CHEM 1211 Sample Packet: Chapter 1

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Introduction:

Many students find the first few chapters covered in a General Chemistry course to be incredibly boring and not relevant. Why do significant figures, accuracy, precision, scientific notation, and dimensional analysis matter anyway? Consider the scenario below:

Megan and Banks have both recently graduated college and were hired at a manufacturing plant. Megan payed close attention to the introductory material in General Chemistry while Banks slept through all of it. In this workshop, we will see multiple examples of Megan succeeding using the knowledge she gained from these early chapters while Banks fails and ultimately loses his job. Don't be like Banks.

Learning Objectives:

- 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.
- Convert numbers into and out of scientific notation.
- Perform temperature conversions between Celsius, Kelvin, and Fahrenheit.
- Describe the difference between precision and accuracy and provide examples of each.
- 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.
- Understand the scientific method and what is included in it.
- Describe the three states of matter, as well as the Kinetic-Molecular Theory of Matter.
- Describe the different classifications of mixtures and provide examples of each.
- Describe the difference between physical and chemical properties and provide examples of each.
- Understand density and how we can apply this physical quantity to various situations.

Significant Figures Introduction:

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 your General Chemistry lab with reading glassware.

- To the right a graduated cylinder. Each mark on the graduated cylinder represents a 0.1 mL increase in volume.
- 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.
- An appropriate guess for the amount of liquid in this graduated cylinder would be 15.02 mL OR 15.03 mL. Either answer would be fine since the digit in the thousandths place is an approximation.

Rules for Significant Figures:

1. All non-zero numbers are significant.
2. Zeroes between two other significant digits are significant, including numbers that include a decimal place (e.g. 2.**0**34, 1**0**14**0**3).
3. Zeroes following a non-zero number that are also to the left of a decimal are significant (83**000**.)
4. In numbers containing a decimal, all zeros at the end of the number are significant (0.023**0**)
    - Combining rules 2 and 4 shows us that the 3 zeros at the end of 0.071**000** are significant
5. 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**0**)
6. Zeroes that occur before any non-zero number are ***not*** significant (**0**.**00**147)

Example: Determine the number of Significant figures contained within each of the following numbers:

1. 0.0400 _____
2. 9.430 _____
3. 0.0067 _____
4. 2.00011 _____
5. 4.335×10^{-22} _____
6. 9600 _____

Rounding following Mathematical Operations:

1. Following any mathematical operation, you must determine how many sig figs will be in your answer using sig fig rules.
2. 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.
3. 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.

Sig Figs in Multiplication and Division Problems:

When multiplying/dividing two or more numbers, in order 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.

(3.2005470) x (30.9) = _____

Sig Figs in Addition and Subtraction Problems:

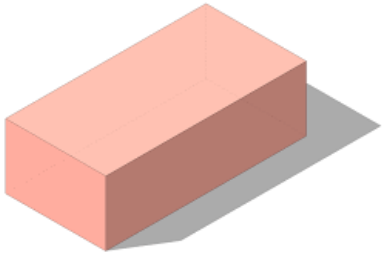
When adding or subtracting two or more numbers, in order to express your answer in the correct number of sig figs, your answer must contain the same number of decimal places as the value in the equation with the fewest decimal places.

(321.1896) + (1.98665) + (0.1) = _____

Example: Record the answer to the problem below to the appropriate number of significant figures:

$$\frac{(0.450)(0.01219)(1.55+25)}{(320/871.0)} = ?$$

Example: A company places a brick order at the manufacturing plant. The bricks needed are determined to be 5.3 m wide and 16.10 m long. What are the perimeter and area of the bricks? Note: Pay attention to significant figures!



a) perimeter: 42.8 m; area: 85.33 m^2.
b) perimeter: 42.8 m; area: 85 m^2.
c) perimeter: 36.7 m; area: 74 m^2.
d) perimeter: 43 m; area: 85.3 m^2.
e) perimeter: 43 m; area: 84.8 m^2.

Why It Matters: Let us take a closer look at the previous example problem. Megan determined the dimensions of the bricks in the problem. Banks had one slight difference in his approach; he determined that the bricks needed to be 16.109 m long instead of 16.10 m. Banks ended up discarding bricks that did not match his specifications exactly to the thousandths place.

