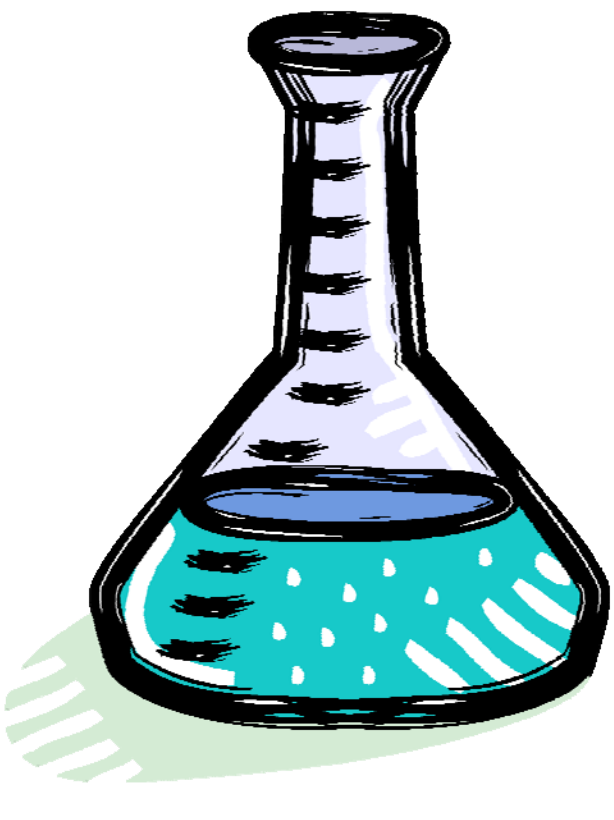
Grossmont College

Chemistry 115
Laboratory Manual



Grossmont College
Periodic Table of the Elements

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	1A	**2A**	**3B**	**4B**	**5B**	**6B**	**7B**	**8B**	**8B**	**8B**	**1B**	**2B**	**3A**	**4A**	**5A**	**6A**	**7A**	**8A**
1	**1** H 1.008																	**2** He 4.003
2	**3** Li 6.941	**4** Be 9.012											**5** B 10.811	**6** C 12.011	**7** N 14.007	**8** O 15.999	**9** F 18.998	**10** Ne 20.180
3	**11** Na 22.990	**12** Mg 24.305											**13** Al 26.982	**14** Si 28.086	**15** P 30.974	**16** S 32.061	**17** Cl 35.453	**18** Ar 39.948
4	**19** K 39.098	**20** Ca 40.078	**21** Sc 44.956	**22** Ti 47.867	**23** V 50.942	**24** Cr 51.996	**25** Mn 54.938	**26** Fe 55.845	**27** Co 58.933	**28** Ni 58.693	**29** Cu 63.546	**30** Zn 65.382	**31** Ga 69.723	**32** Ge 72.640	**33** As 74.922	**34** Se 78.960	**35** Br 79.904	**36** Kr 83.798
5	**37** Rb 85.468	**38** Sr 87.62	**39** Y 88.906	**40** Zr 91.224	**41** Nb 92.906	**42** Mo 95.962	**43** Tc (98)	**44** Ru 101.07	**45** Rh 102.906	**46** Pd 106.42	**47** Ag 107.868	**48** Cd 112.411	**49** In 114.818	**50** Sn 118.710	**51** Sb 121.760	**52** Te 127.60	**53** I 126.905	**54** Xe 131.293
6	**55** Cs 132.905	**56** Ba 137.327	**57** La 138.906	**72** Hf 178.49	**73** Ta 180.948	**74** W 183.84	**75** Re 186.207	**76** Os 190.23	**77** Ir 192.217	**78** Pt 195.078	**79** Au 196.967	**80** Hg 200.59	**81** Tl 204.383	**82** Pb 207.2	**83** Bi 208.980	**84** Po (209)	**85** At (210)	**86** Rn (222)
7	**87** Fr (223)	**88** Ra (226)	**89** Ac (227)	**104** Rf (261)	**105** Db (262)	**106** Sg (266)	**107** Bh (264)	**108** Hs (277)	**109** Mt (268)	**110** Ds (281)	**111** Rg (272)	**112** Cn (285)	**113** Uut (284)	**114** Fl (289)	**115** Uup (288)	**116** Lv (293)	**117** Uus (294)	**118** Uuo (294)

6	**58** Ce 140.116	**59** Pr 140.908	**60** Nd 144.24	**61** Pm (145)	**62** Sm 150.36	**63** Eu 151.964	**64** Gd 157.25	**65** Tb 158.925	**66** Dy 162.500	**67** Ho 164.930	**68** Er 167.259	**69** Tm 168.934	**70** Yb 173.04	**71** Lu 174.967
7	**90** Th 232.038	**91** Pa 231.036	**92** U 238.029	**93** Np (237)	**94** Pu (244)	**95** Am (243)	**96** Cm (247)	**97** Bk (247)	**98** Cf (251)	**99** Es (252)	**100** Fm (257)	**101** Md (258)	**102** No (259)	**103** Lr (262)

rev. March 2013

Locker Number:__________________

Combination: __________________

Grossmont College
Chemistry 115
Laboratory Manual

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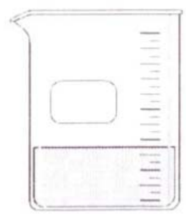
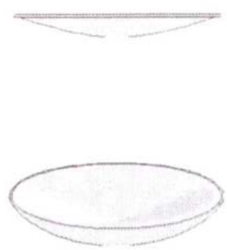
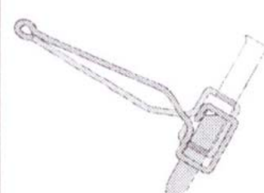
14-0002-162

Name____

Assorted Lab Equipment

Table 0.1 Selected Lab Drawer Equipment

Glass Bottle 	Beaker 	Erlenmeyer Flask 	Graduated Cylinder 
Test Tube 	Watch Glass 	Funnel 	Evaporating Dish 
Crucible Tongs 	Dropper Bulb 	Eye Dropper 	Test Tube Holder 
Thermometer 			

EXPERIMENT 1

EXPONENTIAL NOTATION & SCIENTIFIC NOTATION

Exponential Notation

Scientific study includes measurements and quantities that are tremendously large and some that are inconceivably small. For example, there are approximately 3,350,000,000,000,000,000,000 molecules of sucrose in every teaspoon of table sugar. This is not surprising if one considers that the distance between carbon atoms in a molecule of sucrose is 0.000 000 0015 meters. Both of these numbers are more easily written using scientific notation, 3.35×10^{21} and 1.5×10^{-9}, respectively. The exponential format is of this form, $A \times 10^{n}$, where A is a digit between 1 and 10 and the exponent n is a positive or negative integer. Not only does scientific notation make very large and very small numbers more convenient to write, but it also allows for the recording of measurements and calculations without ambiguous zeroes.

Powers of 10

A power of 10 is the result of 10 raised to some exponential value. For example, "ten squared" is written 10^2. This means that the number 10 is multiplied by itself or 10×10. Ten raised to the third power is 10^3 or $10 \times 10 \times 10$. Ten raised to the nth power is 10^n or $10 \times 10 \times 10 \ldots n$ (this is 10 multiplied by itself *n* times). Since $10^2 = 100$ and $10^3 = 1000$, we can see that the number of zeros is equal to *n*. Thus, $10^5 = 10 \times 10 \times 10 \times 10 \times 10 = 100{,}000$. One million is written in decimal form as 1,000,000 which is one, followed by six zeros; in scientific notation, one million is written as 10^6. So far, our examples have been numbers that are greater than ten, 100, 1000, 100,000 and a million. Numbers that are *less than one* can also be expressed with powers of ten; the difference is that the *exponent will be negative*.

Converting Ordinary Decimal Numbers into Scientific Notation

If a number is ten or greater, we move the decimal point to the left by *n* places to obtain a number between one and ten. That result is then multiplied by a power of 10; actually it is multiplied by ten raised to the *n*th power. Here is an example, the approximate distance between the earth and the sun is 93 million miles or 93,000,000 mi. In scientific notation we write 9.3×10^7 mi.

decimal point moved
seven places to the left

$$93000000.\ \text{mi} = 9.3 \text{ X } 10^{7}\ \text{mi}$$

Decimal point is moved
to obtain a number betwen
1 and 10. (To the right of 1st
non-zero digit.)

If a number is less than one, the decimal point is shifted to the right until, again, we obtain a number between one and ten. Then multiply that number by 10^{-n} where *n* is equal to the number of places that we moved the decimal point. Here is an example. The size of bacteria is about 0.0000322 inches. In scientific notation we write, 3.22×10^{-5} in.

Decimal point moved
five places to the right.

$$0.0000322\ \text{in} = 3.22 \text{ X } 10^{-5}\ \text{in}$$

Decimal point is moved
to obtain a number betwen
1 and 10. (To the right of 1st
non-zero digit.)

Look at some more examples:

1. 450,000 lbs. = 4.5×10^5 lbs.
2. 385.26 h = 3.8526×10^2 h
3. 0.03307 cm = 3.307×10^{-2} cm
4. 0.0005 L = 5×10^{-4} L

Again notice that numbers greater than 10 have positive exponents and numbers less than 1 have negative exponents.

Finally, consider this last example: we usually write 87,200 as 8.72×10^4 but we could also express this same number as 87.2×10^3 or 872×10^2. The only difference is the number of places that we move the decimal point. However, the standard form of scientific notation uses a number between 1 and 10; thus $8.72 \text{ X } 10^4$ is the usual expression for this number.

