The Sagemont School
Chemistry Summer Packet

Welcome to Chemistry. You will study the composition of matter (anything that has mass and takes up space) and the changes it undergoes. In order to prepare to study this material effectively, you should have some background knowledge. This packet contains that information.

This packet should be completed over the summer and returned at the beginning of the school year. It is your first assignment and it will be discussed and graded. You may work with classmates, but you will be held responsible for this material, so be sure you understand it.

I. Introduction

There are 5 divisions of chemistry, and they overlap:

1. Organic chemistry: the study of substances containing carbon – often living, or once living, things
2. Analytical chemistry: the study of the composition of substances – how much of a chemical or substance is present
3. Physical chemistry: the study involving prediction of the behavior of chemicals – why does a substance do what it does under specific conditions or just in general
4. Biochemistry: the study of the chemistry of living organisms – how and why the chemistry of living organisms works and affects lives
5. Inorganic chemistry: the study of substances not containing carbon – nonliving things or things that never lived, and how they interact or react.

Questions:

Which division or divisions of chemistry might be used to examine:

a) The mechanism by which blood clots

b) The amount of a toxic substance found in a water supply

c) The reason a metal melts at a specific temperature

d) The formation of chemical compounds made up of metals

II.The Scientific Method

In order to pursue the study of matter and the changes it undergoes, or to solve other problems in science, it is important to use a logical and predictable technique. The Scientific Method is one such technique. It is a common sense approach applicable to not only scientific questions, but many other, everyday quandaries as well.

There are variations on the Scientific Method, but most of them contain the following steps:

1. **Observations:** Something needs explanation – a problem, a question, an issue.
2. **Hypothesis:** A suggestion which solves the problem, answers the question, explains the issue is offered. It may be termed an "educated guess" because it is based on some logical explanation or suggestion of cause. Often stated as "If ___________, then ___________.
3. **Experimentation:** The testing of the hypothesis, usually repeated trials are necessary to assure that the results of the tests can be accepted as genuine.
4. **Theory:** A theory, or explanation of the results of the experiment is offered. The theory is a possible answer, it cannot be proven to be true; it is possible to disprove a theory.

Some sources will list a fifth step to the Scientific Method:

5. **Scientific Law:** A statement is offered which summarizes the results of experiments and observations. A scientific law does not try to explain the results, it merely states what the results are.

Questions:

1. Given the following five statements, identify which of the five parts of the Scientific Method they describe.

A) Ash trees infested with ash borer insects die.

B) Leaves are falling off my ash trees and there are areas of bark which are destroyed.

C) Ash trees are infested by ash borer insects in Oakland County. It is possible that my trees are also.

D) I will remove the bark in areas of my trees and look for characteristic signs of ash borer disease. I will look at each tree in several spots.

E) If my trees have ash borer disease, they would lose their leaves and their bark would be damaged.

2. Write your own example of an observation or problem, and use the five steps of the Scientific Method to solve it.

III. Dimensional Analysis

Many calculations are required in this course. The technique used to solve most problems is called dimensional analysis. It uses conversion factors in order to derive answers.

For example, suppose the question was “How many inches are there in 3.5 feet?”

There is a defined relationship between feet and inches: 1 foot = 12 inches.

If this relationship was written as a fraction it could be said that

1 foot = 1 or, 12 inches = 1. This is true because 1 foot = 12 inches. 12 inches = 1 foot

Now, returning to my question, and recognizing that multiplying anything by 1 does not change that which is multiplied, I could answer my question using dimensional analysis.

3.5 feet · 12 inches = 42 inches

Units of feet/foot cancel, leaving inches. 1 foot = 12 inches

3.5 feet was multiplied by a conversion factor that allowed feet to be converted to inches.

This method should be used whenever possible.

Always begin these problems with the amount given, above the amount was 3.5 feet. Then use the appropriate conversion factor to convert from the units you have to the units you want. In this example, the desired unit was inches. Remember: the conversion factor is a fraction in which the two components are EQUAL. i.e.: 1 foot = 12 inches.

