2016/2017
JSHS Senior High School
Honors Chemistry Summer Assignment

Student Name:________________

Teacher Name: Mrs. Sousa & Mr. Pickin

Due Date: September 9, 2016 (Friday)

2016 Johnston Senior High School

Honors Chemistry Summer Assignment

Textbook Map Site: Chemsitry-Zumdahl and Zumdahl

http://chemwiki.ucdavis.edu/Textbook_Maps/General_Chemistry_Textbook_Maps/Map%3A_Chemistry_(Zumdahl_and_Zumdahl)

Teacher: Mrs. Sousa Email: jsousa@johnstonschools.org Teacher website: www.sousascience.com

Mr. Pickin Email: spickin@johnstonschools.org

Rationale: Honors Chemistry is a course designed to introduce students to the structure of substances as well as the changes that they undergo at a faster and more rigorous format. Prior to entering the classroom, this course requires basic understanding of several science concepts taught in already taken science courses. To best prepare you for the rigidity of this course, you are required to read the attached sections of text, and complete the accompanying questions. At the beginning of the school year, you will be tested on said topics after review with the teacher. A major component of the Honors Chemistry class for the 2016-2017 school year is designed around the concept of computational thinking. It is important students are able to thin critically, collect data, interpret that data, and communicate that interpretation to others.

Due Date: The readings and questions are due to your Honors Chemistry teacher 9/916, regardless of when you are scheduled for the course. If you have added this course after the summer add/drop period, you are still required to complete the packet for a grade. You will be tested on the material at the start of the school year. This assignment will be collected, graded, and checked for completeness. During class we will discuss and review any questions or concerns regarding this assignment. All of these concepts will be reviewed prior to the quiz/test.

Assignments:

1. Each section below is a topic covered in chemistry. However, you may have covered them in other honors and college preparatory classes. You are to use this packet to practice and familiarize yourself with these concepts. It will be anticipated you have a <u>basic</u> understanding of these concepts entering into the first weeks of the school year in September.
2. Attached you will find the textbook assignments with supporting materials. Thoroughly answer all questions number throughout the sections as well as the section reviews.
3. Also included are notes on recurring and basic topics. You are responsible for knowing the information presented in the packet. Please know that this Honors Chemistry class is designed for students who are willing to challenge themselves. The pace of the class is faster than a college preparatory class and this class is not designed for all students. Please pay careful attention to the course description regarding math placement and other relative science courses.

In the Science Department, we take Academic Integrity seriously, and will expect the same from you. Academic misconduct is any attempt by a student to gain an academic advantage, or to help others do so, through dishonest actions. All work is expected to be your own. You are responsible for understanding these concepts! Any suspicion of academic misconduct will result in a zero on the entire assignment.

Scientific Method: (Please note that this entire S.M. section has been taken from the Science Buddies website)
http://www.sciencebuddies.org/science-fair-projects/project_scientific_method.shtml#overviewofthescientificmethod

Construct a Hypothesis: A hypothesis is an educated guess about how things work:
"If _____*[I do this]* _____, then _____*[this]*_____ will happen."

You must state your hypothesis in a way that you can easily measure, and of course, your hypothesis should be constructed in a way to help you answer your original question. This is a tentative explanation or predication of what is expected!

Variables: The **independent variable** is the one that is changed by the scientist. A good experiment has only one independent variable. As the scientist changes the independent variable, he or she **observes** what happens.

The scientist focuses his or her observations on the **dependent variable** to see how it responds to the change made to the independent variable. The new value of the dependent variable is caused by and depends on the value of the independent variable.

For example, if you open a faucet (the independent variable), the quantity of water flowing (dependent variable) changes in response--you observe that the water flow increases. The number of dependent variables in an experiment varies, but there is often more than one.

Experiments also have **controlled variables**. Controlled variables are quantities that a scientist wants to remain constant, and he must observe them as carefully as the dependent variables. For example, if we want to measure how much water flow increases when we open a faucet, it is important to make sure that the water pressure (the controlled variable) is held constant. That's because both the water pressure and the opening of a faucet have an impact on how much water flows. If we change both of them at the same time, we can't be sure how much of the change in water flow is because of the faucet opening and how much because of the water pressure. In other words, it would not be a fair test. Most experiments have more than one controlled variable. Some people refer to controlled variables as "constant variables."

