CHEM 1211
Chapters 1 & 1R:
Chemical Foundations &
Tools of Quantitative Chemistry

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<u>Section</u>: Chemistry and its Methods

I. **Chemistry** is the branch of science dedicated to studying matter's properties, behavior, and transformations. Often regarded as "the central science," it serves as a bridge connecting various scientific disciplines.

II. **The scientific method** is the path that leads from observations, hypotheses, and experiments into theories or laws. The pathway of the scientific method is **questions or observations → hypothesis → experimental tests → repeatable results that either support or refute the hypothesis.**
   - **a.** An **observation** is something that you notice in nature or an experiment.
     - **i.** **Qualitative observations** consist of *non-numerical* data such as the color of a substance, its physical appearance or smell, etc.
     - **ii.** **Quantitative observations** consist of *numerical* data such as the mass of a substance or the temperature at which a substance melts or boils. Quantitative information must contain a number and a unit.
   - **b.** A **hypothesis** is a tentative explanation or prediction based on observations.
     - **i.** It is important to note that a hypothesis must be falsifiable to be considered scientific and must be testable with an **experiment**.
   - **c.** A **theory (or model)** is a unifying principle that explains a group of facts and the laws based on them. If a hypothesis can explain a large amount of experimental data, it can move from a hypothesis to a theory. Essentially, a theory amplifies a hypothesis and gives predictions.
     - **i.** Theories describe **why/how** a particular situation occurs, or the underlying reasons for them and can be changed if new information is presented.
   - **d.** A **law** is a concise statement or summary of a relation that is always the same under the same conditions and will describe some facet of the natural world.
     - **i.** Laws will describe **what** will occur in a particular situation. Laws summarize a series of related observations, and a mathematical expression will support many laws.

<u>Example</u>: A _________________ summarizes a pattern found in observations and a _________________ is the explanation based on experiments.

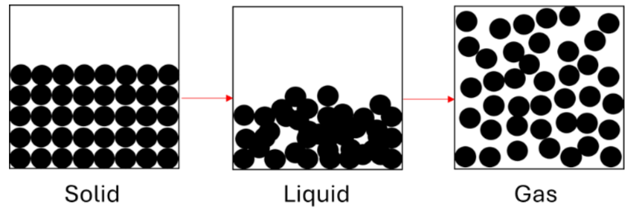
<u>Example</u>: In the lab, you mix two chemicals and notice that the reaction between the two chemicals produces light. Noticing that light is produced is an example of a ______________ and a tentative explanation of this phenomenon that could be further tested would be your ________________.

<u>Example</u>: Describe each statement below as a hypothesis, a theory, or a law.
a) Energy is neither created nor destroyed during a chemical reaction. ________________
b) When a piece of aluminum is placed in hydrobromic acid, bubbles will form. ________________
c) An apple will fall from a tree to the ground because gravity pushes it. ______________

<u>Example</u>: Determine whether each of the following observations is quantitative or qualitative.
a) Science Guyz has a lot of students. ____________________
b) I will go to Science Guyz office hours for 2 hours today. ____________________
c) I was at Science Guyz for a long time today. ____________________
d) It was very hot today as I walked to pick up my Science Guyz packets. ___________________
e) I had five packets to pick up at Science Guyz today. ___________________

Section: Classifications of Matter

I. **Matter** is anything that has mass and occupies space and can exist in three primary **states**: solid, liquid, or gas. Each state is defined by its unique properties.



- **a.** In **solids**, the attractive forces between particles are strong, causing them to be tightly packed in a regular, organized pattern.
  - **i.** Solids are rigid, retain fixed volume and shape, are minimally compressible, and do not expand much upon heating.
  - **ii.** The particles within a solid do contain some energy, which causes the particles to vibrate back and forth about their average positions, however, particles within a solid rarely move past one another.
- **b.** In **liquids**, particles are arranged randomly rather than in a fixed pattern. Liquids are fluid because the particles of liquids are not confined to specific areas, and they can move past each other.
  - **i.** Liquids have no definite shape but have a definite volume.
  - **ii.** Liquids will assume the shape of the container in which they occupy, are minimally compressible and do not expand much upon heating.
- **c.** In **gases**, under ideal conditions, the particles composing the gas are far apart. Gas particles fly about colliding with one another and the walls of their containers. The random motion of gas particles allows a gas to fill the volume of a closed container.
  - **i.** Essentially, the volume of any closed container that contains gas equals the volume of the gas within the container.
  - **ii.** Gases will have no fixed shape or volume, can flow, are highly compressible, and have high expansion upon heating.

Example: Determine whether each statement provided below is **true** or **false**.
a) Liquids are the only state of matter that can flow. __________
b) Gases and liquids are both highly compressible. __________
c) Liquids and gases have definite volumes, but not shapes. __________
d) Liquids and gases will both significantly expand when heated. __________

Example: We have two large containers present: container 1 contains 150 grams of liquid water and container 2 contains 150 grams of water vapor. Use this setup to answer the following questions.
a) Which container contains a greater amount (mass) of water? _____________
b) Which container would have the water take up more space (volume)? _____________
c) Which container would have water in a state of matter where it can flow?_____________

II. The **Kinetic-Molecular Theory of Matter** states that all matter is composed of microscopic particles in constant motion, and the amount of kinetic energy present in the sample determines the state of matter of that sample.

- **a.** **Energy** is the capacity to do work or to transfer heat and the **Law of Conservation of Energy** states that energy cannot be created or destroyed, only converted from one form to another.