While the company that placed the order was satisfied with the bricks that both Megan and Banks produced, the manufacturing company was not content with Banks' performance, as he ended up wasting a lot of product by not considering significant figures in his calculations. This product waste ended up costing the company thousands of dollars. Don't be like Banks.

Errorless Numbers:

Some numbers are exact, meaning they have no error associated with them. Many errorless numbers are those that are defined to have an exact value. Examples: Counting gives errorless numbers, and temperature scales are errorless.

Temperature Conversion Formulas:

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

Example: Convert each of the following temperatures into the desired units:

1. 607 $°C$ to Kelvin: _____

2. 48.3 K to Celsius: _____

3. 340 $°F$ to Kelvin: _____

4. -39.0 $°C$ to Fahrenheit: _____

Scientific Notation:

A number is written in <u>scientific notation</u> when in the form $a \times 10^n$ where $1 \leq |a| < 10$ and n is an integer. Scientific notation is used to make a very large or very small numbers 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.

$$0.\underbrace{0008}07 \Rightarrow 8.07$$

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.

1. 0.0000783 _____
2. 9.3×10^6 _____

<u>Why It Matters:</u> Banks and Megan are sending off very precise specifications for product they need made. Megan reports the length of the product to be 1.01 x 10^{-10} mm and Banks reports the length as 0.000000000101 mm. While both represent the same number, the person inputting the specifications forgot a 0 with the number Banks provided. This ended up in the part Banks needed being too long, wasting time and money for the manufacturing plant. Don't be like Banks.

Accuracy vs Precision:

Accuracy: Accuracy is a measure of how close a value obtained is to the true value.

Precision: A precise measurement is a measurement that yields similar results when repeated in the same manner. Precision values do not have to correspond with the true value.

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

Example: Employees of the manufacturing plant are asked to make measurements regarding the amount of liquid present in a container. The amount of liquid present is known by the company to be 31.5 mL. Their collected measurement data is shown in a table below.

Banks	**Ansley**	**Megan**	**Austin**	**Olivia**
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

1. Which of the employee's measurements is the most accurate? _____

2. Which of the employee's measurements is the most precise? _____

Common SI Base Units Used in Chemistry:

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

The Metric System:

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

Other Useful Conversions:

1 kilometer = 0.62137 mile

1 mile = 5,280 feet

1 meter = 3.281 feet

1 inch = 2.54 cm

1 foot = 12 inches

1 cm^3 = 1 mL

1 pound = 453.59 g = 16 ounces

1 Angstrom (Å) = 1 x 10^{-10} m

Example: Determine how many nanograms (ng) are present in a picogram (pg).

Dimensional Analysis:

In chemistry studies, you will often have to convert from one unit to another. This can be 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.
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 ___."
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. They may be things like conversion factors (1 hour / 60 minutes) or properties like density (grams / liter) or molar mass (grams / mole)
    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 canceled to your final units, multiply across the numerators and denominators, divide the products, and you are finished.

Example: Megan drives for 25 minutes at 55 miles per hour to arrive at her job on time. How many meters has she traveled by the time she arrives at her job?

Example: While working on a project, Megan discovers a 60,000 mL sample of some arbitrary compound that is used to make construction material. The compound is determined to have a density of 14.67 kg per liter. How many 0.2 lb parts of the construction material can a company make with this compound?

Example: The manufacturing company produces a toy car that can reach speeds of 35 miles per hour. Determine how many centimeters per second that the toy car can go.

Example: Megan is moving up in the world and has decided it's time to install a pool. Megan wants the pool to have the following dimensions: 58.7 ft long and 27.0 ft wide and 9.50 ft deep.

1. Once the pool has been completed, how much water will Megan need to fill the pool (in cubic inches)?

2. Convert the cubic inches obtained in the previous example into cubic centimeters.

Chemistry and its Methods:

Hypothesis: A tentative explanation or prediction based on experimental observations.

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.

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.

The Scientific Method: The path that leads from experiments, hypotheses and observations into theories or laws.