SI Units of Measurement

The SI system is an internationally accepted system of measurement. SI stands for the French "System International". It encompasses the metric system as well as specific base units. It is necessary for a chemistry student to know the base units of the SI system.

Whenever these SI units are used, metric prefixes are also used. These should be memorized as well. These prefixes allow for using much larger and much smaller multiples of the base units. Ex: a mega-meter would be 1 x 10^6 meters. A kilogram = 1 x 10^3 grams. A centimeter = 1 x 10^{-2} meters. A micromole = 1 x 10^{-6} moles.

It is possible to convert between units by using **dimensional analysis/conversions/factor-label method**.

Example: How many kilograms are present in 5000 grams?

$$5000 \text{ grams} \cdot \frac{1 \text{ kilogram (kg)}}{1\text{x } 10^3\text{g (1000 g)}} = 5 \text{ kg}$$

Example: How many nanograms are present in 20g?

$$20 \text{ g} \cdot \frac{1 \text{ x } 10^9 \text{ ng}}{1\text{g}} = 20{,}000{,}000{,}000 \text{ ng } (2 \text{ x}10^{10}\text{ng})$$

Scientific Notation

In chemistry we use very large and very small numbers to represent amounts and masses. In order to effectively communicate these numbers without using lots of zeros, we employ scientific notation.

Each value is written as the product of two numbers, one of which is a power of 10. The first number is always a single digit number, which may or may not have any number of decimal places.

Numbers greater than 1 have a positive exponent. Numbers that are less than 1 have a negative exponent.

For example:

- Converting to scientific notation: 36,000 would be written as 3.6×10^{4}
- 23.67 would be written as $2.367 \times 10^{1\ 3}$
- 5069032 would be written as 3.5068032×10^{7}
- 0.003 would be written as $3 \times 10^{-3\ 0}$
- .23999 would be written as 2.3999×10^{-1}
- 0.00000000000000000000000007 would be written as 7×10^{-24}
- Converting from scientific notation to equivalent decimal numbers:
- 3.55×10^{3} would be written as 3550 .
- 6.887×10^{-6} would be written as 0.000006887

Accuracy and Precision

It is impossible for measurements to be perfect. Many sources of error can exist when humans take measurements. All measurements have some amount of uncertainty.

Good precision concerning measurements entails making a series of measurements. Precision is increased with additional measurements. It is also increased with an increased number of decimal places in the measurement. If I were to use two scales to weigh an object, and one scale weighed to the 0.00 place, and the second to the 0.0000 place, I would get more precise measurements with the scale that measured to the 0.0000 place (providing the level of accuracy in both was equal- see below).

If we were to take many measurements of the same object, and the values obtained were very similar, We could state that our measurements were PRECISE. That is, they agreed closely with one another.

If I were weighing a marshmallow on a scale that measured to the hundredth's place and obtained three values of 1.00 g, 1.01 g, and 0.99 g, I could state that I had good precision. My three measurements were close to one another. These measurements were PRECISE. Precision can only be determined if a SERIES of measurements are taken. Good precision would require that the measurements be close in value to one another.

Does that mean that they were also ACCURATE?

Not necessarily. ACCURACY is defined as how close the measurements are to the true value. If the true mass of our marshmallow was indeed 1.00 g, then our measurements were accurate and precise. If, however, the true mass of our marshmallow was 1.50 g, then our measurements were precise, that is, they agreed with one another, but not accurate, as they were far from the true value, which was 1.50 g.

In order to determine accuracy, the true measurement must be known.

Significant Figures

In chemistry the concept of significant figures is always used when discussing or doing calculations with measured amounts. Significant figures apply **ONLY** to measured quantities. The principle behind significant figures is this: if I am measuring any quantity, my measurement will consist of all the digits I am sure of, and the last digit, which is a reasonable estimate.

Example: I use the grocery store scale to weigh my fruit. The scale has divisions of pounds, with 10 divisions between each pound to indicate tenths of pounds. So if I put my apples on the scale and it reads 3.5 pounds and the arrow is not exactly on the 5, but it appears to be exactly between 3.5 and 3.6 pounds, I can estimate that I have 3.55 pounds of apples. I am sure I have 3.5 pounds, and I think I have 3.55 pounds, but the last 5 is really an estimate. I am sure of the 3.5, and reasonably sure of the 3.55, but that last 5 is not certain. Significant figures include all digits we are sure of PLUS the last one, which is a very good estimate. (Estimated digit)

Rules have been established and are used to determine whether digits are significant or not.

