CHEM 1211

Promo Packet (Introduction to Chemistry Material)

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Spring 2020

For hours of operation, important dates and other info, check our regularly updated website:

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Semester plan students will have access to the full version of this packet, which includes practice questions derived from old progress checks.

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<u>Significant Figures</u>

1. All non-zero numbers are significant.
2. Zeroes between two other significant digits are significant, including in numbers that include a decimal place (e.g. 2.**<u>0</u>**34, 1**<u>0</u>**14**<u>0</u>**3).
3. Zeroes following a non-zero number that are also to the left of a decimal are significant (83**<u>000</u>**.)
4. 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
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**<u>0</u>**)
6. Zeroes that occur before any non-zero number are ***not*** significant (**<u>0</u>**.**<u>00</u>**147)

How many Significant figures are contained within each of the following?

1. 0.0300 ____________________
2. 8.450 ____________________
3. 0.0023 ____________________
4. 1.00091 ____________________
5. 3.446 x10^{-22} ____________________
6. 5700 ____________________

<u>Rounding following Mathematical Operations:</u>

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.

<u>Sig Figs in Multiplication and Division Problems:</u>

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.

(2.3004560) x (20.3) = _____________

Addition and Subtraction:

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.

(123.2987) + (2.31556) + (0.1) = ____________

Temperature Conversions

K = °C + 273.15

°C = $\frac{5}{9}(F - 32)$

F = $\frac{9}{5}C + 32$

Errorless Numbers

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. Ex: Counting gives errorless numbers, and temperature scales are errorless.

Practice: Convert 88.5°F to Kelvin

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 = .62137 mile

1 mile = 5,280 feet

1 meter = 3.281 feet

1 inch = 2.54 cm

1 cm^3 = 1 mL

1 pound = 453. 59 g = 16 ounces

Dimensional Analysis

In chemistry studies, you'll 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 unites
    a. Each conversion step is a fraction. They may things like conversion factors (1 hour / 60 seconds) 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, <u>regardless of if they are in consecutive steps.</u>
5. Once your units have canceled to your final units, multiply across the numerators and denominators, divide the products, and you are finished.

Practice: You drive for 30 minutes at 60 miles per hour. How many meters have you traveled?
1 km = 0.621 miles.

Practice: You unearth a 50,000 ml sample of precious metal that has a density of 13.45 kg per liter. How many 0.1 lb earrings can you make with this sample? 1 lb = 453.592 g.

Practice: Convert 35 miles per hour to centimeters per second.

Practice: A swimming pool is 54.7 ft long and 23.0 ft wide and is 8.18 ft deep. How much water is needed to fill the pool, in cubic inches? 1 foot = 12 inches.

Now, convert the previous units of the volume into cubic centimeters:

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Grams/Atoms/Moles Conversions

Avogadro's Number: $6.022 \times 10^{23}\ mol^{-1}$

Avogadro's Number: This is the number of particles (atoms, molecules, etc.…) in one mole of a substance.

Mole: A mole is the amount of a substance (this can be anything) that is equal to the number of particles in 12 g of carbon-12. A mole is equal to the mass in grams of a compound that is equal to the molar mass of the compound.

Dimensional analysis is often used to convert between grams, moles, and atoms.

$$\text{Moles} = \frac{Mass\ (g)}{Molar\ mass\ (\frac{g}{mol})}$$

Molar Mass: Molar mass is the mass, in g/mol, of one mole of a compound. If you are working with a pure element, then the molar mass is found on the periodic table. If you are trying to find the molar mass of a compound, then you add up all the individual elements in the compound.

Practice: Find the molar mass of the following compounds

Ag:

SO_4:

CH_4:

Practice: How many atoms are in 16 g of carbon?

Practice: How many grams of hydrogen are in 1.2 moles of C_2H_6?

<u>Chemistry and its Methods</u>

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

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.

<u>The *Kinetic-Molecular Theory of Matter:*</u> 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 (and vice versa).

<u>States of Matter:</u>

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:** In a liquid, 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. **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.
    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?