Qualitative Information: Consists of non-numerical data such as the color of a substance or its physical appearance.

Quantitative Information: Consists of numerical data such as the mass of a substance or the temperature at which a substance melts or boils.

States of Matter:

Whether a substance is a solid, liquid or gas 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.

1. **Solids:** In a solid, the attractive forces between the particles that compose the solid are stronger than the kinetic energy of the individual particles.
    a. As a result, particles within a solid are packed closely together and are arranged in a regular pattern.
    b. The particles within a solid do have some kinetic energy which causes the particles to vibrate back and forth about their average positions.
    c. However, particles within a solid seldom move past neighboring particles.
    d. Solids retain fixed volume and shape.

2. **Liquids:** The kinetic energy of the particles in a liquid are such that they have begun to overcome their attractive forces.
    a. In a liquid, particles are arranged randomly rather than in a regular pattern.
    b. Liquids are fluid because the particles of liquids are not confined to specific areas and they can move past each other.
    c. Liquids have no definite shape. Instead, they assume the shape of the container in which they occupy.

3. **<u>Gases</u>:** 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.
    a. In a gas, under ideal conditions, the particles composing the gas are far apart.
    b. Gas particles fly about colliding with one another and the walls of the container in which they occupy.
    c. The random motion of gas particles allows a gas to fill the volume of the container in which they reside. Essentially, the volume of any container containing a gas equals the volume of the gas within the container.

1. What state is A in?

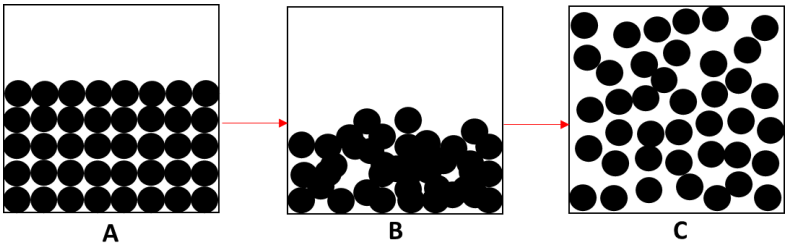
   _____

2. What state is B in?

   _____

3. What state is C in?

   _____



The Kinetic-Molecular Theory of Matter:

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.

Macroscopic, Microscopic, and Submicroscopic/Particulate:

Macroscopic: Refers to properties of matter which are large enough to be seen with the naked eye (phase of matter, and expanding/contracting, for example)

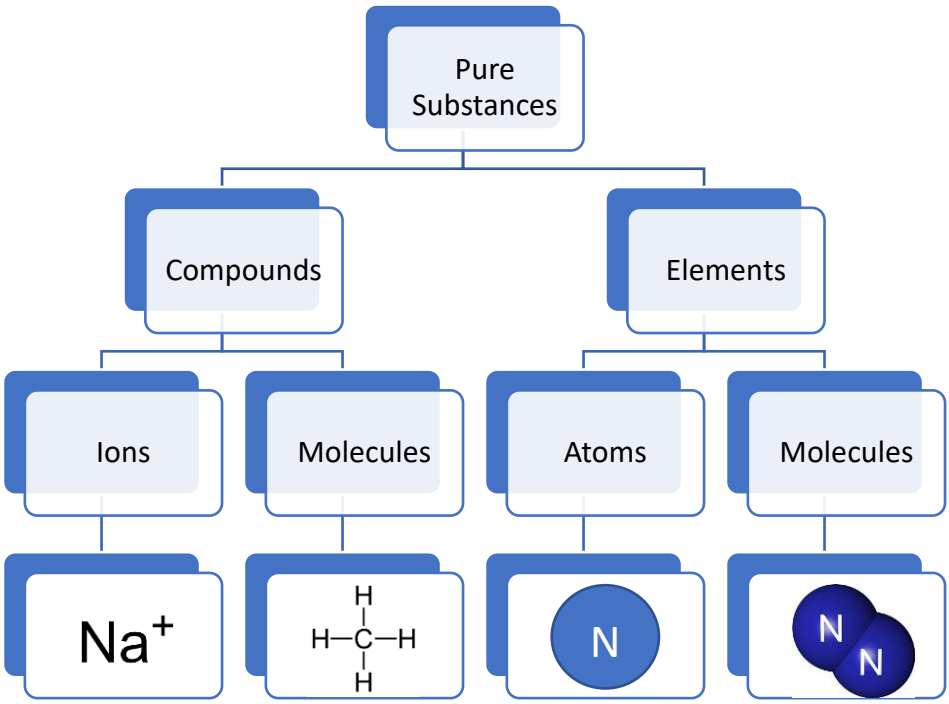
Microscopic: Refers to properties of matter which are very small and require the aid of a machine such as a microscope to view them (arrangement of electrons, for example).

