

AP Chemistry Notes - Chapter 1

Chemistry Notes - Chapter 1, 2, 3, & 4

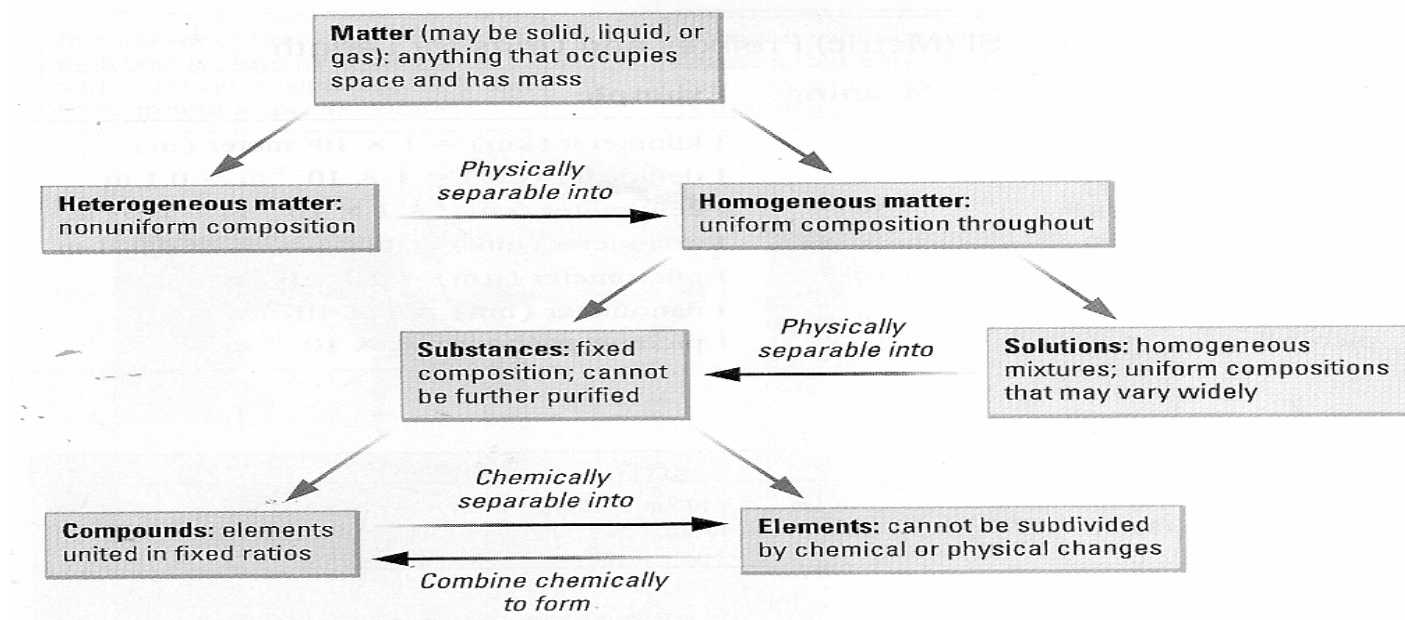
Chemistry is the study of the composition of matter- the stuff things are made of- and the changes matter undergoes. Matter is ultimately composed of chemical elements (atoms) and their compounds. Most matter can exist in one or more states: solid, liquid, and gas.

Areas of Study in Chemistry:

Organic Chemistry	Study of substances containing carbon.
Inorganic Chemistry*	Study of substances NOT containing carbon.
Analytical Chemistry*	Study of the composition of substances.
Physical Chemistry*	Study of the theories and experiments that describe the behavior of substances.
Biochemistry	Study of the chemistry of living things.

I. Matter and Measurement

A. Matter: Elements, Compounds, or Mixtures



(20+ million)

Elements united into fixed ratios

Table salt (NaCl)

Water (H₂O)

Sand (SiO₂)

Sugar (C₁₂H₂₂O₁₁)

Fool's Gold (FeS₂)

117 elements

Cannot be subdivided by chemical or Physical processes

Sodium (Na) Silicon (Si)

Chlorine (Cl) Oxygen (O₂)

Hydrogen (H₂) Iron (Fe)

*When elements become part of a compound, their original properties such as color, hardness, melting point are replaced by the characteristic properties of the new compound. Consider table salt (sodium chloride)

B. Elements and Atoms

117 elements are known - of these only 90 are found in nature; the remainder being created by scientists in a laboratory using techniques of modern physics.

Many elements names often have names and symbols relating to their Latin or Greek origin

<u>Element</u>	<u>Symbol</u>	Latin or <u>Greek</u>
Iron	Fe	Ferrum
Lead	Pb	Plumbum
Tin	Sn	Stannum
Mercury	Hg	Mercurum (Hydragyrum - Greek)
CopperCu		Cuprum
Silver	Ag	Argentum
Potassium	K	Kalium
Sodium	Na	Natrium

Note: Elements can have a one or two letter symbol. Only the first letter of an element's symbol is capitalized.

Elements are listed in the Periodic Table. An atom is the smallest particle of an element that retains the properties (color, hardness, melting point...) of that element.

Element: A substance that CANNOT be changed into a simpler substance (see periodic table)

Compound: A substance that CAN be changed into a simpler substance

Mixture: A physical blend of two or more substances (elements or compounds) that are not chemically combined

Review\Practice

Identify the following as an element, mixture, compound then identify each as a homogeneous or heterogeneous.

	Element, Mixture, Compound	If Mixture Homogeneous or Heterogeneous
1. Oxygen	_____	_____
2. Air	_____	_____
3. blood	_____	_____
4. brass (a blend of copper and zinc)	_____	_____
5. Food coloring	_____	_____
6. Soda	_____	_____
7. Silver	_____	_____
8. Salt Water	_____	_____
9. Oil and Water	_____	_____
10. Candle Wax ($C_{20}H_{42}$)	_____	_____
11. Ice	_____	_____
12. Carbon monoxide (CO)	_____	_____
13. Sugar ($C_{12}H_{22}O_{11}$)	_____	_____

II. The Measures of Science - Observations of Matter

Doing chemistry requires observing elements and compounds undergoing chemical and physical changes.

A. Qualitative Measurements

Observations that involve no measurements or units.

B. Quantitative Measurements

Observations that involve measurements or numbers w/ units.

Classify each observation as qualitative or quantitative:

Blue-Grey color _____
 Melts at 1410°C _____
 Brittle _____

1. SI System- International System of Units

The scientific community has chosen a modified version of the metric system as the standard system

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

Prefixes are used to change SI units by powers of 10. To use the SI units effectively, you should know the meanings of the prefixes.

Prefix	Symbol	Multiplier	Exponential	Example
Femto	f	1/1 000 000 000 000 000	1×10^{-15}	femtosecond (fs)
Pico	p	1/1 000 000 000 000	1×10^{-12}	picometer (pm)
Nano	n	1/1 000 000 000	1×10^{-9}	nanometer (nm)
Micro	~	1/1 000 000	1×10^{-6}	microgram (~g)
Milli	m	1/1 000	1×10^{-3}	milligram (mg)
Centi	c	1/100	1×10^{-2}	centimeter (cm)
Deci	d	1/10	1×10^{-1}	deciliter (dL)
Kilo	k	1000	1×10^3	kilometer (km)
Mega	M	1 000 000	1×10^6	megagram (Mg)
Giga	G	1 000 000 000	1×10^9	gigameter (Gm)
Tera	T	1 000 000 000 000	1×10^{12}	terameter (Tm)

When choosing the best unit and prefix for a measurement you need to choose the correct unit and then the prefix that will give you a measurement that is not extremely large or small.

