What is Chemistry?

- *Chemistry* - descriptive and quantitative study of the properties of matter
  - composition and structure
  - physical and chemical properties
  - transformations (changes in any of the above or energy)
- Chemistry is typically involved in making new materials for society, measuring the amount of matter in something, or determining the physical/chemical properties of matter

Experimentation in Chemistry

Chemistry is an Experimental Science

- **The Experiment** – observation of natural phenomena under controlled conditions such that the results can be duplicated and conclusions made
- **Law** – statement or equation describing the regularity of a fundamental occurrence in nature
- **Hypothesis** – statement of fact governing an observed natural processes testable through experimentation
- **Theory** (Model) – repeatedly tested and observed relationships in nature involving natural phenomena

Scientific Method

- Statement of the problem: statement based on observations.

H_o: The atmosphere is warming from fossil fuel emissions

- Design Experiments to test hypothesis (H_o)
  - How can temperature of troposphere be measured accurately?
  - What is the role of the control? Baseline?
- Collect data from experiment
- Analyze data statistically (relative to control)

Accept or reject hypothesis

- Provide conclusion

Matter: Physical and Chemical Make Up

- **Matter** - Anything which has mass and occupies space

  Mass = quantity of matter

  Space = volume of matter

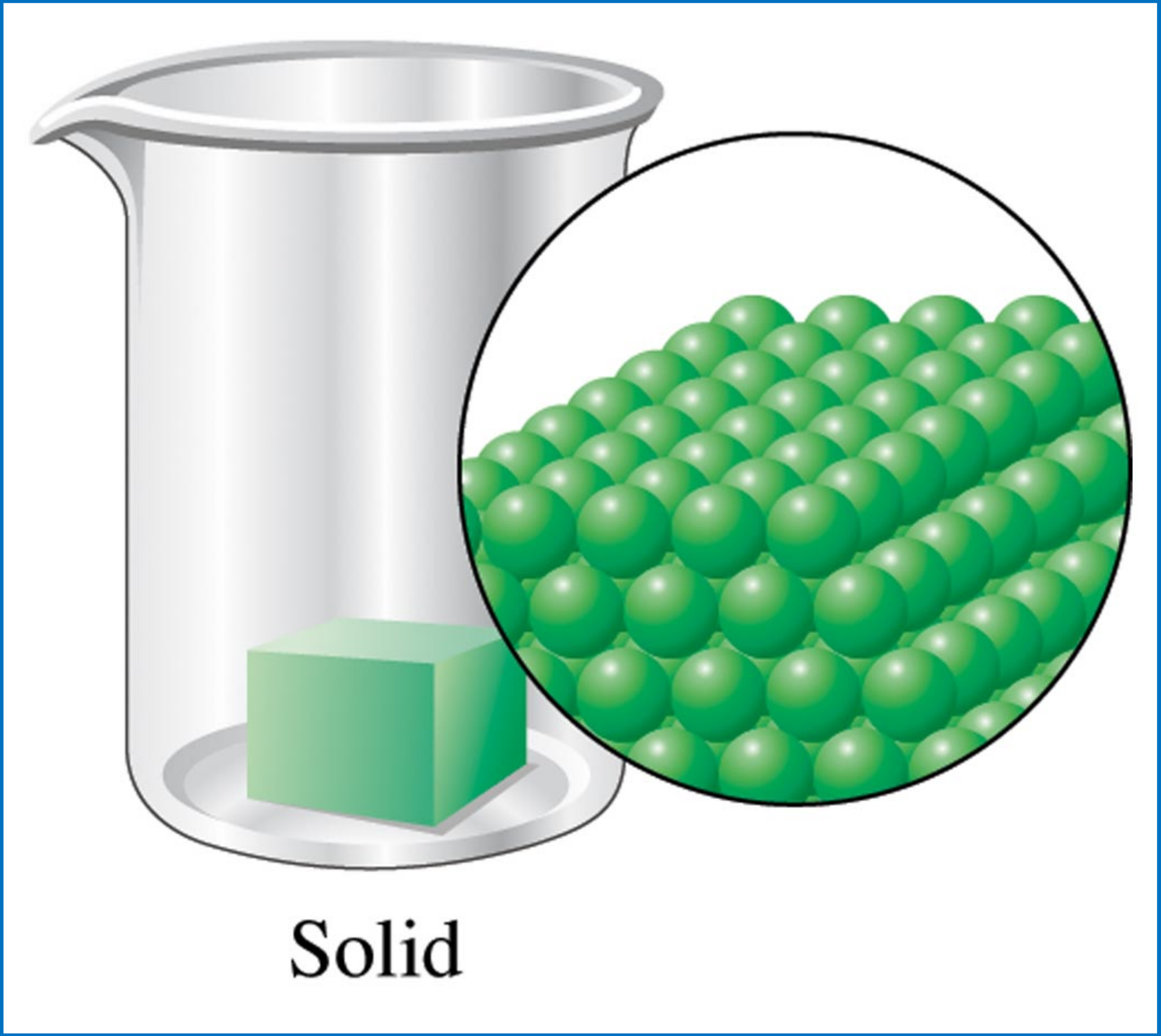
- *Physical Forms Of Matter*

  Solid - Matter with fixed shape and volume; generally incompressible

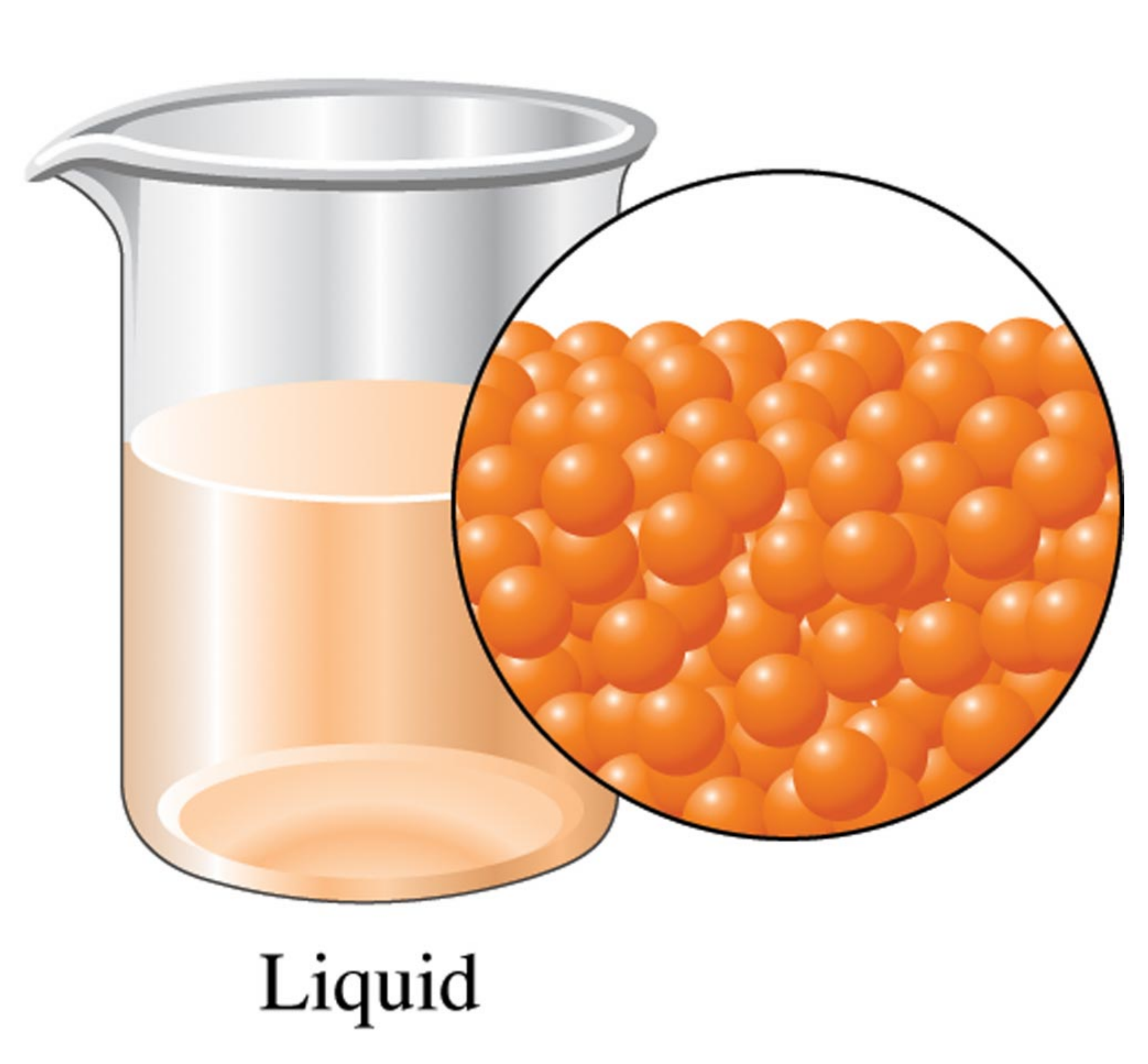
  Liquid- Matter with fixed volume and shape according to container; generally, incompressible

  Gas - Matter which conforms to the shape and volume of its container; compressible