There is one last point to consider when writing any number that is part of a measurement, whether the number is in decimal form or in scientific notation. Every number must include the unit of measurement such as, inches (in), meters (m), milliliters (mL), seconds (s), etc.

Converting Scientific Notation into Decimals

If a number is written in scientific notation and we wish to convert it into a decimal number, we simply reverse the process described above. For example, 5.92×10^5 kg can be written in decimal form as 592,000 kg. To make this conversion, we drop the power of 10 and move the decimal point five places to the right. Remember, a positive exponent means we are dealing with a number greater than 10 so in order to make 5.92 a bigger number we must add zeros which is done by moving the decimal point to the right. Recall that we moved the decimal to the left to express 592 thousand in scientific notation, and the reverse process requires that we move the decimal back to the right.

$592{,}000 \text{ kg} = 5.92 \text{ X } 10^5 \text{ kg}$

decimal point is moved to the left to convert this decimal number into scientific notation.

$5.92000 \text{ X } 10^5 \text{ kg} = 592{,}000 \text{ kg}$

decimal point is moved to the right to convert this scientific notation into a decimal.

The number 9.7×10^{-2} kg is written in decimal form as 0.097 kg. Again, the power of 10 is dropped and the decimal point is moved 2 places to the left.

$0.097 \text{ kg} = 9.7 \text{ X } 10^{-2} \text{ kg}$

decimal point is moved to the right to convert this decimal number into scientific notation

$009.7 \text{ X } 10^{-2} \text{ kg} = 0.097 \text{ kg}$

decimal point is moved to the left to convert this scientific notation into a decimal

Adding and Subtracting Exponential Numbers

To add or subtract exponential numbers, both numbers must have the same exponent. For example, in this problem, the exponents are not the same: $2.3 \times 10^2 + 1.4 \times 10^3$. One of the numbers must have its exponent changed:

$$\begin{array}{r} 2.3 \text{ X } 10^2 \\ + \ 1.4 \text{ X } 10^3 \\ \hline \end{array}$$

These two numbers cannot be added in this form because the exponents are different. It must be transformed as below:

$$\begin{array}{r} 0.23 \text{ X } 10^3 \\ + \ 1.4 \ \ \text{ X } 10^3 \\ \hline 1.63 \text{ X } 10^3 \end{array}$$

Here the upper exponent is changed.

$$\begin{array}{r} 2.3 \text{ X } 10^2 \\ + \ 14 \ \ \text{ X } 10^2 \\ \hline 16.3 \text{ X } 10^2 \text{ or } 1.63 \text{ X } 10^3 \end{array}$$

Here the lower exponent is changed.

$$\begin{array}{r} 230 \\ +1400 \\ \hline 1630 \text{ or } 1.63 \text{ X } 10^3 \end{array}$$

Here both numbers are converted to decimal format.

Now it should be obvious why we occasionally express exponential numbers with a digit number that is not between 1 and 10.

Subtraction follows the same rule:

$$\begin{array}{r} 57.0 \text{ X } 10^2 \\ - \ 4.4 \text{ X } 10^2 \\ \hline 52.6 \text{ X } 10^2 \text{ or } 5.3 \text{ X } 10^2 \end{array} \qquad \begin{array}{r} 5.70 \text{ X } 10^3 \\ + \ 0.44 \text{ X } 10^3 \\ \hline 5.26 \text{ X } 10^2 \text{ or } 5.3 \text{ X } 10^2 \end{array}$$

Greyed out digits are not significant. More in later section.

Later, when we discuss significant figures, it will become apparent why we round up to two digits in our final answer.

Multiplying and Dividing Exponential Numbers

To multiply two exponential numbers, the digits are multiplied and the exponents are added together. Look at this example:

$$\begin{aligned}&\left(2\times10^{4}\right)\left(3\times10^{2}\right)=\left(2\times3\right)\times\left(10^{(4+2)}\right)\\&=6\times10^{6}\end{aligned}$$

For division, the digits are divided and the exponents are subtracted from each other:

$$\frac{3.2\times10^{2}}{1.2\times10^{3}}=\frac{3.2}{1.2}\times10^{(2-3)}=2.7\times10^{-1}$$

Logarithms and Antilogarithms

Logarithms are often used in chemistry to convert negative powers of 10 in concentrations to positive numbers. Logs are especially important in calculating pH and pOH values as will be seen in the acid-base chapter. Antilogarithms undo that process.

$$\log(5.1 \text{ x } 10^{-2}) = -1.292429824 = -1.29$$

$$\text{antilog}(0.422) = 10^{0.422} = 2.64208757 = 2.64$$

Significant Figures

Significant figures are the number of digits used in a measurement. Suppose your class assignment was to find the mass of a 1-cent coin, a penny. There are several types of balances in the room and no further instructions are given as to which balance to use. Each student weighs the same coin on a similar balance and records their measurement on a class data sheet. The values recorded by the students are listed in the table below:

2.900 g	2.843 g
2.860 g	2.850 g
2.844 g	2.890 g

At this point it should be obvious that any measurement involves a certain degree of uncertainty. **Uncertainty** is an indication of how confident we are that our number represents the actual value. What is the actual value? Scientists often state that the actual value can never be known. All we can do is measure a value (mass in this case) and interpret the result as having some degree of uncertainty.

Accuracy and Precision

When we discuss uncertainty, we use the terms **accuracy** and **precision.** In common usage, we use these terms interchangeably but scientists have strict definitions for each. **Accuracy** refers to how well our experimental value compares to the actual value.

Precision refers to how well a number of independent measurements agree with one another. For example, if we measured the mass three times and got numbers that varied greatly, we would say that our precision was poor. Precision also refers to the manner in which the measurement was obtained. A particular measurement is constrained to the precision of the measuring device or technique. For example, if we use a single-pan balance that measures to the nearest tenth of a gram, our precision is ± 0.1 g; we would say that our measurement is precise to within one tenth of a gram. A value of 2.9 g indicates that the actual mass lies within the range of 2.8 to 3.0 grams. In this case, we record the mass as:

5. 2.9 ± 0.1 g

If we use an electronic balance that measures to the nearest 1000th of a gram (as that data shown in the table), then our precision increases to ± 0.001 g and we record the mass as:

6. 2.844 ± 0.001 g

It is understood that the last digit in any measurement is the estimated digit. Sometimes the last digit is called the "doubtful digit" or "unreliable digit".

Scientists communicate the level of uncertainty by the use of significant figures. A mass of 2.9 g contains two significant figures; a mass of 2.8439 g contains 5 significant figures. We can see that the value of 2.8439 g is more precise than 2.9 g, because the greater the number of significant figures, the greater the precision of the measurement. Proper use of significant figures eliminates the need to tell the reader that a single-pan balance was used or an electronic balance was used. Again, it is understood that the last digit is the estimated digit.

One final point about numbers should be mentioned. Scientists refer to numbers as exact or inexact. Exact numbers are numbers that we obtain by counting small groups of objects or numbers obtained by definition. For example, we can count the fingers on our hand and get an exact number (most people have 5). There is no uncertainty in this result, but we cannot count large groups of objects without some degree of uncertainty. For example, the number of stars in our galaxy is not an exact number. Other kinds of exact numbers are the result of mathematical definitions or conversion factors. For example, one kilogram is equal to exactly 1000 grams.

Inexact numbers are those that we have already discussed; all measurements are inexact numbers because they are estimates of the actual value and contain some degree of uncertainty. Some conversion factors can be expressed to varying degrees of precision as well. When we convert English units to the corresponding metric unit, the conversion is not always exact. For example, one English pound expressed in metric units of grams is:

7. 1 lb = 453.59 g

We could also use a less precise value:

8. 1 lb = 454 g

Do we use 5 significant figures (453.59 g) or 3 significant figures (454 g)? We will answer this question when we discuss the guidelines for determining the correct number of significant figures during the course of performing mathematical operations. At this point, it may be useful to add these summary statements:

• The level of uncertainty is determined by the precision of our measurement.

• Precision refers to the reproducibility of a series of measurements.

• Significant figures are used to express the level of precision for a given measurement.

• Most of the numbers that we encounter are inexact numbers.

Rules for Determining the Number of Significant Figures

1. All non-zero digits are significant.

2. Zeros in the middle of a number are significant; 402 has three significant figures.

3. Zeros at the beginning of a number are not significant; these zeros only locate the decimal point. For example, 0.0034 has only two significant figures.

4. Zeros at the end of a number that come after the decimal point are significant. For example, 69.430 g has five significant figures. If the last digit was not significant, it would not have been recorded. This number tells the reader that the measurement was made to the nearest 1/1000th of a gram. It just so happens that the last digit on the balance was zero, but it could have been a 2 or a 7, for example.

5. Zeros at the end of a number that come before the decimal point may or may not be significant. 24,000 L may have 2, 3, 4, or 5 significant figures. We do not know if the zeros are being used to hold the decimal point or if the zeros are part of the measurement. Numbers of this type are the most difficult to deal with since not everyone agrees on how to handle the zeros before a decimal point. Fortunately this ambiguity can be avoided if numbers are expressed in scientific notation. For example, if the number 24,000 L is written as 2.4×10^4 L, we know there are only two significant figures. If the volume measurement had been made to the nearest milliliter, then all digits would be significant and it would be expressed as 2.4000×10^4 L.