- **b. Kinetic energy** is energy because of motion and is dependent on the mass and velocity of an object.
  - **i. Thermal energy** is the form of kinetic energy associated with heat in matter.
- **c. Potential energy** is energy due to position in space or composition. The energy stored in chemical bonds is a type of potential energy.

III. **Physical properties** are properties that can be observed and measured without changing the composition of the substance.

- **a.** Examples of physical properties include color, odor, state of matter, melting point, boiling point, density, solubility, electrical conductivity, malleability, ductility, viscosity, hardness, etc.
- **b. Physical changes** are changes in the physical state (solid, liquid, or gas) or size/shape of a substance.

IV. **Chemical properties** are properties that determine whether and how readily a substance **reacts** (changes into a different substance).

- **a. Chemical changes** are changes that convert one or more substances into one or more different substances. A chemical change will **always** result in a change in composition.
  - **i.** There are other indicators that a chemical change has occurred, including odor production, color change, light production, new product formation, sound production, change in energy, and fizzing or foaming (gas).

Example: Describe the following changes as a physical change or a chemical change.

a) A piece of paper is burned. ________________
b) Paper is balled up. ________________
c) Removing salt from seawater. ______________
d) A metal is attracted to a magnet. ________________
e) Combustion of ethanol. ________________
f) Ice melts. ________________
g) Grilling a steak. ________________
h) Dicing an apple. ________________
i) Lighting a match. ________________
j) Sugar dissolving in water. ________________

Example: Determine which of the statements below represents a **physical property** of some arbitrary substance being explored in the lab. *Select all that apply*.

a) The substance has a density of 2.70 g/mL.
b) The substance becomes flat when pressed down.
c) The substance has a silver color.
d) The substance blackens when heated over a long period.
e) The substance corrodes over time when exposed to air.

V. **Pure substances** have a unique set of physical properties by which they can be recognized. Pure substances *cannot* be separated into two or more different species by any physical technique at ordinary temperatures.

- **a. Elements** are substances that are composed of only a single type of atom.
- **b. Molecules** are two or more atoms (can be the same or different) chemically joined together.
- **c. Compounds** are groups of two or more **different** atoms joined by chemical bonds in fixed ratios.
  - **i.** When atoms bind together to form compounds, the original properties of the elements (color, hardness, melting point, and boiling point) are replaced by the properties of the compound.
  - **ii. In general, all compounds are molecules, but not all molecules are compounds.**

c. **Mixtures** consist of two or more pure substances and can be separated by physical techniques. There are two different types of mixtures that we must consider.
    i. **Heterogeneous mixtures** occur when the components of the mixture are unevenly distributed, or there is not a constant composition throughout. Examples include a bowl of cereal, milk, chicken salad, or a rusted barbell.
    ii. **Homogeneous mixtures** occur when two or more substances, in the same phase, in which the substances are evenly distributed, or we have a constant composition throughout. Homogenous mixtures are also referred to as **solutions**. Examples include wine, gasoline, black coffee, and common sports drinks.

VI. Physical techniques can be used to separate mixtures.
  a. **Filtration** is used when a mixture is made up of both a solid and a liquid. A filter will be used to separate the solid from the liquid (can separate heterogeneous mixtures).
  b. **Distillation** is used based on the difference in how readily two substances in a mixture become gases. Characterized by liquid vaporization followed by condensation and can be used to separate homogenous liquid mixtures.
  c. **Chromatography** involves a system with two phases of matter, known as the mobile phase and the stationary phase. The stationary phase is a solid and the mobile phase is either gas or liquid. Separation occurs when the mobile phase moves over the stationary phase, and the components separate based on the affinity of the solid being used.

Example: In each item below, first classify it as a pure substance or mixture. If it is a pure substance, further classify it as an element or compound. If it is a mixture, further classify it as a homogeneous mixture or a heterogeneous mixture.
a) Solid iodine (I_2). ________________________________
b) Aluminum (Al) foil. ________________________________
c) Aluminum (Al) foil that has been slightly corroded. ________________________________
d) Air. ________________________________
e) A saltwater solution. __________________________
f) Carbon dioxide gas (CO_2). ________________________________
g) Soda with ice cubes. ________________________________
h) Salsa. ________________________________
i) Vodka. ________________________________
j) Copper (Cu) wire. ______________________________
k) Brass metal. ______________________________

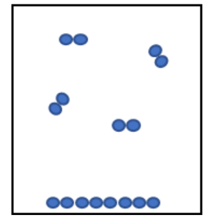
Example: The breakdown of liquid X produces two different liquids that are known to be pure substances. Using this information, answer each of the following statements as **true** or **false**.
a) Liquid X cannot be an element. _______________
b) The products from the breakdown are for sure both elements. _______________

Example: What separation method could be used to separate a mixture of *undissolved* salt crystals in water?

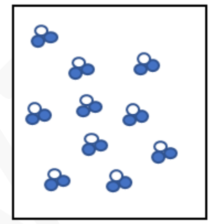
Example: What separation method could be used to separate a mixture of *dissolved* salt crystals in water?

VII. Interpreting models and figures is an important skill you will develop throughout your time in Chemistry. We will walk through a few examples of how to do this.

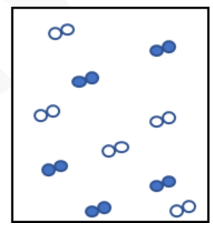
a. The model on the right has two spheres linked together that are the same color. Since all the spheres are the same color and nothing else is present, this is an example of a **pure substance**, more specifically a **molecule** made from the same element. Another interpretation of this model could be a **phase change** since the molecules are close together with a defined shape at the bottom (solid) and closer to the top they are more spread out with no specific shape or volume (gas).