Radius of an atom	1×10^{-12} m	so...	1 pm
Diameter of a virus	1×10^{-6} m	so...	1 ~m
Circumference of Earth	1×10^7 m	so...	10 Mm

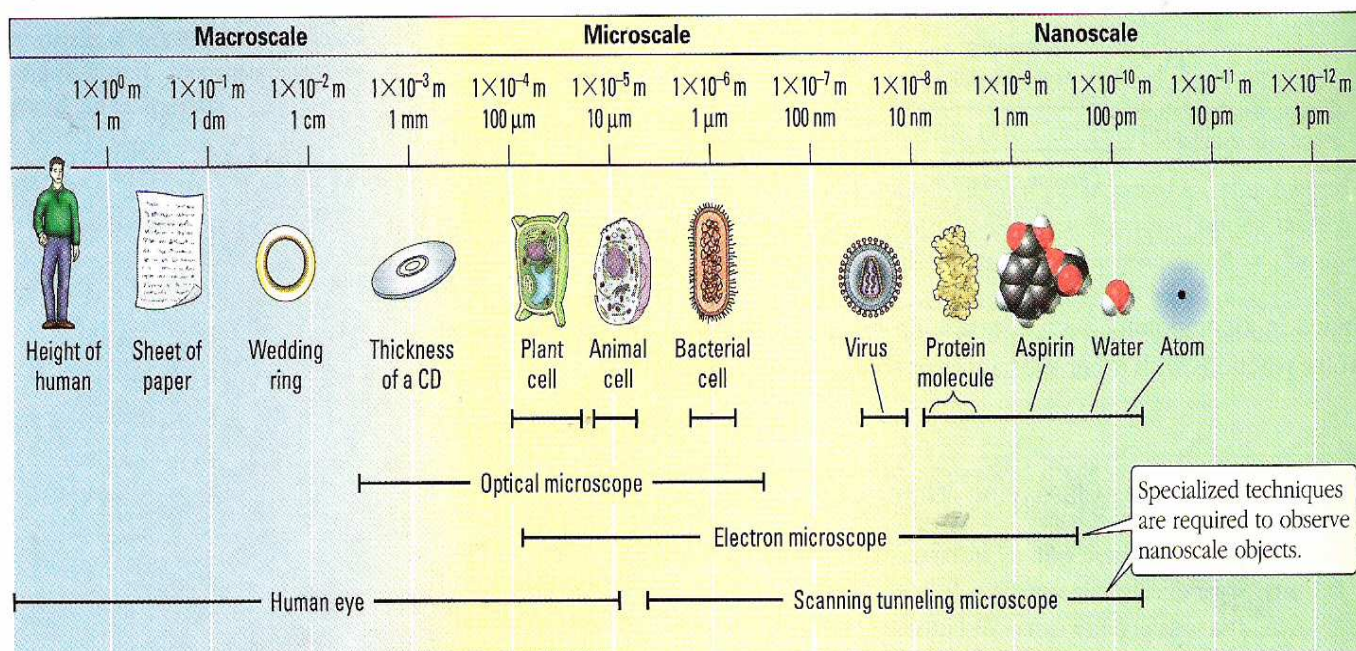


Figure 1.14 Macro, micro, and nano scale.

2. Significant Digits in Measurements

The valid digits (all certain plus one uncertain/estimated) in a measurement are called **significant digits**.

Rules for determining the number of significant figures in a measurement:

1. Non zero numbers are always significant.

2.84 km _____ sig figs

1.8×10^6 mL _____ sig figs

2. All final zeros after a decimal point are significant.

0.760 s _____ sig figs

75.00 kg _____ sig figs

3. Leading zeros used solely as placeholders are not significant.

0.00345 _____ sig figs

0.023 _____ sig figs

4. Zeros between two other significant digits are always significant.

2804 m _____ sig figs

0.003086 _____ sig figs

5. Zeros located at the end of a number and to the left of a decimal point are significant.

20. °C _____ sig figs

3000. K _____ sig figs

20 °C _____ sig figs

3000 K _____ sig figs

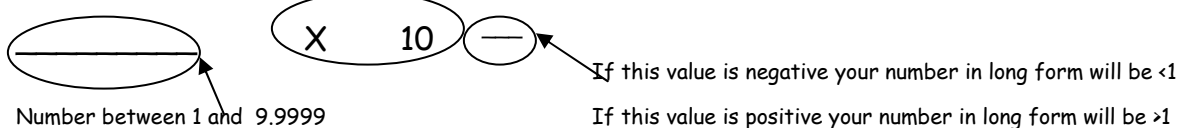
Practice:

Complete the table below

Measurement	Number of Sig. Figs	Measurement	Number of Sig. Figs
2804 m	_____	30.9 V	_____
2.84 km	_____	0.001 010 0 s	_____
0.0029 m	_____	0.106 W	_____
4.6×10^5 m	_____	0.160 A	_____
4.06×10^{-5}	_____	5.20 N	_____
4.060×10^{-5}	_____	30.9 V	_____
250 600 m	_____	5.2234×10^{12} mol	_____
250 600. m	_____	4.2×10^6 g	_____

3. Scientific (Exponential) Notation

Many measurements are very large or very small numbers. Written in this form, the values of the quantities take up too much space, are difficult to read, and are awkward to use in calculations. To work with such numbers, write them in **scientific (exponential)** notation by expressing decimal places as powers of ten.



The average distance from the sun to Mars is 227 800 000 000 m. To write this number in correct scientific notation the decimal is moved to the left 11 places. In scientific notation this distance would be 2.278×10^{11} m.

The mass of a single electron is about 0.000 000 000 000 000 000 000 000 000 911 kg. To write this number in correct scientific notation the decimal is moved to the right 31 places. In scientific notation this mass would be 9.11×10^{-31} m.

Review/Practice:

1. Correct the measurements so that they are in PROPER scientific notation THEN use a prefix rather than scientific notation:

Incorrect Scientific Notation	Correct Scientific Notation	Using Prefix
$128. \times 10^3$ m	_____	_____
0.009320×10^5 K	_____	_____
2345×10^{-2} g	_____	_____
0.562×10^{-5} s	_____	_____

2. Complete the Table

Long Form		Number of Sig figs	Scientific Notation
0.0001267	m	4	1.267×10^{-4} m
10022.5	kg	_____	_____
0.034500	s	_____	_____
_____	A	_____	2.43×10^5 A
_____	mol	_____	1.23450×10^{-3} mol
200	kg	_____	_____
200.	kg	_____	_____
200.0	kg	_____	_____
_____	s	_____	$1,243 \times 10^3$ s
_____	A	_____	5.6×10^4 A

3. Using the five rules of significant figures, use of scientific notation, and use of prefixes - complete the following table:

Long Form	#Sig Figures	Scientific Notation	Unit w/ Prefix
a) 0.004080 g	_____	_____	_____
b) 0.0180 m	_____	_____	_____
c) 3450000 s	_____	_____	_____
d) 0.0000012 m	_____	_____	_____
e) _____	_____	_____	3.3 dm
f) _____	_____	1.67×10^2 g	_____
g) _____	_____	2.33×10^{-3} s	_____
h) _____	_____	_____	2.35 μ s

4. Place the following measurements in order from smallest to largest

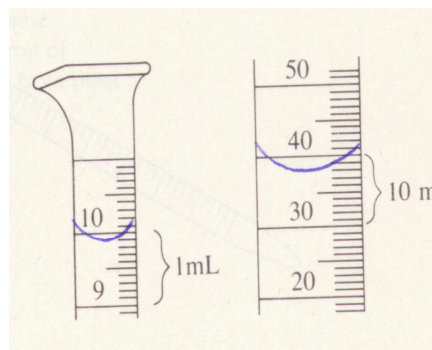
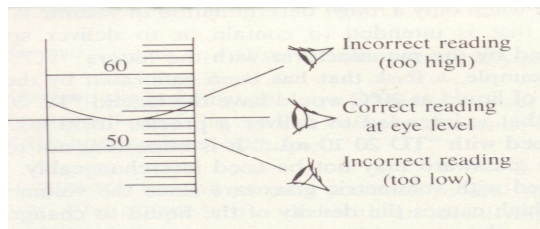
1.0 km, 1.0 Tm, 1.0 cm, 1.0 m, 1.0 nm, 1.0 μ m _____

4. Significant Figures When Measuring with Laboratory Equipment

When recording measurements taken during an experiment, you must always record them with the correct number of significant figures (**all certain plus one uncertain/estimated digit**)

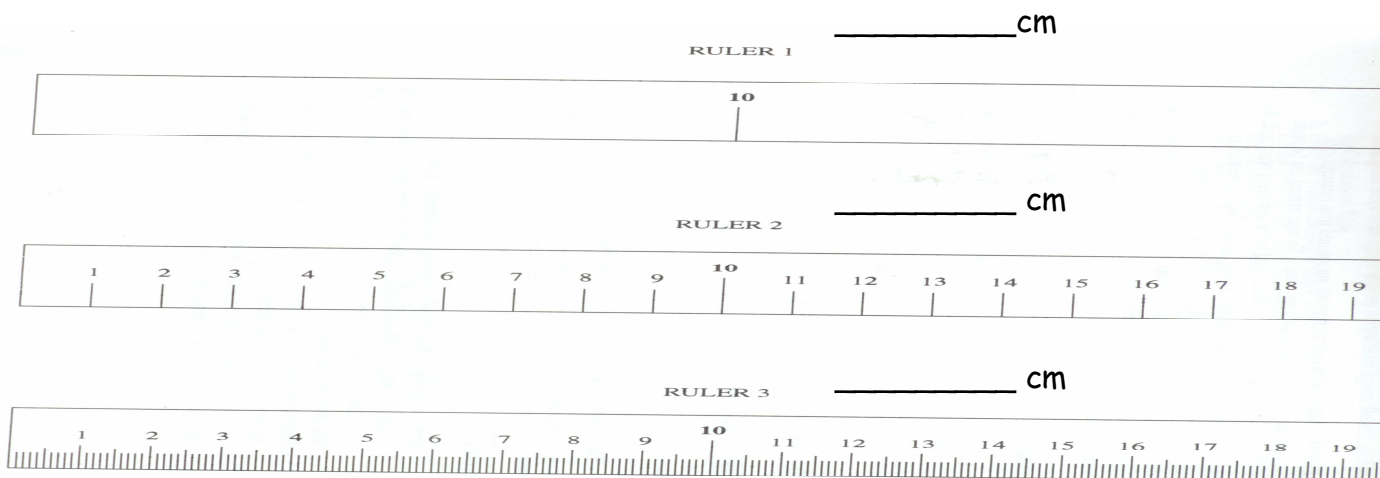
Measuring Volume:

