. When placed into a measuring cylinder it displaces 19.1 mL of water. Identify the metal.

What is the density of CO gas if 0.196 g occupies a volume of 100 mL?

An irregularly-shaped sample of aluminum (Al) is put on a balance and found to have a mass of 43.6 g. The student decides to use the water-displacement method to find the volume. The initial volume reading is 25.5 mL and, after the Al sample is added, the water level has risen to 41.7 mL. Find the density of the Al sample in $g\ cm^{-3}$. (Remember: $1\ mL = 1\ cm^3$.)

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Practice

Gasoline is a non-polar liquid that will float on water. 450 grams of gasoline is spilled into a puddle of water. If the density of gasoline is 0.665 g mL^{-1}, what volume of gasoline is spilled?

A cup of gold colored metal beads was measured to have a mass 425 grams. By water displacement, the volume of the beads was calculated to be 48.0 cm^3. Given the following densities, identify the metal. (Gold: 19.3 g mL^{-1}; Copper: 8.86 g mL^{-1}; Bronze: 9.87 g mL^{-1})

Calculate the <u>mass</u> of a liquid with a density of 3.2 g mL^{-1} and a volume of 25 mL.

Find the volume that 35.2 g of carbon tetrachloride (CCl_4) will occupy if it has a density of 1.60 g mL^{-1}.

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Uncertainty in Measurement

There are two types of numbers used in chemistry:

1) Exact numbers:

a) are those that have defined values

1 kg = 1000 g

1 dozen = 12 objects

b) are those determined by counting

28 students in a class

2) Inexact numbers:

a) measured by any method other than counting

length, mass, volume, time, speed, etc.

An inexact number must be reported to indicate its uncertainty.

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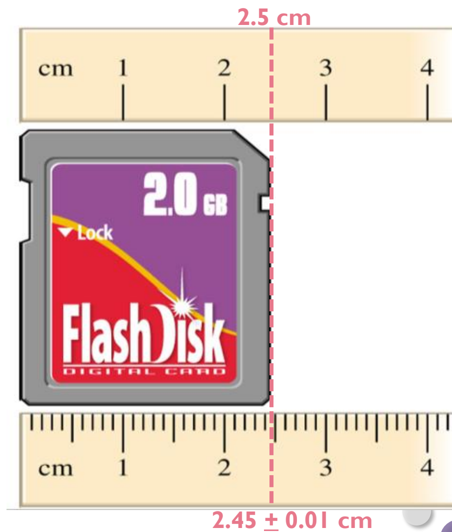
Uncertainty in Measurement

Significant figures **(or digits)** are the meaningful digits in a reported number.

When using the top ruler to measure the memory card, we could estimate 2.5 cm. We are certain about the 2, but we are ***not*** certain about the 5.

When using the bottom ruler to measure the memory card, we might record 2.45 cm. Again, we estimate one more digit than we are certain of.

As a rule, estimate between the markings and the estimated digit is the uncertain digit.



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Significant Figures

The number of significant figures (S.F.) can be determined using the following guidelines:

1) Any nonzero digit is significant.

112.1 — 4 significant figures

2) Zeros between nonzero digits are significant.

305 — 3 significant figures

50.08 — 4 significant figures

3) Zeros to the left of the first nonzero digit are not significant.

0.0023 — 2 significant figures

0.000001 — 1 significant figure

4) Zeros to the right of the last nonzero digit are significant if a decimal is present.

1.200 — 4 significant figures

5) Zeros to the right of the last nonzero digit in a number that does not contain a decimal point may or may not be significant.

100 — 1, 2, or 3 – ambiguous

To avoid ambiguity, use scientific notation…

1 S.F. = 1×10^2

2 S.F. = 1.0×10^2

3 S.F. = 1.00×10^2

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Practice

Determine the number of significant figures in the following measurements:

(a) 123 cm

(b) 25.03 g

(c) 0.0857 kg

(d) 1.106×10^{-7} L

(e) 50.0 mL

(f) 0.1600 m

Number	Sig. Figs.	Comment on Zeros
2.12		
4.500		
0.002541		
0.00100		
500		
500.		
5.0×10^2		
1.05 g		
Dozen Eggs		
0.90×10^{45} L		

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Calculations with Measured Numbers

In addition and subtraction, the answer cannot have more digits to the right of the decimal point than any of the original numbers.

$$\begin{array}{rl} 102.50 & \leftarrow \text{two digits after the decimal point} \\ +\ \ 0.230 & \leftarrow \text{three digits after the decimal point} \\ \hline 102.73 & \leftarrow \text{round to two digits after the decimal point, } \mathbf{102.73} \end{array}$$

$$\begin{array}{rl} 143.29 & \leftarrow \text{two digits after the decimal point} \\ -\ \ 20.1 & \leftarrow \text{one digit after the decimal point} \\ \hline 123.19 & \leftarrow \text{round to one digit after the decimal point, } \mathbf{123.2} \end{array}$$

In multiplication and division, the number of significant figures in the final product or quotient is determined by the original number that has the smallest number of significant figures.

$$\underbrace{(1.4)}_{\text{2 S.F.}}\underbrace{(8.011)}_{\text{4 S.F.}} = 11.2154 \quad \leftarrow \text{fewest significant figures is 2, so round to } \mathbf{11}$$

$$\frac{11.57 \ (\text{4 S.F.})}{305.88 \ (\text{5 S.F.})} = 0.0378252 \quad \leftarrow \text{fewest significant figures is 4, so round to } \mathbf{0.03783}$$

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Calculations with Measured Numbers

Exact numbers can be considered to have an infinite number of significant figures and do not limit the number of significant figures in a result.