Submicroscopic/Particulate: Refers to matter which is so small that it can't be seen, directly, even with the most powerful optical microscope.

Pure Substances:

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 cannot be separated into two or more different species by any physical technique at ordinary temperatures.

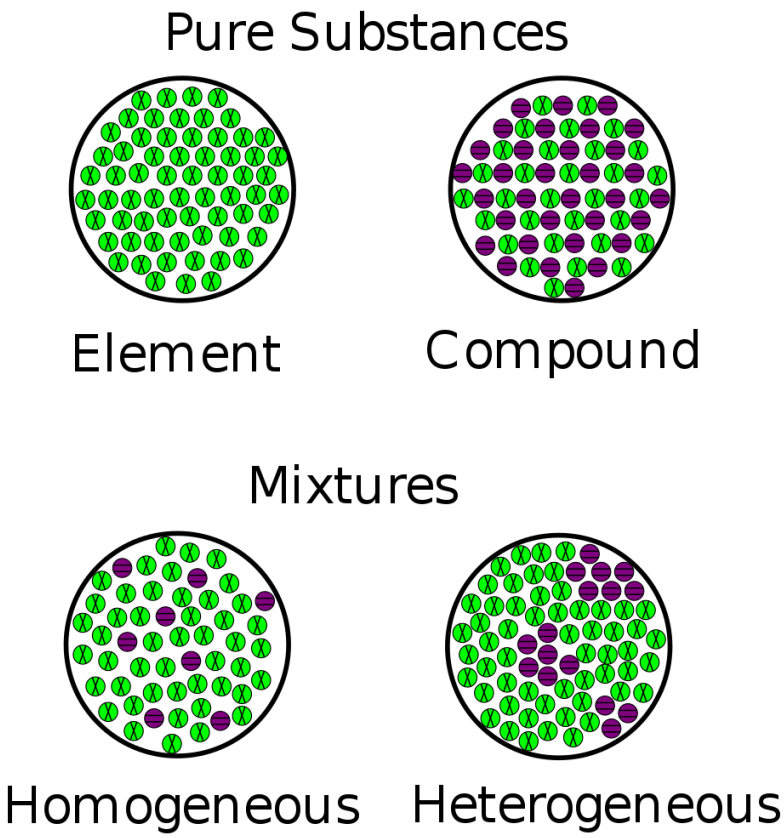
1. **Elements**: Substances which cannot be subdivided by a chemical or physical process.
2. **Molecules**: Formed when two or more atoms (the atoms can be the same or different) join to form chemical bonds in fixed ratios.
3. **Compounds**: Groups of two or more ***different*** atoms joined by chemical 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.
4. **Ions**: Electrically charged atoms or groups of atoms (Na^+ or Cl^-)



Mixtures and Solutions:

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, the components are considered ***purified***.

1. **Heterogeneous Mixture**: A mixture in which the components of the mixture are unevenly distributed. Examples: Sand or milk.
2. **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***. Example: Hydrochloric acid solution.



Example: Determine which type of mixture each one of the following items would be classified as:

1. Vinegar: _____
2. Concrete: _____
3. A dilute solution of copper (II) sulfate in H_2O: _____

Physical Changes and Chemical Changes:

Physical Properties: Properties that can be observed and measured without changing the composition of the substance. Examples: Color, State, Melting Point, Boiling Point, Density, Solubility, Conductivity, Malleability, Ductility, and Viscosity.

