Honors Chemistry **awohlrab@ocsdnj.org** **rcitta@ocsdnj.org**	**Dr. Wohlrab, Professor R. Citta** **Due Dates: Optional but highly recommended**
Topic: Chemistry **Purpose: Introductory Chemistry Concepts**	

Text/Novel(s) & Brief Description; online text "Chemistry"

Approximate Time on Task: 3 hours

Suggested Timeline: Start immediately and finalize at end of August

How It Will Be Assessed :

SO WHAT IS THE SUMMER WORK? The summer work is **<u>highly recommended</u>**. This work covers the first chapters we will be covering in the first few weeks of school. Much of this should be review from previous science course. Some of it will not be familiar but we will go over everything. An answer key will be provided on the first day back to school. Please feel free to email Dr. Wohlrab or Professor Citta at any time during the summer with questions.

1. Memorize common ions and charges. You'll find these at the end of the packet.
2. Memorize the strong acids and bases You'll find these at the end of the packet..
3. Read through the chapter 1 guided notes and complete the practice problems

Online Text Resource:

https://misterchemistry.com/wp-content/uploads/2017/06/Chemistry-the-central-science-.pdf

Chapter 1: Matter and Measurement

Name_______________________

1.1 The Study of Chemistry

Chemistry – the study of the composition of matter and its transformations

- **Scientific Method**

1. **Observation** – statement that describes some we see, hear, taste, feel or smell
2. **Hypothesis** – a tentative explanation of the observed phenomenon.
3. **Experimentation** – collect data to prove/disprove hypothesis.
4. **Theory** – a tested explanation of the behavior of nature or why the phenomenon may occur.
5. **Law** – broad generalization based on the results of many experiments
   Ex.) Law of Gravity, Law of Conservation of Mass, Law of Conservation of Energy

1.2 Classifications of Matter

Matter – anything that has mass and takes up space.

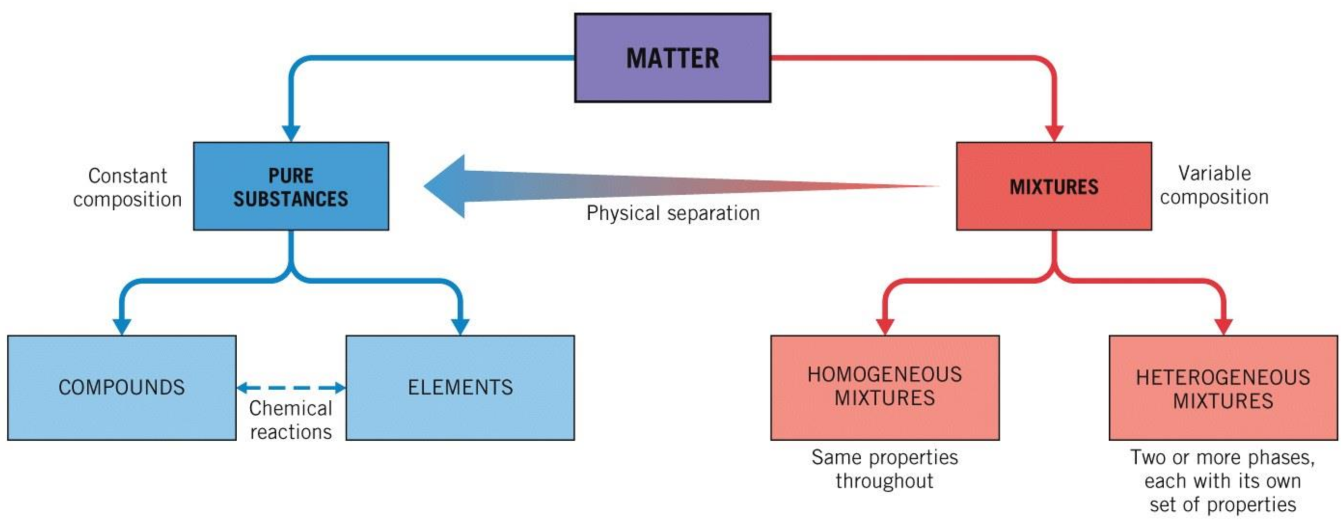
Mass – How much matters an object has.