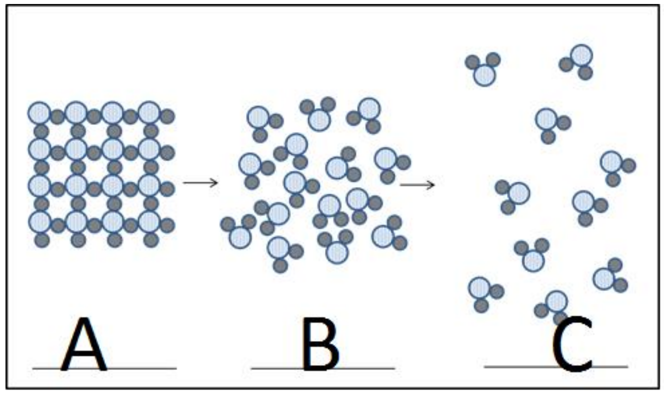
   ____

2. What state is B in?

   ____

3. What state is C in?

   ____



Macroscopic and Submicroscopic/Particulate

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

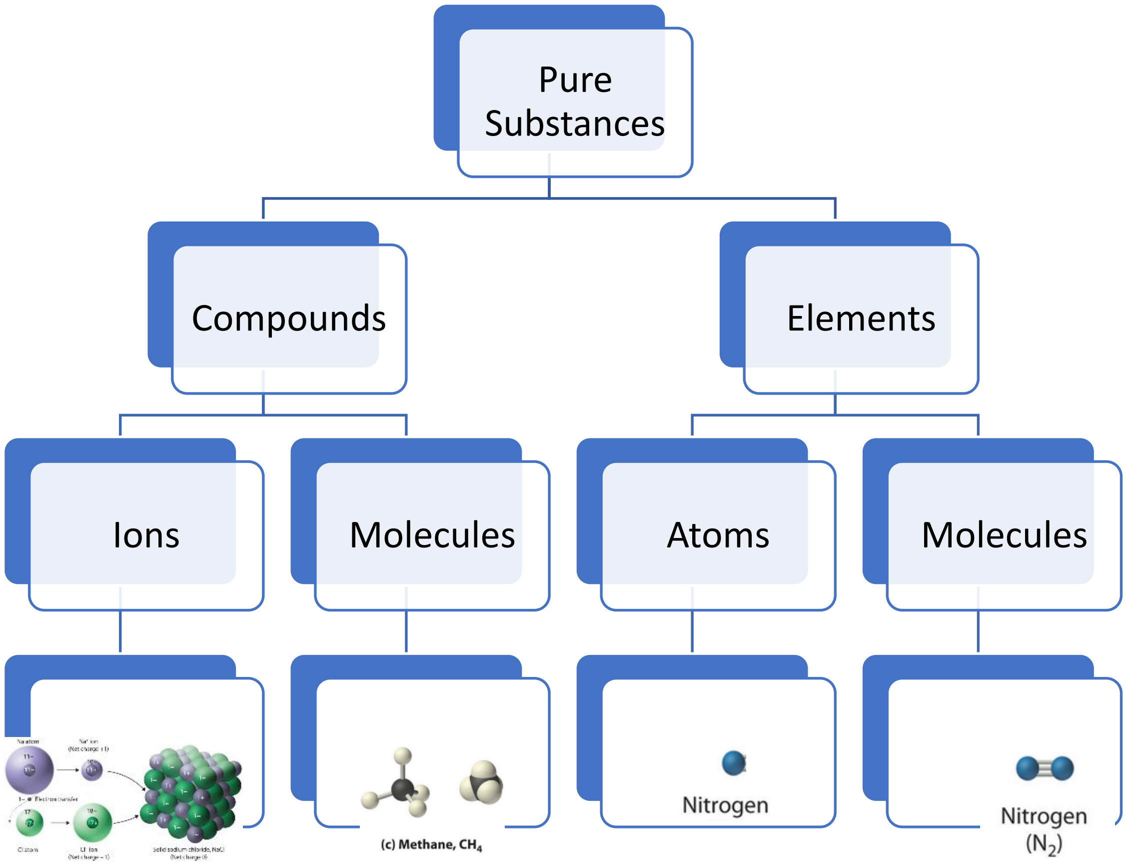
Microscopic: Refers to properties of matter which is very small and requires the aid of a machine such as a microscope to view it (arrangement of electrons, for example).

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^-)



<u>**Mixtures and Solutions**</u>

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 (Ex: Sand or milk).
2. **Homogeneous Mixture**: a mixture of two or more substances, in the same phase, which the substances are evenly distributed. Homogenous mixtures are often called ***solutions***. (ex. Hydrochloric acid solution)

<u>**Chemical Formulas**</u>

Chemical Formula: Describe the ratio of each element present in a compound relative to the other elements present in the compound. Each element is represented by a symbol and any subscript following the symbol of an element describes the quantity of each element that is present in the compound.

Determine the percentages of each element in the compounds shown:

$Pb_3(PO_4)_2$					
Symbol	Atom Name	Quantity	Mass	Total Mass	Percentage
Pb					
P					
O					
Total Mass					

Physical Changes and Chemical Changes

Physical Properties: Properties that can be observed and measured without changing the composition of the substance. Ex: 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 in the mixture will "float" on top of substances that are denser.

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

Practice: A cube is dropped into a 30-liter volume of oil in a barrel. Which of the following would be sufficient to calculate the density of the cube?

a) The volume of oil after adding the cube, and the mass of the oil that the cube displaced
b) The total mass of the cube and the total mass of the oil
c) The volume of oil after adding the cube, and the mass of the barrel and its contents before and after adding the cube
d) The total mass of the barrel and oil, and the final volume of oil after adding the cube

Practice: A bucket sits inside of a larger bucket. The smaller bucket is filled perfectly to capacity. A bowling ball with a volume of 2.3 liters is softly dropped into the smaller bucket, and some of the water flows out into the larger bucket. The mass of the water in the larger bucket is 2,295.4 grams. What is the density of the water?

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.

Practice: Describe the following changes as physical or chemical

a. A piece of paper is burned ______________________________
b. Paper is balled up ______________________________
c. Baking a cake ______________________________
d. Photosynthesis ______________________________
e. Rust forming ______________________________

More on Physical Properties of Substances

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.

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 c***

Practice: Which one of the following lists contains ONLY extensive properties?

a. Mass and Volume
b. Boiling Point and Color
c. Volume, Melting Point, and Color
d. Density and Mass

Practice: At 1.00 °C, a beaker contains 0.450 L of water in its liquid state. What is the volume of the water after it freezes (at 0.00 °C)? The densities of liquid water and ice at 0.00 °C are 1.00 g/mL and 0.917 g/mL, respectively.

