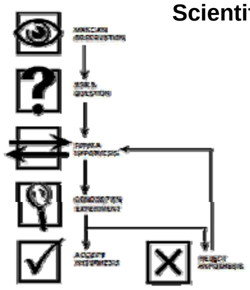
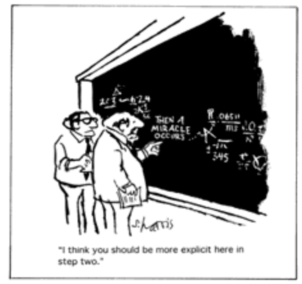
Matter and Measurement

Scientific Method





When our hypothesis successfully predicts what will happen, we designate it as a **scientific law** – a (usually) mathematical description of "here's what will happen."

A **theory** is the *explanation* for a law.

Quantitative vs. Qualitative

Qualitative = Descriptive

Quantitative = Measured

Taking measurements of real world quantities

The precision of the measurement is limited by the measuring device.

You are allowed one digit of uncertainty in any measurement and no more.

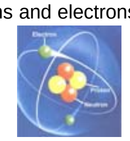
Classification of matter

Pure substances (elements and compounds)

Homogeneous mixtures (solutions)

Heterogeneous mixtures

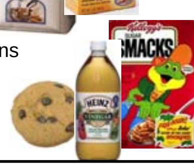
1. All atoms are made from protons, neutrons and electrons



2. All elements are made of atoms

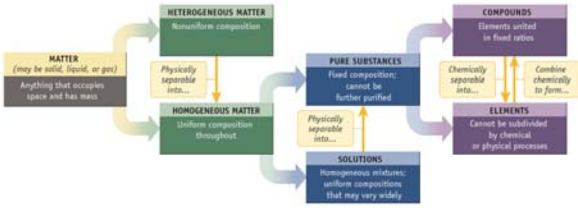
3. All compounds (molecules) are made from elements (chemically bound)

4. All Mixtures are made from combinations of elements and/or compounds (not chemically bound)



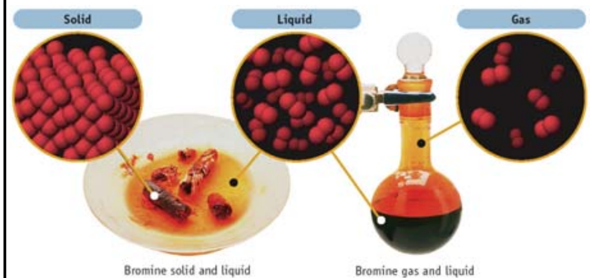
Chicken Noodle Soup	Whole Blood	Salt Water
Heterogeneous mixture	Heterogeneous mixture	Solution
(obvious)	(not so obvious with naked eye observations)	(uniform even under high magnification)

States of Matter



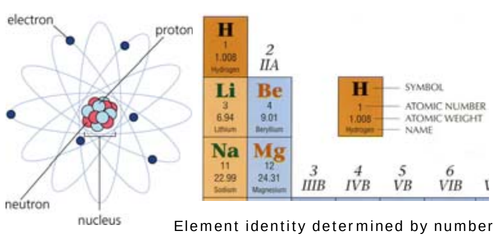
Physical separation could be: distillation, magnetic properties, density, etc.

Kinetic-molecular theory (as applied to states of matter)



Solid: Orderly, closely packed together, crystal, definite shape and volume
Liquid: More random, able to “flow”, definite volume
Gas: Least ordered, wide distances between particles, neither definite shape nor volume

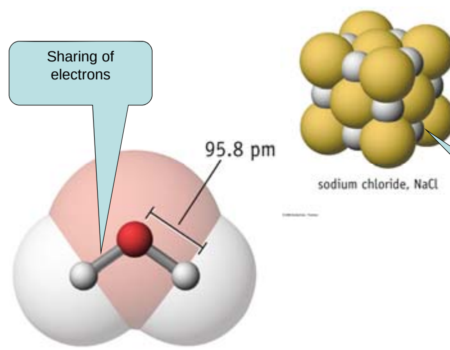
Atoms/elements (uniqueness)