Using Significant Figures in Calculations

Rounding Off Numbers in Addition and Subtraction

When numbers are added or subtracted together, the uncertainty of the result depends on the precision of the least precise number used in the calculation. The least precise number is the number with the fewest decimal places. Because of this rule, the calculated result is rounded off and some digits may be dropped.

For example, if we weighed two coins with a mass of 2.356 g and 2.2 g and then added the masses, we obtain 4.556 g. It should be obvious that the second coin was weighed on a balance with a precision of 1/10th of a gram. It makes no sense to express the final result (4.556 g) to the nearest 1/1000th of a gram since one of the measurements was made with a precision much less than 1/1000th of a gram. In other words, the final result is limited by the least precise measurement, 2.2 g, and therefore, only one digit past the decimal is allowed in the final answer and the remaining digits must be dropped. When we drop the extra digits, the last remaining digit is left unchanged or rounded up by one depending on the value of the dropped digits.

Rules for rounding. If the dropped digit is 5 or greater, the last remaining digit is increased by one; here is an example:

- 2.356 g + 2.2 g = 4.556 g is rounded to one decimal place and the digit is rounded up to give 4.6 g
- 10.432 g + 9.01 g = 19.442 g rounded to two decimal places to give 19.44 g.
- 6.705 g – 2.1 g = 4.605 g rounded to one decimal place to give 4.6 g.
- 536.12 g – 365.994 g = 170.126 g rounded to two decimal places to give 170.13.
- 12.444 + 32.1 + 678.23 + 65 = 787.774 rounded to the nearest ones place to give 788.

Rounding Off Numbers in Multiplication and Division

When numbers are multiplied or divided, the same rule applies; the result cannot be more precise than the least precise number used in the calculation. The least precise number is the number with the fewest number of significant figures, regardless of the position of the decimal point. Rounding is treated in the same manner:

- 6.144 miles / 0.15 hours = 40.96 miles/hr. rounded to two significant figures to give 41 mi/hr.
- 13.6 g/mL x 45.68 mL = 621.248 g rounded to three significant figures to give 621 g.
- (24.7 × 345.24) / 12.441 = 685.4294671 rounded to 3 significant figures to give 685.

Note: For logs the number of significant figures in the data is equal to the number of decimal places in the answer. For antilogs the number of decimal places determines the number of significant figures in the answer.

Summary

- For addition and subtraction, the number of decimal places in the calculated result is equal to the number of decimal places in the least precise number.
- For multiplication and division, the number of significant figures in the calculated result is equal to the number of significant figures in the least precise number.

Report

Name
Date
Section

1. Write each number in scientific notation.

   a. the average height of adult 182.88 cm ____________________

   b. speed of light 186,282 m/s ____________________

   c. length of X rays 0.000 000 009 m ____________________

   d. one pound of water contains 3892000 μL ____________________

2. Convert these numbers to decimal format.

   a. 1.36×10^{1} g/cm^{3} ____________________

   b. 4.39×10^{-7} mm ____________________

   c. 45.4×10^{4} g ____________________

   d. 1.760×10^{3} yd ____________________

3. Perform the following calculations then write your answer in scientific notation.

   a. $(1.2 \times 10^{-3}) \times (6.7 \times 10^{-5}) =$ ____________________

   b. $(3.7 \times 10^{8}) / (7.7 \times 10^{-2}) =$ ____________________

4. Find the log or antilog for these numbers.

   a. $\log 6.7 \times 10^{-4} =$ ____________________

   b. $\log 9.8 \times 10^{3} =$ ____________________

   c. antilog $0.0205 =$ ____________________

   d. antilog $7.6 \times 10^{-7} =$ ____________________

5. How many significant figures are in the following numbers?

    a. 11.9907 ________

    b. 0.4005 ________

    c. 800 ________

    d. 0.0821 ________

    e. 6.022×10^{23} ________

    f. 0.00506 ________

6. Round off each number as indicated:

    a. 0.2178 (round to 2 significant figures) ________

    b. 456.334 (round to 5 significant figures) ________

    c. 0.0010045 (round to 2 significant figures) ________

    d. 273.15 (round to 3 significant figures) ________

    e. 45.7893 (round to 2 significant figures) ________

7. Do the calculations and round to the correct number of significant figures:

    a. 273.15 + 34.6 + 19.233 = ________

    b. 567.1234 - 12.3 = ________

    c. $(6.022 \times 10^{23}) \times 4.6 =$ ________

    d. 90.234 / 56.001 = ________

    e. (23.556 + 21.55) / (34.556 + 43.2)= ________

EXPERIMENT 2

MEASUREMENT AND THE METRIC SYSTEM

Materials and Equipment

Sodium Chloride, NaCl, beakers, stir rod, graduated cylinders, eye dropper, scoopula, ice, centigram and digital balances, ruler, and thermometer.

Introduction

Chemistry, being an experimental science, is fundamentally based on measurements. Therefore it is important to acquire the necessary skills needed to make these measurements and to use them properly.

Unit Prefixes, The S.I. System

The scientific community almost exclusively uses the International System of Units, (Systéme Internationale d'Unités). It is based on the metric system, the decimal system of units and measurements. The basic set of units includes the meter, the gram, and the liter. Factors of 10 are used to express larger or smaller multiples of these units. Prefixes are added to the names of the units to express smaller or larger units. The most familiar of these prefixes are listed in the table below.

Prefix	Decimal Equivalent	Power of 10
Deci	0.1	10^{-1}
Centi	0.01	10^{-2}
Milli	0.001	10^{-3}
Kilo	1000	10^{+3}

Example:

A kilogram is equivalent to 10^3 multiplied by the base unit, the gram. This gives 1000 grams. A millimeter is 10^{-3} the distance of the meter, giving 0.001 meter.

Temperature

The SI unit of temperature is the Kelvin (K). The Kelvin scale is absolute and therefore does not go below zero Kelvin. More commonly we see temperature measured in Celsius (˚C) or Fahrenheit (˚F). Notice that the degree sign is not used for Kelvin temperatures.

Temperature	Boiling Point of Water	Melting Point of Water
Fahrenheit, ˚F	212 ˚F	32 ˚F
Celsius, ˚C	100 ˚C	0 ˚C
Kelvin, (K)	373 K	273 K

Conversion between temperature scales is straightforward using the following equations.

Conversion of Celsius to Fahrenheit. $^{\circ}F = (1.8 \times {}^{\circ}C) + 32$

Conversion of Fahrenheit to Celsius. $^{\circ}C = ({}^{\circ}F - 32)/1.8$

Conversion between Celsius and Kelvin. $K = {}^{\circ}C + 273$ or $^{\circ}C = K - 273$

Measuring Temperature

It is relatively easy to make errors when measuring temperatures. Precision is limited by many factors. Some of these include the calibration of the scale as well as improper placement of the thermometer bulb. To minimize human error involved in thermometer placement the following procedures should be followed:

1. Keep the thermometer away from the container walls.

2. Allow the thermometer enough time to reach equilibrium with the sample involved.

3. If applicable, be sure the sample is adequately mixed.

Mass

Mass is the measurement of the quantity of matter. You are probably familiar with the English unit, pound (lb.) to measure weight. In the sciences we use the gram to measure mass. The gram is the root unit, and the metric prefixes mentioned above are used to modify the quantity.

There are 453.6 g in 1 lb.

Measuring Mass

Chemistry experiments can be performed quite nicely using a 0.01 gram to a 0.0001 gram precision balance. Your instructor will advise you of the balance required for a specific experiment. When measuring mass the following procedures should be adhered to:

1. On an electronic balance be sure it is "zeroed" or "tared" before anything is placed on the pan. The use of balances without a "tare" or "T" button will be demonstrated by your instructor.

2. Never put reagents directly on the pan. Weighing paper, weighing boats or other containers should always be used. Spills should be cleaned up immediately.

3. Never try to make adjustments while using the balance. If something is amiss, inform your instructor.

Volume

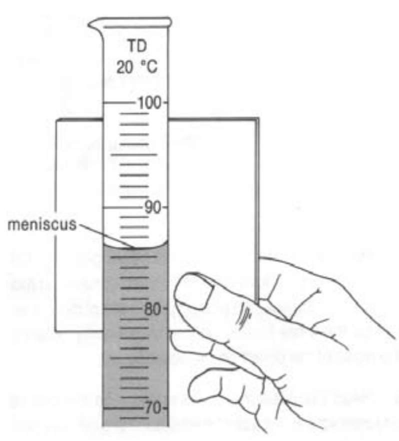
Volume is the measurement of the amount of space that matter occupies. You are probably familiar with the English units for volume such as: fluid ounce (fl. oz.), pint (pt.), quart (qt.), gallon (gal.). Remember, there are 32 fl. oz. in a quart and 4 quarts in a gallon.

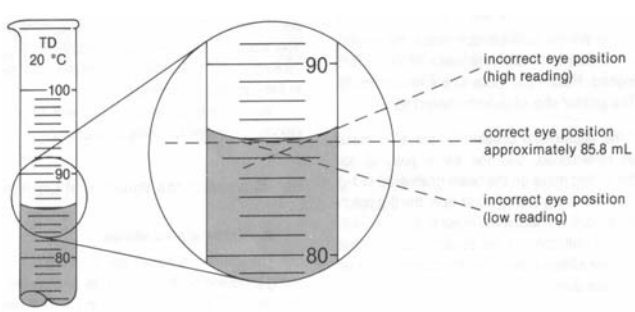
There are 1.06 qt. in 1 L.

