CHEM 1211 Promo Workshop
Chapter 1:
Chemical Foundations

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Learning Objectives: By the end of this workshop, you should be able to:

- Understand the scientific method and what is included in it.
- Compare a hypothesis, theory, and law and know the differences between them.
- Recognize both quantitative and qualitative information and know the differences between them.
- Describe the difference between precision and accuracy and provide examples of each.
- Describe the difference between systematic versus random error.
- Convert numbers into and out of scientific notation.
- Describe what significant figures are and why they are important.
- Determine how many significant figures are present in a number and the rules behind how to determine this.
- Perform operations such as addition, subtraction, division, and multiplication with significant figures in mind.
- Perform multiple operations while considering significant figures.
- Perform temperature conversions between Celsius, Kelvin, and Fahrenheit.
- Describe conceptually the differences between the three temperature scales.
- Describe the SI units used for common properties.
- Understand the metric system, including the prefixes and abbreviations, and be able to convert to various units within the system.
- Understand why we use dimensional analysis and the applications of it in various situations.
- Recognize common conversions for various units.
- Be comfortable with multi-step unit conversions in dimensional analysis.
- Describe density and how we can apply this physical quantity to various situations.
- Describe the different classifications of pure substances and provide relevant examples of each.
- Describe the different classifications of mixtures and provide relevant examples of each.
- Recognize model interpretations of concepts from this chapter (pure substances, mixtures, etc).
- Describe the difference between physical properties/changes and chemical properties/changes and provide examples of each.
- Describe the three states of matter, as well as the Kinetic-Molecular Theory of Matter.
- Define extensive and intensive properties and provide examples of each.
- Describe the difference between mass and weight.
- Perform mathematical operations involving density.

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

I. **The Scientific Method**: The path that leads from experiments, hypotheses and observations into theories or laws. The pathway of the scientific method is **questions → hypothesis → predictions → experimental tests → repeatable results that either support or refute the hypothesis.**

- **a.** **Observation:** Something that you notice in nature or in an experiment. Observations are also referred to as data.
- **b.** **Hypothesis**: A tentative explanation or prediction based on experimental observations. It is important to note that a hypothesis *MUST* be falsifiable to be considered scientific. We only accept a hypothesis when it has withstood *MULTIPLE* experimental tests and it has good explanatory power.
- **c.** **Theory**: 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.
- **d.** **Law**: A concise statement of a relation that is always the same under the same conditions. Laws describe or predict some facet of the natural world. Some hypotheses attempt to explain a behavior that is summarized in a law.
  - **i.** Laws will describe **what** will occur in a particular situation. They summarize a series of related observations
  - **ii.** Many laws (but not all) will be supported by a mathematical expression.
- **e.** **Qualitative Information**: Consists of *non-numerical* data such as the color of a substance or its physical appearance.
- **f.** **Quantitative Information**: Consists of *numerical* data such as the mass of a substance or the temperature at which a substance melts or boils.
  - **i.** Quantitative information *MUST* contain a number and a unit.

<u>Example</u>: A _____ is a mathematical expression that summarizes a pattern found in observations and a _____ is the explanation of the expression based on experiments.

<u>Example</u>: In lab, you mix two chemicals and notice that the reaction between the two chemicals produces light. This 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 option 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 towards the ground. _____.

<u>Example</u>: Determine whether each of the following examples are **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) The density of ice is less than the density of water. _____.
d) It was very hot today as I walked to pick up my Science Guyz packets. _____.
e) It was 95 °F outside as I walked to pick up my Science Guyz packets. _____.

<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 how close measurements are to one another.
  - **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,** and this uncertainty depends on the precision of the device you are using to make a measurement and your skills as an observer.

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

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

Banks	Ansley	Ben	Austin	Kailee
31.0 mL	30.7 mL	29.7 mL	29.1 mL	31.1 mL
31.5 mL	30.5 mL	31.1 mL	30.1 mL	30.7 mL
30.8 mL	30.6 mL	27.6 mL	31.0 mL	29.9 mL

a) Which of the employee's measurements is the most **accurate**? _____

b) Which of the employee's measurements is the most **precise**? _____

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 equal chance of being too low as it does too high.
- **b.** **Systematic Error (or Determinate Error)** occurs in the same direction each time, always being too high or too low.