Using the graduated cylinders below report the correct volume of liquid



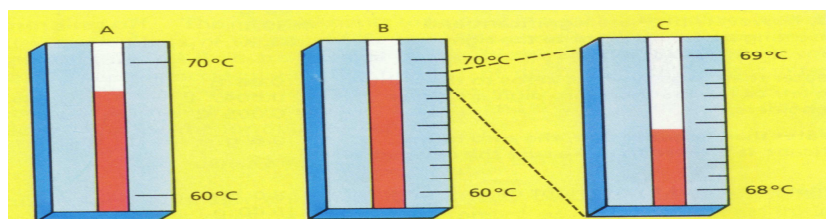
Measuring Length:

Using the rulers below report the length of your pencil



Measuring Temperature:

Using the thermometers below report the correct temperature



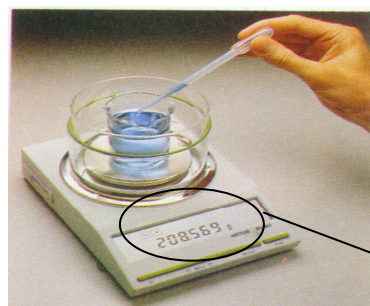
_____ °C

_____ °C

_____ °C

Measuring Mass:

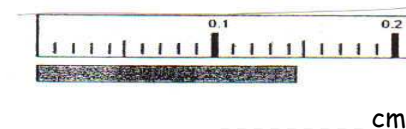
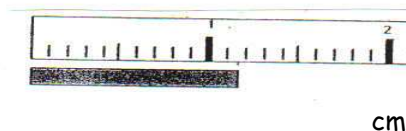
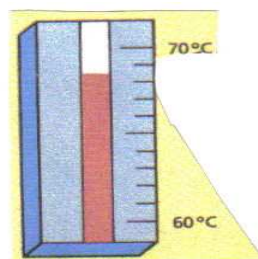
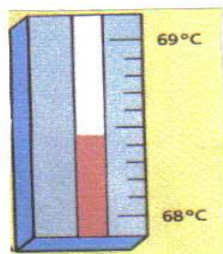
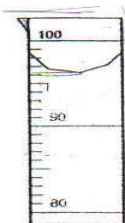
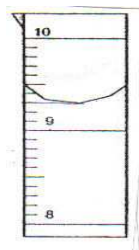
Electronic balances always report all significant digits. This means that the last digit reported by an electronic balance is uncertain/estimated but it is still significant



208.569 g

Review\ Practice:

Read the following instruments. Express your measurement using the correct number of significant figures



_____ mL _____ mL

_____ °C _____ °C

5. Arithmetic with Significant Figures

As mentioned previously, when recording the results of an experiment, you must always record them with the correct number of significant figures. Frequently you will need to add, subtract, multiply, or divide the measurements. **When you perform these arithmetic operations, it is important to remember that the result can never be more precise than the least precise measurement. YOUR CALCULATOR DOES NOT DO THIS FOR YOU!!!**

Rule #1: Addition and Subtraction:

To add or subtract measurements, first perform the operation, then **round off the answer to correspond with the least precise measurement**

$$\begin{array}{r}
 24.686 \quad \text{m} \\
 2.341 \quad \text{m} \\
 + 3.2 \quad \text{m} \\
 \hline
 \end{array}
 \qquad
 \begin{array}{r}
 2.456 \quad \text{s} \\
 - 0.03 \quad \text{s} \\
 \hline
 \end{array}$$

Rule #2: Multiplication and Division

A different rule governs multiplication and division. After performing a calculation, note the measurement with the least number of significant figures and round your final answer to this number of significant figures.

$$3.22 \text{ cm} \times 2.1 \text{ cm} =$$

$$36.5 \text{ m} / 3.414 \text{ s} =$$

Review/Practice

1. Round each measurement to three significant figures.

a) 98.473 L _____ b) 12.17 °C _____ c) 57.048 m _____

d) 4323.34 s _____ e) 4327 s _____ f) 0.0007635211 L _____

(Use Sci. Notation)

2. Complete the table

Report answers to the correct number of significant figures and include correct units with each answer):

	Long Form	Scientific Notation
1. 6.201 cm + 7.4 cm + 0.68 cm + 12. cm =	_____	_____
2. 1.6 km + 0.0162 km + 1.2 km =	_____	_____
3. 10.8 g - 8.264 g =	_____	_____
4. 475 m - 0.4168 m =	_____	_____
5. 131 cm x 2.3 cm =	_____	_____
6. 5.761 N x 6.20 m =	_____	_____
7. 13.78 g / 11 mL =	_____	_____
8. 3.1416 cm / 12.4 s =	_____	_____
9. (1.68) ($\frac{23.56 - 2.3}{1.248 \times 10^3}$) =	_____	_____
10. (6.2 x 10 ¹⁸ m) (4.7 x 10 ⁻¹⁰ m) =	_____	_____
11. (6.5 x 10 ⁵ kg) / (3.4 x 10 ³ m ³) =	_____	_____

3. The following operations were completed on a calculator. The answer the calculator provided is provided for you. Please round the provided answer to the correct number of significant figures with the correct units.

	Calculator Answer	Rounded to Correct num. of SF's
a) 3.46 cm + 104.5 cm + 0.346 cm =	_____	_____
b) 2.384 g - 1.5 g =	_____	_____
c) 9.40 mm x 2.6 mm =	_____	_____
d) 1.50 g / 2 cm ³ =	_____	_____
e) 21.50 g / (4.06cm x 1.8 cm x 0.905cm) =	_____	_____

6. Unit Conversions (dimensional analysis)

Many time the measurements you collect in laboratory are not the units you desire. This will require you to convert a measurement to another unit using **dimensional analysis**.

$\text{Measurement in original unit} \times \frac{[\text{new unit}]}{\text{original unit}} = \text{Measurement in new units}$ <p style="text-align: center;"> ↑ Conversion factor (s) </p>

a. Converting Units Within the Metric System

Example 1: Convert 456 to kg

$$456 \text{ g} \quad \frac{1 \text{ kg}}{1000\text{g}} \quad =$$

Example 2: Convert 2.1 m to cm

$$2.1 \text{ m} \quad \frac{\text{cm}}{\text{m}} =$$

Example 3: Convert 76.2 pm to mm

$$76.2 \text{ pm} \quad \frac{\text{m}}{\text{pm}} \quad \frac{\text{mm}}{\text{m}} =$$

b. Converting Ratios

Example 1: Convert 65 mi/hr to m/s

$$65 \frac{\text{mi}}{\text{hr}} \quad \frac{\text{km}}{\text{mi}} \quad \frac{\text{m}}{\text{km}} \quad \frac{\text{hr}}{\text{min}} \quad \frac{\text{min}}{\text{s}} = \text{m/s}$$

Example 2: Convert 42.3 g/mL to mg/L

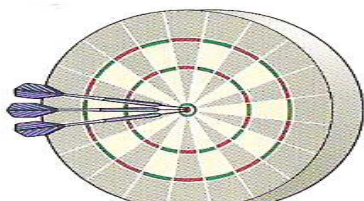
$$42.3 \frac{\text{g}}{\text{mL}} \quad \frac{\text{mg}}{\text{g}} \quad \frac{\text{mL}}{\text{L}} = \text{mg/L}$$

c. Converting Volumes and Areas

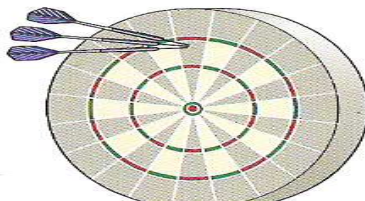
Example 1: Convert 6.2 m² to mm²

Example 2: Convert 6.34 x 10⁻⁸ cm³ to m³

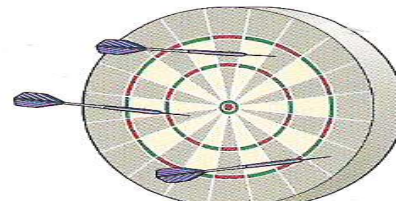
C. Accuracy and Precision of Laboratory Measurements