Remember: 1 milliliter (mL) is equivalent to 1 cubic centimeter (cm^3).
1 mL = 1 cm^3

Measuring Volume

It is not useful to use beakers or Erlenmeyer flasks for volume measurements. The measurement marks on these pieces are only approximate volumes. Therefore volume measurements are usually made using graduated cylinders. When reading a graduated cylinder the bottom of the curved surface, the meniscus, of the liquids is the point where the reading is taken (See figures below.). Graduated cylinders are usually read to the 0.1 milliliters.





Precision and Accuracy

The precision of a measurement allows us to estimate its reproducibility. Precision is described in terms of deviation, (observed value – average value). Whereas accuracy is the extent to which a measurement agrees with the believed to be true value. Accuracy is described in terms of error, (observed value – true value). However, not all measurements can be compared to a known value. For a small number of measurements significant figures account for precision.

Significant Figures

When a measurement is made to the highest precision possible for the measuring instruments, the digits in the measurement are significant and are called significant figures. Significant figures in a measurement include all of the certain digits in a measurement plus one "doubtful" digit.

It is often unnecessary to measure exact quantities of substances while conducting experiments. For example, the procedure might state, "Weigh approximately 3 grams of potassium chlorate". This indicates to the experimenter that the measured quantity of this salt should be about 3 grams plus or minus a small amount. In this case a measured quantity of 2.8 grams to 3.2 grams is sufficient. It is waste of time to weigh exactly 3.0, 3.00 or 3.00 grams when the instructions call for "about 3 grams".

On the other hand, it is sometimes necessary to measure an amount of substances precisely within the acceptable range. For instance, if the instructions direct you to weigh about 3 grams to the nearest 0.001 g, this does not mean the amount must be 3.000 grams. Rather, the amount should be between 2.8 grams and 3.2 grams and the measurement recorded to the nearest 3 decimal places. For example, 2.956 g would be acceptable. Always record every significant digit offered by your measuring device.

Procedure

Part 1: Temperature and Mass

Record all temperatures to the nearest 0.1°C.

1. Fill a large beaker (400 mL) halfway with tap water. Follow the procedure for temperature readings. Read and record the temperature of the water.

2. Fill a small beaker (150 mL) halfway with tap water. Set up a ring stand with a ring and wire gauze. Adjust the height of the ring so that the hottest part of the flame from the Bunsen burner reaches the bottom of the beaker. Heat the water to boiling. Read and record the temperature of the boiling water.

3. Fill a medium sized beaker (250 mL) about one quarter full with tap water and add about 100 milliliters of ice. Without stirring, place your thermometer on the bottom of the beaker. Read and record the temperature. Next, stir the mixture for about a minute. Hold the thermometer the proper way and record the temperature reading. Save the ice water for step 4.

4. Weigh an empty, clean, and dry 100 mL beaker (If you don't have a 100 mL beaker, anything close will suffice.) to the nearest 0.01 g (If you use the electronic balance you will be able to measure it to the nearest 0.001 g). Record the mass of the beaker. Add a small amount of NaCl to the beaker and re-weigh the beaker. Do this until there is approximately 6 grams of sodium chloride in the beaker. Transfer this NaCl to the ice water. Stir the mixture for about a minute. Add more ice if necessary. Read and record the temperature.

Part 2: Volume

Use the appropriate graduated cylinder for the following volume measurements to maximum precision, usually 0.1 milliliters. Be sure to read the volume at the meniscus.

1. Fill a test tube to the brim with tap water and measure the volume. Record your data.

2. Measure 5.0 mL of water in a graduated cylinder and pour it into a test tube. With a ruler, measure the height in centimeters and mark it with a grease pencil.

3. Measure 10.0 mL of water the same way as in the preceding step. Again, mark the height with a grease pencil.

4. You will find that this is a convenient way to estimate volumes of 5 milliliters and 10 milliliters by observing the height of liquids in test tubes.

Part 3: Drops in a mL

You will need you 10 mL graduated cylinder, and you should record your volumes to 0.01 mL.

1. Place approximately 5 mL of water into your 10 mL graduated cylinder. Record this volume to a precision of 0.01 mL.

2. Add water with a dropper, dropwise while counting the drops until there is approximately 8 mL of water in the graduated cylinder. Record the number of drops and the final volume of water in the graduated cylinder.

3. Calculate the number of drops per mL.

Name	
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Report

Part I: Temperature and Mass	
Temperature of room-temperature water (°C)	
Convert the temperature of room-temperature water to °F and K.	
Temperature of boiling water	
Temperature of ice water before stirring	
Temperature of ice water after stirring for 1 min.	
Mass of empty beaker	
Mass of beaker and NaCl	
Mass of NaCl (*This salt will now be added to the ice water to determine how it affects the temperature*).	
Temperature of ice water with salt added	
Show mass of salt calculations here.	

Part 2: Volume	
Volume of test-tube	
Height of 5 mL of water in test-tube	
Height of 10 mL of water in test-tube	
Part 3: Drops in a mL	
Initial volume of water in graduated cylinder	
Final volume of water in graduated cylinder	
Volume of water added to graduated cylinder	
Number of drops water added to graduated cylinder	
Drops/mL	
Show set-up and calculations for determining volume of water added to graduated cylinder and drops/mL here.	

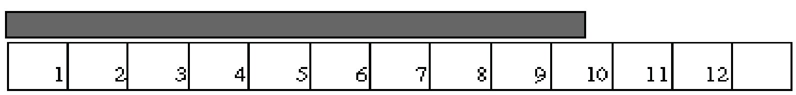
Problems

Name

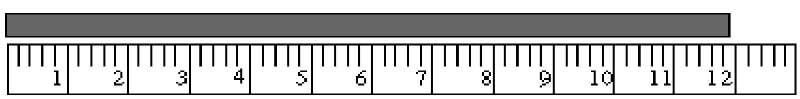
1. How many seconds in 8.0 years?

2. Calculate the number of mm in 2.0 miles.

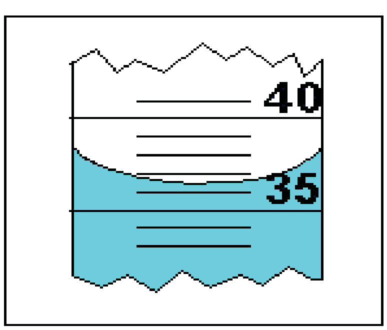
3. Record each of the following to the correct precision:

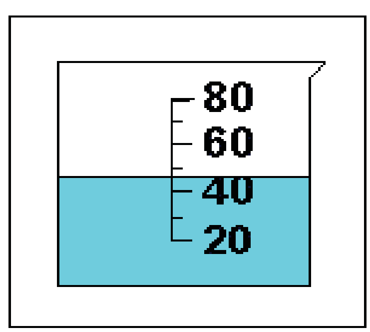


cm



cm





mL	mL

4. Poultry should not be stored above 40 ˚F. If your refrigerator is at 1.5 ˚C, will your chicken spoil?

5. A 50 mL graduated cylinder is filled with 23.0 mL of water. Water is then added to the cylinder dropwise until 36.0 mL of water are in the cylinder. If it took 265 drops to increase the volume in the cylinder to 36.0 mL, how many drops are in 1 mL?

6. A parsec is the distance that light travels in 3.26 years. How many kilometers does light travel in 1.0 parsec if light travels at a velocity of 3.00×10^8 m/s?

EXPERIMENT 3

DENSITY

Materials and Equipment

Balance, graduated cylinder, beaker, watch glass, small rubber stoppers, metal cylinder and unknown liquid.

Introduction

The density of a substance is a physical property that requires measurements of mass and volume. The ratio of the mass per unit volume is density, d and is written as d = m/V. Using the metric system, the ratio for solids and liquids is expressed in g/cm^3 or g/mL. Since the density of gases is much less than the densities of liquids and solids, the density of gases is usually expressed in grams per liter, g/L.

Matter expands and contracts according to temperature. For liquids and solids, the change in volume is very small, but for gases, a temperature change results in a proportional change in the volume. Therefore, gas density is dependent upon temperature; for liquids and solids, the small effect on volume is often ignored.

It is important to understand how density plays a role in everyday life. For example, wine makers measure densities to determine sugar content in the fermentation process and auto mechanics measure the density of the acid in car batteries to determine charge.

Procedure

All weighings and volumes should be read to the highest precision. This is usually 0.1 mL for volume, and 0.1 g to 0.001 g for mass. The precision to which you measure mass depends on the type of balance called for in the experiment. Follow all safety procedures throughout the experiment.

Part 1: Density of a Rubber Stopper

Obtain a small rubber stopper, it must be small enough to fit inside a 50 mL graduated cylinder. Weigh and record the mass of the dry stopper. Use tap water to fill your graduated cylinder to approximately 25 mL. Read and record this volume to the nearest 0.1 mL remembering to read the volume at the bottom of the meniscus. Carefully submerge the rubber stopper in the graduated cylinder. Read and record the new volume. What is the volume of the rubber stopper? Show your calculation. What is the density of the rubber stopper? Once again, show your calculation. Use the graduated cylinder containing the water and rubber stopper for Part 2.