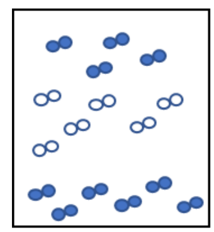
b. The model to the right has three spheres linked together with two of them being the same color and one being a different color. Since the spheres are not all the same color, this is an example of a **compound**. Understand that this is still an example of a **pure substance** since there is nothing else present besides that one compound.



c. The model to the right has two different sets of spheres linked together that are different colors. Since we have these two separate sets of spheres that are different colors, we know that we are looking at a **mixture**. Defining this mixture further, we can see that there is an even distribution of the different colored spheres, meaning that we are looking at a **homogeneous mixture**.

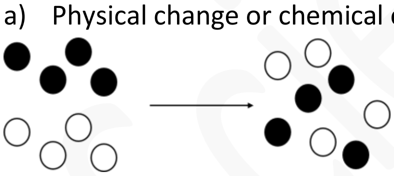


d. The model to the right has two different sets of spheres linked together that are different colors. Since we have these two separate sets of spheres that are different colors, we know that we are looking at a **mixture**. Defining this mixture further, we can see that there is not an even distribution of the different colored spheres (they are layered), meaning that we are looking at a **heterogeneous mixture**.

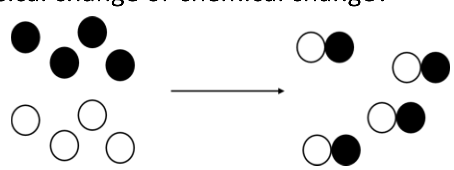


__Example__: Identify whether the models below depict chemical changes or physical changes.

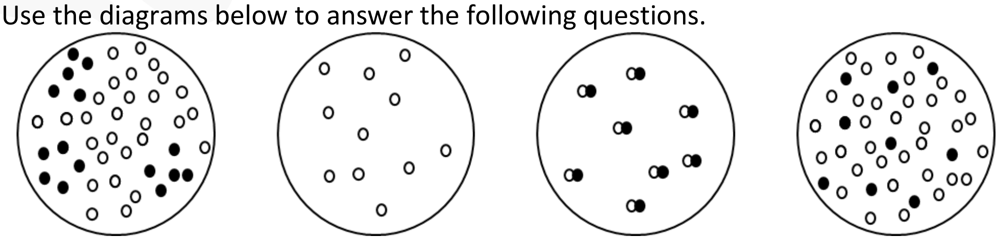
a) Physical change or chemical change?



b) Physical change or chemical change?



__Example__: Use the diagrams below to answer the following questions.



I II III IV

a) Which diagram represents a pure substance that is an element? ______________
b) Which diagram represents a pure substance that is a compound? ______________
c) Which diagrams represent pure substances? ______________
d) Which diagrams represent mixtures? ______________

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<u>Section</u>: The Periodic Table of the Elements

I. The **periodic table** is used to organize the different elements. You must be familiar with the organization of this table.

a. You are generally expected to know the names, spelling, and symbols of all elements in the first five rows of the periodic table. You are also expected to the names, spelling, and symbols of certain elements in rows six and seven, which will be indicated.

H 1 1.01																	He 2 4.00
Li 3 6.94	Be 4 9.01											B 5 10.81	C 6 12.01	N 7 14.01	O 8 16.00	F 9 19.00	Ne 10 20.18
Na 11 22.99	Mg 12 24.31											Al 13 26.98	Si 14 28.09	P 15 30.97	S 16 32.06	Cl 17 35.45	Ar 18 39.95
K 19 39.10	Ca 20 40.08	Sc 21 44.96	Ti 22 47.87	V 23 50.94	Cr 24 51.99	Mn 25 54.94	Fe 26 55.85	Co 27 58.93	Ni 28 58.69	Cu 29 63.55	Zn 30 65.38	Ga 31 69.72	Ge 32 72.63	As 33 74.92	Se 34 78.97	Br 35 79.90	Kr 36 83.80
Rb 37 85.47	Sr 38 87.62	Y 39 88.91	Zr 40 91.22	Nb 41 92.91	Mo 42 95.95	Tc 43 98.91	Ru 44 101.07	Rh 45 102.91	Pd 46 106.42	Ag 47 107.87	Cd 48 112.41	In 49 114.82	Sn 50 118.71	Sb 51 121.76	Te 52 127.60	I 53 126.90	Xe 54 131.29
Cs 55 132.91	Ba 56 137.33	57-71	Hf 72 178.49	Ta 73 180.95	W 74 183.84	Re 75 186.21	Os 76 190.23	Ir 77 192.22	Pt 78 195.09	Au 79 196.97	Hg 80 200.59	Tl 81 204.38	Pb 82 207.2	Bi 83 208.98	Po 84 [209]	At 85 [210]	Rn 86 [222]
Fr 87 [223]	Ra 88 [226]	89-103	Rf 104 [267]	Db 105 [268]	Sg 106 [269]	Bh 107 [270]	Hs 108 [269]	Mt 109 [277]	Ds 110 [281]	Rg 111 [282]	Cn 112 [285]	Nh 113 [286]	Fl 114 [290]	Mc 115 [290]	Lv 116 [293]	Ts 117 [294]	Og 118 [294]

