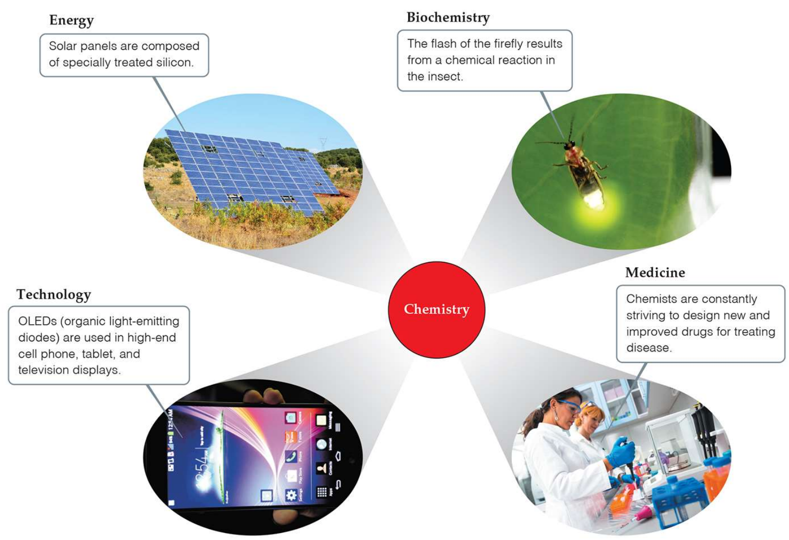
Chapter 1
Introduction: Matter and Measurement

Learning outcomes:

- **Distinguish among elements, compounds, and mixtures.**
- **Identify symbols of common elements.**
- **Identify common metric prefixes.**
- **Demonstrate the use of significant figures, scientific notation, and SI units in calculations.**
- **Attach appropriate SI units to defined quantities, and employ dimensional analysis in calculations**

- Chemistry is the study of **matter**, its properties, composition, and structure and the changes it undergoes.
- It is central to our fundamental understanding of many science-related fields.

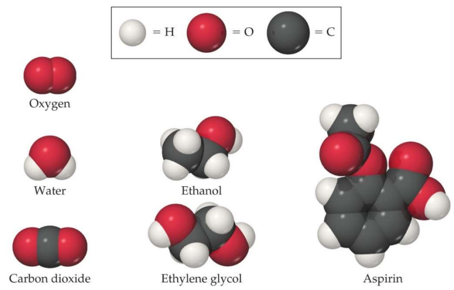


Atomic and Molecular Perspective

Matter – Anything that has mass and occupies space.

Atom – The smallest stable building block of matter.

Molecule – Groups of atoms held together with a specific connectivity and shape.



Composition - the types of atoms that are present in a compound and the ratio of these atoms (for example H_2O, C_2H_6O).

Structure - how atoms are connected (bonded) to each other, how far apart they are, and the shape of the molecule.

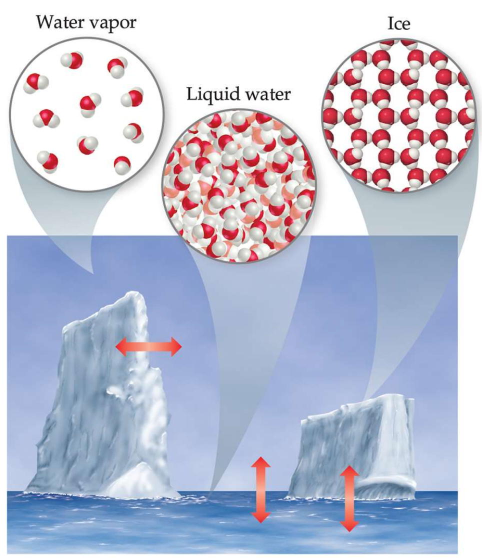
Methods of Classification of Matter

State of Matter - physical state is gas, liquid, or solid.

Composition of Matter - element, compound, or mixture

States of Matter

1) **Gas** (vapor) – has no fixed volume or shape, uniformly expands to fill its container, compressible, flows readily, diffusion occurs rapidly.
2) **Liquid** - has a distinct volume independent of its container, assumes the shape of the portion of the container it occupies, not significantly compressible, diffusion occurs but slower than a gas.
3) **Solid** - has both a definite shape and definite volume, not significantly compressible, diffusion occurs extremely slowly.