(a) Good accuracy
Good precision



(b) Poor accuracy
Good precision



(c) Poor accuracy
Poor precision

In laboratory we will often do several quantitative measurements of on a sample of matter (i.e. density, melting point, boiling point)

The **PRECISION** of a set of measurements is how well several measurements of the same quantity agree. The precision of a set of measurements is often expressed by the **average deviation and/or percent deviation**. That is we calculate the difference between each experimental result and the average result. The differences, each expressed as a positive quantity, are averaged, and the experiments results are reported as the average value (\bar{x}) the average deviation.

$$\% \text{ deviation} = \frac{\text{avg deviation}}{\text{avg. of experimentally determined values}} \times 100\%$$

If percent deviation is less than 20% you are accurate

ACCURACY is the agreement of a measurement with the accepted value of a quantity. Measurements can be precise but not accurate. The accuracy of a set of measurements is often expressed by the percent error. If percent error is less than 20% you are accurate

$$\% \text{ error} = \frac{|\text{avg of experimentally determined values} - \text{accepted value}|}{\text{accepted value}} \times 100\%$$

If percent error is less than 20% you are accurate

Example:

Pam makes four measurements of the diameter of a coin using a micrometer. Max measures the same coin using a simple plastic ruler. The true measurement is 27.00 mm. They report the following results:

<u>Pam</u>	<u>Deviation</u>	<u>Max</u>	<u>Deviation</u>
28.246 mm		28.9 mm	
28.244		23.0	
28.246		30.8	
28.248		21.1	
Avg: _____		Avg: _____	

a) Calculate each students **average deviation?** **percent deviation?** Which students results were most precise?

b) Calculate each student's **percent error**? Which student was most accurate?

Practice:

Indicate whether each statement below refers to precision (p) or accuracy (a)

- _____ a) This term refers to how close your measurement (or average measurement) is to the true value.
_____ b) This term refers to how close your measurements are to one another.
_____ c) The results of your percentage error calculation gives you an idea of the _____ of your measurement(s).
_____ d) The results of your average deviation calculation gives you an idea of the _____ of your measurement(s).

2. a) Calculate the mean and average deviation for the series of density measurements on samples of zinc below. Please make proper use of significant figures in your calculations

Trial #	Density (g/cm ³)	Abs Deviation From The Mean
1	7.76	_____
2	7.82	_____
3	7.65	_____
Average	_____ ±	_____

b) The true value for the density of zinc is 7.30 g/cm³. What was the percent error of the measurements above?

III. Properties of Matter

A. Chemical Properties

The ability of a substance to undergo chemical reactions to form new substances.

B. Physical Properties

A quality of a substance that can be observed or measured without changing the substance's chemical composition. Just as you identify your friends by their physical properties; height, weight, eye color, hair color ... Chemical substances are also identified by their physical properties. Different chemical substances clearly differ in properties that allow us to classify and identify substances of the world. These physical properties may depend on the amount of substance present (**extensive**) or may not depend on the amount of substance present (**intensive**). **Intensive properties are useful in identifying unknown substances.**

Practice:

Three primary sources can be used to identify physical and chemical properties of known elements and compounds are ...

- Handbook of Chemistry and Physics
- The Merck Index
- Material Safety Data Sheet (MSDS)
-

Table 2.1

Physical Properties of Some Common Substances						
Substance	Formula	State	Color	Melting point (°C)	Boiling point (°C)	Density (g/cm ³)
Neon	Ne	gas	colorless	-249	-246	0.0009
Oxygen	O ₂	gas	colorless	-218	-183	0.0014
Chlorine	Cl ₂	gas	greenish-yellow	-101	-34	0.0032
Ethanol	C ₂ H ₅ OH	liquid	colorless	-117	78	0.789
Mercury	Hg	liquid	silvery-white	-39	357	13.5
Bromine	Br ₂	liquid	red-brown	-7	59	3.12
Water	H ₂ O	liquid	colorless	0	100	1.00
Sulfur	S	solid	yellow	113	445	2.07
Sucrose	C ₁₂ H ₂₂ O ₁₁	solid	white	185	d*	1.59
Sodium chloride	NaCl	solid	white	801	1413	2.17

* d = decreases on heating

Review/Practice

1) (Using the table on page 12) List (a) physical property(ies) that can be used to distinguish Ethanol and Water.

2) (Using the table on page 12) List (a) a physical property(ies) that can be used to distinguish sodium chloride and sulfur.

Common physical properties of matter:

<u>Property</u>	<u>Intensive or Extensive Physical Property</u>
Mass	_____
Volume	_____
Color	_____
State of Matter	_____
*Melting Point	_____
*Boiling Point	_____
*Density	_____
Solubility	_____
Electrical Conductivity	_____
Malleability	_____
Ductility	_____
Viscosity	_____

* These intensive properties can be used to identify pure substances like elements and compounds because every element and compound has a unique values.

Important Physical Property -- Density

Density is the ratio of the mass of an object to its volume. This property is useful in identifying an unknown substance. The density of a substance relates the mass and volume of a substance. If you know any two of the three quantities you can solve for the third.

$$D = \frac{\text{Mass}}{\text{Volume}}$$

Sample Calculations:

The Handbook of Chemistry and Physics lists the density of mercury as 13.534 g/cm^3 (at 20°C).

a) What is the mass of 24 cm^3 (or 24 milliliters, mL) of mercury?

b) What is the volume of 65.5 g of mercury?

c) An unknown metal has a mass of 2.361 g and is $2.35 \text{ cm} \times 0.134 \text{ cm} \times 1.05 \text{ cm}$. What is the identity of the element ($d_{\text{Ni}} = 8.91 \text{ g/cm}^3$; $d_{\text{Ti}} = 4.50 \text{ g/cm}^3$; $d_{\text{Zn}} = 7.14 \text{ g/cm}^3$; $d_{\text{Sn}} = 7.23 \text{ g/cm}^3$)

d) The Handbook of Chemistry and Physics reports that the density of Zinc (Zn) is 7.50 g/cm^3 .

i) What is the **mass** of 10.0 cm^3 of Zn? ii) What would be the **volume** of 50.0 g of Zinc?

iii) What is the density of Zinc in g / L (using unit conversion method)

Temperature Dependence of Physical Properties

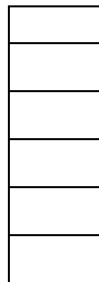
The numerical values of the physical properties of matter are often affected by temperature. Density is an important example of this.

Temperature (°C)	Density of Water (g/cm ³)
0 (ice)	0.917
0 (liquid water)	0.99984
2	0.99994
3.98	0.999973 (Max Density)
4	0.99997
10	0.9970

Solid water (ice) is less dense than liquid water so it floats. Because the density of materials changes with temperature, it is important to report temperature when doing density studies. Different substances will separate from one another based on their different densities.

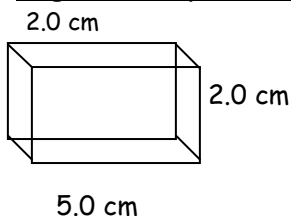
Three liquids along with two solids are placed into a cylinder. If they arranged themselves from top to bottom according to their densities what would the cylinder look like?

Oak Wood	0.71 g/mL
Solid Gold	19.3 g/mL
Water	1.00 g/mL
Gasoline	0.67 g/mL
Mercury (liquid)	13.6 g/mL



Calculating Density of Substances in Laboratory

Regular Shaped Object:



Mass = 10.00 g

Volume = _____

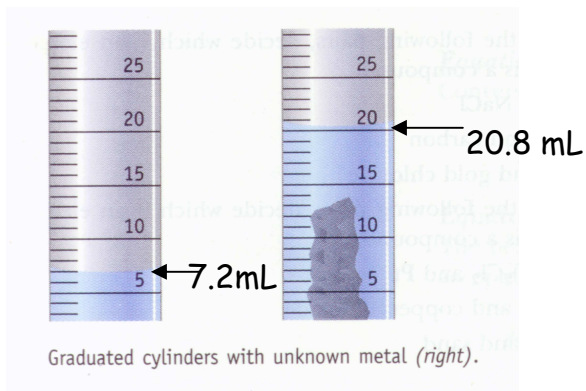
Density: _____

Mini Challenge Activity:

A piece of aluminum foil that is _____ cm by _____ cm has a mass of _____ g. The true density of the aluminum is 2.70 g / cm³. How thick is the foil?

Irregular Shaped Object

A sample or unknown irregularly shaped metal is placed into a graduated cylinder containing water. The mass of the sample is 37.5 g, and the water levels before and after adding the sample to the cylinders is shown below. What is the density and what type of metal is the unknown?