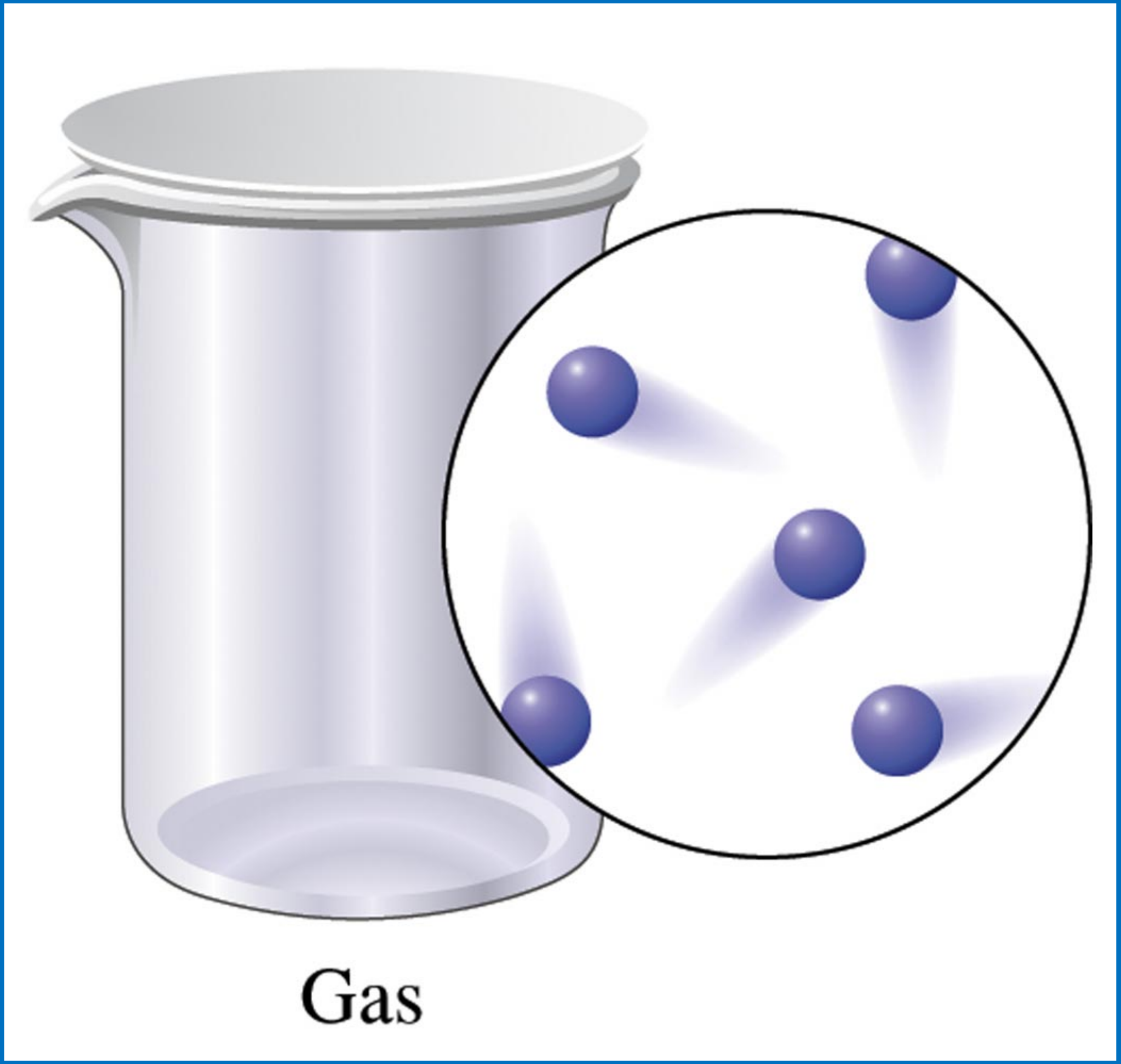
Molecular Representation of A Solid



Molecular Representation of A Liquid



Molecular Representation of A Gas



Physical & Chemical Properties

- **Physical Change** – a change in which the form of matter does not change identity
- **Physical Property** – an observed characteristic whereby the chemical form remains intact (e.g., mp, bp, density, color, refractive index)
- **Chemical Change** – change in which matter changes from one form to another or to other forms, through a chemical reaction
- **Chemical Property** - Any of a material's properties that becomes evident during a chemical reaction; that is, any quality that can be established only by changing a substance's chemical identity

Law of Conservation of Mass

- Modern chemistry emerged in 18th century upon the advent of the analytical balance; provide accurate mass measurements
- Antoine Lavoisier - French chemist who used balance measurements to show weighing substances before and after change that mass is conservative
- Law of conservation of mass - total mass remains constant during a chemical change

Law of Conservation of Mass

The total mass of the substances does not change during a chemical reaction

180 g glucose + 192 g oxygen $\rightarrow$ 264 g CO_2 + 108 g H_2O

372 g before $\rightarrow$ 372 g after

Practice Problem

Aluminum powder burns in oxygen to produce a substance called aluminum oxide (Al_2O_3)

A sample of 2.00 g aluminum is burned in oxygen and produces 3.78 g of aluminum oxide

How many grams of oxygen were used in this reaction?

Solution:

mass of aluminum + mass oxygen = mass aluminum oxide

2.00 g aluminum + x g oxygen = 3.78 g aluminum oxide

X = 3.78 g - 2.00 g = 1.78 g (Oxygen)

Elements and Compounds

- **Substance** – type of matter than cannot be further separated by physical processes
- **Element** – substance that cannot be decomposed into simpler substances
- **Compound** – a substance formed when two or more elements are combined
  - Compounds obey the