La 57 138.91	Ce 58 140.12	Pr 59 140.91	Nd 60 144.24	Pm 61 [145]	Sm 62 150.36	Eu 63 151.96	Gd 64 157.25	Tb 65 158.93	Dy 66 162.50	Ho 67 164.93	Er 68 167.26	Tm 69 168.93	Yb 70 173.05	Lu 71 174.97
Ac 89 [277]	Th 90 232.04	Pa 91 231.04	U 92 238.03	Np 93 [237]	Pu 94 [244]	Am 95 [243]	Cm 96 [247]	Bk 97 [247]	Cf 98 [251]	Es 99 [252]	Fm 100 [257]	Md 101 [258]	No 102 [259]	Lr 103 [262]

b. Elements **cannot** be decomposed into simpler substances by chemical or physical techniques and most elements exist as monatomic particles in nature (single atoms), however, some elements will exist as molecules in their most stable form, and you should commit these molecules to memory.

i. Diatomic Elements: Hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine (**H**ave **N**o **F**ear **Of** **I**ce **C**old **B**eer).

ii. Sulfur and phosphorus also exist as molecules (S_8 and P_4, respectively).

<u>**Example**</u>: Write the name of each element for each elemental symbol provided below.

a) K: ______________________ b) W: ______________________
c) Sb: ______________________ d) Cu: ______________________
e) Ag: ______________________ f) Au: ______________________
g) Hg: ______________________ h) Na: ______________________
i) Sn: ______________________ j) Pb: ______________________
k) Mn: ______________________ l) Mg: ______________________
m) Sr: ______________________ n) Se: ______________________
o) B: ______________________ p) Br: ______________________
q) Fe: ______________________ r) P: ______________________

<u>**Example**</u>: Provide the chemical symbol for the following examples.

a) Carbon = ______ b) Calcium = ______ c) Cobalt = ______
d) Chromium = ______ e) Chlorine = ______ f) Copper = ______

<u>**Example**</u>: Which elemental symbols are **not** matched up with the correct name? <u>*Select all that apply*</u>.

a) Mn = Magnesium b) Br = Barium c) Pb = Lead d) Sb = Tin e) Sr = Strontium

<u>Section</u>: Scientific Notation & Units of Measurement

I. A number is written in **scientific notation** when in the form $a \times 10^n$ where $1 \le |a| < 10$ and n is an integer. Scientific notation makes a large or small number more compact by writing them as a product of the power of 10.

a. In putting a large number into scientific notation, move the decimal place to the **left** until $1 \le |a| < 10$. With $a \times 10^n$, n will be a positive number equal to the number of times you moved the decimal to the left.

b. In putting a small number into scientific notation, move the decimal place to the **right** until $1 \le |a| < 10$. With $a \times 10^n$, n will be a negative number equal to the number of times you moved the decimal to the right.

<u>Example</u>: Convert the following numbers into or out of scientific notation.

a) 0.0000783 = ____________________ b) 985512 = ____________________

c) 9.31×10^6 = ____________________ d) 7.55×10^{-3} = ____________________

II. **The International System of Units, or SI**, is the scientific system for measurements. Most measurements in Chemistry are made in SI units and it is important to know the SI unit used for each property shown to the right.

Property	Unit Used
Mass	kilogram (kg)
Length	meter (m)
Time	second (s)
Temperature	kelvin (K)
Amount of Substance	mole (mol)
Electric Current	ampere (A)
Luminous Intensity	candela (cd)

III. **Common SI Prefixes (Commit to Memory!):**

Prefix	Abbreviation	Meaning	Example
Exa-	E	10^{18}	1 exabyte = 1 x 10^{18} bytes
Peta-	P	10^{15}	1 petabyte = 1 x 10^{15} bytes
Tera-	T	10^{12} (trillion)	1 terahertz = 1 x 10^{12} hertz
Giga-	G	10^{9} (billion)	1 gigahertz = 1 x 10^{9} hertz
Mega-	M	10^{6} (million)	1 megaton = 1 x 10^{6} tons
Kilo-	k	10^{3} (thousand)	1 kilogram = 1 x 10^{3} grams
Hecto-	h	10^{2} (hundred)	1 hectopascal = 1 x 10^{2} pascals
Deka-	da	10^{1} (ten)	1 dekameter = 1 x 10^{1} meters
Base Unit	-	10^{0}	________________________
Deci-	d	10^{-1} (tenth)	1 decimeter = 1 x 10^{-1} meters
Centi-	c	10^{-2} (hundredth)	1 centimeter = 1 x 10^{-2} meters
Milli-	m	10^{-3} (thousandth)	1 millimeter = 1 x 10^{-3} meters
Micro-	µ	10^{-6} (millionth)	1 micrometer = 1 x 10^{-6} meters
Nano-	n	10^{-9} (billionth)	1 nanometer = 1 x 10^{-9} meters
Pico-	p	10^{-12} (trillionth)	1 picometer = 1 x 10^{-12} meters
Femto-	f	10^{-15}	1 femtometer = 1 x 10^{-15} meters
Atto-	a	10^{-18}	1 attometer = 1 x 10^{-18} meters