Volume: _____

- $d_{Mg} = 1.74 \text{ g/cm}^3$
- $d_{Fe} = 7.87 \text{ g/cm}^3$
- $d_{Ag} = 10.5 \text{ g/cm}^3$
- $d_{Al} = 2.70 \text{ g/cm}^3$
- $d_{Cu} = 8.96 \text{ g/cm}^3$

% Error: _____

Graphical Methods

- a) 44. Figure 2-14 shows the mass of three substances for volumes between 0 and 60 cm³.
- What is the mass of 30 cm³ of each substance?
 - If you had 100 g of each substance, what would their volumes be?
 - In one or two sentences, describe the meaning of the steepness of the lines in this graph.

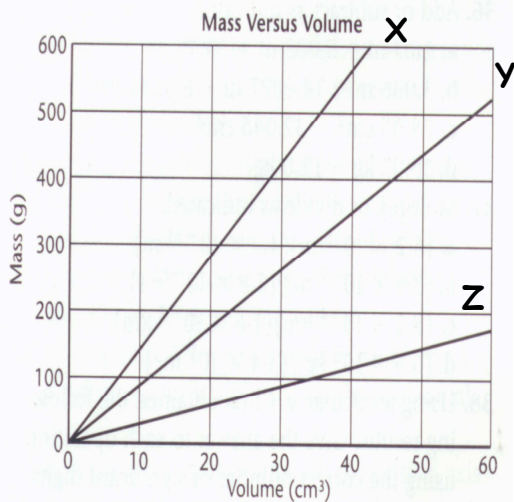


FIGURE 2-14

- X: _____ g

Y: _____ g

Z: _____ g
- X: _____ mL

Y: _____ mL

Z: _____ mL

c)

d) What is the density (slope) of each substance?

- X = _____ g/cm³
- Y = _____ g/cm³
- Z = _____ g/cm³

IV. Temperature and Heat of Matter

A. Temperature and its Units

Temperature is the property of matter that determines whether heat energy can be transferred from one body to another and the direction of that transfer. Temperature is also a measure of the motion (kinetic energy) of the particles that make up a sample of matter. Heat is what moves from matter at a higher temperature to matter that is at a lower temperature. Heat will continue to flow until both samples of matter are at the same temperature.

Aluminum Pellet Air (25°C)
50°C

In the U.S. everyday temperatures are reported using the Fahrenheit (°F) scale, but the **Celsius scale** is used in most other countries. The scientific community has adopted the **Kelvin scale (SI Unit)**. In our class we will be using either the Celsius or Kelvin scales.

$$K = ^\circ C + 273$$

**Absolute Zero	F.P. Water	B.P. Water
↙ -273°C	32°F 0°C	212°F 100°C
0 K	273 K	373 K
Solid Water	Liquid Water	Gaseous Water

****At 0 K, absolute zero**, all particles in a sample of matter loses ALL KINETIC ENERGY. The particles do not move!!!!

Review/Practice:

A sample of water is at 200 K. What is its temperature in °C? What state is the water in?

Surgical instruments may be sterilized by heating to 170°C for 1.5 hours. Convert 170°C to kelvins?

B. Heat(q) and its Units

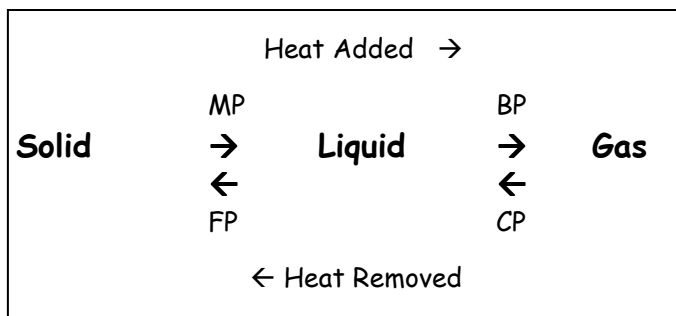
Heat is a form of energy that can be transferred from one body to another because of temperature difference. A transfer will occur from the body at higher temperature to a body that is at a lower temperature until both bodies are at the same temperature. The **joule** is the SI unit of heat.

IV. Matter Can Undergo Physical and Chemical Changes

A. Physical Changes

Changes in physical properties are called **physical changes**. In a physical change the identity of a sample doesn't change even though its shape, size, or physical state (solid liquid, or gas) has changed. Chemical bonds are not broken, that is atoms are not rearranged) during physical changes. Physical changes usually involve the addition or removal of heat from a substance (element, compound, or mixture).

Examples: Melting, Boiling, Subliming, Freezing, **Dissolving**, Cutting a piece of copper wire



B. Chemical Changes

When one or more chemical substances (**reactants**) are changed to one or more DIFFERENT chemical substances (**products**) a **chemical change**, or **chemical reaction** has occurred. At a particulate level a chemical change produces a new arrangement of atoms without a gain or loss in the number of atoms of each kind. The particles (atoms, molecules, or ions) present after the reaction, however, are different from those present before the reaction. Chemical change is represented by using a chemical formula called a **chemical equation**.

Indications that a chemical change has occurred include:

1. Energy (Heat or light) is released or absorbed.
2. Change in color or odor
3. Production of gas or solid (precipitate) from a liquid
4. Chemical change is often irreversible.

Review/Practice Questions:

1) Classify the following changes as **chemical** or **physical**:

___ Water boils

___ A firefly emits light

___ A metal chair rusts

___ Salt dissolves in water

___ Milk spoils

___ bending of wire

___ Cutting of grass

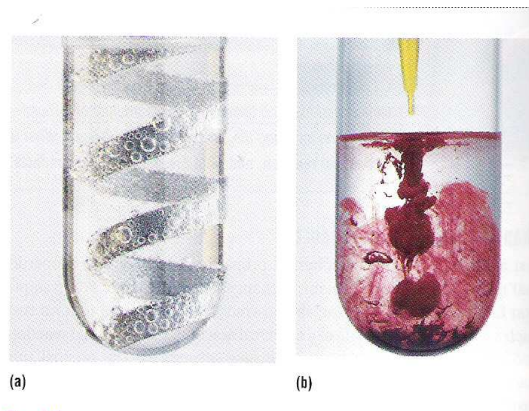
___ Burning coal

___ Dry ice subliming

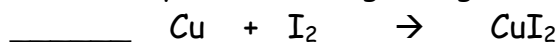
2) Classify the following properties as **chemical** or **physical**:

___ a) a blue-gray color ___ b) brittle ___ c) insoluble in water

___ d) melts at 1410°C ___ e) reacts vigorously with fluorine ___ f) gasoline burns



3) Classify the following changes as **chemical** or **physical**:



_____ Ice Melts

_____ An alka-seltzer tablet is added to water and dissolved (bubbles form)

_____ Table salt (NaCl) is dissolved in water

4) In each case, identify the underlined property as a physical or a chemical property. Give a reason for your choice.

_____ (a) Dry Ice sublimes (changes directly from a solid to a gas) at -78°C .

_____ (b) Methanol (methyl alcohol) burns in air with a colorless flame.

_____ (c) Sugar is soluble in water.

_____ (d) Hydrogen peroxide, H_2O_2 , decomposes to form oxygen, O_2 , and water, H_2O .

In each case, identify the underlined property as a physical or chemical property. Give a reason for your choice.

_____ (a) The normal color of the element bromine is red-orange.

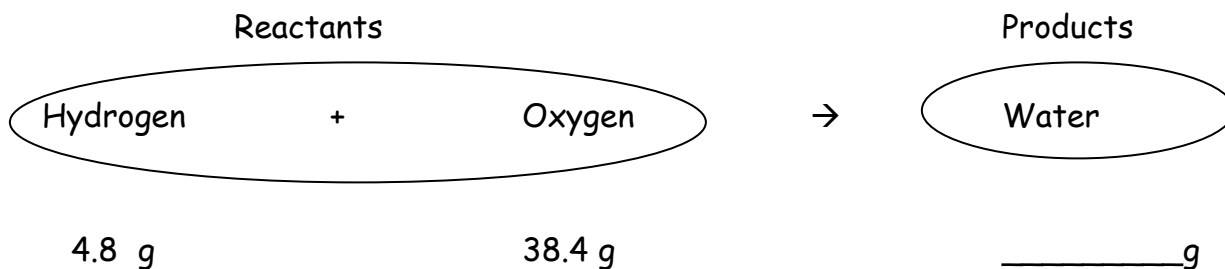
_____ (b) Iron is transformed into rust in the presence of air and water.

_____ (c) Dynamite can explode.

_____ (d) Aluminum metal, the shiny "foil" you use in the kitchen, melts at 660°C .

C. Law of Conservation of matter

In any physical or chemical change, mass is neither created or destroyed; it is conserved. In other words the total mass of reactants is equal to the total mass of products.