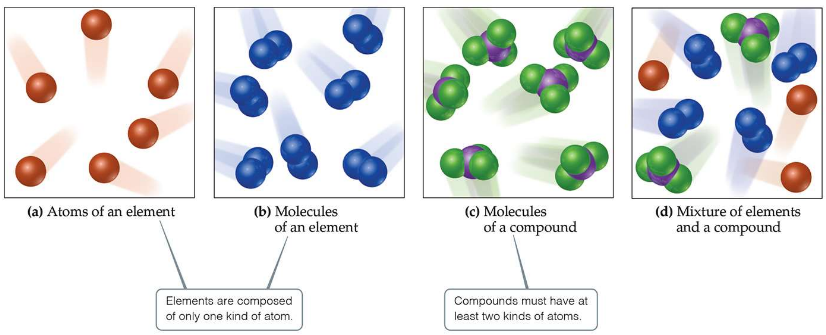
Elements, Compounds & Mixtures

Pure Substance Matter that has a fixed composition and distinct properties. All substances are either elements or compounds.

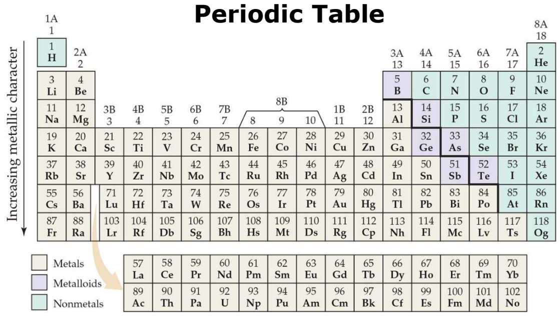
Elements All atoms are the same kind, elements have only one type of atom. e.g. oxygen (O_2), gold (Au), silicon (Si) and diamond (C).

Compounds Contains more than one type of atom, but all molecules (or repeat units) are the same, e.g. water (H_2O), ethanol (C_2H_6O), quartz (SiO_2), sodium chloride (NaCl).

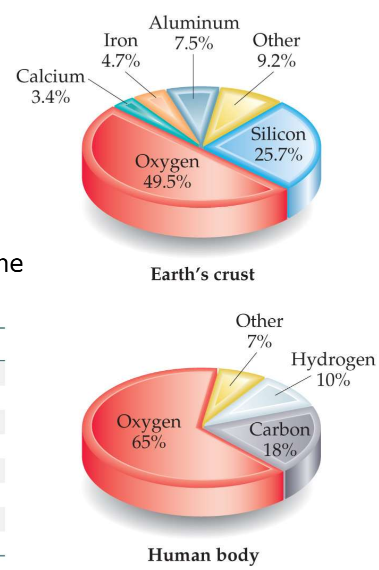
Mixture Have variable composition and can be separated into component parts by physical methods. Mixtures contain more than one kind of molecule, and their properties depend on the relative amount of each component present in the mixture.



Periodic Table



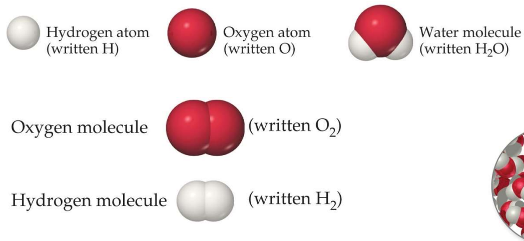
Relative abundances of elements in the Earth's crust and human body.

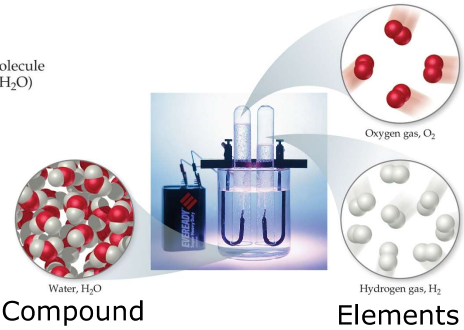


Elements are represented as symbols with one or two letters; the first is always capitalized.