Law of Definite Proportions

A pure compound contains

constant proportions of elements by mass

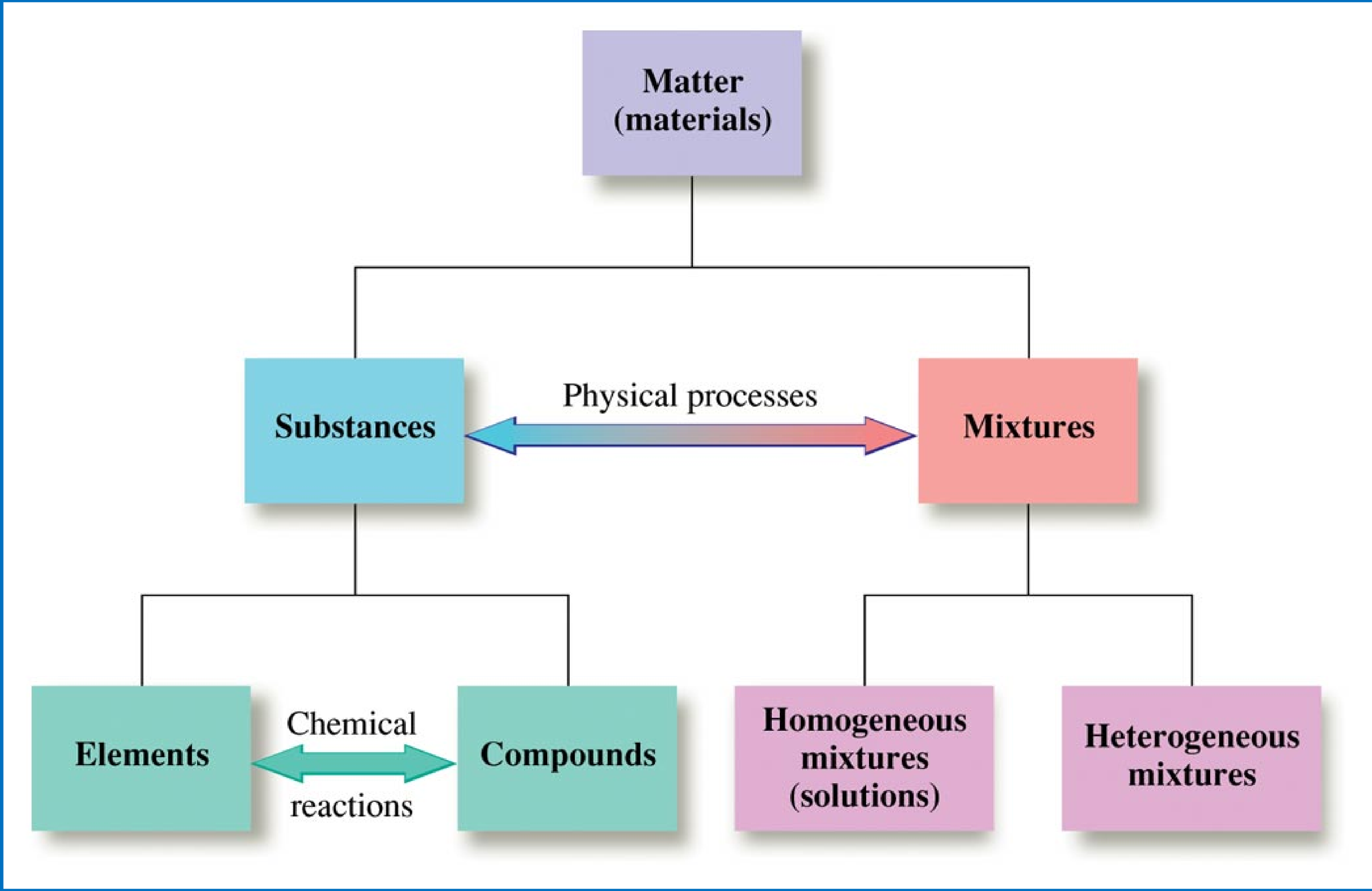
Mixtures

- **Mixture** –material that can be separated into two or more substances
  - **Heterogeneous mixture** – mixture that is divided among parts with distinctly different physical properties
  - **phase** – part of mixture with uniform properties
  - **Homogeneous mixture** – mixture with no visible boundaries and uniform physical properties throughout

Mixtures

- Mixtures separated by:
  - **Filtration**: Mixture consists of a solid and liquid; liquid separated by filtration.
  - **Chromatography**: Separates mixtures by distributing components between a mobile and stationary phase.
  - **Distillation**: Liquid mixture is boiled; components in the mixture boil off at different temperatures.

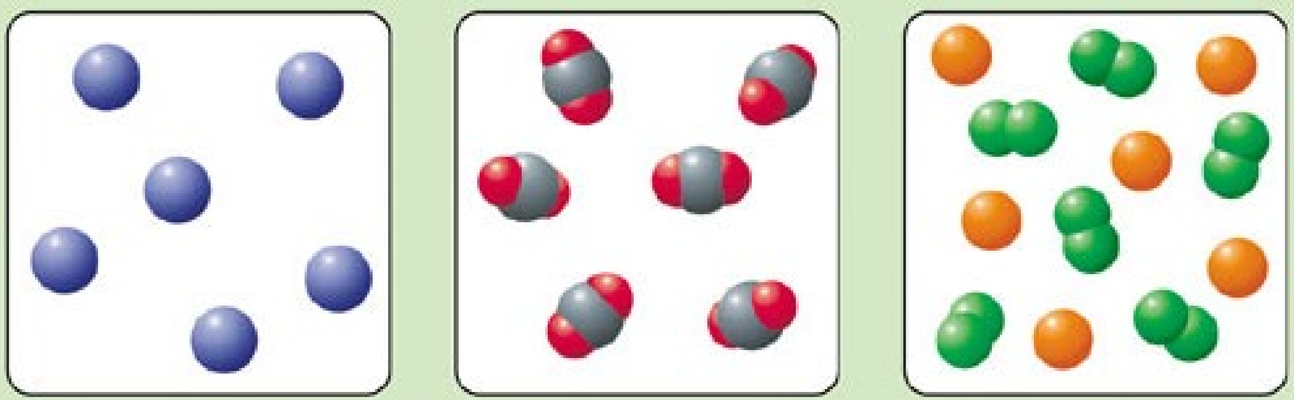
Relationships Among Elements, Compounds and Mixtures



Concept Check

Elements, Compounds, Mixtures

Matter can be represented as being composed of individual units. For example, the smallest individual unit of matter can be represented as a single circle, •, and chemical combinations of these units of matter as connected circles, ••, with each element represented by a different color. Using this model, place the appropriate label—element, compound, or mixture—on each container.



Element | Compound | Mixture

Accuracy and Precision

- Analytical data must be defined in terms of its accuracy, precision and uncertainty

  - **Accuracy** – closeness of the result to the real value (often not known); must have reference standard to determine it

  - **Precision** – reproducibility of repeated measurements of the same sample; often defined by a sample standard deviation

  - **Uncertainty** - error in a measurement; often expressed as a standard deviation

Uncertainty

- In measurements involving a 50-mL buret, the uncertainty is normally ±0.02 mL for any reading
- The first uncertain digit fixes the sig figs in the result. Buret measurements cannot be made past 2 decimal places

25.639 mL (incorrect – too many decimal places

25.6 mL (incorrect – too few decimal places

25.64 mL (± 0.02) mL

First uncertain digit corresponds to <u>last</u> significant figure

Accuracy and Precision

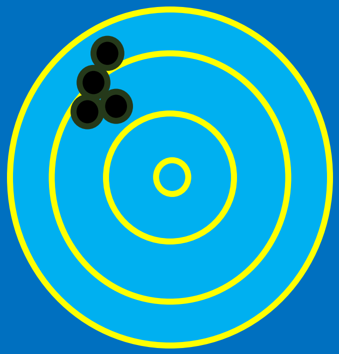
- In a series of laboratory measurements of the chemical composition of a sample, the following results were obtained as the mean ± std dev for 10 replicates of sample analysis