Weight – Force with which an object is attracted by gravity

- Mass vs. Weight

Ex.) Consider a brick on the moon and on earth.
Weight on moon = 1/6 weight on earth
Mass on moon = mass on earth

Classification of Matter by Properties



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Elements – substances that cannot be decomposed into simpler substances by chemical reactions.

- Composed of only one type of atom
- Made of identical atoms
- Can exist as monoatomic (Ar), diatomic (H_2), or polyatomic (S_8) in nature
- Shown on periodic table as symbols

Ex.)

Element	**Symbol**	**Latin name**
Potassium	K	Kalium
Tin	Sn	Stannum
Iron	Fe	Ferrum
Lead	Pb	Plumbun
Sodium	Na	Natrium
Copper	Cu	Cuprum
Silver	Ag	Argentum
Mercury	Hg	Mercurum

Compounds – Formed from two or more atoms of different elements combined (bonded) in a fixed proportion.

Ex.) carbon dioxide (CO_2) water (H_2O)
table salt (NaCl) sand (SiO_2)

- Have different properties that the elements that compose them
- Can be broken down into elements by some chemical changes

Mixtures – consist of varying amounts of two or more pure substances (elements or compounds) and can be separated by physical means.

- **Homogeneous mixtures** – aka solutions, have variable compositions and are made of two or more substances. Have visibly indistinguishable parts.

  Ex.) Liquid solution = sugar water
  Solid solution = Brass
  Gas solution = air

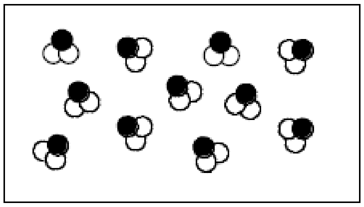
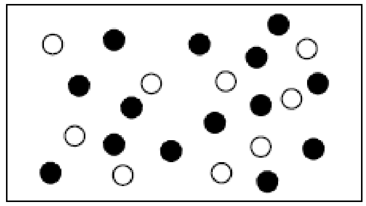
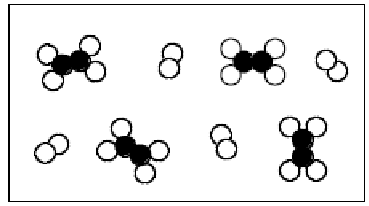
- **Heterogeneous mixtures** – consists of two or more phases. Have visibly distinguishable parts.

  Ex.) Salad dressing, soda, ice and water

Separation of mixtures

- **Filtration** - separation of solids from liquids
- **Distillation** – separation of liquids based on differences in boiling points
- **Chromatography** – separation of dissolved solids based on physical properties of the molecules

Practice problem 1) Classify each the following pictures as an element, compound, mixture of elements, mixture of compounds, or mixture of elements and compounds.

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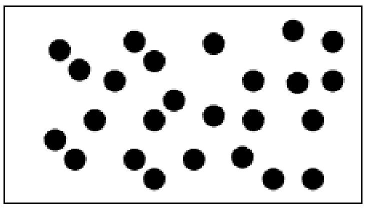 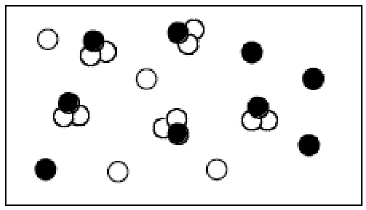 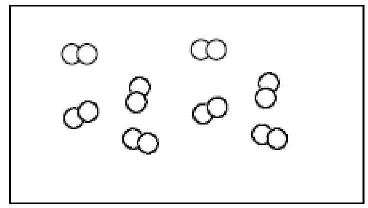

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Practice problem 2) Classify each form of matter

	Pure Substance	Element	Compound	Heterogeneous Mixture	Homogeneous Mixture
Chicken Soup					
Ice (H_20)					
Liquid soap					
Soda					
Gold					
Wood					
Air					
An orange					
Salt water					
Table salt (NaCl)					
Aluminum					

<u>Classification of Matter by State</u>

1. **Solid**s
   - fixed shape and volume
   - molecules close together and vibrating about fixed points
   - incompressible
   - have restricted motion

2. **Liquids**
   - fixed volume, indefinite shape
   - molecules close together and vibrating , but able to flow
   - incompressible to slightly compressible

3. **Gases**
   - indefinite shape and volume
   - molecules vibrate, rotate and translate and are independent of each other
   - VERY far apart
   - highly compressible

   Vapor – the gas phase of a substance that is normally a solid or liquid at room temperature

   Fluid – that which can flow; gases and liquids

1.3 Properties of Matter

Chemical Properties – describes behavior of matter that leads to a new substance.