TABLE 1.1 Some Common Elements and Their Symbols

Carbon	C	Aluminum	Al	Copper	Cu (from *cuprum*)
Fluorine	F	Bromine	Br	Iron	Fe (from *ferrum*)
Hydrogen	H	Calcium	Ca	Lead	Pb (from *plumbum*)
Iodine	I	Chlorine	Cl	Mercury	Hg (from *hydrargyrum*)
Nitrogen	N	Helium	He	Potassium	K (from *kalium*)
Oxygen	O	Lithium	Li	Silver	Ag (from *argentum*)
Phosphorus	P	Magnesium	Mg	Sodium	Na (from *natrium*)
Sulfur	S	Silicon	Si	Tin	Sn (from *stannum*)





Elements can interact with other elements to form compounds, and compounds can be decomposed into elements.

The elemental composition of a compound is always the same, which is known as the Law of Constant Composition (or Law of Definite Proportions).

TABLE 1.2 Comparison of Water, Hydrogen, and Oxygen

	Water	Hydrogen	Oxygen
State[a]	Liquid	Gas	Gas
Normal boiling point	100 °C	−253 °C	−183 °C
Density[a]	1000 g/L	0.084 g/L	1.33 g/L
Flammable	No	Yes	No

[a] At room temperature and atmospheric pressure.

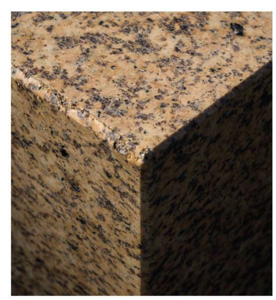
Homogeneous & Heterogeneous Mixtures

Heterogeneous Mixture - non-uniform.

Chocolate Chip Cookie – Chocolate, Dough, etc.

Concrete – Cement, Rocks, etc.

Nachos – Chips, cheese, jalapeños, salsa, etc.



(a)

Homogeneous Mixture – uniform throughout, also called a *solution*.

Air – principle components include O_2, N_2 & CO_2

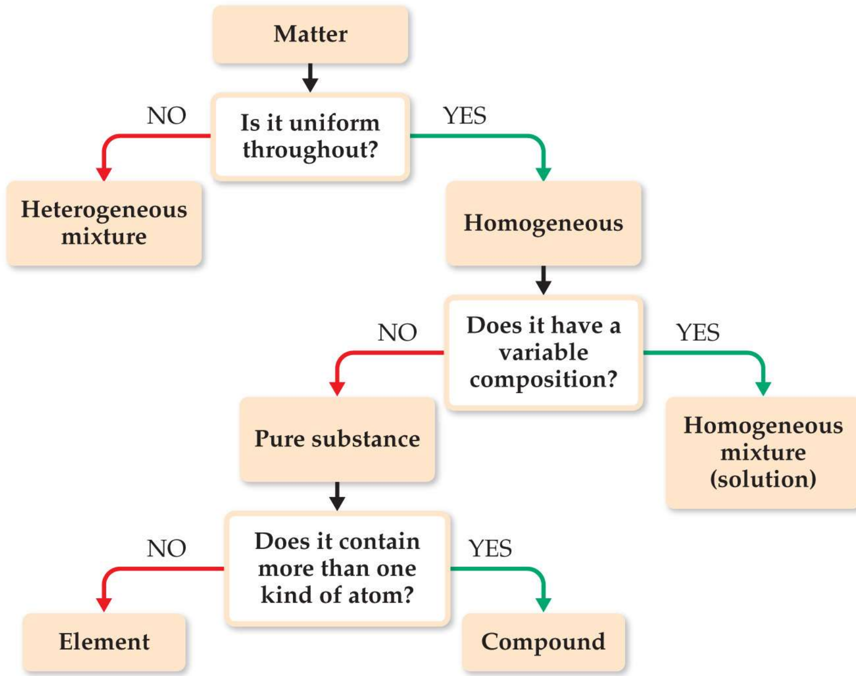
Vodka – principle components are ethanol and water

Brass – solid solution of Cu and Zn

Ruby – solid solution of Al_2O_3 and Cr_2O_3

Composition of a mixture can vary.