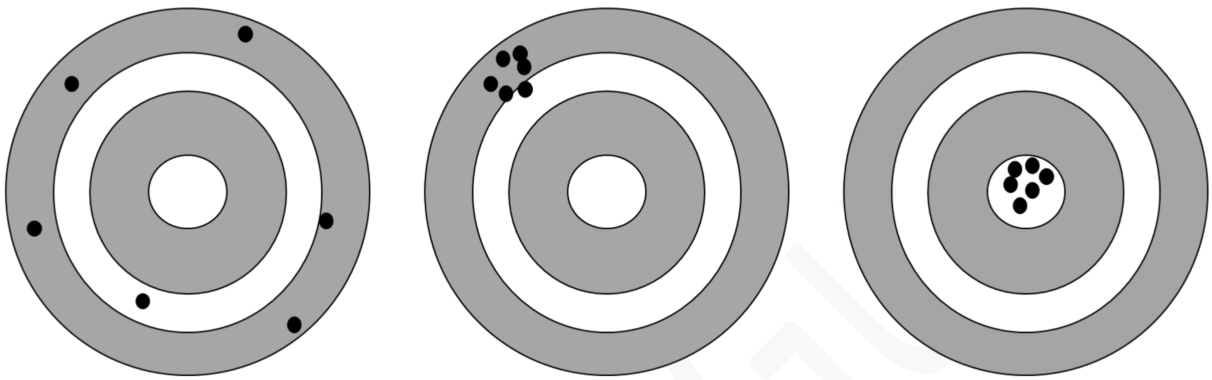
<u>Example</u>: Provide the correct abbreviations for the following units of measurement.

a) micrometer: _____ b) nanogram: _____ c) milliliter: _____ d) centimeter: _____

<u>Section</u>: Accuracy vs Precision

I. An ideal instrument will provide a measurement that is both accurate and precise.

- **a. Accuracy** is a measure of how close a value obtained is to the true value.
- **b. Precision** refers to the degree of agreement among several measurements of the same quantity.
  - **i.** It is possible for measurements to have high precision, yet low accuracy.
- **c.** It is important to point out that *ALL* measurements have some degree of uncertainty**.**



Low precision, low accuracy **High precision, low accuracy** **High precision, high accuracy**

<u>Example</u>: A group of students in lab are asked to make several measurements regarding the amount of liquid present in a graduated cylinder. The amount of liquid present is known by the TA to be 31.5 mL. Their collected measurement data is shown in the table.

Banks	Emma	Connor	Austin	Kailee
31.2 mL	30.7 mL	29.7 mL	29.1 mL	41.2 mL
31.5 mL	30.5 mL	31.1 mL	30.1 mL	41.5 mL
31.8 mL	30.6 mL	27.6 mL	31.0 mL	41.8 mL

a) Which of the student's measurements are the most **accurate**? ____________________

b) Which of the student's measurements are the most **precise**? ____________________

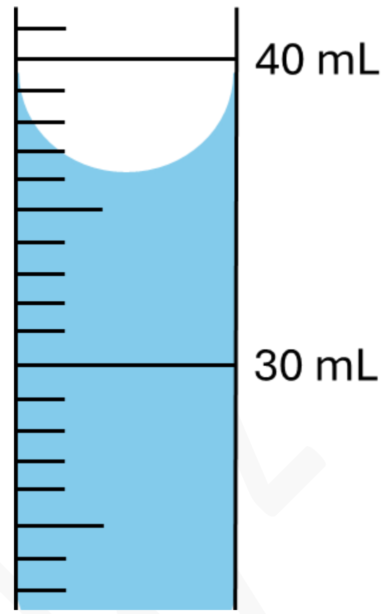
<u>Example</u>: You are playing Pin the Tail on the Donkey, but you are only given a single tail to try and pin in the correct spot. Do you need to be accurate, precise, or accurate and precise when pinning this single tail on the donkey? Explain.

II. Errors in measurements may occur due to both random and systematic errors.

- **a. Random error (or indeterminate error)** tells us that a measurement has an equal chance of being too low as it does being too high.
- **b. Systematic error (or determinate error)** occurs in the same direction each time, always being too high or too low.

<u>Example</u>: A student in the lab weighs the amount of an intermediate produced in a four-step synthesis reaction. Based on calculations, there should be approximately 10.55 grams of intermediate produced, but after four trials of the experiment, the student obtained the following amount of the intermediate, all weighed by the same scale in the lab: 12.25 grams, 12.26 grams, 12.26 grams, and 12.27 grams. Using what you have learned about accuracy and precision, explain what may be causing these results.

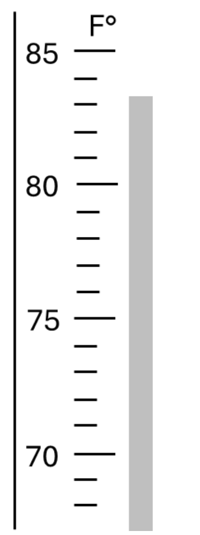
Section: **Significant Figures**



I. **Significant figures** involve the numbers in a measured quantity or value that are known to be correct and one digit that is an estimation. A common application of significant figures is with reading glassware in the lab.

- **a.** To the right is a graduated cylinder. Each mark on the graduated cylinder represents a 1 mL increase in volume.
- **b.** With the graduated cylinder to the right, it is known for sure that the true liquid measurement lies somewhere between 36 mL and 37 mL (always read from the meniscus).
- **c.** We are confident in the value of the ones and tens place, but the digit in the tenth place is an approximation.
- **d.** An appropriate guess for liquid in this graduated cylinder would be something along the lines of 36.1 mL or 36.2 mL. Either of these answers would be acceptable since the digit in the tenth place is an approximation.
- **e.** This graduated cylinder is one example of a measurement tool but understand that different measurement tools will provide a different number of significant figures.

Example: A thermometer filled with mercury is shown to the right. What is the temperature reading for this thermometer, considering significant figures?