The conversion factor can also be the number required for each, such as: each student requires 2 pencils to take a test. How many pencils are needed if 15 students are taking a test?

15 students · 2 pencils = 30 pencils

Units of “student” cancel, leaving 1 student pencils

Questions:

Use dimensional analysis to solve:

1. How many eggs are there in 6.2 dozen?
2. Each automobile made by Ford requires 4 tires. How many tires are needed to manufacture 259 cars?
3. Each student in chemistry requires 3 beakers in her lab drawer. How many beakers are required for a class of 7 students?

IV. 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.

QUANTITY MEASURED SI UNIT SI SYMBOL

Length: Meter m

Mass: Kilogram kg

Amount: Mole mol

Temperature: Kelvin K (not °K)

Time: Second s

Volume: Cubic meter (informally, liter) m3

Pressure: Pascal Pa

Density: Grams per cm3 or grams per mL g/cm3 or g/mL

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.

For example, a megameter would be 1 x 10^6 (1,000,000) meters.

A kilogram = 1 x 10^3 grams. (1,000)

A centimeter =1 x 10^-2 meters. (.001) A micromole = 1 x 10^-6 moles. (.0000001)

PREFIX SYMBOL MEANING

Mega M X 10^6 (a million times larger)

Kilo k X 10^3 (a thousand times larger)

Deci da X 10^-1 (ten times smaller)

Centi c X 10^-2 (a hundred times smaller)

Milli m X 10^-3 (a thousand times smaller)

Micro µ X 10^-6 (a million times smaller)

Nano n X 10^ billion times smaller)

Pico p X 10^-12(a trillion times smaller)

It is possible to convert between units by using dimensional analysis.

Example: How many kilograms are present in 5000 grams?
5000 grams · 1 kilogram (kg) = 5 kg
1x 10^3
1kg = (1000 g)
Example: How many nanograms are present in 20 g?
20 g · 1 x 10^9
ng = 20,000,000,000 ng
1 ng = (1,000,000,000g)

Questions:
How many grams are present in 29 kg?
How many millimeters are present in 2 meters?

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

For example:
Converting to scientific notation:
36,000 would be written as 3.6 x 10^4

23.67 would be written as 2.367 x 10^1
35069032 would be written as 3.5068032 x 10^7
0.003 would be written as 3 x 10^-3
0.23999 would be written as 2.3999 x 10^-1
0.000000000000000000000007 would be written as 7 x 10^-24

Converting from scientific notation to equivalent decimal numbers:

3.55 x 10^3 would be written as 3550

6.887 x 10^-6 would be written as 0.000006887

Questions:

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 x 10^6 _______________ 4.9 x 10^-2_______________ 1.9734 x 10^5 _____________

5.55050 x 10^-9 ____________________ 7 x 10^3_______________ 9.97 x 10^-1 _________

VI. 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.

Questions:

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. 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 the board was accurately measured by Home Depot? _______________ Can I say that my measurement = 4 inches was precise? ______________

3. In question # 2, 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 ½ pound weight. Which will be more precise if both are accurate? ____

VII. 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.
SO: significant figures include all digits we are sure of PLUS the last one, which is a very good estimate.

Questions:

Which digit of the following measurements is NOT significant, but a very good estimate?

2.379 inches ___________ 19.334 m ____________ 890.5 miles_______________

In addition, there are rules that determine whether digits are significant or not. Below are the Rules:

Rule 1: All non-zero digits and zeros BETWEEN non-zero digits, are significant.

Example: 245 has 3 significant figures (sig figs)… 2405 has 4 sig figs.

Rule 2: Zeros at the END of a number and to the right of a decimal point ARE always significant.

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

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

Example: 0.0071 has 2 sig figs (the 7 and 1) The three zeros to the left of the 71 are Not significant because they are merely place holders.

Rule 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 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., there would be 4 sig figs because the decimal point appearing after the 7100. makes the two zeros significant.]