Ex.) Flammability, Combustibility, Reactivity with metals, oxidation potential

Physical Properties – can be observed without changing the composition of the substance

Ex.) Color, density, melting point, volume, solvation

Intensive properties – independent of sample size

Ex.) Color, temperature, texture, density, melting point

Extensive properties – dependent of sample size

Ex.) Volume, mass

Practice problem 3) Identify the following properties as chemical or physical.

	Chemical	Physical
Aluminum metal is grey		
Silver tarnishes		
Ice melts at 273 K		
Calcium metal reacts with HCl		
Salt dissolves readily in water		
Flammability		
Good conductor of heat		
Density		
Color		

Practice problem 4) Identify the following properties as intensive or extensive

	Intensive	Extensive
Mass		
Density		
Melting point		
Color		
Volume		
Length		

Changes in Matter

Chemical Change – a process that results in the formation of a new substance

- evidence of a chemical change
  - formation of new solid, liquid, or gas,
  - temperature change
  - color change

Physical Change – a process that results in no new substance but may change the state

Practice problem 5) Identify the following changes as chemical or physical.

	Chemical	Physical
Gasoline ignites		
Tearing paper		
Formation of clouds		
Digestion of food		
Sugar dissolving in water		
A rock is crushed		
Paper burns		
Iron rusts		
Evaporation of ethyl alcohol		
Mild spoils		

1.4 Measurement

Measurements

Quantitative - Numerical data, measured with an instrument

Ex.) Boiling point, volume, mass

Qualitative – Non numerical observations

Ex.) Color, loudness, formation of a gas

SI Units - Standard system of units (Metric) used in scientific and engineering measurements

- Base units - seven units from which all other metric units are derived

Measurement	Base Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electric Current	ampere	A
Temperature	kelvin	K
Amount of Substance	mole	mol
Luminous Intensity	candela	cd

Derived Units – all physical quantities will have units derived from the seven base units.

Measurement	Symbol
Area = length x width	m^2
Volume = length x Width x Height	m^3
Velocity = distance/time	m/s
Density = mass/volume	kg/m^3

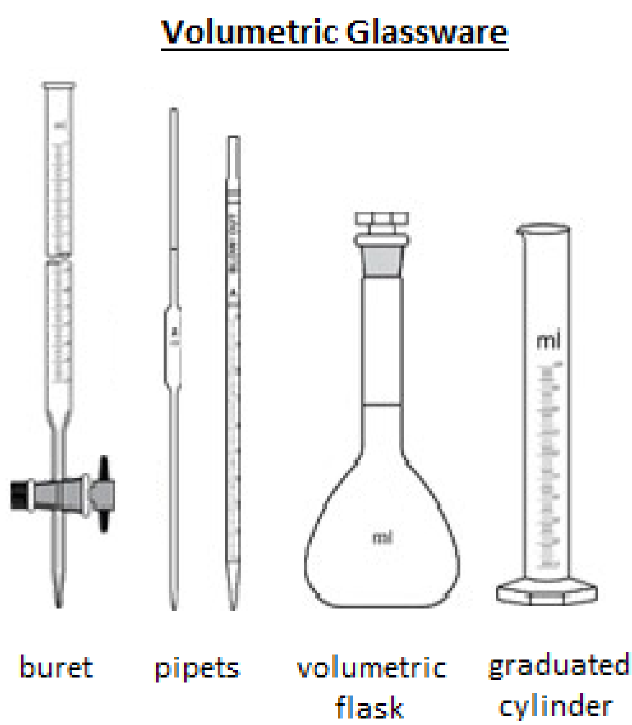
Decimal Multipliers

- Convenient way of expressing very large or small measurements
- Base units = meter, liter, second, gram, etc.