II. **Rules for Significant Figures**:

- **a.** All non-zero numbers are significant.
- **b.** Zeroes between two other significant digits are significant, including numbers that have a decimal place (7.**0**51, 6**0**71**0**1).
- **c.** Zeroes following a non-zero number that is also to the left of a decimal are significant (72**000**.).
- **d.** In numbers containing a decimal, all zeros at the end of the number are significant (0.032**0**).
  - **i.** Combining rules c and d shows us that the three zeros at the end of 0.071**000** are significant.
- **e.** Zeroes that do not have either a decimal point or non-zero digit to the right of them are "trailing" zeros and are **not** significant (32**0**).
- **f.** Zeroes that occur before any non-zero number are **not** significant (**0**.**00**147).

III. **Exceptions to Traditional Rules:**

- **a.** Counting numbers have an infinite number of significant figures because they represent exact quantities and cannot be made more precise.
  - **i.** For instance, counting 12 pencils or stating that a molecule consists of 3 atoms are examples of exact values.
- **b.** Known conversions have an unlimited number of significant figures.
  - **i.** For instance, 1 foot = 12 inches is an example of an errorless conversion.

Example: Determine the number of significant figures contained within each of the following values.

a) 0.0900: _________ b) 6.230: _________ c) 0.0076: _________

d) 4.00028: _________ e) 4.335 x 10^{-22}: _________ f) 5600: _________

g) 0.12000 inches: _________ h) 1.20 feet: _________ i) 5600 pieces of paper: _________

j) 5462.0 kilometers: _________ k) 5462.00 kilometers: _________

IV. Rounding Following Mathematical Operations:

- **a.** Following any mathematical operation, you must determine how many significant figures will be in your final answer.
- **b.** Once you have determined the correct number of significant figures, you must round your answer to the correct number of significant figures by looking at the first non-significant digit in your answer.
  - **i.** If this number is 5 or greater, you will round your last significant digit up, otherwise, you will round the last significant digit down.

V. Significant Figures in Multiplication and Division Problems:

- **a.** When multiplying or dividing two or more numbers, your answer must contain the same number of significant figures as the value in your equation with the fewest number of significant figures.

Example: Perform the operation below and report your answer to the correct number of significant figures.

(3.2005470) x (30.9) = ____________

VI. Significant Figures in Addition and Subtraction Problems:

- **a.** When adding or subtracting two or more numbers, your answer must contain the same number of decimal places as the value in the equation with the fewest decimal places.

Example: Perform the operation below and report your answer to the correct number of significant figures.

(321.1896) + (1.98665) + (0.1) = ____________

VII. Significant Figures with Multiple Operations:

- **a.** You will often find yourself performing operations that involve both addition/subtraction rules and multiplication/division rules. Perform operations in order as dictated by PEMDAS and carry the exact numbers that you obtain throughout the operation. We will only consider significant figures at the end.

Example: Perform the operation below and report your answer to the correct number of significant figures.

a) (2.9 x 4.719) + 12.710 = ____________

b) (30.0031 + 0.3) (6.211 – 6.185)/ (5.233 x 10^{-2}) = ____________

<u>Section</u>: Dimensional Analysis

I. In Chemistry and other STEM courses, you will often need to convert from one unit to another unit. This is sometimes a multistep process that can become tedious. **Dimensional analysis** provides a strategy to organize these conversions.

1. When unit conversions are needed, first establish what your final units must be in.
2. Consider the numerator/denominator relationship of the final unit.
   a. For example, miles per hour requires miles in the numerator and hours in the denominator (miles/hour). So, the result of your dimensional analysis conversions should reflect this.
3. Choose a value as a starting point for your conversion.
4. Set up a series of conversion steps that allow you to cancel all unwanted units and leave you only with the final desired units. Units can be canceled when the same unit can be found in the numerator of one step and the denominator of another step (or vice versa), *regardless of whether they are in consecutive steps.*
5. Once your units have been canceled to your final units, multiply the numerators and denominators, divide the products, and you are finished. Always make sure you report your final answer to the correct number of significant figures.

<u>Example</u>: How many picograms are in a milligram?

<u>Example</u>: How many angstroms (Å) are in 442 nm? <u>Note</u>: 1 Å = 1 x 10^{-10} m

<u>Example</u>: Convert 5.10 meters per second to micrometers per hour.

<u>Example</u>: Kailee is traveling to Atlanta. One gallon of gas will allow her to travel for 8.5 miles. If it takes her 95 minutes to get to Atlanta and she travels at an average speed of 65 miles/hour, how many gallons of gas does she use?

<u>Example</u>: You are a medical assistant and need to administer medication to help a patient with hip pain. The recommended dosage of this medication is 6.71 mg/kg of body mass. What would be the dosage in milligrams for a 225-lb individual? <u>Note</u>: 1 lb = 453.59 g.

<u>Example</u>: Austin is moving up in the world and has decided it is time to install a pool at his house. Austin wants the pool to have the following dimensions: 37.2 ft long, 15.0 ft wide, and 7.92 ft deep. Once the pool has been completed, how much water will Austin need to fill the pool to the top (in cubic inches)?

<u>Example</u>: Convert the cubic inches obtained in the previous problem into milliliters. <u>Note</u>: 1 inch = 2.54 cm and 1 cm^3 = 1 mL.