Element identity determined by number of protons in the nucleus (given by the atomic number)

Ionic compound vs. molecule (covalent compound)

Sharing of electrons



Alternating charges (ions) Electrostatic attracion

Physical vs. chemical properties (Table 1.1 p21)

Chemical Properties: What a substance reacts or does not react with

Table 1.1 Some Physical Properties

Property	Using the Property to Distinguish Substances
Color	Is the substance colored or colorless? What is the color and what is its intensity?
State of matter	Is it a solid, liquid, or gas? If it is a solid, what is the shape of the particles?
Melting point	At what temperature does a solid melt?
Boiling point	At what temperature does a liquid boil?
Density	What is the substance's density (mass per unit volume)?
Solubility	What mass of substance can dissolve in a given volume of water or other solvent?
Electric conductivity	Does the substance conduct electricity?
Malleability	How easily can a solid be deformed?
Ductility	How easily can a solid be drawn into a wire?
Viscosity	How easily will a liquid flow?

Extensive vs. intensive properties

- **Intensive** - Properties that do not depend on the amount of the matter present.
  - **Color**
  - **Odor**
  - **Luster** - How shiny a substance is.
  - **Malleability** - The ability of a substance to be beaten into thin sheets.
  - **Ductility** - The ability of a substance to be drawn into thin wires.
  - **Conductivity** - The ability of a substance to allow the flow of energy or electricity.
  - **Hardness** - How easily a substance can be scratched.
  - **Melting/Freezing Point** - The temperature at which the solid and liquid phases of a substance are in equilibrium at atmospheric pressure.
  - **Boiling Point** - The temperature at which the vapor pressure of a liquid is equal to the pressure on the liquid (generally atmospheric pressure).
  - **Density** - The mass of a substance divided by its volume $D = \frac{m}{V}$
- **Extensive** - Properties that do depend on the amount of matter present.
  - **Mass** - A measurement of the amount of matter in a object (grams).
  - **Weight** - A measurement of the gravitational force of attraction of the earth acting on an object.
  - **Volume** - A measurement of the amount of space a substance occupies.
  - **Length**

SI units and prefixes

Table 1.2 Some SI Base Units

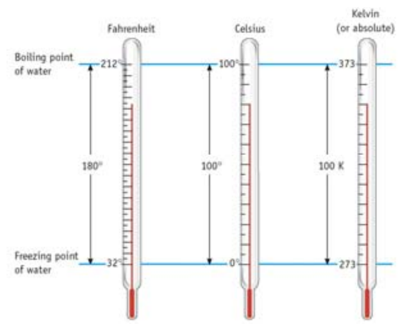
Measured Property	Name of Unit	Abbreviation
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Amount of substance	mole	mol
Electric current	ampere	A

Remember L and mL are not SI base units

Table 1.3 Selected Prefixes Used in the Metric System

Prefix	Abbreviation	Meaning	Example
mega-	M	10^6 (million)	1 megaton = 1×10^6 tons
kilo-	k	10^3 (thousand)	1 kilogram (kg) = 1×10^3 g
deci-	d	10^{-1} (tenth)	1 decimeter (dm) = 1×10^{-1} m
centi-	c	10^{-2} (one hundredth)	1 centimeter (cm) = 1×10^{-2} m
milli-	m	10^{-3} (one thousandth)	1 millimeter (mm) = 1×10^{-3} m
micro-	μ	10^{-6} (one millionth)	1 micrometer (μm) = 1×10^{-6} m
nano-	n	10^{-9} (one billionth)	1 nanometer (nm) = 1×10^{-9} m
pico-	p	10^{-12}	1 picometer (pm) = 1×10^{-12} m
femto-	f	10^{-15}	1 femtometer (fm) = 1×10^{-15} m

Temperature scales and conversions



$^oC = 5/9\ (^oF - 32)$

$^oF = 9/5\ ^oC + 32$

$K = {}^oC + 273.15$

Precision, accuracy and experimental error (percent error)