Prefix	Symbol	Equivalent of Base Value
tera	T	10^{12} = 1000000000000
giga	G	10^{9} = 1000000000
mega	M	10^{6} = 1000000
Kilo	k	10^{3} = 1000
hecto	h	10^{2} = 100
deka	da	10^{1} = 10
deci	d	10^{-1} = 0.1
centi	c	10^{-2} = 0.01
milli	m	10^{-3} = 0.001
micro	µ	10^{-6} = 0.000001
nano	n	10^{-9} = 0.000000001
pico	p	10^{-12} = 0.000000000001

Common Laboratory Units

1. **Length**
   - SI unit is meter (m)
   - Commonly use centimeter (cm)
2. **Volume**
   - SI unit is liter (L)
   - Commonly use milliliter (mL)
   - Measured with volumetric glassware
3. **Mass**
   - SI unit is kg
   - Commonly use gram (g)
   - Measured with balance



4. Temperature

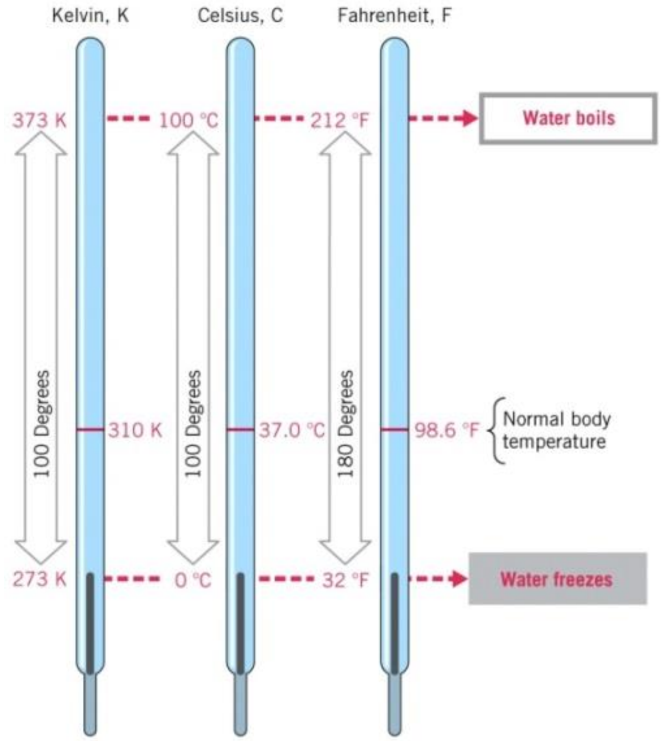
- SI unit is K
- Commonly use degrees Celsius (°C)
- Measured with thermometer
- **Absolute Zero** – zero point on Kelvin scale, lowest possible temperature at which all atomic/molecular vibrations cease.

- Notice that °C and K units are the same size.
- °F are the smallest temperature units.
- There are 9 °F units for every 5 °C units
- Temperature conversions

$$T_F = \left(\frac{9}{5}\right) T_C + 32$$

$$T_K = T_C + 273$$

$$T_C = \frac{5}{9}(T_F - 32)$$



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Practice example 6) Complete the following temperature conversions.

a) Convert 26 °C to °F.

b) Convert 98.6 °F to °C

c) Convert 37 °C to K

Density

Density: is the ratio of the mass of an object to its volume.

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

- 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.
- Common units:
  - Liquids and solids: g/ml
  - Gases : g/L

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

a) What is the mass of 24.0 cm^3 (or 24 mL) of mercury?

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

Practice problem 7) The Handbook of Chemistry and Physics reports that the density of Zinc is 7.50g/cm^3.

a) What is the mass of 10.0 cm^3 of Zn?