Significant Figure (digit) Rules:

1. All non-zero digits and zeros BETWEEN non-zero digits **are significant**. (we call these captive zero's)

Example: 245Liters has 3 sig figs, 2405 Liters has 4 sig figs.

2. Zeros at the END of a number and to the right of a decimal point **are always significant**.

Example: 23.0g has 3 sig figs., 23.00 g has 4 sig figs., 23.000g has 5 sig figs.

3. Zeros appearing to the left of all non-zero digits (leading zeros) **are NOT significant** if they just act as placeholders.

Example: 0.0071mL has 2 sig figs (the 7 and 1) the three zeros to the left of the 71 **are NOT significant** because they are merely placeholders.

4. Zeros appearing to the right of non-zero digits where NO DECIMAL POINT is present **are NOT significant**. These are merely placeholders.

Example 7100 mg has 2 sig figs, the 7 and the 1. The two zeros are merely placeholders.

There is no decimal point. (If the number had been written as 7100. mg there would be 4 sig figs because the decimal point appearing after the 7100. mg makes the two zeros significant.]

5. In two cases, there are an unlimited number of sig figs., i.e; they cannot be quantified. These two cases are examples.

a) When objects are counted (you cannot have 5.3 toes on your left foot)

b) When numbers are used as conversion factors, i.e.; 12 inches/ 1 ft. Here sig figs do not enter the equation. These can also be called **exact numbers!**

Using Significant Figures in Calculations

When making measurements, and using those measurements in a calculation, the result of the calculation can **NEVER** have more accuracy than the measurement with the least amount of accuracy.

So, in order to obey this rule of Sig Figs, it is necessary to define how we will do calculations

ADDITION OR SUBTRACTION If the calculation involves addition or subtraction; the answer can have no more decimal places than the measurement with the LEAST number of decimal places.

Example: 12.1 inches + 2.1 inches + 3.03 inches = According to our calculator, the answer to this question is 17.23 inches. BUT: using sig fig rules we would give the answer as ~ 17.2 inches. Yes, the calculator says 17.23 inches, but sig fig rules say we can only have an answer with ONE DECIMAL PLACE, because our least precise measurements, the 2.1 and 12.1 inch measurements only have one decimal place. OUR ANSWER CAN ONLY HAVE ONE DECIMAL PLACE if ADDING OR SUBTRACTING.

Do all the calculations on the calculator, and then ROUND for significant figures. If adding or subtracting, put all the digits in the calculator. Calculate the answer, and then round off to the correct number of sig figs.

MULTIPLICATION OR DIVISION When multiplying or dividing, and using sig figs, the final answer can have only as many sig figs as the measurement with the LEAST number of sig figs.

Example: 6.3 cm x 2.2 cm = 13.86 cm2 according to my calculator. Using sig figs, my answer would be ~ 14. I must round my answer to 2 sig figs because each measurement in the problem, the 6.3 and 2.2 cm, each only have 2 sig figs.

Example: 6.66 cm / 3.2 cm = 2.08125 cm2 according to my calculator. Using sig figs, my answer is rounded to 2.1. The measurement in the problem with the least number of sig figs is the 3.2 cm. Thus, my answer cannot have more than 2 sig figs, because 3.2 has 2 sig figs.

Qualitative and Quantitative Measurements

Qualitative measurements are descriptive – "It is really cold today." This sentence describes the temperature without stating an exact temperature.

Quantitative measurements are specific – "It is 10 degrees below zero today." This sentence tells us exactly how cold it really is. It quantifies, or gives an exact quantity, to the temperature.

Both qualitative and quantitative measurements have value.

Temperature Conversions

Temperature measurements are important in chemistry. We will employ two temperatures scales (Metric), the Celsius and Kelvin scales in Chemistry. Conversion between the two is much simpler than any Fahrenheit (English- Not Metric) conversions you have learned in other science classes. **Forget Fahrenheit**. It is not used in chemistry. The Celsius scale is named after the Swedish astronomer Anders Celsius.