Part 2: Density of a Metal Cylinder

Obtain a metal cylinder from your instructor and record the cylinder number on your data sheet. Read and record the mass of the metal cylinder. Record the initial volume of water and stopper in the graduated cylinder (it may be the final volume of water and stopper from Part 1). Slightly tilt the graduated cylinder containing the tap water and rubber stopper. Carefully slide and submerge your metal cylinder. The rubber stopper will act as a bumper and prevent the graduated cylinder from breaking. Read and record the new volume. Return the metal cylinder. Calculate the volume and density of the metal cylinder. Show your calculations.

Part 3: Density of an unknown liquid

Weigh and record the mass of a clean dry 150 mL beaker and watch-glass cover to the highest precision of the balance. Place about 25-35 mL of the unknown liquid in a 50 mL graduated cylinder. Read and record the volume to the nearest 0.1 milliliter. Carefully transfer as much liquid as possible to the clean, dry 150 mL beaker, cover with the watch glass and reweigh. Repeat the procedure with a different volume of the same liquid. Calculate the density of your unknown liquid for each trial. Be sure to show your calculations. Take an average of the two density determinations that you made.

Report

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Part I: Density of a Rubber Stopper	Trial 1	Trial 2
Stopper Type		
Mass of rubber stopper		
Initial volume of water in cylinder		
Final volume of water in cylinder		
Volume of rubber stopper		
Density of rubber stopper		
Average density of rubber stopper		
Show sample calculations here.		

Part 2: Density of a Metal Cylinder	Trial 1	Trial 2
Number stamped on metal cylinder		
Mass of metal cylinder		
Initial volume of water in graduated cylinder		
Final volume of water in graduated cylinder		
Volume of metal cylinder		
Density of metal cylinder (g/mL)		
Average density of metal cylinder (g/mL)		
Show sample calculations here.		

Name

Part 3: Density of Unknown Liquid	Trial 1	Trial 2
Mass of beaker and cover		
Mass of beaker cover and liquid		
Mass of liquid		
Volume of liquid		
Density of liquid		
Average density of liquid		
Show sample calculations here.		

Post Lab Questions

1. A stone is found to have a volume of 34.9 cm^3. It has a mass of 30.8 g. What is the density of the stone?

2. A graduated cylinder was filled to 30.0 mL with a liquid. An unknown metal cylinder having a mass of 60.27 g was submerged in the in the liquid. The final volume of the liquid level was 45.2 mL.

   a. Calculate the density of the unknown metal cylinder.

   b. The density of the liquid is 0.899 g/mL. What is the mass of the liquid in the graduated cylinder?

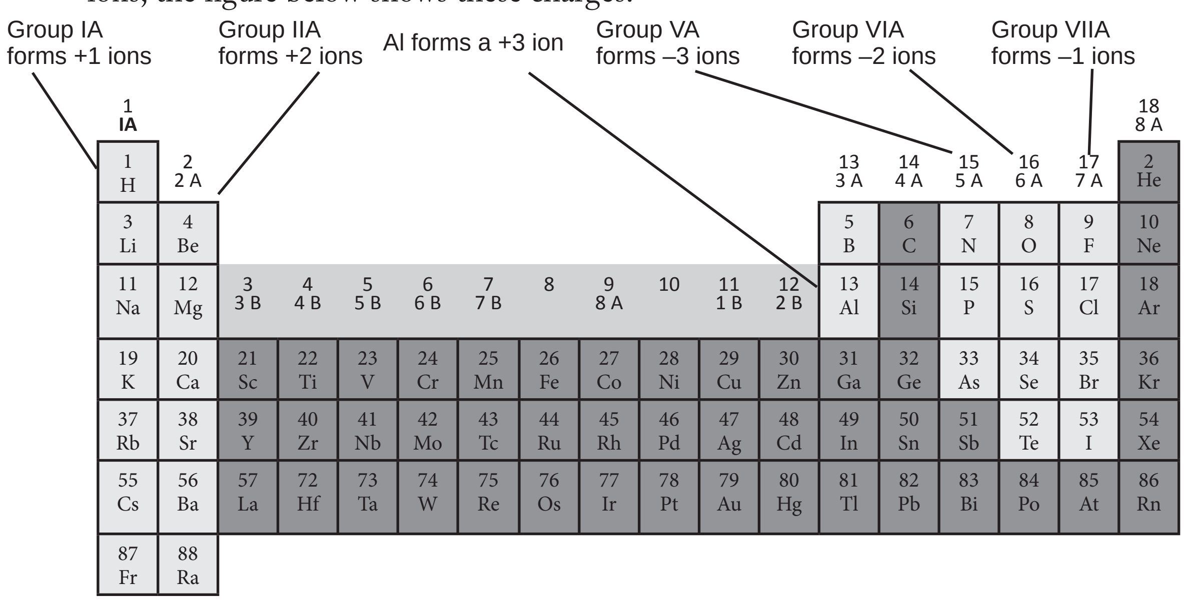
3. The density of gasoline is 0.6775 g/cm^3. Calculate the volume of 200.0 grams of gasoline.

EXPERIMENT 4

CHEMICAL NOMENCLATURE

Binary Ionic Compounds

All ionic compounds have a net charge of zero. This means that all the positive charge brought by the cation(s) must be exactly equaled by that contributed by the anion(s). This is a fundamental principle in ionic compounds. Take calcium bromide as an example. Two Br^- are required to balance the single Ca^{2+} ion, yielding the formula $CaBr_2$. The periodic table can help you remember the charges on certain monatomic ions; the figure below shows these charges.



In all ionic compounds the positive ion (cation) appears first in the formula, and the negative ion (anion) appears second. In writing the names of binary ionic compounds, the cation is named first, and the anion is named second with the suffix -ide added to the name. Look at the examples that follow:

Lithium nitride The two ions are Li^+ and N^{3-}. Three lithium cations are needed to balance the negative charge of one nitride ion. 3 × (+1) = +3, which just balances the nitride ion. The formula then is Li_3N

Magnesium fluoride The ions are Mg^{2+} and F^-. Two fluoride ions (2 × (–1) = –2) are required to balance the charge of one magnesium ion. The formula is MgF_2.

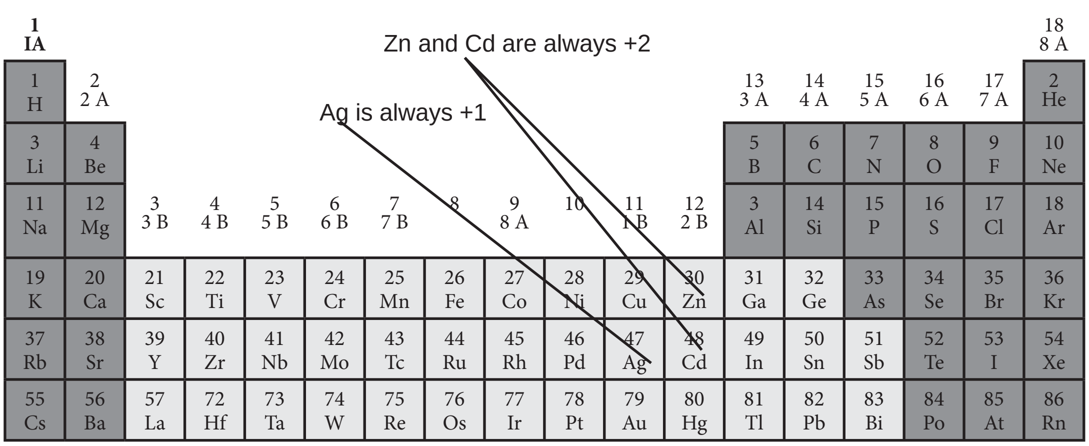
Aluminum oxide The ions are Al^{3+} and O^{2-}. The smallest number that both 2 and 3 will go into evenly is six. Therefore, two aluminum ions (2 × (+3) = +6) and three oxide ions (3 × (–2) = – 6) are the least number that can be combined to cancel out all of the charge. The formula is then Al_2O_3.

Transition Metals (Stock System)

Most transition metals do not always acquire the same charge upon the formation of an ionic bond. For instance, iron (Fe) sometimes forms the Fe^{2+} and other times forms Fe^{3+}. This would give the chlorides, $FeCl_2$ and $FeCl_3$, respectively. Remember, the charge on all compounds is zero, therefore the Fe^{2+} ion requires two –1 chloride ions and the Fe^{3+} ion requires three –1 chloride ions. Following the rules above, the name for each of these compounds would be iron chloride. Each of these compounds, however, has unique properties and there needs to be a method to differentiate these by name.

To show the difference between these compounds we include a roman numeral in the name to show the charge on the transition metal. Therefore $FeCl_2$ is called iron(II) chloride and $FeCl_3$ is called iron(III) chloride. The roman numeral DOES NOT refer to the number of anions, in this case chlorides. It refers to the charge on the transition metal cation. Iron(II) chloride means that the charge on the iron is +2; it does not mean that there are two chlorides.

Those elements shown below require a roman numeral when writing their name. Notice that Zn, Cd, and Ag are exceptions to this rule.