Rule 5: In two cases, there are an unlimited number of sig figs., i.e, they cannot be quantified. These two cases are a) when objects are counted (you cannot have 5.3 toes on your left foot) and b) when numbers are used as conversion factors, i.e.; 12 inches/ 1 ft. Here sig figs do not enter the equation.

VIII. 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 we are ADDING OR SUBTRACTING.

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

Questions:
Do the following calculations and round for significant figures:
a) 8.7 g + 15.43 g + 19 g =
b) 853.2 L – 642.333 L =

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.

Questions:

Do the following calculations and round for significant figures:

a) 4.32 cm x 1.7 cm =

b) 38.742 kg ÷ 0.421 kg =

c) 5.40 m x 3.21 m x 1.871 m =

d) 38.0 in ÷ 2.121 in =

IX. 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.

Questions:

Give an example of a qualitative measurement you might make in everyday life.

Give an example of a quantitative measurement that you might make in everyday life.

X. Temperature Conversions

Temperature measurements are important in chemistry. We will employ two temperatures scales, the

Celsius and Kelvin scales. Conversion between the two is much simpler than any Fahrenheit 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°C, and water boils at 100°C. Easy to remember and IMPORTANT.

Room temperature is between 20°C and 25°C, and body temperature is 37°C.

The Kelvin scale is named after Lord Kelvin, and is designated as just Kelvins. <u>Not degrees Kelvin.</u>

Kelvin degrees are the same size as Celsius degrees, so that an increase of 1°C is the same Magnitude of temperature increase as an increase of 1 K.

0°C = 273 K SO water freezes at 273 K

100oC = 373 K SO 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°C = 283 K 298 K = 25°C.

Questions:

Do the following conversions:

4 K = _______°C   170°C = ________K   87 K = _______°C   222°C = ____________ K

Helpful Websites to Aid in Completing Chemistry/Honors Chemistry 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.

II. Scientific method:
http://phyun5.ucr.edu/~wudka/Physics7/Notes_www/node5.html
http://www.selu.edu/Academics/Education/EDF600/Mod3/sld001.htm

III. Dimensional analysis (also called unit analysis)
http://physics.nist.gov/cuu/Units/units.html
http://www.chem.tamu.edu/class/fyp/mathrev/mr-da.html

IV. Metric system and SI Units:
http://lamar.colostate.edu/~hillger/everyday.htm
http://physics.nist.gov/cuu/Units/
http://physics.nist.gov/cuu/Units/units.html

V. Scientific notation:
http://janus.astro.umd.edu/astro/scinote.html
http://www.nyu.edu/pages/mathmol/textbook/scinot.html
http://www.ieer.org/clssroom/scinote.html

VI. Accuracy and precision:
http://elchem.kaist.ac.kr/vt/chem-ed/data/acc-prec.htm
http://www.fordhamprep.com/gcurran/sho/sho/lessons/lesson22.htm

VII. Significant figures:
http://science.widener.edu/svb/tutorial/sigfigures.html
http://chemed.chem.purdue.edu/genchem/topicreview/bp/ch1/sigfigs.html
http://lectureonline.cl.msu.edu/~mmp/applist/sigfig/sig.htm

IX. Qualitative and quantitative measurement:
http://www.guidestarco.com/Qualitative-and-Quantitative-Survey-Research.HTM

X. Temperature conversion:

http://www.onlineconversion.com/temperature.htm

http://www.convert-me.com/en/convert/temperature

Kelvin temperature scale:

http://www.infoplease.com/ce6/sci/A0827335.html

http://lamar.colostate.edu/~hillger/temps.htm

http://www.factmonster.com/ce6/sci/A0827335.html

Video Resources:

Tyler Dewitt You Tube Channel:

https://www.youtube.com/channel/UCj3EXpr5v35g3peVWnVLoew

Bozeman Science You Tube Channel: https://www.youtube.com/user/bozemanbiology/videos

Professor Dave Explains: https://www.youtube.com/channel/UC0cd_-e49hZpWLH3UIwoWRA