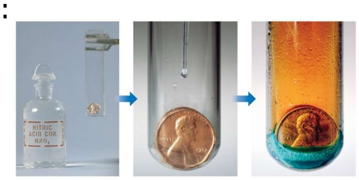
(b)



Chemical and Physical Properties

Physical Properties Some properties can be readily measured with our senses, e.g. odor and color, instruments are needed to measure other properties, such as electrical resistivity, hardness, melting point, boiling point, density, mass, volume, etc.

Chemical Properties Describe the reactivity of a substance toward other substances. Examples include:



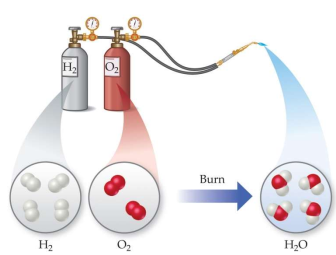
Ethanol burns in air (reacts with oxygen)
Sodium reacts vigorously with water,
Corrosion of metal parts (rust),
Trinitrotoluene (TNT) is explosive.

Physical changes are changes in matter that do *not* change the composition of a substance.

– Examples include changes of state, temperature, and volume.

Chemical changes result in new substances.

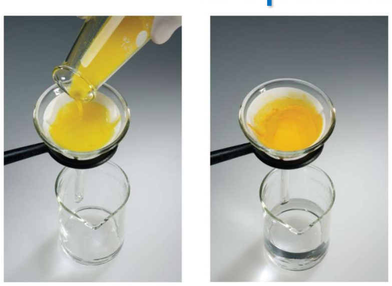
– Examples include combustion, oxidation, and decomposition.



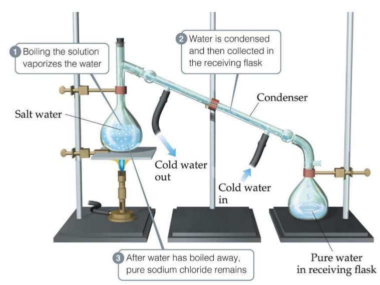
Properties of Matter

- Intensive Properties:
  - Independent of the amount of the substance that is present.
    - Density, boiling point, color, etc.
- Extensive Properties:
  - Dependent upon the amount of the substance present.
    - Mass, volume, energy, etc.

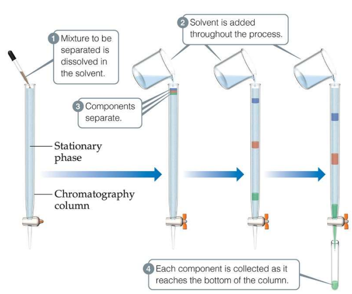
Separation of Mixtures



Filtration



Distillation



Chromatography

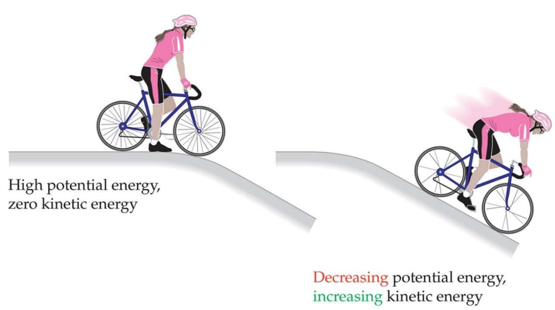
Energy

- **Energy** is the capacity to do work or transfer heat.
- **Work** is the energy transferred when a force exerted on an object causes a displacement of that object.
- **Heat** is the energy used to cause the temperature of an object to increase.
- **Force** is any push or pull on an object.

(a)

(b)

Fundamental Forms of Energy



- **Kinetic energy** is the energy of motion.
  – Its magnitude depends on the object's mass and its velocity:
    KE = $\frac{1}{2}mv^2$
- **Potential energy** of an object depends on its relative position compared to other objects.
  - Potential energy also refers to the composition of an object, including the energy stored in chemical bonds.

One of the goals in chemistry is to related the energy changes in the macroscopic world to the kinetic or potential energy of substances at the molecular level.