You must work "backwards" to determine the charge on the transition metal cation. Take the compound VS as an example. The name for this compound is not simply vanadium sulfide. Notice that vanadium is one of the elements that requires a roman numeral. This means that you need to determine the charge on the vanadium ion so that you can use the correct roman numeral. To do this look at the anion. This anion is the sulfide ion that has a –2 charge (If you don't know that sulfide has a –2 charge, then refer to "Binary Ionic Compounds"). Remember that all compounds have a charge of zero, so if this is true for this compound, then the charge on the vanadium must be +2. Therefore, the name of this compound is vanadium(II) sulfide. Notice that there is no space between the name of the cation and the roman numeral, and that the roman numeral is placed inside parentheses. Look at the examples below.

copper(II) oxide	CuO
nickel(IV) sulfide	NiS_2
chromium(III) fluoride	CrF_3

Ternary Ionic Compounds

When naming and writing formulae for compounds that include polyatomic ions (ternary ionic compounds) the rules for binary compounds apply to the polyatomic case. Treat the polyatomic ion as a single unit, and enclose it in parentheses to indicate more than one. Make sure the net charge is zero! You will need to memorize the ions and their charges found in the following table:

–1 Ions		**–2 Ions**		**–3 Ions**	
acetate	$C_2H_3O_2^-$	carbonate	CO_3^{2-}	phosphate	PO_4^{3-}
chlorate	ClO_3^-	chromate	CrO_4^{2-}	phosphite	PO_3^{3-}
chlorite	ClO_2^-	dichromate	$Cr_2O_7^{2-}$		
cyanide	CN^-	oxalate	$C_2O_4^{2-}$		
hydrogen carbonate (bicarbonate)	HCO_3^-	sulfate	SO_4^{2-}		
hydroxide	OH^-	sulfite	SO_3^{2-}		
hypochlorite	ClO^-				
nitrate	NO_3^-				
nitrite	NO_2^-				
perchlorate	ClO_4^-				
permanganate	MnO_4^-				

+1 Ion	
ammonium	NH_4^+

A general trend for polyatomic ions is as follows.:

Polyatomic ion	Number of oxygen atoms	Most Common	Example ion
per...ate	most oxygen atoms	4	ClO_4^- perchlorate ion
...ate	next most oxygen atoms	4/3	ClO_3^- chlorate ion
...ite	second least oxygen atoms	3/2	ClO_2^- chlorite ion
hypo...ite	least oxygen atoms	2/1	ClO^- hypochlorite ion

Below are a few examples.

$Ca(ClO_3)_2$

The cation name is unchanged from that of the element and so is calcium. Notice that the formula for the anion is enclosed in parentheses. This is necessary when the formula contains more than one of the polyatomic ion. The anion name, from the list of seven, is chlorate and so the compound name is calcium chlorate.

$KClO_2$

The cation, K^+ is of course potassium. The anion, ClO_2^-, corresponds to the "ite" ion, ClO_2^- except that it contains one less oxygen atom, therefore, it must be called chlorite. The compound's name then is potassium chlorite.

Zinc phosphate

The cation zinc is Zn^{2+} and the anion is PO_4^{3-}. The formula for the cation is determined directly from the Periodic Chart. In order to make a neutral compound, three zinc cations (total charge = +6) and two phosphate anions (total charge = –6) are needed. Again note the use of the parentheses around the polyatomic anion to show the presence of two anions. The completed formula then is then $Zn_3(PO_4)_2$

Binary Covalent Compounds

This method is restricted to compounds made up of two nonmetals. As in all binary compound names, the name of the first element is unchanged (except for the prefix) and the second element ends in "-ide." The number of atoms of each type is denoted by using the Greek prefix for that number. These nonmetal compounds are bonded together covalently; they do not have ions and so there are no cations and anions. Therefore, you will notice that on the practice sheets and on the computer exercises no provision is made to write out the ion formulas.

The prefixes and their corresponding numbers are listed in the table below. Know and be able to use them in naming and writing formulas of compounds. Examples follow.

mono	1	hexa	6
di	2	hepta	7
tri	3	octa	8
tetra	4	nona	9
penta	5	deca	10

N_2O_5 dinitrogen pentoxide

N_2O dinitrogen monoxide

SF_6 sulfur hexafluoride

CO carbon monoxide

Hydrates

Several ionic compounds contain what is called water of crystallization. When the substance is crytallized from water, some of the water is brought along into the crystal. In order to show that the water is part of the compound yet has some type of its own independent existence, the number of waters of crystallization are shown by suffixing "• n H_2O" to the formula for the compound. For example: $CuSO_4 \cdot 5\ H_2O$ called copper(II) sulfate pentahydrate. Notice that the ionic compound is name normally and the number of waters present is suffixed to this name by using a prefix for the number of waters followed by the word hydrate.

Acids

Compounds that contain hydrogen in the cation position act are acids when dissolved in water. Compounds such as C_2H_6 and $Ca(OH)_2$ contain hydrogen but they are not acidic. This can be seen in that the hydrogens do not occupy the cation position in the formula. There are two main classes of acids, (i) binary which contain hydrogen and one other element and (ii) ternary acids which contain hydrogen and two other elements. (Some acids contain more than three elements but they are named by the rules for the ternary acids). For binary acids, you can easily recognize the formula of the acid as it will have one or more hydrogens in the cation position and the "(aq)" immediately following the formula showing that the substances is dissolved in water. The ternary compounds are almost always named as acids whether or not there is an "(aq)" attached to the formula.

Formula	Compound Name	Acid Name
HCl *(aq)*	hydrogen chloride	hydrochloric acid
HBr *(aq)*	hydrogen bromide	hydrobromic acid
H_2S *(aq)*	hydrogen sulfide	hydrosulfuric acid
HCN *(aq)*	hydrogen cyanide	hydrocyanic acid

Binary Acids

Consider some examples. HCl, HBr and H_2S all have acidic properties when dissolved in water. In their pure state at standard temperature and pressure (STP) of 0 °C and 1 atm they are gases. The following table shows their names as both the compound and the acid when dissolved in water.

To derive the name for a binary acid from its compound name, drop the "-gen" from the hydrogen, change the "-ide" to "-ic" on the anion name, put these two pieces together to form one word and then add the word acid.

H_2S (aq) does not quite fit the pattern, however. The "sulfide" becomes "sulfuric" not sulfic as the rule suggests. The acid is then named hydrosulfuric acid.

Ternary Acids

Ternary acids are composed of a polyatomic oxyanion such as SO_4^{2-}, and hydrogen ions. They are named somewhat differently than the binary acids. The name of ternary acids is derived from the name of the oxyanion. If the name of the oxyanion ends in "-ate" the acid name is obtained by changing the ending of the anion from "-ate" to "-ic" and adding the word acid. Unlike the binary acid names, there is no "hydro" in the name at all. If the oxyanion ends in "-ite" then the ending is changed to "-ous" and the balance of the procedure remains the same. This process is shown in the following diagram for HNO_3, "hydrogen nitrate." It should be noted that many of these "compounds" do not exist in the pure state, but only as acids in aqueous solution.

Acid Formula	Oxyanion	Anion Name	Modified Anion Name	Acid Name
HNO_3	NO_3^-	Nitrate	nitric	nitric acid
H_2SO_3	SO_3^{2-}	sulfite	sulfurous	sulfurous acid
$HClO$	ClO^-	hypochlorite	hypochlorous	hypochlorous acid
HIO_4	IO_4^-	periodate	periodic	periodic acid

Note again that sulfuric acid does not follow the rules exactly as it did not for hydrosulfuric acid. This irregularity can soon be learned and should present no problem. As you can also see, the rules work just as well with the "per-" and "hypo-" compounds. Some of the acids such as $HClO_{(aq)}$ cannot be obtained in the pure state and exist only in aqueous solutions. Others such as $H_2SO_{4\,(l)}$, and $HC_2H_3O_{2\,(l)}$ are stable and exist in the pure state.

Remember, "-ate" ions make "-ic" acids and "-ite" ions make "-ous" acids.

Common Naming Conventions (Latin)

Note that some of the transition metals have Latin names for the cations. Most of these have fallen into disuse, but you may see them, especially for copper and iron. Some of these names are listed below. Use the Latin name in place of the stock name for the cation. For example, $FeBr_3$ can be named either iron(III) bromide or ferric bromide.

Ion Formula	Stock Name	Common(Latin) Name
Cr^{2+}	chromium(II)	chromous
Cr^{3+}	chromium(III)	chromic
Co^{2+}	cobalt(II)	cobaltous
Co^{3+}	cobalt(III)	cobaltic
Cu^{+}	copper(I)	cuprous
Cu^{2+}	copper(II)	cupric
Fe^{2+}	iron(II)	ferrous
Fe^{3+}	iron(III)	ferric
Pb^{2+}	lead(II)	plumbous
Pb^{4+}	lead(IV)	plumbic
Hg_2^{2+}	mercury(I)	mercurous
Hg^{2+}	mercury(II)	mercuric
Sn^{2+}	tin(II)	stannous
Sn^{4+}	tin(IV)	stannic

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Prelab Questions

Directions

Name the following monoatomic ions:

1. Na^{+} ________________
2. Mg^{2+} ________________
3. Ag^{+} ________________
4. P^{3-} ________________
5. Cu^{2+} ________ or ________
6. O^{2-} ________________
7. C^{4-} ________________
8. Fe^{3+} ________ or ________
9. Cl^{-} ________________
10. Hg_2^{2+} ________ or ________

Identify the following compounds as ionic compounds, covalent compounds, or acids:

1. NaCl ________________
2. CO_2 ________________
3. HBr *(aq)* ________________
4. FeN ________________
5. P_2O_5 ________________
6. hydroiodic acid ________________
7. calcium fluoride ________________
8. tetraiodine nonoxide ________________
9. manganese(III) carbide ________________
10. silver phosphide ________________

Directions

Name the following polyatomic ions:

1. NH_4^+ ________________________________
2. $C_2H_3O_2^-$ ________________________________
3. ClO^- ________________________________
4. OH^- ________________________________
5. CO_3^{2-} ________________________________
6. NO_3^- ________________________________
7. NO_2^- ________________________________
8. PO_4^{3-} ________________________________
9. ClO_4^- ________________________________
10. SO_4^{2-} ________________________________

Identify the following compounds as ionic compounds, covalent compounds, hydrate, or acids:

1. NH_4Cl ________________________________
2. $C_4H_8O_2$ ________________________________
3. $HBrO_3$ *(aq)* ________________________________
4. $Fe(NO_2)_2$ ________________________________
5. $Na_2SO_3 \cdot 7\ H_2O$ ________________________________
6. silver dichromate ________________________________
7. chromium(VI) hydroxide ________________________________
8. sulfurous acid ________________________________
9. sodium phosphite ________________________________
10. ammonium nitrate ________________________________

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Binary Nomenclature

Directions

This is an example of the types of compounds that you will need to be able to name. These examples are only binary compounds; the next worksheet will be ternary compounds!