<u>Section:</u> Scientific Notation

I. A number is written in **scientific notation** when in the form $a \times 10^n$ where $1 \leq |a| < 10$ and n is an integer.

a. Scientific notation is used to make a large or small number with lots of zeros (like 18,300,000,000) more compact by writing them as a product of a power of 10.

<u>Practice</u>: Write 2,000.0 in scientific notation form.

1. Move the decimal so that the number has a value (absolute) between 1 and 10.
   In this example, we must move the decimal three places to the **left**.

$$2\underbrace{000.0} \Rightarrow 2.0000$$

2. Multiply the value you are left with by 10^n where n is the number of places the decimal was moved. If the decimal had to be moved to the left, make n positive, and if the decimal was moved to the right, make n negative. In this example, we moved the decimal 3 space to the left, so n is positive 3. Thus, in scientific notation, the value is.

$$\Longrightarrow 2.0000 \times 10^3$$

<u>Practice</u>: Write 0.000807 in scientific notation form.

1. We must move the decimal 4 places to the right to get a number between 1 and 10.

$$0.\underbrace{000807} \Rightarrow 8.07$$

2. Multiply by 10^{-4} (where the 4 is negative since we moved the decimal to the **right**) to get the value in scientific notation:

$$\Longrightarrow 8.07 \times 10^{-4}$$

<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} _____

<u>Section</u>: 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 not known for sure. A common application of significant figures happens nearly every day in lab with reading glassware.

- **a.** To the right is a graduated cylinder. Each mark on the graduated cylinder represents a 0.1 mL increase in volume.
- **b.** With the graduated cylinder to the right, it is known for sure that the true liquid measurement lies between 15.0 mL and 15.1 mL. We are confident in the value of the ones and tenths place, but the digit in the hundredths place is an approximation.
- **c.** An appropriate guess for liquid in this graduated cylinder would be 15.<u>02</u> mL OR 15.<u>03</u> mL. Either answer would be fine since the digit in the hundredths place is an approximation.

Figure 1 Work by "Horia Varlan"/Flickr https://creativecommons.org/licenses/by/2.0/legalcode

II. **Rules for Significant Figures**:

- **a.** All non-zero numbers are significant.
- **b.** Zeroes between two other significant digits are significant, including numbers that include a decimal place (e.g. 2.**<u>0</u>**34, 1**<u>0</u>**14**<u>0</u>**3).
- **c.** Zeroes following a non-zero number that are also to the left of a decimal are significant (83**<u>000</u>**.)
- **d.** In numbers containing a decimal, all zeros at the end of the number are significant (0.023**<u>0</u>**)
  - Combining rules 2 and 4 shows us that the 3 zeros at the end of 0.071**<u>000</u>** 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. (e.g. 32**<u>0</u>**)
- **f.** Zeroes that occur before any non-zero number are ***not*** significant (**<u>0</u>**.**<u>00</u>**147)

III. **Exceptions to traditional rules:**

- **a.** Counting numbers have an unlimited number of significant figures, meaning that there is no way to make them more precise than they already are. Examples of this would be counting 12 pencils or saying that a molecule is made of three atoms.
- **b.** Known conversions also fall in this category. For example, 1 foot = 12 inches is an example of an errorless conversion.

<u>Example</u>: Determine the number of Significant figures contained within each of the following numbers:

a) 0.0900 _____
b) 6.230 _____
c) 0.0076 _____
d) 4.00028 _____
e) 4.335×10^{-22} _____
f) 5600 _____
g) 3 Science Guyz Packets _____

IV. Rounding following Mathematical Operations:

- **a.** Following any mathematical operation, you must determine how many sig figs will be in your answer using sig fig rules.
- **b.** Once you have deteremined the correct number of sig figs, you must round your answer to the correct number of significant figures by looking to the first non-significant digit in your answer.
- **c.** If this number is 5 or greater, you will round your last significant digit up, otherwise you will round you answer down to the first significant digit to the left of the non significant digit.

V. Sig Figs in Multiplication and Division Problems:

- **a.** When multiplying/dividing two or more numbers, to express your answer in the correct number of sig figs, your answer must contain the same number of sig figs as the value in your equation with the fewest number of sig figs.