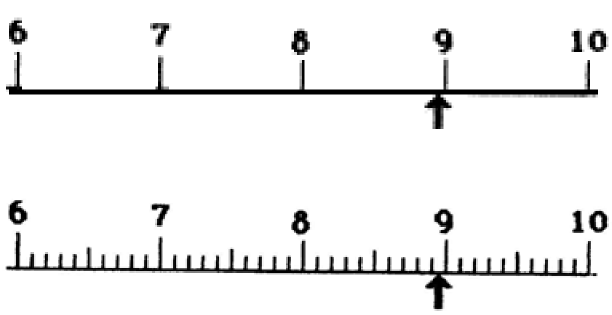
b) What would be the volume of 50.0 g of Zinc?

Practice problem 8) A sheet of aluminum metal that is 5.00 cm by 5.00 cm has a mass of 5.40 g. The true density of the aluminum is 2.70 g/cm3. How thick is the sheet of metal?

Error/Uncertainty in Measurement

- All measurements have uncertainty
- A measurement should be taken one decimal place past the scale on the instrument
- **Significant figures** - For a measurement, all the digits you know for sure PLUS one estimated digit

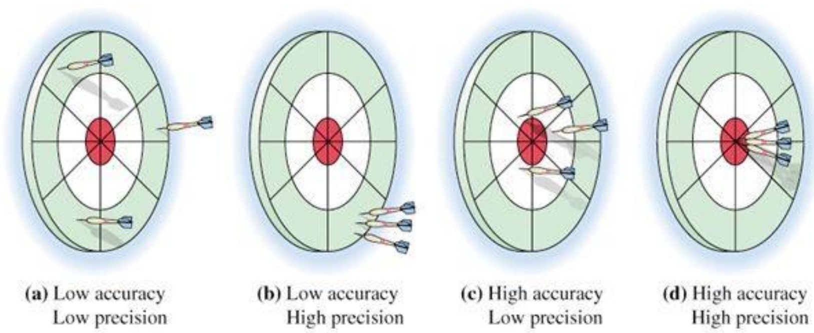
Example) Provide the correct scale reading and the number of significant figures.
Remember the scale reading must include an estimate. If the scale increments are in the "one's" decimal place, the measurements must be recorded to the "10ths" decimal place.



Scale Reading	Significant Figures
8.9	2
8.95	3

Precision and Accuracy

Precision: how well measured quantities agree with each other.
Accuracy: how well measured quantities agree with the "true value."



- **% deviaton:** expresses the precision of a set of measurements

$$\% \text{ deviation } = \frac{|\text{ average} - \text{experimental }|}{\text{average}} \times 100$$

- **% error:** expresses the accuracy of a set of measurements

$$\% \text{ error } = \frac{|\text{ actual} - \text{experimental }|}{\text{actual}} \times 100$$

Rules for Counting Significant Figures

1. All non-zero digits are significant.

   2.84 km ___3___ sig figs
   1.8×10^6 mL ___2___ sig figs

2. Imbedded zeros are significant.

   5005 mL ___4___ sig figs

3. Trailing zeros are significant only when a decimal point is specified

   20. ^{0}C ___2___ sig figs
   3000. K ___4___ sig figs
   20 ^{o}C ___1___ sig figs
   3000 K ___1___ sig figs
   0.760 s ___3___ sig figs
   75.00 kg ___4___ sig figs

4. Leading zeros are never significant.

   0.00345 ___3___ sig figs
   0.02300 ___4___ sig figs

<u>Rules for Significant figures in math operations</u>

Addition/ Subtraction: round the answer to the same number of decimal places as the measurement with the least number of decimal places.

Multiplication/ Division: round the answer to the same number of sig. figs. as the measurement with the least number of sig. figs.

- Note: For multistep calculations, remember your order of operations (PEMDAS) and to apply the sig.fig. rules at each step.

Example) Solve the expression.

2.8723 x 1.6 = 4.59568 → round to 2 sig figs → 4.6

Example) Solve the expression.

3.247	3 decimal places
41.36	2 decimal places
+125.2	1 decimal places
169.807	→ round to 1 decimal place → 169.8

Example) Solve the expression.

(55.41 – 54.79)/55.32 =
2dp 2dp

First, calculate the term in parentheses. Apply the addition/subtraction rule. You're looking for the least number of decimal places. Notice each has 2 decimal places, there for the difference must also have 2 decimal places

0.62/55.32 =
2sf 4sf

Finally solve the expression using the multiplicaiton/division rule and round your answer to the appropriate number of sig figs. Here you're looking for the least number of sig figs.