Example

Three pennies each have a mass of 2.5 g. What is the total mass?

In calculations with multiple steps, round at the end of the calculation to reduce any rounding errors.

If the 1st digit to be dropped is < 5, round down

If the 1st digit to be dropped is > 5, round up

If the dropped digit is 5, we round up or down to yield a value for the retained digit that is even.

Do not round after each step!

Compare the following:

Rounding after each step	Rounding at end
1) 3.66 × 8.45 = 30.9	1) 3.66 × 8.45 = 30.93
2) 30.9 × 2.11 = 65.2	2) 30.93 × 2.11 = 65.3

In general, keep at least one extra digit until the end of a multistep calculation.

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Summary– Significant Figures

Functions	Rule
Addition and Subtraction	Least number of **decimals**
Multiplication and Division	Least number of **digits**

$$x = \frac{(1.278 + 0.3480)}{1.08881}$$

$$x = \frac{(8.278 + 2.34)}{1.08}$$

583.00 ÷ 83

(57.6 * 3) ÷ (34 * 3.00 * 87.507)

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Accuracy and Precision

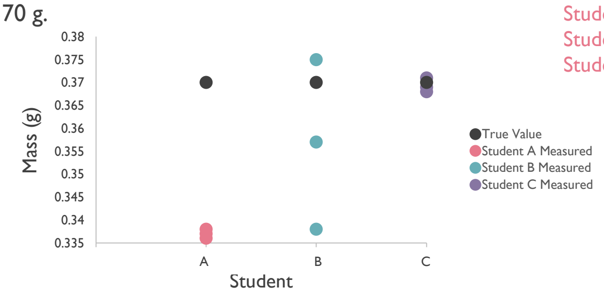
Accuracy tells us how close a measurement is to the true value.
Precision tells us how close a series of replicate measurements are to one another.


Good accuracy/ good precision


Poor accuracy / good precision


Poor accuracy / poor precision

Three students were asked to find the mass of an aspirin tablet. The true mass of the tablet is 0.370 g.



Student A: Precise but not accurate
Student B: Neither precise nor accurate
Student C: Both precise and accurate

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Dimensional Analysis – Tracking Units

A **conversion factor** is a fraction in which the same quantity is expressed one way in the numerator and another way in the denominator.

For example, 1 in = 2.54 cm, may be written: $\frac{1\text{ in}}{2.54\text{ cm}}$ or $\frac{2.54\text{ cm}}{1\text{ in}}$

The use of conversion factors in problem solving is called **dimensional analysis** or the factor–label method.

Example
Convert 12.00 inches to meters. $\left(\frac{12.00\text{ in}}{}\right)$

What conversion factor will cancel inches and give us centimeters?

$$\left(\frac{12.00\text{ in}}{}\right)\left(\frac{2.54\text{ cm}}{1\text{ in}}\right) \quad \text{or} \quad \left(\frac{12.00\text{ in}}{}\right)\left(\frac{1\text{ in}}{2.54\text{ cm}}\right)$$

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Common Conversion Factors

Length	Volume	Mass
1 m = 1.0936 yd	1 L = 1.0567 qt	1 kg = 2.2046 lb
1 in. = 2.54 cm (exact)	1 qt = 0.94635 L	1 lb = 453.59 g
1 km = 0.62137 mi	1 ft^3 = 28.317 L	1 (avoirdupois) oz = 28.349 g
1 mi = 1609.3 m	1tbsp = 14.787 mL	1 (troy) oz = 31.103 g

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Example

Convert 20.0 milligrams of gold into pounds.

Start with what you are given. Look for a pathway: compare the units. What units are shared between the conversion factors?

To find the path you must find the necessary conversion factors derived from the equality statements. From the textbook we know: 1 g = 1000 mg and 1 lb = 453.6 g.

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Example

An average adult has 5.2 L of blood. What is the volume of blood in pints?

If someone weighs 175 lbs. What is that in kilograms. (1 lb = 453.6 g)

A child requires a 5 ml dose of medicine each day. How many days would a gallon of this medicine last? Note that each gallon has 3.7854 L.

The moon is 384,403 km from the earth. Estimate how many quarters laid end to end will it take to reach the moon if a quarter has a diameter of 2.3 cm.

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Example

Convert 14.62 in^3 to cubic centimeters, given that there are 2.54 centimeters for every inch.

A chair lift at the Divide ski resort in Cold Springs, WY is 4806 feet long and takes 9 minutes. A.) What is the average speed in miles per hour (There are 5280 ft per every mile)? B.) How many feet per second does the lift travel?

An average human heart beats 60 times per minute. If an average person lives to the age of 75, how many times does the average heartbeat in a lifetime?

A car travels 14 miles in 15 minutes. A.) How fast is it going in miles per hour? How fast is it going in meters per second (1.60933 km = 1 mile)?

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Example

How many ng are there in 5.27×10^{-13} kg?

A. 5.27×10^{-16} ng
B. 5.27×10^{-7} ng
C. 5.27×10^{-4} ng
D. 5.27×10^{-1} ng

If the density of an object is 2.87×10^{-4} lbs per cubic inch, what is its density in g mL^{-1}? There are 453.592 grams in each pound and 2.54 centimeters in every inch.

A. 2.49×10^{-7} g/mL
B. 1.13×10^{-4} g/mL
C. 7.95×10^{-3} g/mL
D. 5.13×10^{-2} g/mL

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