  If a result is known to be 9.7, describe the accuracy and precision in the context of each group of measurements

Case A:	9.5 ± 5.2	2nd most accurate	–	least precise
Case B:	7.6 ± 0.2	least accurate	–	most precise
Case C:	9.8 ± 0.3	most accurate	–	2nd most precise

Practice Problem

The figure below represents the bulls eye target for an archer. The black dots represent where the archer's arrows hit:



How can this archer best be described?

a. Accurate

b. precise

c. accurate and precise

d. neither accurate nor precise

Ans: b precise

Practice Problem

A lab instructor gives a sample of amino-acid powder to each of four students. They weigh the samples:

I: 8.72g, 8.74g, 8.70g   II: 8.56g, 8.77g, 8.83g

III: 8.50g, 8.48g, 8.51g   IV: 8.41g, 8.72g, 8.55g

The true value is 8.72g

a. Calculate average mass of each

$$I_{avg} = \frac{8.72g + 8.74g + 8.70g}{3} = 8.7200 = 8.72g$$

$$II_{avg} = \frac{8.56g + 8.77g + 8.83g}{3} = 8.7200 = 8.72g$$

$$III_{avg} = \frac{8.50g + 8.48g + 8.51g}{3} = 8.4967 = 8.50g$$

$$IV_{avg} = \frac{8.41g + 8.72g + 8.55g`3}{3} = 8.5600 = 8.56g$$

Practice Problem

b. Which set is the most accurate?

Ans: Sets I & II are closest to the true value (8.72g)

c. Which set is most precise?

Ans:

Compute range of each set

$I_{rng} = 8.74g - 8.70g = 0.04g$

$II_{rng} = 8.83g - 8.56g = 0.27g$

$III_{rng} = 8.51g - 8.48g = 0.03g$ **(most precise)**

$IV_{rng} = 8.72g - 8.41g = 0.31g$

d. Which set combines best accuracy and precision

Ans: Set I (8.72g & 0.04g)

Significant Figures

- **Significant figure** is a number derived from a measurement or calculation that indicates all relevant digits.

The final digit is <u>uncertain</u>

- Every measurement has a reporting limit (detection limit)
- The greater the number of digits usually indicates the higher the precision

Significant Figures (Con't)

1. All nonzero numbers are significant.
2. Zeros in between nonzero numbers are significant.
3. Trailing zeros (zeros to the right of a nonzero number) that fall AFTER a decimal point are significant.
4. Trailing zeros BEFORE a decimal point are not significant unless indicated w/ a bar over them or an explicit decimal point.
5. Leading zeros (zeros to the left of the first nonzero number) are not significant.

Significant Figures (Con't)

Examples: Determine the number of significant figures in each number below

1.12	(3)
0.00345	(3)
0.0300	(3)
125.999	(6)
1.00056	(6)
1000	(?, 1-4, no decimal point)
1000.	(4)

Scientific Notation

- Number of significant figures can be stated unequivocally by using **Scientific Notation**
  - In scientific notation, a number is represented by the form

$$A.bcd\ldots \times 10^{n}$$

A = A 1 digit number to the left of the decimal point (1-9)

bcd = The remaining significant figures

n = an integer that indicates how many powers of 10 the number must be multiplied by to restore the original value; n can be negative (-) or positive (+)

Practice Problem

Express each of the numbers below in terms of scientific notation and indicate no. of significant figs

12.45	1.245×10^{1}	(4 sig figs)
127	1.27×10^{2}	(3 sig figs)
0.0000456	4.56×10^{-5}	(3 sig figs)
1000	1.0×10^{3}	(2 sig figs)
131,000.0	1.310000×10^{5}	(7 sig figs)

Scientific notation removes any ambiguity in significant figures – Note Example #4

Significant Figures in Calculations

- **Multiplication and Division** – the result of multiplication or division is limited by the number with the least sig figs.

5.02	×	89.665	×	0.10	=	45.0118	=	45
(3 sig. figures)		(5 sig. figures)		(2 sig. figures)				(2 sig. figures)

12.4 x 3.1 =

12.4 x 3.1 = 38.44 = 38 (no more than 2)

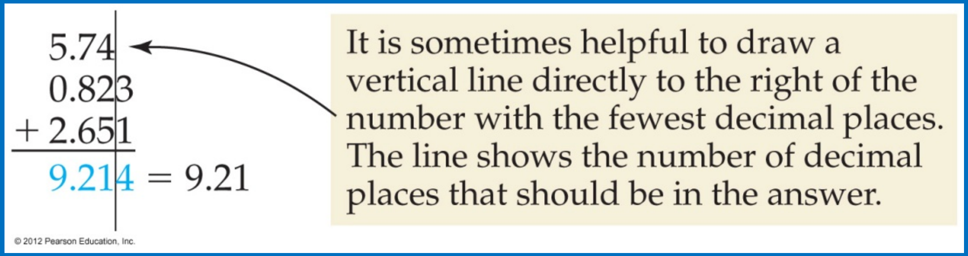
5.892	÷	6.10	=	0.96590	=	0.966
(4 sig. figures)		(3 sig. figures)				(3 sig. figures)

144 ÷ 2.6781 =

144 ÷ 2.6781 = 53.76946343 = 53.8 (3 maximum)

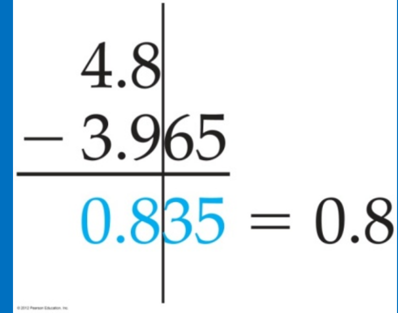
Significant Figures in Calculations

- **Addition and Subtraction** – The answer has the same number of PLACES as the quantity carrying the fewest places. *Note that the number of sig figs could increase or decrease.