Numbers and Units in Chemistry

Major role in quantifying:

- Units of measurement
- Quantities that are measured and calculated
- Uncertainty in measurement
- Significant figures
- Dimensional analysis
  (e.g. 1 inch = 2.54 cm)

TABLE 1.3 SI Base Units

Physical Quantity	Name of Unit	Abbreviation
Length	Meter	m
Mass	Kilogram	kg
Temperature	Kelvin	K
Time	Second	s or sec
Amount of substance	Mole	mol
Electric current	Ampere	A or amp
Luminous intensity	Candela	cd

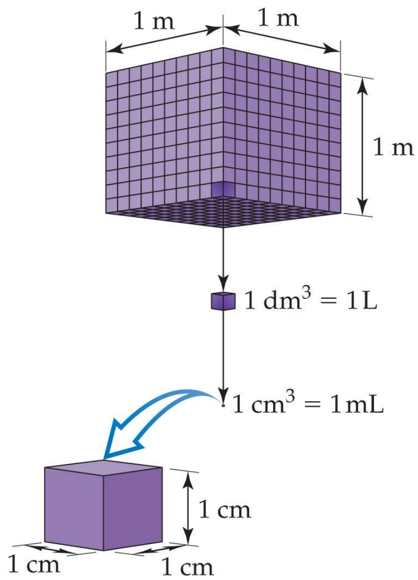
Metric System Prefixes

TABLE 1.4 Prefixes Used in the Metric System and with SI Units

Prefix	Abbreviation	Meaning	Example	
Peta	P	10^{15}	1 petawatt (PW)	$= 1 \times 10^{15}$ watts[a]
Tera	T	10^{12}	1 terawatt (TW)	$= 1 \times 10^{12}$ watts
Giga	G	10^{9}	1 gigawatt (GW)	$= 1 \times 10^{9}$ watts
Mega	M	10^{6}	1 megawatt (MW)	$= 1 \times 10^{6}$ watts
Kilo	k	10^{3}	1 kilowatt (kW)	$= 1 \times 10^{3}$ watts
Deci	d	10^{-1}	1 deciwatt (dW)	$= 1 \times 10^{-1}$ watt
Centi	c	10^{-2}	1 centiwatt (cW)	$= 1 \times 10^{-2}$ watt
Milli	m	10^{-3}	1 milliwatt (mW)	$= 1 \times 10^{-3}$ watt
Micro	μ[b]	10^{-6}	1 microwatt (μW)	$= 1 \times 10^{-6}$ watt
Nano	n	10^{-9}	1 nanowatt (nW)	$= 1 \times 10^{-9}$ watt
Pico	p	10^{-12}	1 picowatt (pW)	$= 1 \times 10^{-12}$ watt
Femto	f	10^{-15}	1 femtowatt (fW)	$= 1 \times 10^{-15}$ watt
Atto	a	10^{-18}	1 attowatt (aW)	$= 1 \times 10^{-18}$ watt
Zepto	z	10^{-21}	1 zeptowatt (zW)	$= 1 \times 10^{-21}$ watt

[a]The watt (W) is the SI unit of power, which is the rate at which energy is either generated or consumed. The SI unit of energy is the joule (J); $1\,J = 1\,kg \cdot m^2/s^2$ and $1\,W = 1\,J/s$.
[b]Greek letter mu, pronounced "mew."

Volume

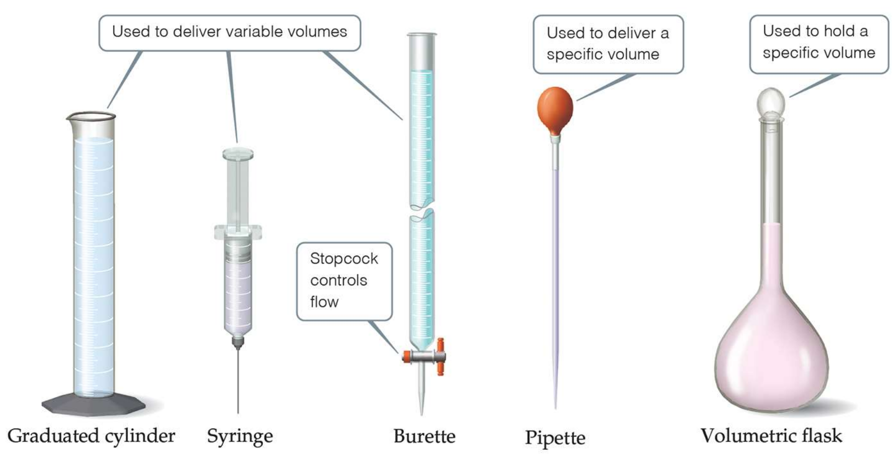


Volume is not a base unit for SI; it is a derived unit from length (m × m × m = m^3).