Consider the following explosive chemical reaction and determine the mass of water produced (Oklahoma City Bomb)



When powdered iron (Fe) is left exposed to air, it rusts (forms Fe_2O_3). Explain why the rust weighs more than the original iron. **Does this go against the law of conservation of mass?**

D. Energy and Physical and Chemical Changes of Matter

Chemical and physical changes are *ALWAYS* accompanied by energy changes. Energy is always either released (**exothermic reaction; feels warm; heat exits**) or absorbed (**endothermic reaction; feels cold; heat enters**) during chemical and physical changes.

Exothermic Chemical Reaction:

Hydrogen + Oxygen → Water + **Heat**

Endothermic Chemical Reaction:

limestone + **Heat** → lime + carbon dioxide
(calcium carbonate) (calcium oxide)

Photosynthesis (very important!!!)

Carbon Dioxide + Water + Light Energy → Glucose + Oxygen

Exothermic Physical Reaction

Liquid water → Ice + **Heat**

Endothermic Physical Reaction

Ice + **Heat** → Liquid water

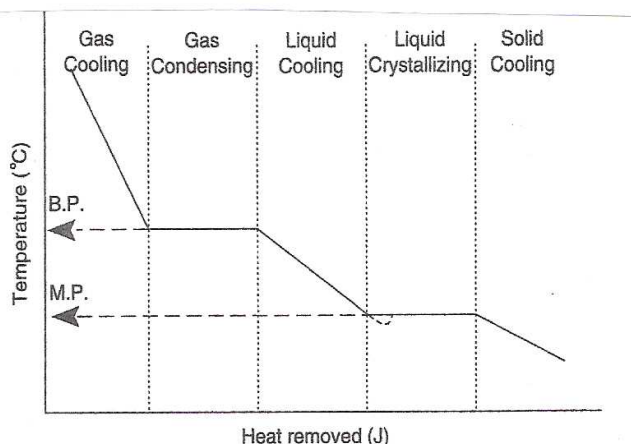
Review\Practice:

Classify the following as Chemical or Physical changes and then as Exothermic or Endothermic Process

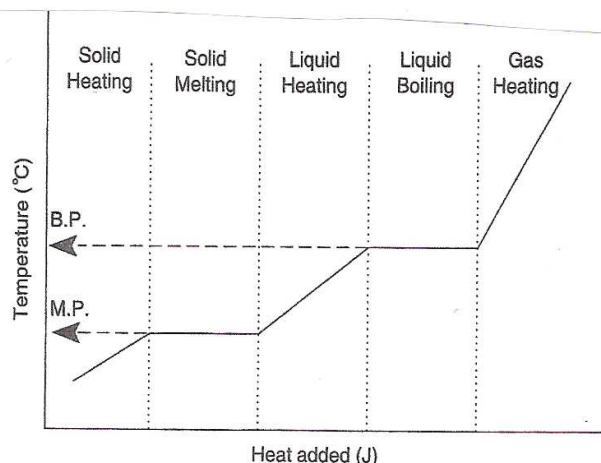
	Chemical or Physical	Endothermic or Exothermic
1. Solid silicon melts	_____	_____
2. Solid KBr is dissolved in water and the solution feels warm	_____	_____
3. Natural gas (CH ₄) is burned in a furnace	_____	_____
4. Water is boiled in a tea kettle	_____	_____
5. Gaseous water condenses into liquid water in a radiator	_____	_____
6. Fe + S + Heat → FeS	_____	_____
7. I ₂ (g) → I ₂ (s) + Heat	_____	_____

Phase Changes of Matter

Heating Curve of Matter:



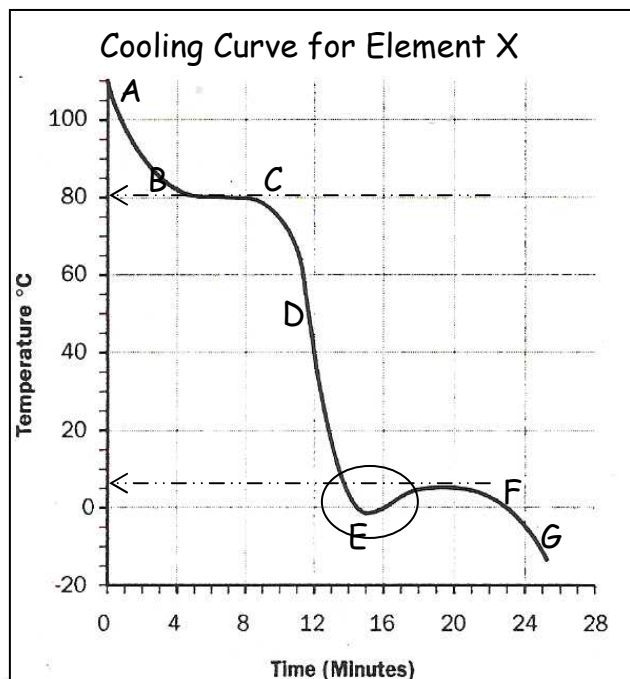
Cooling Curve of Matter:



Melting: The motion of the particles in a solid (KE) are great enough to overcome attraction between particles in solid state, they spread out into a more random liquid arrangement of particles. All pure compounds, pure elements, and mixtures have unique MP and BP.

Boiling: The motion of the particles in a liquid (KE) are great enough to overcome attraction between particles in liquid state, they spread out into a extremely random gas arrangement of particles. All pure compounds, pure elements, and mixtures have unique MP and BP.

Review \ Practice Reading Heating or Cooling Curve



The following questions relate to the Graph to the left

1) Is this a heating or cooling for a substance?

2) What is the substances Freezing Pt: _____°C

3) What is the substances Boiling Pt: _____°C

4) Identify the state of matter of X at

Point A: _____ Point E: _____

Point B: _____ Point F: _____

Point C: _____ Point G: _____

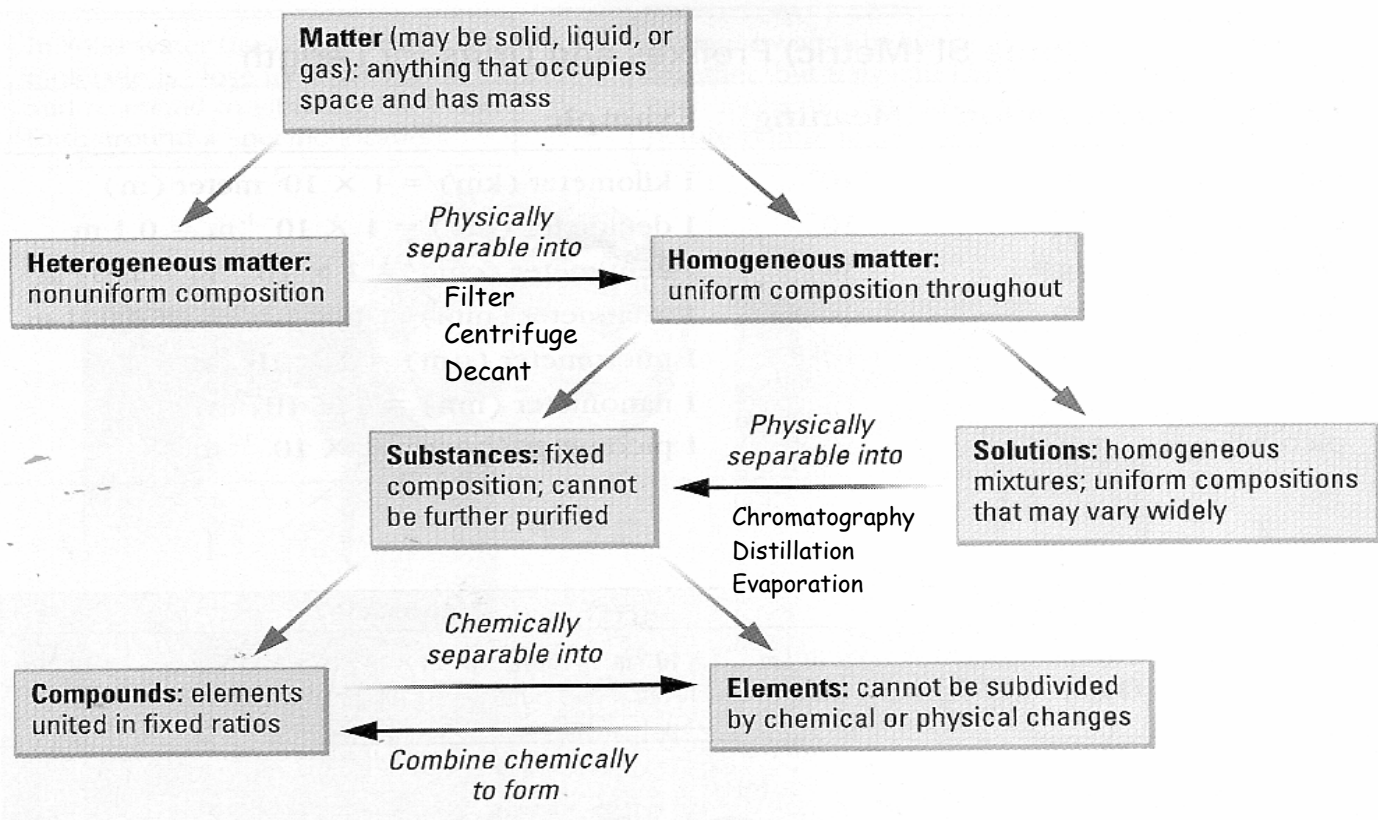
Point D: _____

5) What change of state is occurring from points b to c: _____

6) What change of state is occurring from points b to c: _____

7) What process is occurring at the circled region on the curve? _____

B. Classifying Matter Based on Composition



(20+ million)