12.43 + 3.1 =

12.43 + 3.1 = 15.53 = 15.5 (1 decimal place)



144 - 2.6781 =

144 - 2.6781 = 141.3219 = 141 (no more than 3)

- Addition and subtraction operations could involve numbers without decimal places.
- The general rule is: “The number of significant figures in the result of an addition/subtraction operation is limited by the least precise number.”

Significant Figures in Calculations

- **Exact Numbers** – exact numbers are numbers known without uncertainty (because they are not derived from measurement), and they have <u>no influence on the significant figures in the result</u>

12.43 x 12 (exact) =

12.43 x 12 (exact) = 149.16 = 149.2 (4 sig figs)

144.22 ÷ 3 (exact) =

144.22 ÷ 3 (exact) = 48.07333333 = 48.073 (5 sig fig)

Mixed Operations

- In calculations involving both addition/subtraction and multiplication/division, we evaluate in the proper order, keeping track of sig figs.
- DO NOT ROUND IN THE MIDDLE OF A CALCULATION!!
- Carry extra digits and round at the end.
- *e.g.* $3.897 \times (782.3 - 451.88) = ?$

Practice Problem

How many significant figures should be reported for the difference between 235.7631 and 235.57?

a. 1 b. 2 c. 3 d. 5 e. 7

Ans: b

235.7631 - 235.57 = 0.19 (2 sig figs)

235.57 is less precise than 235.7631

Rounding

- **Rounding** is the process of dropping non-significant digits in a calculation and adjusting the last digit reported
  - If the number following the last sig fig is 5 or greater, add 1 to the last digit reported and drop all digits that follow
  - If the last sig fig is <5, simply drop all digits farther to the right

14.225**8**	to 5 sig figs =	14.226
3.441**1**	to 4 sig figs =	3.441
7.7**52237**	to 2 sig figs =	7.8

Practice Problem

- Carry out the following calculation, paying special attention to sig figs, rounding, and units

$$[(1.84 \times 10^2 \text{ g})(44.7 \text{ m/s})^2] / 2$$

Ans: $1.8382 \times 10^5 = 1.84 \times 10^5 \text{ g}\bullet\text{m}^2/\text{s}^2$

3 Significant figures

Note: Assumes "2" is an <u>exact</u> number

Measurements and Units

- Measurements are reported in a variety of units, or dimensions. Units are somewhat standardized globally in the form of the International System (metric units) called SI units.
- Units are often associated with prefixes that make them more convenient to use and report.
- The most common prefixes include:

tera- = 10^{12}	giga- = 10^{9}
mega- = 10^{6}	kilo- = 10^{3}
deci- = 10^{-1}	centi- = 10^{-2}
milli- = 10^{-3}	micro- = 10^{-6}
nano- = 10^{-9}	pico- = 10^{-12}

SI Base Units

Quantity	Unit	Symbol
Length	Meter	m
Mass	Kilogram	Kg
Time	Second	s
Volume	Liter	l
Temperature	Kelvin	K
Pressure	Pascal	Pa
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd
Energy	Joule	J

Practice Problem

The earth's surface is 5.10×10^8 km^2

Its crust has a mean thickness of 35 km

The crust has a mean density of 2.8 g/cm^3

The two most abundant elements in the crust are:

Oxygen (conc: 4.55×10^5 g/metric ton)

Silicon (conc: 2.72×10^5 g/metric ton)

The two least abundant elements in the crust are:

Ruthenium (conc: 1×10^{-4} g/metric ton)

Rhodium (conc: 1×10^{-4} g/metric ton)

What is the total mass of each of these elements in the earth's crust? (1 metric ton = 1000 kg)

Con't

Practice Problem (Con't)

Mass of elements in earth's crust

Mass = Depth * Area * Density
(Volume)

$$\text{Mass of the crust} = (35\ \text{km})(5.10\times10^{8}\ \text{km}^{2})\left(\frac{(1000\ \text{m})^{3}}{(1\ \text{km})^{3}}\right)\left(\frac{(1\ \text{cm})^{3}}{(0.01\ \text{m})^{3}}\right)\left(\frac{2.8\ \text{g}}{1\ \text{cm}^{3}}\right)\left(\frac{1\ \text{kg}}{1000\ \text{g}}\right)\left(\frac{1\ \text{t}}{1000\ \text{kg}}\right)$$

$$= 4.998\times10^{19}\ \text{t} \qquad \text{Note: Intermediate answer; precision not applicable}$$

$$(4.998\times10^{19}\ \text{t})\left(\frac{4.55\times10^{5}\ \text{g Oxygen}}{1\ \text{t}}\right) = 2.2741\times10^{25} = 2.3\times10^{25}\ \text{g Oxygen}$$

Note: 2 sig figs

$$(4.998\times10^{19}\ \text{t})\left(\frac{2.72\times10^{5}\ \text{g Silicon}}{1\ \text{t}}\right) = 1.3595\times10^{25} = 1.4\times10^{25}\ \text{g Silicon}$$

$$(4.998\times10^{19}\ \text{t})\left(\frac{1\times10^{-4}\ \text{g Ruthenium}}{1\ \text{t}}\right) = 4.998\times10^{15} = 1.4\times10^{15}\ \text{g Ruthenium (and Rhodium)}$$

Temperature

- Temperature is normally quantified in any of three common units: kelvins, Celsius and Fahrenheit

K (kelvin)	=	absolute scale
Celsius (^oC)	=	water based scale
Fahrenheit (^oF)	=	mercury based scale

$$0^oC = 32^oF = 273.15^oK$$

$$100^oC = 212^oF = 373.15^oK$$

- Common temperature inter-conversions

$$K = {^oC} + 273.15$$

$$^oF = ({^oC} \times 1.8) + 32 \quad \text{or} \quad ({^oC} \times 9/5) + 32)$$

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

Temperature

- Temperature Scales – Conversions
  - Derivation of General Conversion formula

^{o}X = degrees on "X" scale

^{o}Y = degrees on "Y" scale

Let ΔT_X = Sample temperature (T_x) minus freezing point ($T_{x\ fp}$) on "X" scale