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

(3.2005470) x (30.9) = _____

VI. Sig Figs in Addition and Subtraction Problems:

- **a.** When adding or subtracting two or more numbers, to express your answer in the correct number of sig figs, 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. Sig Figs 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. Consider the examples below:

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>: Common SI Base Units Used in Chemistry

I. **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 provided below:

Property	**Unit Used**
Mass	Kilogram (kg)
Length	Meter (m)
Time	Second (s)
Temperature	Kelvin (K)
Amount of Substance	Mole (mol)

II. **The Metric System (Commit to Memory!):**

Prefix	**Abbreviation**	**Meaning**	**Example**
Giga-	G	10^9 (billion)	1 gigahertz = 1 x 10^9 Hz
Mega-	M	10^6 (million)	1 megaton = 1 x 10^6 tons
Kilo-	k	10^3 (thousand)	1 kilogram (kg) = 1 x 10^3 g
Dec-	d	10^{-1} (tenth)	1 decimeter = (dm) 1 x 10^{-1} m
Centi-	c	10^{-2} (hundredth)	1 centimeter = (cm) 1 x 10^{-2} m
Milli-	m	10^{-3} (thousandth)	1 millimeter = (mm) 1 x 10^{-3} m
Micro	µ	10^{-6} (millionth)	1 micrometer = (µm) 1 x 10^{-6} m
Pico-	p	10^{-12}	1 picometer = (pm) 1 x 10^{-12} m
Femto-	f	10^{-15}	1 femtometer = (fm) 1 x 10^{-15} m

III. **Other Useful Conversions:**

a. Several useful conversions are provided below. Depending on when you are taking this class you may be expected to memorize all conversions, or your instructor may give you certain ones. Ask your instructor for their expectations on the conversions!

1 kilometer = 0.62137 mile	1 mile = 5,280 feet
1 meter = 3.281 feet	1 inch = 2.54 cm
1 cm^3 = 1 mL	1 Angstrom (Å) = 1 x 10^{-10} m
1 pound = 453.59 g = 16 ounces	1 mile = 1760 yards

<u>Example</u>: The mass of a cat is most appropriately measured using the _____ metric unit.
a) Grams. b) Milligrams. c) Kilograms. d) Decigrams. e) Centigrams.

<u>Example</u>: The length of a cat is most appropriately measured using the _____ metric unit.
a) Meter. b) Millimeter. c) Kilometer. d) Decimeter. e) Centimeter.

Section: Dimensional Analysis

I. In Chemistry and other STEM courses, you will often have to convert from one unit to another. 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 must reflect this.
3. Choose a value as a starting point for your conversion. This is the first blank in "I am converting ____ to ____." It is usually helpful to start with a given value that only has one unit associated with it if possible.
4. Set up a series of conversion steps that allow you to cancel all unwanted units and leave you only with the final units.
   a. Each conversion step is a fraction. Examples include conversion factors (such as 1 hour / 60 seconds) or properties like density (grams / liter).
   b. Units can be canceled when they are found in the numerator and denominator of a dimensional analysis calculation, *regardless of if they are in consecutive steps.*
5. Once your units have been canceled to your final units, multiply across the numerators and denominators, divide the products, and you are finished.

Example: How many picograms are in a milligram?

Example: How many millimeters are in 4.20 centimeters?

Example: Convert 5.10 meters per second to micrometers per hour.

Example: Kailee is travelling 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 a speed of 65 miles/hour, how many gallons of gas does she use?

Example: 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? Note: 1 lb = 453.59 g.

Example: Ben is moving up in the world and has decided it is time to install a pool. Ben wants the pool to have the following dimensions: 58.7 ft long and 27.0 ft wide and 9.50 ft deep. Once the pool has been completed, how much water will Ben need to fill the pool (in cubic inches)?

Example: Convert the cubic inches obtained in the previous problem into cubic centimeters. Note: 1 inch = 2.54 cm.

<u>Section</u>: Temperature Scales

I. **Temperature** is a physical quantity representing the manifestation of thermal energy. It is how we associate something as being 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). There are no negative values associated with this scale and the ° symbol is not used in the scale.
  - **i.** The Kelvin scale is the only temperature scale that is an absolute temperature scale.
- **d.** The three formulas you need to commit to memory 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.** Pay special attention to the parenthesis placement is the Celsius versus Fahrenheit formulas!