0.62/55.32 = 0.0112075199 = 0.011

Practice problem 10) Indicate the number of significant figures present in each of the following measurements.

Measurement	Number of Sig. Figs	Measurement	Number of Sig. Figs
2804 m		30.9 V	
2.84 km		0.0010100 s	
0.0029 m		0.106 W	
4.6×10^{5} m		0.160 A	
4.06×10^{-5} g		5.20 N	
4.060×10^{-5} g		30.9 V	
250 600 m		5.2234×10^{12} mol	
250 600. M		4.2×10^{6} g	

Practice problem 11) Solve and use the appropriate rules round to the appropriate number of sig figs.

	Calculated Answer	# of Sig. Figs.
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) =		

Scientific (Exponential) Notation

- A convenient way of expressing very large or small numbers. To work with such numbers, write them in scientific (exponential) notation by expressing decimal places as powers of ten.

$\underline{\qquad\qquad}\ \times\ 10^{\underline{\quad}}$

between 1 and 9.9999

If this value is negative, the number in long form will be < 1
If this value is positive, the number in long form will be > 1

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

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

Practice problem 12) Complete the following table.

Long Form	Scientific Notation	# Sig Figs
0.0001267 m	1.267×10^{-4} m	4
10022.5 kg		
0.034500 s		
	2.43×10^{5} A	
	1.23450×10^{-3} mol	
200 kg		
200. kg		
	1.243×10^{3} s	
	5.6×10^{4} A	

Using Calculators Correctly

- Even if you've set up a problem correctly, you may still get it wrong if it's not entered into the calculator correctly. 2 Tips to avoid errors

  - Use the EE botton when entering numbers in scientific notation.

    6.022×10^{23} → 6.022EE23

  - If you prefer not to use the EE button, you must use parenthesis to "glue" the coefficient to the exponential term

    6.022×10^{23} → (6.022*10^23)

Example) Solve the following problem

$$3.85 \times \frac{9.11 \times 10^{-31}}{1.60 \times 10^{-19}} \times \frac{3}{25} \times \frac{6.022 \times 10^{23}}{35.4527} \times \frac{55.85}{96500} = \_\_\_\_\_\_\_\_\_\_$$

Consider four different methods of calculator entry.

Version 1

```
3.85*9.11*10^-31
/1.60*10^-19*3/2
5*6.022*10^23/35
.4527*55.85/9650
0
  2.58599493E-31
```

Version 2

```
(3.85*9.11E-31*3
*6.022E23*55.85)
/(1.60E-19*25*35
.4527*96500)
     25859949.32
```

Version 3

```
3.85*9.11*10^-31
*3*6.022*10^23*5
5.85
  3.538867398E-4
Ans/(1.60*10^-19
*25*35.4527*9650
0)
     25859949.32
```

Version 4

```
3.85*9.11*10^-31
*3*6.022*10^23*5
5.85
  3.538867398E-4
1.60*10^-19*25*3
5.452*96500
    1.3684472E-11
3.538867398E-4/1
.3684472E-11
     25860459.93
```

Version one is wrong. The student did not use the parenthesis to "glue" the coefficient to the base.

Version two is correct. Version three and four are also correct, however, version two only required 60 keystrokes, while version three required 70 and version four 91. The fewer keystrokes, the less likely it is that you will make a mistake. The slight difference in the answers for version two/three and version four is due to the fact that the numbers in two and three were carried in the memory of the calculator and not entered by the student.

1.6 Dimensional Analysis

- Based on the use of conversion factors and unit cancelations

Conversion factors: an fraction that tells you how many of one unit equals how many of another unit.