Elements united into fixed ratios

Table salt (NaCl) Water (H_2O)
 Sand (SiO_2) Sugar ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$)
 Fool's Gold (FeS_2)

117 elements

Cannot be subdivided by chemical or Physical processes

Sodium (Na) Silicon (Si)
 Chlorine (Cl) Oxygen (O_2)
 Hydrogen (H_2) Iron (Fe)

Element: A substance that CANNOT be changed into a simpler substance (see periodic table)

Compound: A substance that CAN be changed into a simpler substance

Mixture: A physical blend of two or more substances (elements or compounds) that are not chemically combined

Review\Practice:

1. Classify the mixture as homogenous or heterogeneous then identify the best means of separating the mixture into pure substances (Chromatography, Filtration, or Distillation)

Mixture	Homogeneous or Heterogeneous	Separation Technique
Sand (SiO_2) and water	_____	_____
Salt and water (Δ Boiling Points)	_____	_____
Food Coloring	_____	_____

2. Separating Mixtures. Match the mixtures on the left with the separation method for separating the mixture on the right:

_____ Chromatography

_____ Filtration

_____ Distillation

_____ Evaporation

_____ Centrifuge

_____ Decant

a) Used to separate homogeneous mixture based on BP differences of pure substances making up a mixture

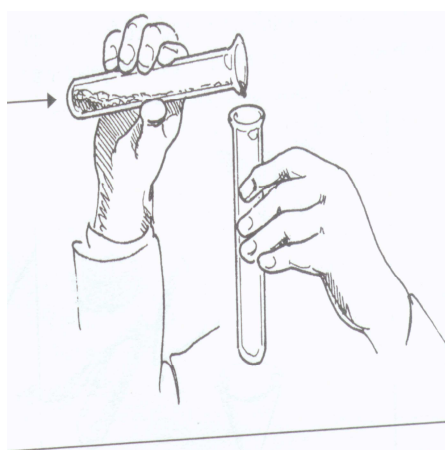
b) Used to separate homogeneous mixture based on differences in the pure substances in a mixture's ability to adhere to a solid substrate.

c) Used to separate heterogeneous mixture based on differences in the size or density of the pure substances making up the mixture

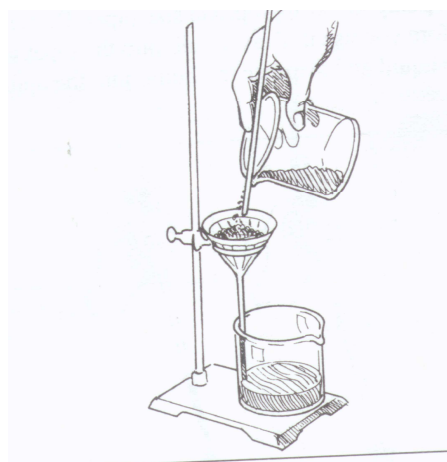
Separating Heterogeneous Mixtures (filtration, centrifuge, decant)

Example: Sand & Water (SiO_2 and H_2O)

Decanting



Filtering



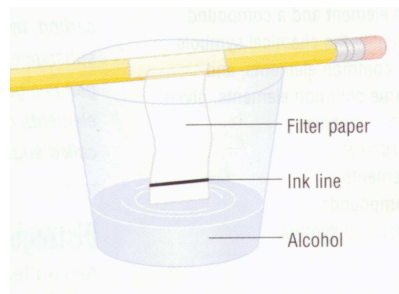
Centrifuge



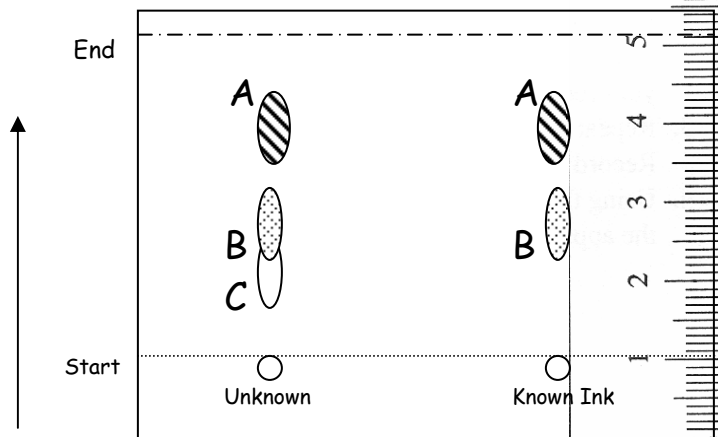
Separating Homogeneous Mixtures (solutions) of Matter

Chromatography -

A process of separating a homogeneous solution (like ink) of closely related compounds by allowing the components of a solution to adsorb ("stick") to a solid substrate (paper or glass fiber) while a mobile phase (alcohol, water, etc...) allows the components of the mixture to migrate up the solid substrate. Because the components "stick" with differing degrees of tenacity the components of the mixture separate into colored layers. The compounds found closest to the bottom of the chromatograph have not moved up the solid substrate quickly because they have "stuck" to the solid with a greater tenacity. The components of the mixture will each be assigned an R_f value = distance pure compound migrated / distance solvent migrated.



Solvent
Flow



- 1) How many pure compounds make up the unknown ink? _____
- 2) How many pure compounds make up the known ink? _____
- 3) a) Do either of the inks have an pure compounds in common? _____
b) How do you know? _____

4) Calculate the Retention Factor, R_f value for compound A _____

5) a) Which pure compound (A, B, or C) greatest adhesion to the chromatography paper? _____

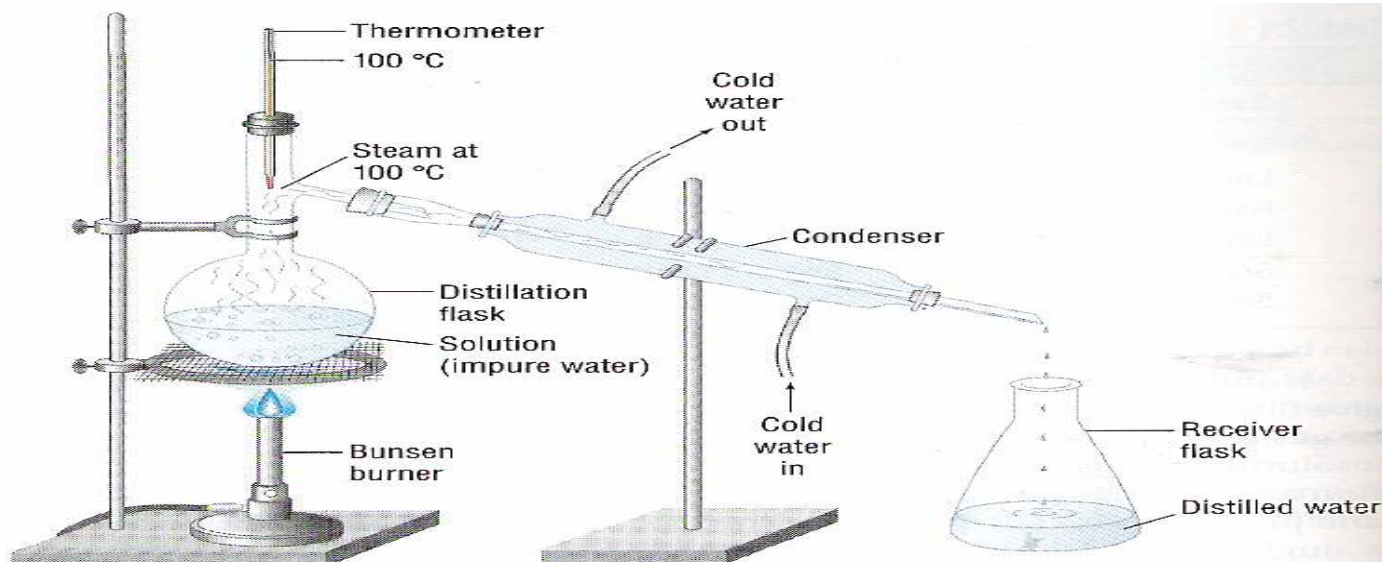
b) Which pure compound (A, B, or C) least adhesion to the chromatography paper? _____