K = °C + 273.15	**°C =** $\frac{5}{9}(°F - 32)$	**°F =** $\frac{9}{5}°C + 32$

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

a) 607 $°C$ to Kelvin: _____

b) 48.3 K to Celsius: _____

c) 340. $°F$ to Kelvin: _____

d) -39.1 $°C$ to Fahrenheit: _____

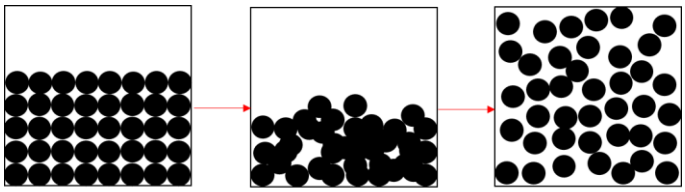
<u>Example</u>: Which of the statements provided below regarding the temperature scales is/are known to be **true**? *<u>Select all that apply</u>*.

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.

<u>Section</u>: States of Matter

I. Whether a substance is a solid, liquid or gas, this refers to the **state** of the substance. The state of a substance depends on how the individual particles which make up a substance interact with one another.

II. **The Kinetic-Molecular Theory of Matter** states that as matter gains energy, its temperature increases. Increased temperature reflects an increase in the average kinetic energy of the particles. As this kinetic energy increases, matter eventually transforms from the solid phase to liquid, and eventually gas. Therefore, the gas phase contains the most kinetic energy, and solid the least.



- **a. Solids:** In a solid, the attractive forces between the particles that compose the solid are stronger than the kinetic energy of the individual particles. As a result, particles within a solid are packed closely together and are arranged in a regular pattern. The particles within a solid do have some kinetic energy which causes the particles to vibrate back and forth about their average positions. However, particles within a solid seldom move past neighboring particles.
  - **i.** Solids retain fixed volume and shape and are not compressible.
- **b. Liquids:** The kinetic energy of the particles in a liquid are such that they have begun to overcome their attractive forces. In a liquid, particles are arranged randomly rather than in a regular 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. Instead, they assume the shape of the container in which they occupy and are not compressible.
- **c. Gases:** In a gas, the kinetic energy of the particles that compose the gas is such that the individual particles of the gas have completely overcome their attractive forces of the individual particles. In a gas, under ideal conditions, the particles composing the gas are far apart. Gas particles fly about colliding with one another and the walls of the container in which they occupy. The random motion of gas particles allows a gas to fill the volume of the container in which they reside.
  - **i.** Essentially, the volume of any container containing gas equals the volume of the gas within the container.
  - **ii.** Gases ARE compressible.

<u>Example</u>: Determine whether each statement provided below is **true** or **false**.

a) Only liquids can flow.

b) Gases and liquids are both compressible.

c) Liquids and gases have definite volumes, but not shapes.

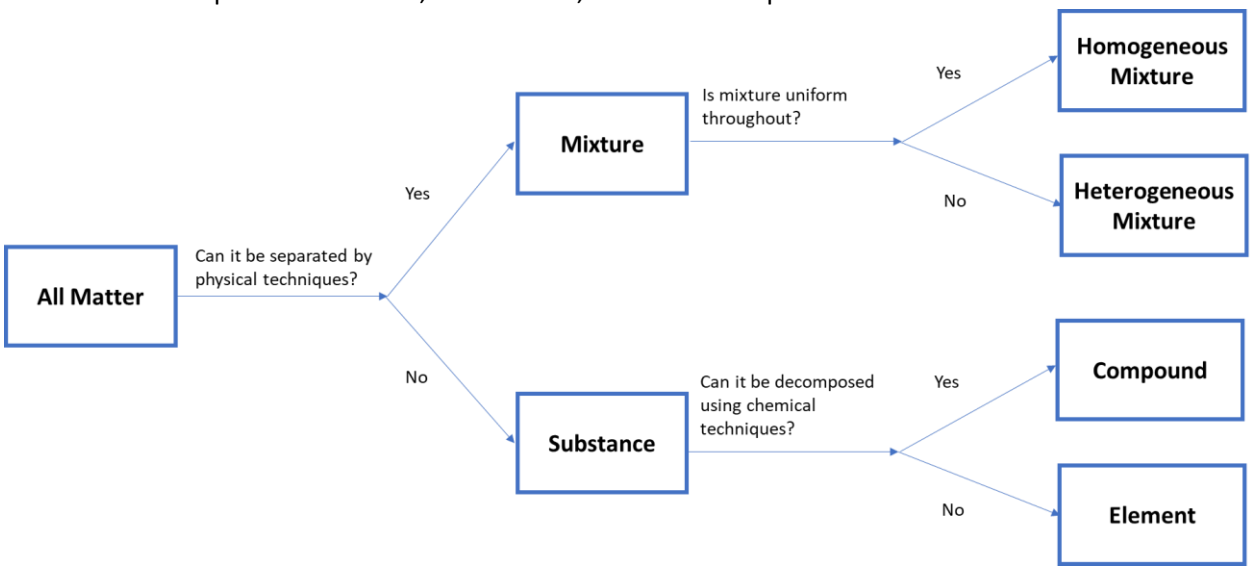
<u>Section</u>: Pure Substances and Mixtures