- Water freezes at 0.0^{o}C, and water boils at 100.0^{o}C. Easy to remember and IMPORTANT.
- Room temperature is between 20.0 ^{o}C and 25.0 ^{o}C, and body temperature is approximately 37.0 ^{o}C.

The Kelvin scale is named after Lord Kelvin, and is designated as just Kelvins. Not degrees Kelvin. Kelvin degrees are the same size as Celsius degrees, so that an increase of 1^{o}C is the same magnitude of temperature increase as an increase of 1 K. Absolute Zero is -273.14 ^{o}C. At absolute zero all particle movement stops for most substances.

Temperature Conversion Table

from **Celsius**	to **Kelvin**	$T_K = t_C + 273.15$
	to **Fahrenheit**	$t_F = (1.8 \cdot t_C) + 32$
from **Fahrenheit**	to **Kelvin**	$T_K = \frac{t_F - 32}{1.8} + 273.15$
	to **Celsius**	$t_C = \frac{t_F - 32}{1.8}$
from **Kelvin**	to **Celsius**	$t_C = T_K - 273.15$
	to **Fahrenheit**	$t_F = 1.8 \cdot (T_K - 273.15) + 32$

0^{o}C = 273 K water freezes at 273 K

100^{o}C = 373 K water boils at 373 K

To convert from Celsius to Kelvin, simply add 273 to the Celsius temperature.

To convert from Kelvin to Celsius, simply subtract 273 from the Kelvin temperature.

Example: 10^{o}C = 283 K 298 K = 25^{o}C

Helpful websites to aid in completing packet: Here are some websites that may be helpful. It is not required that you view them unless you feel you need more help understanding the concepts.

1) http://chemwiki.ucdavis.edu/Homework_and_Exercises/Exercises%3A_General_Chemistry/Exercises%3A_Zumdahl_and_Zumdahl
2) https://www.khanacademy.org/math/pre-algebra/rates-and-ratios/metric-system-tutorial/v/converting-within-the-metric-system
3) http://chemwiki.ucdavis.edu/Textbook_Maps/General_Chemistry_Textbook_Maps/Map%3A_Chemistry%3A_The_Central_Science_(Brown_et_al.)/01._Introduction%3A_Matter_and_Measurement

APPLICATION: ***Complete the following mathematical problems and questions.***

1. How many grams are present in 29 kg?
2. How many millimeters are present in 2 meters?
3. Do the following conversions: (temperature conversion)
4. 4 K = ______ ^{o}C
5. 170^{o}C = _______ K
6. 87 K = ______ ^{o}C
7. 222^{o}C = ___________ K

Convert the following in scientific notation:

45 ____________________ 699 ____________________ 2884 ______________________

28395____________________ 0.344________________ 45.83 ____________________

302.555__________________ 0.00043502 _____________________

Convert the following from scientific notation to their equivalent decimal numbers:

3.692×10^{6} _______________ 4.9×10^{-2}__________________ 1.9734×10^{5} _________________

5.55050×10^{-9} _________________ 7×10^{-3}_________________ 9.97×10^{-1}________________

1. Three measurements are made of the mass of a new automobile. The first is 3500 lbs., the second is 3502 lbs., and the third is 3501 lbs. The actual mass is known to be 3500 lbs. So, are these measurements of mass accurate? ___________ Are they precise?_____________
2. I buy a 2 x 4 board from Home Depot. I measure it with an extremely fine measuring tool and find the width of the board to be 4 inches. Can I say that Home Depot accurately measured the board? _______________ Can I say that my measurement = 4 inches was precise?______________
3. In the above question, how would it be possible to increase the precision of my measurement?
4. I have a choice of two scales to use to weigh myself. The first will weigh in pounds, the second scale in pounds with marks between each to indicate the 1/2-pound weight. Which will be more precise if both are accurate?
5. Which digit of the following measurements is NOT significant, but a very good estimate?
6. 2.379 inches ___________ 19.334 m ____________ 890.5 miles _______________
7. Do the following calculations and round when necessary for the correct significant figures (Do Not Forget your Units!): a) 8.7g+15.43g +19g= b) 853.2 L – 642.333 L =
8. Give an example of a qualitative measurement you might make in everyday life.
9. Give an example of a quantitative measurement that you might make in everyday life.