The most commonly used metric units for volume are the liter (L) and the milliliter (mL).

- A liter is a cube 1 decimeter (dm) long on each side.
- A milliliter is a cube 1 centimeter (cm) long on each side, also called 1 cubic centimeter (cm × cm × cm = cm^3).

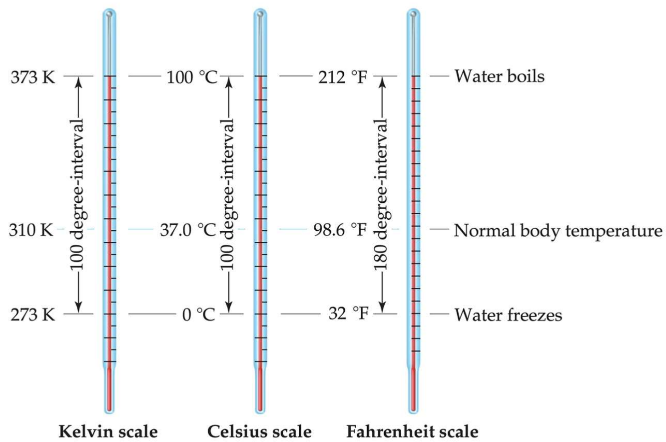
Glassware for Measuring Volume



Uncertainty in Measurements - Different measuring devices have different uses and different degrees of precision.

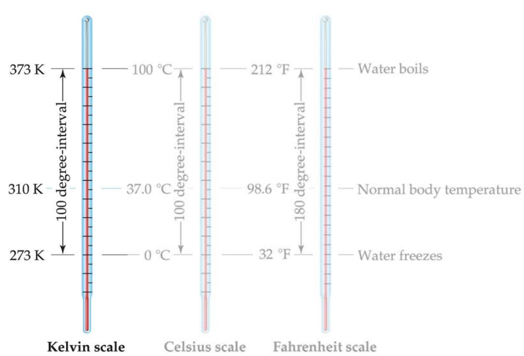
Temperature Scales

Temperature – the "hotness and coldness" of an object.



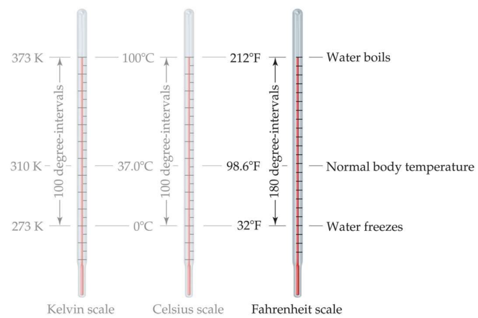
Heat flows spontaneously from an object with a higher temperature to an object with a lower temperature.

Temperature



- The Kelvin is the SI unit of temperature.
- It is based on the properties of gases.
- There are no negative Kelvin temperatures.
- K = °C + 273.15

Temperature



- The Fahrenheit scale is not used in scientific measurements.

°F = 9/5(°C) + 32

°C = 5/9(°F − 32)

The '9/5', '5/9', and '32' are *exact* numbers and do not influence significant figures.

Density

$$\text{Density} = \text{d} = \frac{\text{mass}}{\text{volume}} = \frac{\text{m}}{\text{V}}$$

TABLE 1.5 Densities of Selected Substances at 25 °C

Substance	Density (g/cm^3)
Air	0.001
Balsa wood	0.16
Ethanol	0.79
Water	1.00
Ethylene glycol	1.09
Table sugar	1.59
Table salt	2.16
Iron	7.9
Gold	19.32

Example: A piece of unknown metal with a right rectangular prism shape has a width of 3.2 cm , a length of 17.1 cm and height of 4.0 cm. Its mass is 1.5 kg. Calculate the density of the metal in g/cm^3.