I. **Pure Substances** have a unique set of physical properties by which they can be recognized (ex: melting point, boiling point, and density). Additionally, pure substances *<u>cannot</u>* be separated into two or more different species by any physical technique at ordinary temperatures.

- **a. Elements**: Substances which cannot be subdivided by a chemical or physical process.
- **b. Compounds**: Groups of two or more *different* atoms joined by chemical bonds in fixed ratios. 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.
- **c. Molecules** are two or more atoms (can be same or different) chemically joined together.
  - i. **All compounds are molecules, but not all molecules are compounds!**

II. **Mixtures** consist of two or more pure substances that can be separated by physical techniques. Mixtures can be categorized as homogenous or heterogeneous. When a mixture is separated into its individual components (typically through a lab technique known as **filtration**), the components are considered purified.

- **a. Heterogeneous Mixture**: A mixture in which the components of the mixture are *unevenly distributed.* Examples include a bowl of cereal, milk, a salad, sand, and mud.
- **b. Homogeneous Mixture**: A mixture of two or more substances, in the same phase, in which the substances are *evenly* distributed. Homogenous mixtures are often called **solutions**. Examples include wine, black coffee, and common sports drinks.



<u>Example</u>: You begin to boil some water for pasta! You are not paying attention and accidentally evaporate all the water, leaving behind only salt crystals that were not previously visible in the water. This serves as an example of a _____.

<u>Example</u>: 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.

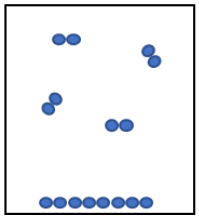
b) The products from the breakdown are for sure both elements.

<u>Section</u>: Pure substances, Mixtures and Model Interpretation

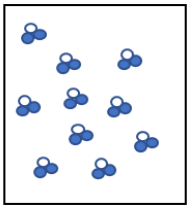
I. The definitions of types of pure substances and mixtures are straight forward at a glance; however, your exam is not going to be simple recall of definitions, but rather application of the different definitions.

II. Interpreting models and figures is an important skill you will develop throughout your time in Chemistry. An instructor will give you a figure or a model and ask you what it is an example of. We will walk through a few examples of this below.

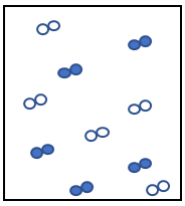
a. The model on the right has two spheres linked together that are the same color. Since all the spheres are the same color and nothing else is present, this is an example of a **pure substance**, more specifically a **molecule** made from the same elements. We know this since the spheres linked together are the same color. 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 not specific shape or volume (gas).



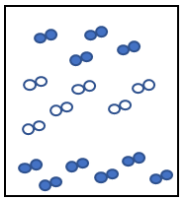
b. The model to the right has three spheres linked together with two of them being the same color and one being a different color. Since the spheres are not all the same color, this is an example of a **compound**. Understand that this is still an example of a **pure substance** since there is nothing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



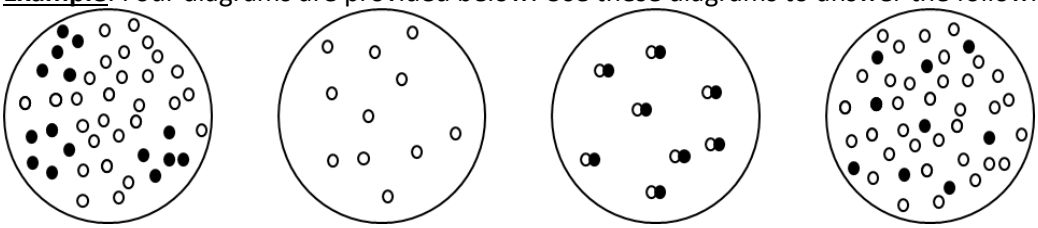
c. The model to the right has two different sets of spheres linked together that are different colors. Since we have these two separate sets of spheres that are different colors, we know that we are looking at a **mixture**. Defining this mixture further, we can see that there is an even distribution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



<u>Example</u>: Four diagrams are provided below. Use these diagrams to answer the following questions.