c) What pure compound (A, B, or C) is found in the unknown but not in the known ink? _____

d) What is the R_f factor for pure compound C? _____

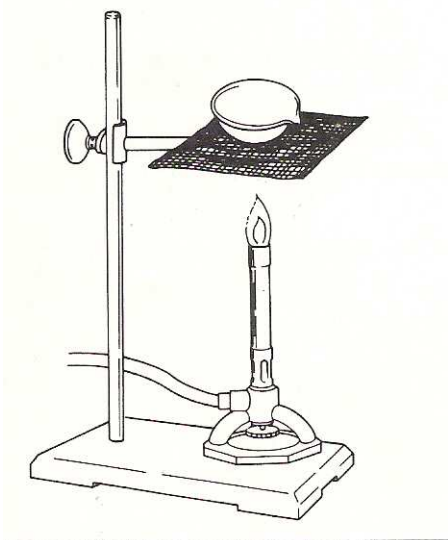
Distillation -

A process of separating a homogeneous solution of closely related compounds based on their large differences in boiling points. A process by which a homogeneous mixture can be separated into its components. In the process a mixture is heated and the component with the lowest boiling point is driven from the mixture as a vapor or gas. This vapor then condenses on the inside of a condensing tube and the purified liquid with the lowest boiling point drips out the far end of the condensing tube. The component of the mixture with the much higher boiling point remains as a liquid in the round bottom flask.



A solution of impure water is being distilled. As the solution boils, the water turns into steam, leaving the impurities behind in the distillation flask. As the steam passes through the water-cooled condenser, it turns to liquid. The distilled water is collected in the receiver flask.

Evaporation -



Practice:

1. Classify the mixture as homogenous or heterogeneous then identify the best means of separating the mixture into pure substances (Chromatography, Filtration, or Distillation)

Mixture	Homogeneous or Heterogeneous	Separation Technique
Sand (SiO_2) and water	_____	_____
Salt and water (Δ Boiling Points)	_____	_____
Food Coloring	_____	_____

2. Separating Mixtures. Match the mixtures on the left with the separation method for separating the mixture on the right:

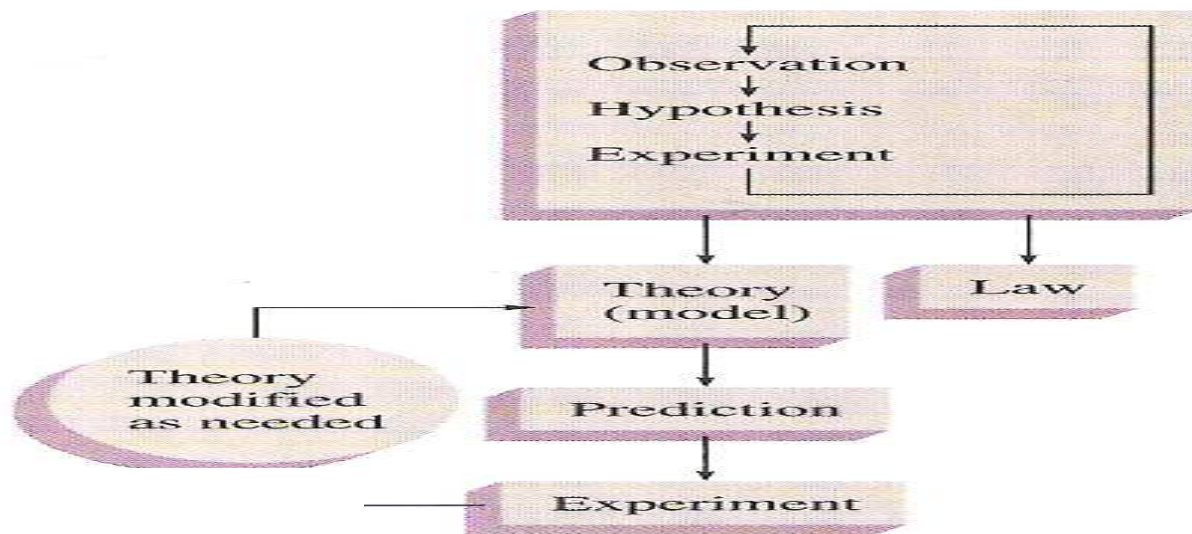
_____ Chromatography	a) Used to separate homogeneous mixture based on BP differences of pure substances making up a mixture
_____ Filtration	b) Used to separate homogeneous mixture based on differences in the pure substances in a mixture's ability to adhere to a solid substrate.
_____ Distillation	c) Used to separate heterogeneous mixture based on differences in the size or density of the pure substances making up the mixture
_____ Evaporation	
_____ Centrifuge	
_____ Decant	

Lab Practical:

Devise a way, using physical methods of separating mixtures previously discussed in the notes, to separate and quantify the mass of each pure component in a 10.0 g mixture of iron (Fe), sand (SiO_2), and salt (NaCl), into its pure components (you may also use water, H_2O).

VI. The Scientific Method and Visualizing Data

A. The SCIENTIFIC METHOD



B. VISUALIZING DATA

One of the most important skills to master in science is to determine how changing one variable in an experiment affects another. The variable the experimenter changes or manipulates is the **independent variable**. The variable that changes in response to the independent variable is called the **dependent variable**. The value of the dependent variable "depends" on the independent variable. **When graphing data the independent variable is plotted on the x-axis and the dependent variable is plotted on the y-axis.**

When Graphing Data follow the following steps:

1. Identify the independent and dependent variables in your experiment. The independent variable is plotted on the x-axis and the dependent variable is plotted on the y-axis.
2. Determine the range of the independent variable to be plotted.
3. Decide whether the origin (0,0) is a valid data point.
4. Spread the data out as much as possible. Let each division on the graph paper stand for a convenient unit.
5. Number and label (including what is being plotted and units) the horizontal axis.
6. Repeat steps 2-5 for the dependent variable.
7. Plot the data points on the graph.
8. Draw the "best fit" straight line or smooth curve that passes through as many data point as possible. Do not use a series of straight lines that "connect the dots."
9. Give a graph a title that clearly tells what the graph represents.
10. Give the equation for the line of best fit a R^2 using Excel or your calculators regression functions.

Dependent Variable vs Independent Variable

Linear Relationship

Direct Proportion

$Y = mx + b$



Non-Linear Relationship

Parabolic (quadratic)

$y = ax^2 + bx + c$



Hyperbolic (inverse)

$y = a / x$ or $xy = a$



Practice 1:

A sample of gas is placed in a sealed syringe. The sample has a constant mass and is kept at a constant temperature. The pressure on the gas is increased over a series of intervals by the scientist and the volume is then measured.

Pressure (atm)	Volume (cm ³)
1.01	22.4
1.49	14.9
1.99	11.2
2.51	8.95
3.01	7.46

Dependent Variable _____

Independent Variable _____

Describe the Resulting Curve



- a) Plot the data given in the table and draw curve that best fits all points.
- b) What is the relationship between Volume and Pressure of a gas?
- c) What is the equation relating Volume and Pressure of a gas?

Practice 2:

A ball is dropped from a 50.0 m building and the scientist measures the distance the ball has fallen each second for 5 seconds.

Time (s)	Distance (m)
1	5
2	20
3	44
4	80
5	122

Dependent Variable _____

Independent Variable _____

- a) Plot the data given in the table and draw curve that best fits all points.
- b) Describe the Resulting Curve: _____
- c) What is the relationship between Time and Distance of the ball?
- d) What is the equation relating Time and Distance of the ball?



Practice 3:

Pete and Sally wanted to investigate how the mass of sulfur is related to its volume. They took various masses of sulfur and measured their volumes. The following data was obtained:

<u>Mass (g)</u>	<u>Volume (cm³)</u>
23.5	11.4
60.8	29.2
115	55.5
168	81.1

Dependent Variable _____

Independent Variable _____

a) Plot the data given in the table and draw curve that best fits all points.

b) Describe the Resulting Curve: _____

c) What is the relationship between mass and volume the sulfur sample? _____

d) What is the equation relating mass and volume the sulfur sample? _____

e) Does the slope of the graph represent the sulfur's density? Explain

f) What is the density of the sulfur sample?

g) The true density of pure sulfur is 2.07 g/cm³. What is the percent error of Pete and Sally's density determination?