Let ΔT_Y = Sample temperature (T_y) minus freezing point ($F_{y\ fp}$) on "Y" scale

$\Delta T_{x\ ref}$ = Reference Boiling Point ($T_{x\ bp}$) minus Freezing Point ($T_{x\ fp}$) on "X" scale

$\Delta T_{y\ ref}$ = Reference Boiling Point ($T_{y\ bp}$) minus Freezing Point ($T_{y\ fp}$) on "Y" scale

Temperature

The ratio of a temperature on one scale and its equivalent on another scale is the same as the ratio of boiling point minus freezing point on one scale and its equivalent on the other scale

$$\frac{\Delta T_x}{\Delta T_y} = \frac{\Delta T_{x\,(ref)}}{\Delta T_{y\ ref}} \quad \Rightarrow \quad \left(\frac{(T_x) - (T_{x\,fp})}{(T_y) - (T_{y\,fp})} = \frac{(T_{x\ bp}) - (T_{x\,fp})}{(T_{y\,bp}) - (T_{y\,fp})} \right)$$

$$\text{Ex. } 35^0\,C \text{ on Centigrade (X) scale } = \ ? \text{ on Fahrenheit (Y) scale}$$

$$(T_{x\ bp}) = 100.0; \quad (T_{x\ fp}) = 0.0 \qquad (T_{y\,bp}) = 212.0; \quad (T_{y\,fp}) = 32.0$$

$$(T_x) = 35.0^0\,C; \quad (T_y) = ?^0\,F$$

$$\frac{35.0^0\,C - 0.0^0\,C}{T_y - 32.0^0\,F} = \frac{100.0^0\,C - 0.0^0\,C}{212.0^0\,F - 32.0^0\,F} \qquad \frac{35.0^0\,C}{T_y - 32.0^0\,F} = \frac{100.0^0\,C}{180.0^0\,F} \qquad \frac{35.0^0\,C}{T_y - 32.0^0\,F} = \frac{5^0\,C}{9^0\,F}$$

$$\text{Rearrange:} \quad T_y = 35.0^0\,C\left(\frac{9^0\,F}{5^0\,C}\right) + 32.0^0\,F = 63.0^0\,F + 32.0^0\,F = 95.0^0\,F$$

Volume & Density

- Volume is the amount of 3-D space matter occupies, and is described as length-cubed

1 mL = 1 cm^3

1 dm = 10 cm

1 L = 1 dm^3 = 1000 cm^3 = 1000 mL

- **Density** (d) is mass per unit volume

d = mass (g)/volume (mL)

D = g/mL = g/cm^3

Practice Problem

The density of 1.59 mL of a solution is 1.369 g/mL. What is the mass of the solution?

$$d = m / v$$

$$m = d \times v$$

$$m = 1.369\ g/\cancel{mL} \times 1.59\ \cancel{mL}$$

$$m = 2.18\ g \quad \text{(3 sig figs)}$$

Practice Problem

The volume of a 30.0% (by mass) sodium bromide solution is 150.0 mL

The density of the solution is 1.284 g/mL

What is the mass of solute in this solution?

Mass = Density (g/mL) x Volume (mL)

m = d x v

m = 1.284 g/~~mL~~ x 150.0 ~~mL~~ sol'n = 192.6 g sol'n

The solute is 30% (by mass) of the solution

192.6 g sol'n x 30.0/100 = 57.8 g (3 sig figs)

Practice Problem

An empty vial weighs 55.32 g

a. If the vial weighs 185.56 g when filled with mercury (d = 13.53 g/cm^3), what is its volume?

$$\mathbf{m = 185.56g - 55.32g = 130.24g \ (mass\ of\ mercury)}$$

$$\mathbf{v = \frac{m}{d} = \frac{130.24\,g}{13.53\,\frac{g}{cm^3}} = 9.926\,cm^3 \ (Vol\ of\ mercury,\ also\ Vol\ of\ flask)}$$

b. How much would the vial weigh if filled with water?

(density of water – 0.997 g/cm^3)

$$\mathbf{55.32g + (9.626\,cm^3 \times 0.997\,g/cm^3) = 55.32g + 9.597g = 64.92\,g}$$

Dimensional Analysis
Unit Factor Calculations

- **Dimensional Analysis** – method of calculation where units are canceled to obtain the result
- The fundamental parameter in dimensional analysis is the conversion factor
- A **conversion factor** is a ratio used to express a measured quantity in different units
- A conversion factor used in dimensional format converts one unit to another
- Conversion factors can be strung together indefinitely in the calculation of a result

Conversion Factors

- The ratio (3 feet/1 yard) is called a

  conversion factor

  - The conversion-factor method may be used to convert any unit to another, provided a conversion equation exists

    3 feet = 1 yard

    3 feet/1 yard = 1 yard/3 feet = 1

  - Relationships between certain U.S. units and metric units are given in Table 1.5 of text

Dimensional Analysis

- Dimensional analysis is the method of calculation in which one carries along the units for quantities
- Suppose you simply wish to convert 20 yards to feet

$$20\ \cancel{\text{yards}} \times \frac{3\ \text{feet}}{1\ \cancel{\text{yard}}} = 60\ \text{feet}$$

- Note that the "yard" units have cancelled properly to give the final unit of feet