I. **II.** **III.** **IV.**

a) Which diagram(s) represent(s) an element? _____
b) Which diagram(s) represent(s) a compound? _____
c) Which diagram(s) represent(s) a pure substance? _____
d) Which diagram(s) represent(s) a mixture? _____

<u>Section</u>: Physical Changes and Chemical Changes

I. **Physical Properties**: Properties that can be observed and measured *without changing the composition of the substance.*

- **a.** Examples of physical properties include color, state of matter, melting point, boiling point, density, solubility, conductivity, malleability, ductility, and viscosity.

II. **Physical Changes**: Changes in the physical state (solid, liquid or gas) or size/shape of a substance.

III. **Chemical Properties**: Properties which determine whether and how readily a substance **reacts** (changes into a different substance).

IV. **Chemical Changes**: Changes which convert one or more substances into one or more different substances.

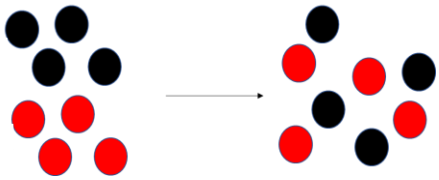
- **a.** A chemical change will **always** result in a change in composition.
- **b.** There are other things that occur that indicate a chemical change has occurred, including odor production, color change, light production, new product formation, sound production, change in energy, and fizzing or foaming (gas).

<u>Example</u>: Describe the following changes as physical or chemical.

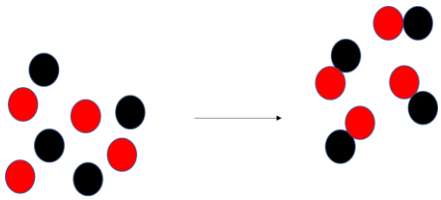
a) A piece of paper is burned _____
b) Paper is balled up _____
c) Removing salt from sea water _____
d) Removing carbon from carbon dioxide_____
e) Ice melting _____
f) Grilling a steak _____
g) Dicing an apple _____
h) Lighting a match _____

<u>Example</u>: Several models are provided below. Identify whether the model is depicting a chemical change or a physical change.

a) Physical or chemical?



b) Physical or chemical?



V. **More on Physical Properties of Substances:** Now that we have defined the difference between physical and chemical properties, there are some additional aspects of physical properties that we need to discuss.

- **a.** 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.
  - **i.** Examples include density, odor, color, luster, malleability, ductility, conductivity, temperature, hardness, melting point, and boiling point.
- **b.** **Extensive Properties**: An extensive property is a physical property of a system that changes with the size of the sample measured (it is additive).
  - **i.** Examples include mass, volume, length, or total charge.

VI. **The Difference Between "Mass" and "Weight": Mass** describes the quantity of matter contained within a thing/object. The standard unit for mass is the kilogram, however we will often discuss mass in terms of grams. **Weight** on the other hand has to do with the force of gravity on an object and is proportional to the mass of an object.

a. If you were to take your weight on a scale on the Earth versus the moon, your weight would be significantly less on the moon since the gravitational force is less. Your mass, however, would be identical on the Earth as on the moon.

Example: For each of the following, determine if the statement provided is **true** or **false**:

a) True/False: An intensive property can be determined by dividing two extensive properties.
b) True/False: The value of an intensive property can change.
c) True/False: Density, boiling point, and melting point are all considered intensive properties.

VII. **Density** is a common physical property that a pure substance or a homogenous mixture can demonstrate. Density is represented, mathematically, by dividing the mass of the pure substance by its volume.

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

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.

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 volume of the object used in the displacement. A walk through of a practice problem involving fluid displacement and density can be seen in the following practice problem:

Example: There is a small pellet of some metal that has a mass of 3.45 g. The pellet is placed into a beaker that contains 20.00 mL of water. The metal pellet is then submerged into the water, which causes the water level in the beaker to rise to 21.28 mL. What is the density of the metal pellet?