1. Al_2O_3 ______________________
2. S_2O_3 ______________________
3. $CdBr_2$ ______________________
4. NiF_2 ______________________
5. HCl *(aq)* ______________________
6. N_2O_4 ______________________
7. SnF_2 ______________________
8. NO ______________________
9. CoS_3 ______________________
10. PbO ______________________
11. PBr_3 ______________________
12. $FeCl_3$ ______________________
    or Latin ______________________
13. ZnI_2 ______________________
14. Mn_3P_2 ______________________
15. $CuCl_2$ ______________________
    or Latin: ______________________
16. S_2F_{10} ______________________
17. HI *(aq)* ______________________
18. $CrBr_3$ ______________________

1. calcium oxide ______
2. chromium(III) oxide ______
3. iodine heptafluoride ______
4. sodium oxide ______
5. silver chloride ______
6. nickel(II) bromide ______
7. chromium(III) carbide ______
8. stannic sulfide or tin(IV) sulfide ______
9. cadmium sulfide ______
10. copper(I) phosphide ______
11. zinc fluoride ______
12. hydrofluoric acid ______
13. carbon tetrachloride ______
14. cupric bromide or copper(II) bromide ______
15. iron(III) chloride ______
16. carbon disulfide ______
17. aluminum fluoride ______
18. mercury(I) chloride ______
19. ferrous oxide or iron(II) oxide ______

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Ternary Nomenclature

Directions

This review represents examples of ternary ionic compounds. Where the formula is given, give the name,where the name is given give the formula.

1. K_2CO_3 ____________________
2. NH_4NO_2 ____________________
3. $Zn_3(PO_4)_2$ ____________________
4. $KClO_3$ ____________________
5. $CaSO_4$ ____________________
6. $Pb(NO_3)_2$ ____________________
7. $Cd(ClO)_2$ ____________________
8. $Sn(C_2H_3O_2)_2$ ____________________

   or Latin: ____________________
9. $HNO_{3\,(aq)}$ ____________________
10. $Co(ClO_2)_2$ ____________________
11. $NaHCO_3$ ____________________
12. $Ni_3(PO_4)_2$ ____________________
13. $CuSO_4$ ____________________

    or Latin: ____________________
14. Ag_2CO_3 ____________________
15. $HClO_{4\,(aq)}$ ____________________
16. $Fe(OH)_3$ ____________________

    or Latin: ____________________
17. Na_2SO_3 ____________________

1. silver nitrate ______________________
2. magnesium hydroxide ______________________
3. zinc sulfate ______________________
4. ammonium hydroxide ______________________
5. magnesium perchlorate ______________________
6. potassium permanganate ______________________
7. manganese(II) carbonate ______________________
8. sodium sulfate ______________________
9. cadmium dichromate ______________________
10. calcium cyanide ______________________
11. magnesium bicarbonate ______________________
12. lead(II) acetate ______________________
13. sulfuric acid ______________________
14. tetraphosphorus hex(a)oxide ______________________
15. calcium phosphite ______________________
16. ammonium chromate ______________________
17. silver phosphate ______________________
18. mercury(II) oxalate ______________________
19. ammonium carbonate ______________________

EXPERIMENT 5

ESTIMATING THE CALORIC CONTENT OF PEANUTS

Introduction

Natural gas is composed primarily of methane (CH_4). When we use natural gas to cook or heat our homes we burn it in air to produce the heat. We call this type of reaction of a compound with O_2 a **combustion reaction**. Energy is produced in the form of heat when methane reacts with O_2, to form carbon dioxide and water vapor.

$$CH_{4\,(g)} + 2\,O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\,H_2O_{(g)} + \text{heat}$$

A chemical reaction that produces energy is called an **exothermic** reaction, and a reaction that absorbs energy is called an **endothermic** reaction. Exothermic reactions produce heat and warm their surroundings; all combustion reactions are exothermic. There are many familiar examples of combustion reactions, including the burning of wood.

Endothermic reactions absorb heat, thus cooling their surroundings. "Cold Packs" that are used to treat some sports injuries are an example of an application of an endothermic reaction. Perhaps the most remarkable example of a combustion process is digestion.

Foods produce the same amount of energy, whether they are burned outside the body or broken down by the metabolic processes involved in digestion. Therefore, one method scientists use to determine the total energy content of food is to burn the food in a special container and measure the heat produced. This special container is called a calorimeter. Like a speedometer that measures speed, and a thermometer that measures temperature, a calorimeter measures heat (calories).

The S.I. unit in which energy is measured is called the joule (J). Often another unit is used when measuring the energy content of food; this unit is called the calorie (cal). There are 4.184 joules in 1 calorie. More often the kilocalorie (kcal) is used; remember, there are 1000 calories in 1 kilocalorie. One calorie is the energy required to raise the temperature of 1 g of water by 1 °C. It is important to note that there is a difference between calories as defined above and the calories recorded on food packaging. It is common to list the energy content in foods as dietary calories (Cal). Notice that this abbreviation uses a capital "C". 1 dietary calorie (1 Cal) is equal to 1 kilocalorie. That is, when a food label for a soft drink shows that one serving contains 280 Cal, it actually contains 280 kcal or 280,000 calories.

Not all of the energy in food is available to our body. On average, only 97% of ingested carbohydrates, 95% of fats, and 92% or proteins are broken down and absorbed by our intestines. The average net or physiologic energy values for the major energy-yielding nutrients are listed in Table 5.1. The caloric values listed on food labels reflect physiologic, not total, energy values. This value is also referred to as food energy. The most common method for determining the approximate available energy content (Cal) of a food is to determine the number of grams each of fat, protein, and carbohydrate in the food, and then multiply each mass by the appropriate physiologic energy value per gram.

Physiologic energy values

Nutrient	Cal/g
Fat	9.0
Protein	4.0
Carbohydrate	4.0

Caloric Content of Nuts

Peanuts are an excellent source of protein, energy, vitamins, and minerals, are low in salt and saturated fats (no cholesterol), and contain dietary fiber. One pound of peanut butter contains the same food energy as 1 gallon of milk, 32 eggs, or 2.5 lb of steak. The table below lists the general nutrient composition of peanuts.

The lipid portion of peanuts is commonly known as peanut oil, and is responsible for approximately 50% of the peanut's mass, and approximately 75% of the food energy provided by a peanut. We can indirectly obtain an estimate of the total food energy by directly measuring the energy released from the chemical reaction between the oil mol-

Nutrient	Amount in 100 g Edible Portion
water	7.0 g
protein	25.9 g
lipids	48.4 g
carbohydrates	16.4 g
dietary fiber	8.1 g
crude fiber	2.1 g
ash	2.3 g
Total Food Energy = 562 Cal	

ecules and O_2. This value is an estimate as not all energy is found in the oil. Since most of it is, however, it is a useful measure.

Burning Peanut Oil

Peanuts burn because the oil is readily combustible. The peanuts used in this experiment will light readily with a match. The average amount of energy produced from fat or oil is 9.4 Cal/g. We can estimate the amount of energy released by the complete combustion of the oil in one of these peanuts using the following equation:

$$\text{heat released, Cal} = (\text{mass of oil combusted, g})\left(\frac{\text{heat released, Cal}}{1\text{ g oil}}\right)$$

In this experiment you will estimate the amount of energy released by the combustion of the oil in one sample peanut. You will measure the energy by using it to heat a small sample of water. You can determine the amount of heat used to raise the temperature because the relationship between temperature change and energy input is known for water. This relationship is known as the **specific heat capacity.**

The specific heat capacity, c, is a measurement of the amount of heat, q, that must be absorbed by a substance to raise 1 g of the substance 1 ˚C. The larger the specific heat capacity, the more energy it takes to raise the temperature of a substance. The specific heat capacity of water is **1 cal/g ˚C** or 1 **Cal/kg ˚C**. Summarized by:

$$q = mc(T_f - T_i)$$

This means that the input of 1 calorie of heat will raise 1 g of water 1 ˚C. This also means that 1 Calorie (dietary) will raise 1 kg of water 1 ˚C. Notice the difference between the two. 1 Calorie (capital "C") will heat 1000 g of water 1 ˚C, whereas 1 calorie (lower case "c") will raise 1 g of water 1 ˚C. You can use the specific heat capacity of water to determine the amount of heat released when the peanut is burned. We will assume that all of the energy released by the peanut will enter the water in the calorimeter. Calculations are show on the next page.

Procedure

Constructing a Simple Calorimeter

Most of the calorimeter will be assembled for you. Refer to your instructor for proper placement and set-up.