Energy

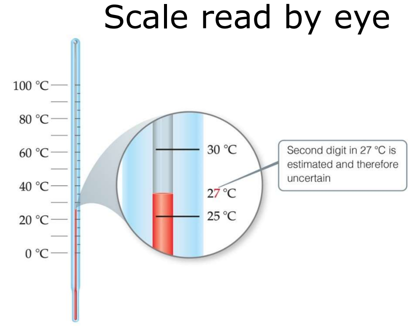
- The unit of energy: Joule (J). It is a derived unit, $kg \cdot m^2/s^2$:
  - Kinetic energy: $KE = \frac{1}{2} m v^2$
  - If the object is 2 kg, and it moves at 1 m/s, it will posses 1 J of kinetic energy:
  - $1\ J = \frac{1}{2}\ (2\ kg)\ (1\ m/s)^2$ OR: $1\ J \equiv 1\ kg \cdot m^2/s^2$
- The kJ is commonly used for chemical change.
- Historically, the calorie was used: 1 cal = 4.184 J
- This calorie is NOT the nutritional Calorie. That one is a kcal.
- 1 nutritional Calorie = 1 Cal = 1000 cal

Numbers in Chemistry

- **Exact** numbers are counted or given by definition. For example, there are 12 eggs in 1 dozen and 3 feet in 1 yard.
- **Inexact** (or **measured**) numbers depend on how they were determined. Scientific instruments have limitations (*equipment errors*) and individuals can read some instrumentation differently (*human errors*).

Digital Reading

Scale read by eye



The last digit measured is considered *reliable*, but not *exact*.

Precision and Accuracy

- Precision is a measure of how closely individual measurements to agree with one another.
- Accuracy refers to how closely individual measurements agree with the correct "true" value.

Significant Figures

- The term significant figures refers to digits that were measured.

1.03×10^4 g	(three significant figures)
1.030×10^4 g	(four significant figures)
1.0300×10^4 g	(five significant figures)

- When rounding calculated numbers, we pay attention to significant figures so we do not overstate the accuracy of our answers.

Significant Figures

1. Zeros between non-zero numbers are always significant.
2. Zeros at the beginning of a number are never significant, merely indicate the position of the decimal point.
3. Zeros at the end of the number after a decimal place are significant if the number contains a decimal point.
4. Zeros at the end of a number before a decimal place are ambiguous (*e.g.* 23,800 g), unless a decimal point is written at the end (i*.e.* 23,800. g). Assume the zeros are insignificant, unless there is a decimal point. Avoid ambiguity by using scientific notation.

Significant Figures

How many significant figures are present in each of the measured quantities?

0.0012

108

900.0

3.0012

0.002070

4.80×10^{-3}

4.800×10^{-3}

Rounding*

After determining the appropriate number of significant figures, round off your final answer.

1. If the first digit you drop is greater than 5, add 1 to the last digit you keep. You are rounding up.
2. If the first digit you drop is less than 5, do nothing to the digits you keep. You are rounding down.
3. If the digit you drop is 5, and there are no following digits, *round down*. If there are digits following the 5, *round up*.

*You may receive a different rule #3 from your lab instructor.

Significant Figures & Calculations

Addition and Subtraction

Line up the numbers at the decimal point and the answer cannot have more decimal places than the *measurement* with the fewest number of decimal places.

This number limits	20.42	⟵	two decimal places
the number of significant	1.322	⟵	three decimal places
figures in the result ⟶	83.1	⟵	one decimal place
	104.842	⟵	round off to one decimal place (104.8)

Addition and Subtraction

- The absolute uncertainty can be no smaller than the least accurate number.
- $$\begin{array}{r} 12.04 \\ -\ 10.4 \\ \hline 1.64 \end{array} \qquad \mathbf{1.6}$$
- The answer should have no more decimal places than the least accurate number.

Multiplication and Division

The answer cannot have more significant figures than the measurement with the fewest number of significant figures.

$$\mathbf{3121} \times \mathbf{12} = \mathbf{37}452 = \mathbf{3.7} \times 10^4$$

\# sig. digits **4** **2** **2**

Know the number of appropriate digits throughout, round at the end.