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 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. Rank the three globes in terms of increasing volume:

hydrogen 1 **H** 1.0079																		helium 2 **He** 4.0026
lithium 3 **Li** 6.941	beryllium 4 **Be** 9.0122												boron 5 **B** 10.811	carbon 6 **C** 12.011	nitrogen 7 **N** 14.007	oxygen 8 **O** 15.999	fluorine 9 **F** 18.998	neon 10 **Ne** 20.180
sodium 11 **Na** 22.990	magnesium 12 **Mg** 24.305												aluminium 13 **Al** 26.982	silicon 14 **Si** 28.086	phosphorus 15 **P** 30.974	sulfur 16 **S** 32.065	chlorine 17 **Cl** 35.453	argon 18 **Ar** 39.948
potassium 19 **K** 39.098	calcium 20 **Ca** 40.078		scandium 21 **Sc** 44.956	titanium 22 **Ti** 47.867	vanadium 23 **V** 50.942	chromium 24 **Cr** 51.996	manganese 25 **Mn** 54.938	iron 26 **Fe** 55.845	cobalt 27 **Co** 58.933	nickel 28 **Ni** 58.693	copper 29 **Cu** 63.546	zinc 30 **Zn** 65.39	gallium 31 **Ga** 69.723	germanium 32 **Ge** 72.61	arsenic 33 **As** 74.922	selenium 34 **Se** 78.96	bromine 35 **Br** 79.904	krypton 36 **Kr** 83.80
rubidium 37 **Rb** 85.468	strontium 38 **Sr** 87.62		yttrium 39 **Y** 88.906	zirconium 40 **Zr** 91.224	niobium 41 **Nb** 92.906	molybdenum 42 **Mo** 95.94	technetium 43 **Tc** [98]	ruthenium 44 **Ru** 101.07	rhodium 45 **Rh** 102.91	palladium 46 **Pd** 106.42	silver 47 **Ag** 107.87	cadmium 48 **Cd** 112.41	indium 49 **In** 114.82	tin 50 **Sn** 118.71	antimony 51 **Sb** 121.76	tellurium 52 **Te** 127.60	iodine 53 **I** 126.90	xenon 54 **Xe** 131.29
caesium 55 **Cs** 132.91	barium 56 **Ba** 137.33	57-70 *	lutetium 71 **Lu** 174.97	hafnium 72 **Hf** 178.49	tantalum 73 **Ta** 180.95	tungsten 74 **W** 183.84	rhenium 75 **Re** 186.21	osmium 76 **Os** 190.23	iridium 77 **Ir** 192.22	platinum 78 **Pt** 195.08	gold 79 **Au** 196.97	mercury 80 **Hg** 200.59	thallium 81 **Tl** 204.38	lead 82 **Pb** 207.2	bismuth 83 **Bi** 208.98	polonium 84 **Po** [209]	astatine 85 **At** [210]	radon 86 **Rn** [222]
francium 87 **Fr** [223]	radium 88 **Ra** [226]	89-102 * *	lawrencium 103 **Lr** [262]	rutherfordium 104 **Rf** [261]	dubnium 105 **Db** [262]	seaborgium 106 **Sg** [266]	bohrium 107 **Bh** [264]	hassium 108 **Hs** [269]	meitnerium 109 **Mt** [268]	ununnilium 110 **Uun** [271]	unununium 111 **Uuu** [272]	ununbium 112 **Uub** [277]		ununquadium 114 **Uuq** [289]				

*Lanthanide series	lanthanum 57 **La** 138.91	cerium 58 **Ce** 140.12	praseodymium 59 **Pr** 140.91	neodymium 60 **Nd** 144.24	promethium 61 **Pm** [145]	samarium 62 **Sm** 150.36	europium 63 **Eu** 151.96	gadolinium 64 **Gd** 157.25	terbium 65 **Tb** 158.93	dysprosium 66 **Dy** 162.50	holmium 67 **Ho** 164.93	erbium 68 **Er** 167.26	thulium 69 **Tm** 168.93	ytterbium 70 **Yb** 173.04
**Actinide series	actinium 89 **Ac** [227]	thorium 90 **Th** 232.04	protactinium 91 **Pa** 231.04	uranium 92 **U** 238.03	neptunium 93 **Np** [237]	plutonium 94 **Pu** [244]	americium 95 **Am** [243]	curium 96 **Cm** [247]	berkelium 97 **Bk** [247]	californium 98 **Cf** [251]	einsteinium 99 **Es** [252]	fermium 100 **Fm** [257]	mendelevium 101 **Md** [258]	nobelium 102 **No** [259]

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