Density: 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. Density was used, before modern analytical methods, to determine the identity of an unknown substance. In liquid and gas mixtures, substances that are less dense will "float" on top of substances that are denser.

$$\boldsymbol{Density = \frac{mass}{volume}}$$

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:

Practice: 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?

1. First subtract the change in water level to determine the volume of the metal pellet:

$$21.28 \text{ mL} - 20.00 \text{ mL} = 1.28 \text{ mL}$$

2. We are given the mass of the metal pellet, which is 3.45 g. We also know that the formula for density is mass/volume. We can now determine the density of the metal pellet:

$$3.45\text{g} / 1.28 \text{ mL} = \underline{2.70 \text{ g/mL}}$$

Example: A commonly used metal in the 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:

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

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

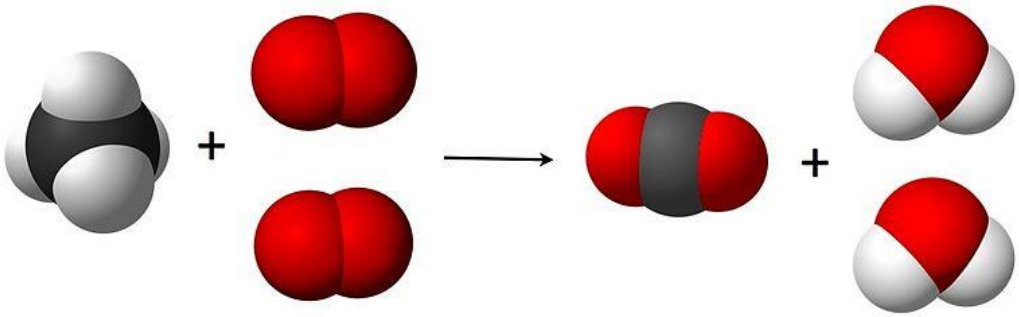
Chemical Changes: Changes which convert one or more substance into one or more different substances. Factors that indicate a chemical change has occurred include odor production, color change, gas production, new product formation, and light and sound production

Example: Describe the following changes as physical or chemical.

1. Burning a paper chem exam _____
2. Cutting up a paper chem exam _____
3. Baking some brownies after a chem exam _____
4. Rust formation you notice after on a car after a chem exam _____

More on Physical Properties of Substances:

Law of Conservation of Matter: For any system closed to all transfers of matter, the mass of the system must remain constant over time; the mass of a system cannot change if mass is not added or removed from the system.



$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

Intensive Properties: 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: Density, Odor, Color, Luster, Malleability, Ductility, Conductivity, Hardness, Melting Point, and Boiling Point.

Extensive Properties: An extensive property is a physical property of a system that changes with the size of the sample measured (it is additive). Examples: Mass, Volume, Length or, Total Charge. (***Think: As you extend [increase] the quantity of the substance, the value of the of the substance's property changes)***

Example: In your chemistry lab, you are exploring several properties of some unknown substance. These properties include:

I. Density
II. Boiling point
III. Temperature
IV. Mass
V. Melting Point

In the properties you investigated, which are considered intensive properties?

a) I, II, and V.
b) I, II, III, IV, and V.
c) I, II, III, and V.
d) III and IV.
e) II and V.

Example: A reaction vessel in the manufacturing plant contains 6.0 g of hydrogen gas and 23.0 g of oxygen gas. How many grams of water will be produced in the complete reaction of these two gases?

a) 6.00 g.
b) 23.0 g.
c) 17.0 g
d) 29.0 g.
e) Additional information is needed.

Why It Matters: The above example is the COMPLETE reaction of hydrogen and oxygen, meaning that ALL reactant is made into product. These are ideal results; often in the real world some reactant is lost when being converted into product. We often look for the most efficient way to perform a reaction. Megan, due to her accuracy in measurements, consistently has a higher efficiency in her reactions than Banks, who is much less careful with his measurements. Don't be like Banks.

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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]

Science Guyz, LLC: 250 W Broad St STE 102 Athens, GA 30601 Phone: 706-723-8436 Email: tutor@scienceguyz.com
Website: www.scienceguyz.com