Mixed Operations

Determine accuracy in the <u>same order</u> as the mathematical operations, # of significant digits in **blue**

•but, retain *at least one* additional digit past the significant figures in combined operations, so rounding doesn't affect results…

-keep track of the proper significant figures to use at the end.

$$d = \frac{m}{v} = \frac{\overset{\boxed{3}}{2.79\text{ g}}}{\underset{\boxed{3}}{8.34}\text{ mL} - \underset{\boxed{3}}{7.58}\text{ mL}} = \frac{\overset{\boxed{3}}{2.79\text{ g}}}{\underset{\boxed{2}}{0.76}\text{ mL}}$$

$$d = \underset{\boxed{2}}{3.7}\text{ g/mL}$$

Evaluate each expression to the correct number of significant figures.

(a) $4.184 \times 100.620 \times (25.27 - 24.16) =$

(b) $\dfrac{8.925 - 8.904}{8.925} \times 100\% \quad =$

(c) $\dfrac{9.6 \times 100.65}{8.321} + 4.026 \quad =$

(d) $320.75 - (6102.1/3.1) \quad =$

(e) $[(853.6\times10^4) - (6.967\times10^2)]\times3.6810 =$

Retain at least one additional digit past the significant figures in *combined* operations, so rounding doesn't affect result…

-keep track of the proper significant figures for the final answer.

Dimensional Analysis

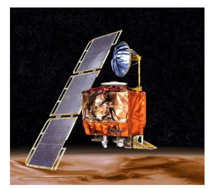
Units are multiplied together or divided into each other along with the numerical values.

- Keep track of <u>both</u> numerical values and units.

$$\text{Number of centimeters} = (8.50\ \cancel{\text{in.}})\ \frac{2.54\ \text{cm}}{1\ \cancel{\text{in.}}} = 21.6\ \text{cm}$$

Desired unit

Given unit



http://www.wired.com/2010/11/1110mars-climate-observer-report/

Conversions: Two or More Factors

What is the mass in g, of 1.00 gal of H_2O? The density of water is 1.00 g/mL.   1 L = 1.057 qt, 1 gal = 4 qt

Conversions Involving Volume

Express a volume of 1.250 L in mL and cm^3

$$(1.250\ \mathrm{L})\times\left(\frac{1\ \mathrm{mL}}{1\times10^{-3}\ \mathrm{L}}\right)=1{,}250.\ \mathrm{mL}$$

$$(1.250\ \mathrm{L})\times\left(\frac{1000\ \mathrm{mL}}{1\ \mathrm{L}}\right)=1{,}250.\ \mathrm{mL}$$

$$(1.250\ \mathrm{L})\times\left(\frac{1000\ \mathrm{cm}^3}{1\ \mathrm{L}}\right)=1{,}250.\ \mathrm{cm}^3$$

Express a volume of 1,250. cm^3 in m^3.

The prefix *centi* is 10^{-2}, 1 cm = 1 × 10^{-2} m for length.

Volume involves cubed units, create a conversion:

$$\left(\frac{1\times10^{-2}\ \mathrm{m}}{1\ \mathrm{cm}}\right)\left(\frac{1\times10^{-2}\ \mathrm{m}}{1\ \mathrm{cm}}\right)\left(\frac{1\times10^{-2}\ \mathrm{m}}{1\ \mathrm{cm}}\right)=\left(\frac{1\times10^{-6}\ \mathrm{m}^3}{1\ \mathrm{cm}^3}\right)$$

Use the conversion to express the volume in m^3:

$$(1250.\ \mathrm{cm}^3)\times\left(\frac{1\times10^{-6}\ \mathrm{m}^3}{1\ \mathrm{cm}^3}\right)=1.250\times10^{-3}\ \mathrm{m}^3$$

A magnesium anode rod in a hot water heater prevents corrosion. Magnesium has a density of 1.74 g/cm^3. Assume a solid cylindrical rod with diameter 1.00 inch. How long of a section (in cm) must we cutoff to obtain 1.00 kg of the rod?