Neither precise nor accurate

Precise but not accurate

Both precise and accurate

Scientific (exponential) notation

Prefix	Symbol	Factor
tera	T	10^{12}
giga	G	10^{9}
mega	M	10^{6}
kilo	k	10^{3}
hecto	h	10^{2}
deci	d	10^{-1}
centi	c	10^{-2}
milli	m	10^{-3}
micro	μ	10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}
femto	f	10^{-15}

$10^0 = 1$

$10^1 = 10$

$10^2 = 100$

$10^3 = 1000$

$10^n = \underbrace{10 \times 10 \times 10 \ldots \times 10}_{n\ \text{times}}$ = 1 with n zeros

$10^{-1} = \frac{1}{10} = 0.1$

$10^{-2} = \frac{1}{100} = 0.01$

$10^{-3} = \frac{1}{1000} = 0.001$

Make sure to enter exponential notation correctly on your calculator!

E.g. [5][.][3][EE][6] = 5.3×10^6

Significant figures

Example	Number of Significant Figures
1.23	3; all nonzero digits are significant.
0.00123 g	3; the zeros to the left of the 1 (the first significant digit) simply locate the decimal point. To avoid confusion, write numbers of this type in scientific notation: thus, $0.00123 = 1.23 \times 10^{-3}$.
2.040 g	4; when a number is greater than 1, *all zeros to the right of the decimal point are significant.*
0.02040 g	4; for a number less than 1, only zeros to the right of the first nonzero digit are significant.
100 g	1; in numbers that do not contain a decimal point, "trailing" zeros may or may not be significant. *The practice followed in this book is to include a decimal point if the zeros are significant.* Thus, 100. is used to represent three significant digits, whereas 100 has only one significant digit. To avoid confusion, an alternative method is to write numbers in scientific notation because all digits are significant when written in scientific notation. Thus, 1.00×10^2 has three significant digits, whereas 1×10^2 has only one significant digit.
100 cm/m	Infinite number of significant digits. This is a *defined quantity*. Defined quantities do not limit the number of significant figures in a calculated result.
$\pi = 3.1415926$	The value of certain constants such as π is known to a greater number of significant figures than you will ever use in a calculation.

Computations

Addition/Subtraction:

The result of an addition and/or subtraction may contain only as many decimal digits as the number with the fewest decimal digits in the computation.

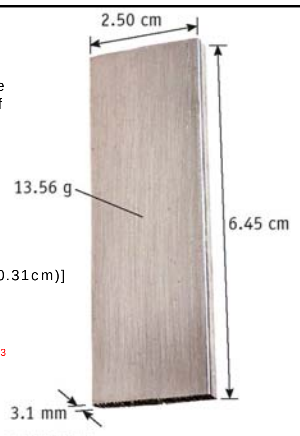
Multiplication/Division:

The result of a multiplication and/or division may contain only as many significant digits and the number with the fewest total significant digits in the computation.

When performing multiple computations, try to avoid intermediate rounding as this tends to compound the errors in your result.

Dimensional analysis

Determine the density of the bar to the correct number of significant digits in units of kilograms/m^3.



Answer:

[13.56g / (2.50cm x 6.45cm x 0.31cm)]

x (1 kg / 1,000g)(100cm /1m)^3

= 2712.67816... = 2.7×10^3kg/m^3

Example:

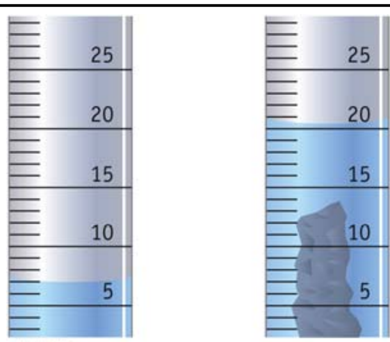
Perform the following computation to the correct number of significant digits:

(2.763 + 4.7 – 1.113)/6.775

Answer:

6.35 / 6.775 = .93726937... = 9.4×10^{-1}

Given the water displacement of the object, if the density of the sample is 1.75kg/dm^3, determine the mass (in grams) of the object to the correct number of significant digits.



Answer: 1.75 kg/dm^3 (1dm^3/1,000mL)(1,000g/1kg)(20.3mL – 7.0mL)

= 23.275g = 23.3g

Graphing