<u>Section</u>: Temperature Scales

I. **Temperature** is a physical quantity that reflects the presence of thermal energy and determines whether something feels hot or cold. There are three different temperature scales that you need to become familiar with and how to convert between these different scales.
   - **a.** **The Celsius (°C) scale** is defined by assigning 0 °C to the freezing point of pure water and 100 °C to the boiling point of pure water.
   - **b.** **The Fahrenheit (°F) scale** is defined by assigning 32 °F to the freezing point of pure water and 212 °F to the boiling point of pure water.
   - **c.** **The Kelvin (K) scale** uses the same size unit as the Celsius scale but assigns 0 to the lowest possible temperature (**absolute zero**).
     - **i.** There are no negative values associated with this scale and the "°" symbol is not used.
     - **ii.** The Kelvin scale is the only temperature scale that is an absolute temperature scale.
   - **d.** The formulas you will primarily be working with are provided below. On your exam, be ready to convert between the three scales and/or be asked a conceptual question about the scales.
     - **i.** When converting between the different temperature scales, watch out for significant figures!

$$\mathbf{K = {}^{\circ}C + 273.15} \qquad\qquad \mathbf{{}^{\circ}C = \frac{5}{9}({}^{\circ}F - 32)}$$

<u>Example</u>: Convert each of the following temperatures into the desired unit.

a) 608 $°C$ to Kelvin: __________

b) -48.1 $°C$ to Fahrenheit: ___________

c) 85 K to Fahrenheit: __________

d) 155 $°F$ to Kelvin: __________

<u>Example</u>: Which of the statements provided below regarding the temperature scales is/are **true**? *<u>Select all that apply</u>*.

a) The boiling point of water on the Celsius scale is 212 degrees.
b) The freezing point of water in both the Fahrenheit and Celsius scales is 0 degrees.
c) Negative values are possible in the Celsius, Fahrenheit, and Kelvin scales.
d) The Celsius degree is larger than the Fahrenheit degree.
e) The size of the temperature unit is the same for the Celsius and Kelvin scales.

Section: Density

I. **Density** is an intensive property representing the relationship between the mass of a substance and the amount of space that it takes up. Density is a quantity represented, mathematically, by dividing the mass of the pure substance by its volume.

$$\boxed{\boldsymbol{Density = \frac{mass\ (g)}{volume\ (mL\ or\ cm^3)}}}$$

- **a.** Density was used, before modern analytical methods, to determine the identity of an unknown substance.
- **b.** In liquid and gas mixtures, substances that are less dense in the mixture will "float" on top of substances that are denser.
  - **i.** For example, if ethanol (density is 0.789 g/mL) and water (density is 1.00 g/mL) were poured into a beaker, the ethanol would float to the top and the water would sink to the bottom after some time.
- **c.** A common application of density problems involves fluid displacement. The displacement of a fluid before an object is added to it and after can be used to determine the volume of the object used in the displacement.
- **d.** An **intensive property** is a bulk property, meaning that it is a physical property of a system that does not depend on the size or the amount of sample in the system. Examples include density, odor, color, luster, malleability, ductility, conductivity, temperature, hardness, melting point, boiling point, average atomic mass, etc.
- **e.** An **extensive property** is a property of a system that changes with the size of the sample measured (it is additive). Examples include mass, volume, and length.

Example: A commonly used metal in a manufacturing plant is known to have a density of 11.25 g/mL. If a 0.400 kg cube sample of the metal were taken, what would the volume of the metal be (in cm^3)?

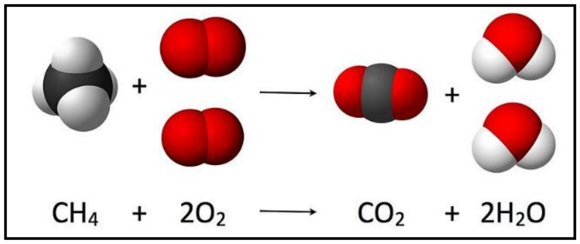
Example: There is a small pellet of some metal that has a mass of 3.45 grams. The pellet is placed into a beaker that contains 20.00 mL of water. The metal pellet is then submerged in the water, which causes the water level in the beaker to rise to 21.28 mL. Determine the density of the metal pellet (in g/cm^3).

Example: A student places a 4.56 gram aluminum cube (density = 2.70 g/cm^3) into a graduated cylinder filled with some water and sees that the volume is 75.56 mL. Determine the initial volume reading (in mL) of the graduated cylinder.

Example: There are three globes on display at your favorite local tutoring company, Science Guyz. These globes have the same mass but have increasing density with globe 1 having the smallest density and globe 3 having the largest density. Rank the three globes in terms of **increasing** volume.

Section: Fundamental Laws in Chemistry & Dalton's Atomic Theory

II. The Law of Conservation of Mass: For any system closed to transfers of matter, the mass of a system cannot change if mass is not added or removed from it. Mass cannot be created or destroyed.



Example: Hydrogen gas and oxygen gas are known to react to form water. Use this setup to answer the following questions.

a) A reaction vessel contains 2.02 grams of hydrogen gas and 16.00 grams of oxygen gas. How many grams of water will be produced in the complete reaction of these two gases?

b) In a separate experiment, it was observed that 36.04 grams of water were produced. Assuming all the oxygen gas and hydrogen gas were reacted to form the water, how much hydrogen gas was used to form the 36.04 grams of water, given that 32.00 grams of oxygen gas were used?

III. The Law of Definite Proportions (Law of Constant Composition): Joseph Proust originally proposed this law and states that a given chemical species will always contain the same elements in the same proportion by mass, no matter the sample size.