SCIENTIFIC NOTATION & UNIT ANALYSIS

Change the following to Scientific Notation (maintain the number of significant figures):

1. 5.280 mg =___________
2. 2,000 L = ___________
3. 15 sec =___________
4. 6,589,000 moles =___________
5. 70,400,000,000 =___________
6. .00263 µm =___________
11. 2,560 ml = ___________
12. .0009 g = ___________
13. 8,900,000 g =___________
14. .0920 L = ___________
15. 6,300. mm = ___________
16. .90 mg =___________

7. .00589 Mm= ____________

8. .006 nm=_____________

9. .400 Kg= ____________

10. .08060 m =_____________

17. 250. Kg= _____________

18. .006087 mg = _____________

19. 500,000 mm =_____________

20. .0000000105 μm =_____________

Make the following Metric System conversions using “unit analysis” (you may use scientific notation):

1. 100 mg _____________ = _____________ g

2. 20 cm _____________ = _____________ m

3. 50 L _____________ = _____________ kL

4. 22 g _____________ = _____________ cg

5. 825 cm _____________ = _____________ km

6. 2,350 kg _____________ = _____________ g

7. 19 mL _____________ = _____________ cL

8. 52 km _____________ = _____________ m

9. 36 m _____________ = _____________ cm

10. 18 cm _____________ = _____________ mm

11. 6 g _____________ = _____________ mg

12. 4,259 mg _____________ = _____________ g

1) Construct a table with the headings “Solid,” “Liquid,” and “Gas.” For any given substance, state what you expect for each of the following:

a) the relative densities of the three phases
- the physical shapes of the three phases
- the volumes for the same mass of compound
- the sensitivity of the volume of each phase to changes in temperature
- the sensitivity of the volume to changes in pressure

2) Classify each substance as homogeneous or heterogeneous and explain your reasoning.

a) platinum
b) a carbonated beverage
c) bronze
d) wood

e) natural gas
f) Styrofoam

3) Classify each substance as homogeneous or heterogeneous and explain your reasoning.
a) snowflakes
b) gasoline
c) black tea
d) plastic wrap
e) blood
f) water containing ice cubes

4) Classify each substance as a pure substance or a mixture and explain your reasoning.
a) seawater
b) coffee
c) 14-karat gold
d) diamond
e) distilled water

5) Classify each substance as a pure substance or a mixture.
a) cardboard
b) caffeine
c) tin
d) a vitamin tablet
e) helium gas

6) Classify each substance as an element or a compound.
a) sugar
b) silver
c) rust
d) rubbing alcohol
e) copper

7) Classify each substance as an element or a compound.
a) water
b) iron
c) hydrogen gas
d) glass
e) nylon

8) What techniques could be used to separate each of the following?
a) sugar and water from an aqueous solution of sugar
b) a mixture of sugar and sand
c) a heterogeneous mixture of solids with different solubilities

9) What techniques could be used to separate each of the following?
a) solid calcium chloride from a solution of calcium chloride in water
b) the components of a solution of vinegar in water
c) particulates from water in a fish tank

10) Calculate the volume of 10.00 g of each element and then arrange the elements in order of

decreasing volume. The numbers in parentheses are densities**. Density=mass/volume**

a) copper (8.92 g/cm3)
b) calcium (1.54 g/cm3)
c) titanium (4.51 g/cm3)
d) iridium (22.85 g/cm3)

Complete the following table.

Density (g/cm3)	**Mass (g)**	**Volume (cm3)**	**Element**
3.14	79.904		Br
3.51		3.42	C
	39.1	45.5	K
11.34	207.2		Pb
	107.868	10.28	Ag
6.51		14.0	Zr

1) Gold has a density of 19.30 g/cm3. If a person who weighs 85.00 kg (1 kg = 1000 g) were given his or her weight in gold, what volume (in cm3) would the gold occupy? Are we justified in using the SI unit of mass for the person’s weight in this case?
2) An irregularly shaped piece of magnesium with a mass of 11.81 g was dropped into a graduated cylinder partially filled with water. The magnesium displaced 6.80 mL of water. What is the density of magnesium?
3) The density of copper is 8.92 g/cm3. If a 10.00 g sample is placed in a graduated cylinder that contains 15.0 mL of water, what is the total volume that would be occupied?
4) At 20°C, the density of fresh water is 0.9982 kg/m3, and the density of seawater is 1.025 kg/m3. Will a ship float higher in fresh water or in seawater? Explain your reasoning.