Ex.) $\frac{1\text{ dollar}}{10\text{ dimes}}, \frac{12\text{ inches}}{1\text{foot}}, \frac{365\text{ days}}{1\text{ year}}, \frac{5280\text{ feet}}{1\text{ mile}}, \frac{12\text{ eggs}}{1\text{ dozen}}, \frac{1760\text{ yards}}{5280\text{ feet}}, \frac{5280\text{ feet}}{1760\text{ yards}}.$

Dimensional analysis: a method for converting a measurement in one unit into a measurement in a different unit.

$$\textbf{Measurement in original unit } \times \underbrace{\left(\frac{\textbf{new unit}}{\textbf{original unit}}\right)}_{\textbf{Conversion factor(s)}} = \textbf{Measurement in new unit}$$

Ex.) Convert 2.1 m to cm

$$2.1\text{ m} \times \frac{100\text{cm}}{1\text{ m}} = 210\text{ cm}$$

Ex.) Convert 4.5 cm^3 to in^3

$$4.50\text{ cm}^3 \times \frac{(1\text{in})^3}{(2.54\text{cm})^3} = 4.50\text{ cm}^3 \times \frac{(1)^3\text{in}^3}{(2.54)^3\text{cm}^3} = 0.275\text{ in}^3$$

Practice problem 13)

a) Convert 76.2 pm to mm

$$76.2\text{ pm} \times \frac{\_\_\_\_\text{ m}}{\_\_\_\_\text{ pm}} \times \frac{\_\_\_\_\text{ mm}}{\_\_\_\_\text{ m}} = \_\_\_\_\text{ mm}$$

b) Convert 65 mi/hr to m/s

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

c) Convert 0.325m^2 to in^2

$$0.325\text{ m}^2 \times \frac{\_\_\_\_}{\_\_\_\_} \times \frac{\_\_\_\_}{\_\_\_\_} = \_\_\_\_\text{ in}^2$$

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
($CaCO_3$) (CaO)

Photosynthesis (very important!): Carbon Dioxide + Water + Light Energy → Glucose + Oxygen

Exothermic Physical Reaction: Liquid water → Ice + **Heat**

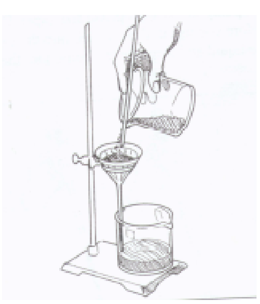
Endothermic Physical Reaction: Ice + **Heat** → Liquid water

Practice problem 14) Classify the following as Chemical or Physical changes and then as Exo or Endothermic

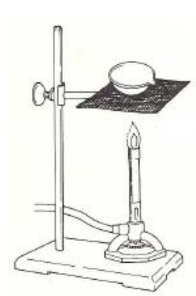
Process	Chemical or Physical	Endo or Exothermic
Solid silicon melts		
KBr dissolves in water and feels warm		
Natural gas burns in the a furnace		
Water is boiled in kettle		
Gaseous water condenses into liquid water		
Fe + S + Heat → FeS		
I_2 (g) → I_2 (s) + Heat		

Separation of Mixtures

Heterogeneous mixtures of solids in liquids (filtration, evaporation, centrifuge, decant)



Filtration



Evaporation

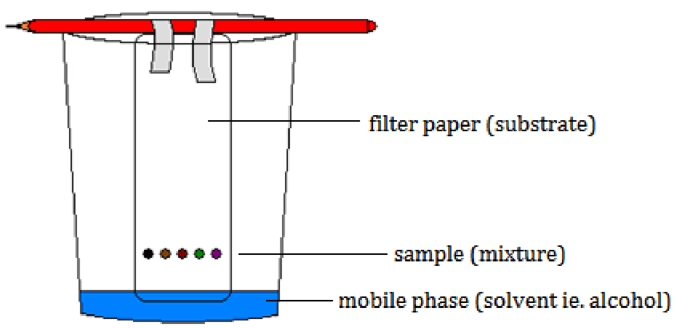
Centrifuge

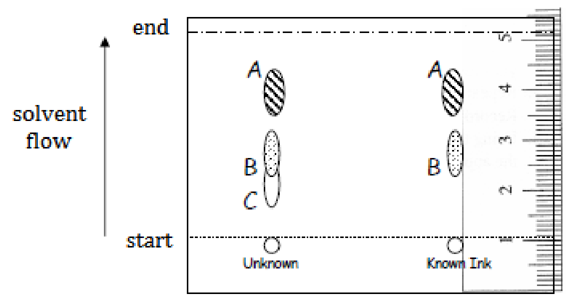
Decanting

Homogeneous mixtures

- **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 Rf value.