- **a.** Water (H_2O) is commonly used to illustrate the Law of Definite Proportions.
  - **i.** The fact that all samples of water (no matter how much water is in the sample) are made of ~11.21% hydrogen and ~88.79% oxygen illustrates this law.
- **b.** Let us look at three different samples of carbon dioxide (CO_2).

Mass of Sample (in g)	Mass of Carbon in Sample (in g)	Mass of Oxygen in Sample (in g)
10.00 g	2.729 g	7.271 g
30.00 g	8.187 g	21.813 g
70.00 g	19.103 g	50.897 g

Example: Two separate samples of compounds that contain only carbon and oxygen are provided to the right. Using the data, determine if either sample could be CO_2.

Mass of Sample (in g)	Mass of Carbon in Sample (in g)	Mass of Oxygen in Sample (in g)
80.00 g	21.83 g	58.17 g
130.0 g	55.74 g	74.26 g

<u>Example</u>: A sample of water contains 137.32 grams of oxygen and 17.38 grams of hydrogen. If a second sample of water has a total mass of 95.23 grams, what is the mass of hydrogen (in grams) in the second sample?

<u>Example</u>: A sample of some compound X contains 6.00 grams of hydrogen, 10.23 grams of oxygen, and 15.69 grams of carbon. If a second sample contains a total mass of 235.6 grams, what mass of oxygen is present in this second sample?

<u>Example</u>: A sample of ethane gas contains 36.45 grams of carbon and 9.19 grams of hydrogen. A different sample of ethane gas contains 76.20 grams of carbon. Determine the number of grams of hydrogen contained in the second sample of ethane gas.

IV. Dalton used the Law of Conservation of Mass and the Law of Definite Proportions as the basis for his atomic theory, known today as **Dalton's Atomic Theory**. You must be familiar with the postulates of this theory.

1. Matter is composed of small particles known as atoms, which are the smallest units of an element (they are indivisible) that can participate in a chemical change or reaction.
2. An element consists of only one type of atom, which means that atoms of a given element are identical. Atoms of one element differ in properties from atoms of all other elements.
3. A compound consists of atoms of two or more different elements combined in a small, whole-number ratio. A specific compound will always have the same relative numbers and types of atoms.
4. Atoms are neither created nor destroyed during a chemical change; chemical reactions simply involve changes in the way that atoms are bound together.
    - **a.** It is important to note that not all of Dalton's Atomic Theory holds true today.
        - **i.** We now know that atoms are made up of even smaller units, namely protons, neutrons, and electrons.
        - **ii.** All atoms of the same element are not necessarily identical. For example, isotopes of elements exist, which we will discuss in a later section.

<u>Example</u>: Select the statement(s) below that is/are **true** regarding Dalton's Atomic Theory. *<u>Select all that apply</u>*.

a) The relative numbers and types of atoms are consistent in a given compound.
b) Atoms have the ability to be transformed into atoms of a different element.
c) An element is composed of small divisible particles known as atoms.
d) Atoms of a specific element will have identical properties that differ from the properties of all other elements.
e) Compounds will form when atoms of the same elements combine with each other in small whole-number ratios.

V. **The Law of Multiple Proportions:** States that when two elements can form a series of compounds, the ratio of the masses of the second element which combines with a fixed mass of the first element can always be reduced to small whole numbers. Originated as a prediction from Dalton's Atomic Theory. It is easiest to understand this law by looking at an example.

a. Let us look once again at carbon dioxide (CO_2). We observe the mass ratio below:

$$\frac{2.66\ g\ O}{1.00\ g\ C}$$

b. Now, let us look at the molecule carbon monoxide (CO), which has one fewer oxygen atom than carbon dioxide. We observe the mass ratio below:

$$\frac{1.33\ g\ O}{1.00\ g\ C}$$

c. Looking at the compounds in the carbon and oxygen series (CO and CO_2), the masses of oxygen that combine with 1.00 grams of carbon is a small whole number ratio.

$$\frac{2.66}{1.33} = 2$$

Example: Based on what we have discussed, what do you predict the ratio will be for 1.00 g of carbon to oxygen for the carbonate anion (CO_3^{2-})? Test your prediction with the data provided below.

Compound	Mass of Sample (in g)	Mass of Carbon in Sample (in g)	Mass of Oxygen in Sample (in g)
CO	130.00 g	55.74 g	74.26 g
CO_2	80.00 g	21.83 g	58.17 g
CO_3^{2-}	95.00 g	19.01 g	75.99 g

Example: Determine which pair of compounds below could be used to illustrate the Law of Multiple Proportions. *Select all that apply.*

a) NH_4Br and NH_4.
b) $ZnCl_2$ and ZnI_2.
c) NO and NO_2.
d) H_2O and HCl.
e) N_2O_4 and N_2O.

Example: We are given two different compounds composed of gold and chlorine that have the following masses of chlorine for every 65.000 grams of gold: 8.993 grams and 17.986 grams. Are these observations consistent with the Law of Multiple Proportions? Why or why not?

1 H 1.01																	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 51.99	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.97	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc 98.91	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57-71	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.09	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po [209]	85 At [210]	86 Rn [222]
87 Fr [223]	88 Ra [226]	89-103	104 Rf [267]	105 Db [268]	106 Sg [269]	107 Bh [270]	108 Hs [269]	109 Mt [277]	110 Ds [281]	111 Rg [282]	112 Cn [285]	113 Nh [286]	114 Fl [290]	115 Mc [290]	116 Lv [293]	117 Ts [294]	118 Og [294]

57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm [145]	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97
89 Ac [277]	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np [237]	94 Pu [244]	95 Am [243]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]	103 Lr [262]