Examples of Common Conversion Factors

Length	Mass	Volume
1 in = 2.54 cm	1 lb = 0.4536 kg	1 qt = 0.9464 L
1 yd = 0.9144 m	1 lb = 16 oz	4 qt = 1 gal
1 mi = 1.609 km	1 oz = 28.35 g	1 dm^3 = 1 L
1 mi = 5280 ft	1 tonne = 10^3kg	1 m^3 = 35.3 ft^3
	1 tonne = 2,204.6 lbs	1 L = 1000 mL
	1 ton (US) = 2000 lbs	1 mL = 1 cm^3

Unit Conversion Example

- Sodium Hydrogen Carbonate (baking soda) reacts with acidic materials such as Vinegar to release Carbon Dioxide gas. Given an experiment calling for 0.348 kg of Sodium Hydrogen Carbonate, express this mass in milligrams.

$$0.348\ \text{kg} \times \frac{10^3\ \text{g}}{1\ \text{kg}} \times \frac{10^3\ \text{mg}}{1\ \text{g}} = 3.48 \times 10^5\ \text{mg}$$

Unit Conversion

- Mass of element in compound
- Ex. If 84.2 g of Pitchblend contains 71.4 g Uranium, find the mass (kg) of uranium in 102 kg of Pitchblend

$$\text{Mass Uranium} = \text{Mass Pitchblend} \times \frac{\text{Mass Uranium in Pitchblend}}{\text{Mass Pitchblend}}$$

$$\text{Mass U} = 102\,\text{kg Pitchblend} \times \frac{71.4\,\text{kg Uranium}}{84.2\,\text{kg Pitchblend}} = 86.5\,\text{kg U}$$

Practice Problem

6. An empty Erlenmeyer flask weighs 241.3 grams

   When filled with Water (d = 1.00 g/cm^3), the flask and its contents weigh 489.1 g

   a. What is the volume of the flask

$$\text{mass of water(m)} = 489.1\text{ g} - 241.3\text{ g} = 247.8\text{ g}$$

$$\text{volume(v)} = \frac{m}{d} = \frac{247.8\text{ g}}{\frac{1\text{g}}{1\text{cm}^3}} = 247.8 = 248.\text{ cm}^3$$

   b. How much does the flask weigh when filled with Chloroform ($CHCl_3$) (d = 1.48 g/cm^3)?

$$\text{mass (m)}(CHCl_3) = v * d = (247.8\text{ cm}^3)\left(\frac{1.48\text{ g}}{\text{cm}^3}\right) = 366.744\text{ g}$$

$$(\text{flask} + CHCl_3) = 241.3 + 366.744\text{ g} = 608.044\text{ g} = 608\text{ g}$$

Practice Problem

A fictitious unit of length called the "zither" is defined by the relation 7.50 cm = 1.00 zither. A 100.0 m distance (1 m = 100 cm) would be described as:

a. 133 zither

b. 266 zither

c. 750 zither

d. 1.330×10^3 zither

e. $7.5e^4$ zither

Ans: d

$$= 100.0 \text{ m} \times 100 \text{ cm/m} \times 1 \text{ zither}/7.5 \text{ cm}$$

$$= 1.330 \times 10^3 \text{ zither}$$

Practice Problem

Grunerite has a tensile strength of 3.5×10^2 kg/mm^2

The tensile strengths of Aluminum and Steel No. 5137 are 2.5×10^4 lb/in^2 and 5.0×10^4 lb/in^2, respectively

Calculate the cross-sectional area (in mm^2) of wires of Aluminum and of Steel No. 5136 that have the same tensil strength as a fiber of Grunerite with a cross-sectional area of 1.0 um^2

Calculate the mass of Grunerite that can be held up by 1.0 um^2 of grunerite :

$$(1.0\ \text{um}^2)\left(\frac{(1\times10^{-6}\ \text{m})^2}{(1\ \text{um})^2}\right)\left(\frac{(1\ \text{mm})^2}{1\times10^{-3}\ \text{m})^2}\right)\left(\frac{3.5\times10^2\ \text{kg}}{1\ \text{mm}^2}\right) = 3.5\times10^{-4}\ \text{kg}$$

Calculate the area of aluminum required to match that mass :

$$(3.5\times10^{-4}\ \text{kg})\left(\frac{2.205\ \text{lb}}{1\ \text{kg}}\right)\left(\frac{1\ \text{in}^2}{2.5\times10^4\ \text{lb}}\right)\left(\frac{(2.54\ \text{cm})^2}{(1\ \text{in})^2}\right)\left(\frac{(10\ \text{mm})^2}{(1\ \text{cm})^2}\right) = 1.9916\times10^{-5} = 2.0\times10^{-5}\ \text{mm}^2$$

Calculate the area of Steel 5137 required to match that mass :

$$(3.5\times10^{-4}\ \text{kg})\left(\frac{2.205\ \text{lb}}{1\ \text{kg}}\right)\left(\frac{1\ \text{in}^2}{5.0\times10^4\ \text{lb}}\right)\left(\frac{(2.54\ \text{cm})^2}{(1\ \text{in})^2}\right)\left(\frac{(10\ \text{mm})^2}{(1\ \text{cm})^2}\right) = 9.9580\times10^{-6} = 1.0\times10^{-5}\ \text{mm}^2$$

Equation Summary

Common temperature inter-conversions

K = ^{o}C + 273.15

^{o}F = (^{o}C x 1.8) + 32 or (^{o}C x 9/5) + 32)

^{o}C = 5/9 x (^{o}F – 32)

Temperature Scale conversion

$$\frac{\Delta T_x}{\Delta T_y} = \frac{\Delta T_{x\,(ref)}}{\Delta T_{y\ ref}} \qquad \left(\frac{(T_x) - (T_{x\ fp})}{(T_y) - (T_{y\ fp})} = \frac{(T_{x\ bp}) - (T_{x\ fp})}{(T_{y\ bp}) - (T_{y\ fp})} \right)$$