Rf = distance pure compound migrated/distance solvent migrated.





Practice Problem 15) Use the chromatography data (above right) to answer the following questions.

a) How many pure compounds make up the unknown ink? ____

b) How many pure compounds make up the known ink? _____

c) Do either of the inks have an pure compounds in common? How do you know?

d) Calculate the Retention Factor, Rf value for compound A

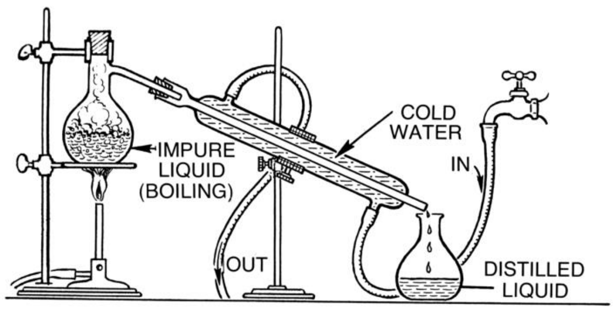
e) a) Which pure compound (A, B, or C) greatest adhesion to the chromatography paper? ____

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

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

h) What is the Rf 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. 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.



Practice problem 16)

a) 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 and water	______________	______________
Salt and water	______________	______________
Food Coloring	______________	______________
Alcohol and water	______________	______________

b) Separating Mixtures. Match the separation methods on the left with the descriptions 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
d) Used to separate homogeneous or heterogeneous mixtures based on BP differences of the components in the mixture.

Polyatomic Ion List

+1

ammonium	NH_4^+
hydronium	H_3O^+

-1

Acetate	$C_2H_3O_2^-$, or CH_3COO^-
bromate	BrO_3^-
bromite	BrO_2^-
chlorate	ClO_3^-
chlorite	ClO_2^-
cyanide	CN^-
dihydrogen phosphate	$H_2PO_4^-$
hydrogen carbonate	HCO_3^- (bicarbonate)
hydrogen sulfate	HSO_4^- (bisulfate)
hydrogen sulfite	HSO_3^- (bisulfite)
hydroxide	OH^-
hypochlorite	ClO^-
hypobromite	BrO^-
hydrogen sulfide	HS^-
iodate	IO_3^-
nitrate	NO_3^-
nitrite	NO_2^-
permanganate	MnO_4^-
perchlorate	ClO_4^-
thiocyanate	SCN^-
cyanate	CNO^-

-2

carbonate	CO_3^{-2}
chromate	CrO_4^{-2}
dichromate	$Cr_2O_7^{-2}$
hydrogen phosphate	HPO_4^{-2}
oxalate	$C_2O_4^{-2}$
peroxide	O_2^{-2}
sulfate	SO_4^{-2}
sulfite	SO_3^{-2}
thiosulfate	$S_2O_3^{-2}$

-3

phosphate	PO_4^{-3}
phosphite	PO_3^{-3}
arsenate,	AsO_4^{-3}
borate	BO_3^{-3}

-ite is one less oxygen than the –ate
Hypo- is one less oxygen than the –ite
Per- is one more oxygen than the –ate
Hydrogen can be added to -2 or -3 ions to make a "new ion" i.e. $H_2PO_4^{-1}$ is dihydrogen phosphate (note the – charge went up 1 for each H^+ added)

Strong Acids

HCl - hydrochloric acid

HBr - hydrobromic acid

HI - hydroiodic acid

HNO_3 - nitric acid

$HClO_3$ - chloric acid

$HClO_4$ - perchloric acid

H_2SO_4 - sulfuric acid

Strong Bases

LiOH - lithium hydroxide

NaOH - sodium hydroxide

KOH - potassium hydroxide

RbOH - rubidium hydroxide

CsOH - cesium hydroxide

$Ca(OH)_2$ - calcium hydroxide

$Sr(OH)_2$ - strontium hydroxide

$Ba(OH)_2$ - barium hydroxide