Burning the Peanut Oil

1. Using your graduated cylinder carefully measure 200.0 mL of room temperature water into the can. Make sure that you record the amount of water used on your data sheet.

2. Place the lid on the calorimeter and insert the thermometer into the can. Record the temperature of the water in the calorimeter on you data sheet.

3. Select a whole peanut for your measurement, and record the brand, type of peanut, serving size, and calories (total calories not calories from fat).

4. Weigh the peanut on the balance and record this value on your data sheet. Make sure that you use weighing paper on the balance.

Remember to make sure the balanced is tared or zeroed (i.e. the balance reads 0.000 g or 0.0000 g).

5. Position the peanut on the support stand and place it under the can. Using a match, ignite the peanut.

This may take a while.

6. Immediately after the peanut stops burning, carefully stir the water in the can, and determine the temperature of the water. Be sure to record this temperature on your data sheet.

7. Allow the peanut residue to cool and weigh it on the balance. Record this value on your data sheet.

8. Remove the thermometer assembly from the can and pour the water down the drain. Clean the bottom of the can with a damp paper towel.

9. Repeat steps 1-8 using a second peanut of the same brand and type.

Calculations

1. Calculate the total energy released by the peanut. The following equation is used to determine how much energy was transferred to the water. The mass of water is determined using the density of water (1.00 g/mL).

$$\text{energy released from peanut, cal} = (\text{mass of water heated, g})\left(\frac{1.0\text{ cal}}{\text{g }^{o}C}\right)(\text{temperature increase, }^{o}C)$$

2. The following equation can be used to calculate the amount of energy released per gram of peanut burned.

$$\text{energy released, cal/g} = \left(\frac{\text{total heat released, cal}}{\text{mass of peanut combusted, g}}\right)$$

3. Use the package to determine the accepted amount of energy released in Cal/g. Use the following equation to determine the Cal/g of accepted energy content.

$$\text{accepted amount of energy released Cal/g} = \left(\frac{\text{number of calories per serving}}{\text{mass of serving, g}}\right)$$

4. Calculate the percent difference between your value and the value given on the label using the equation below.

$$\text{percent difference} = \left[\frac{\begin{pmatrix}\text{expt avg amount}\\ \text{of energy released}\end{pmatrix} - \begin{pmatrix}\text{accepted amount}\\ \text{of energy released}\end{pmatrix}}{(\text{accepted amount of energy released})}\right] \times (100\%)$$

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Experiment 5	Trial 1	Trial 2	Trial 3 (if necessary)
Brand and type of peanut			
Serving size, g			
Calories/serving			
Initial mass peanut & weighing paper			
Mass of weighing paper			
Initial mass of peanut			
Mass of peanut residue & weighing paper			
Mass of weighing paper			
Mass of peanut residue (left over material)			
Mass of peanut consumed by combustion			
Show calculation mass calculations here (inital mass of peanut and mass of peanut consumed).			
Final temp. of water			
Initial temp. of water			
Change in water temp			
Show calculation for temperature change here.			
Volume of water in calor-meter, mL			
Mass of water in calorim-eter, g			
Show calculation of mass of water here.			

	Trial 1	Trial 2	Trial 3 (if necessary)
Energy released from peanut, cal			
Show calculation here.			
Amount of energy released per g of peanut consumed by combustion, cal/g			
Show calculation here.			
Average amount of energy released per g of peanut burned, cal/g			
Show calculation here.			
Convert your result from above to dietary calories, Cal/g			
Show calculation here.			
Accepted amount of energy, Cal/g			
Show calculation here.			
Percent difference			
Show calculation here.			

Questions

1. What assumptions, approximations, or generalizations do we make about peanut oil in this experiment?

2. Briefly discuss the sources of experimental error inherent in this experiment. This is not necessarily the error introduced by careless measuring, but that error due to the particular technique.

3. Based on your data, how many grams of peanuts would need to be ingested in order to provide enough energy for a 150 lb person to golf for two hours (see table below)? Assume that your data from the burning of the peanut is a measurement of all of the energy available.

Activity	Cal/kg person min
Bicycling (racing)	0.127
Bicycling (slowly)	0.042
Cross-Country Skiing (level)	0.099
Driving a Car	0.015
Playing Golf	0.065
Standing Relaxed	0.008
Jogging (fast)	0.173

EXPERIMENT 6

PERCENT COMPOSITION OF POTASSIUM CHLORATE

Introduction

The percentage composition of a compound is the percentage by mass of each element in the compound. For instance, the compound water (H_2O) is composed of 11%, by mass, hydrogen and 89% oxygen. For every 100 grams of water there are 89 grams of oxygen and 11 grams of hydrogen. In order to calculate the percentage composition of a compound we need to know the mass of each of the component elements in a particular sample, or the empirical formula of the compound. In this experiment we will be determining the percent composition of $KClO_3$ by both methods. We will use the calculation derived from the known empirical formula to check our experimental results.

Calculating the Percent Composition from the Empirical Formula

If the formula of a compound is known, the percentage composition can be calculated from the molar mass and the total mass of each element in the compound. The molar mass of a compound is determined by adding up the atomic masses of all the atoms making up the formula. For example, sodium chlorate is composed of Na, Cl, and O. When using the empirical formula to calculate percent composition, it is useful to assume a convenient quantity of sample. Remember, the percent composition of a compound is a constant value; water will always have a percent composition of 89% oxygen and 11% hydrogen. If it did not, then it would not be water. For this reason, we assume that we have 1 mole of the compound. For our example, sodium chlorate ($NaClO_3$)

Element	Atomic Mass (g/mol)
1 Na	23.0
1 Cl	+ 35.5
3 O	3(16.0) = 48.0
Molar Mass	106.5

we know that there is 1 mole of sodium, 1 mole of chlorine, and 3 moles of oxygen in one mole of sodium chlorate.

Calculating % Composition for $NaClO_3$ from its Empirical Formula

1 mole of $NaClO_3$ contains 1 mole Na, 1 mole Cl, and 3 moles O. Calculate the mass of each component as shown below.

$$1 \text{ mol Na} \times \frac{23.0 \text{ g Na}}{1 \text{ mol Na}} = 23.0 \text{ g Na}$$

$$1 \text{ mol Cl} \times \frac{35.5 \text{ g Cl}}{1 \text{ mol Cl}} = 35.5 \text{ g Cl}$$

$$3 \text{ mol O} \times \frac{16.0 \text{ g O}}{1 \text{ mol O}} = 48.0 \text{ g O}$$

So in this one mole of sodium chlorate there are 23.0 g of Na, 35.5 g of Cl, and 48.0 g of O.

Next, to find the percent composition for each of these elements we divide their mass by the mass of one mole of the compound, sodium chlorate.

This calculation yields the theoretical percent composition.

$$\frac{23.0 \text{ g Na}}{106.5 \text{ g NaClO}_3} \times 100\% = 21.6\% \text{ Na}$$

$$\frac{35.5 \text{ g Cl}}{106.5 \text{ g NaClO}_3} \times 100\% = 33.3\% \text{ Cl}$$

$$\frac{48.0 \text{ g O}}{106.5 \text{ g NaClO}_3} \times 100\% = 45.1\% \text{ O}$$

Calculating the Percent Composition from Experimental Data

In order to calculate the percent composition of a compound using experimental data, we need to know the mass of a sample of the compound as well as the mass of the constituent elements. For our $NaClO_3$ example let us consider the following reaction.

$$2 \text{ NaClO}_{3\,(s)} \longrightarrow 2 \text{ NaCl}_{(s)} + 3 \text{ O}_{2\,(g)}$$

If a sample of sodium chlorate is heated, it decomposes into sodium chloride and oxygen. This allows us to calculate the quantity of oxygen in a sample of sodium chlorate.

Calculating% Composition form Experimental Data

A 1.04 g sample of $NaClO_3$ is heated and decomposed into NaCl and O_2 according to the reaction above. The mass of the sample after heating is 0.572 g. Calculate the percent composition of O in $NaClO_3$.

1st calculate the mass of oxygen lost.

$$\begin{array}{rl} 1.04 & \text{g } NaClO_4 \\ -\ 0.572 & \text{g residue} \\ \hline 0.47 & \text{g O} \end{array}$$

Now we know that of the original 1.04 g of $NaClO_3$, 0.468 g of that sample was oxygen. The 0.572 g of residue is most likely NaCl, which we can run further tests to confirm. We can calculate the percent oxygen composition as follows.:

$$\frac{0.47\text{ g O}}{1.04\text{ g sample}} \times 100\% = 45\%\text{ O}$$

This experiment will focus on obtaining the following data:

1. Mass of original sample ($KClO_3$).
2. Mass lost when the sample was decomposed.
3. Mass of the residue solid (KCl).

Using these data we can perform a similar calculation as outlined above in order to calculate the percent composition of potassium chlorate. It is important to point out that we will not break the percentage composition into each element. We will find the percent composition of oxygen and the percent composition due to potassium and chlorine combined. You will find the following equations helpful in other required calculations.

• ***Percentage KCl in sample:***

$$\text{Percent KCl} = \left(\frac{\text{Mass of Residue}}{\text{Mass of Original Sample}}\right) \times 100\%$$

• ***Theoretical value for percent age KCl in $KClO_3$:***

$$\text{Theoretical \% KCl in } KClO_3 = \left(\frac{\text{Molar Mass KCl}}{\text{Molar Mass } KClO_3}\right) \times 100\%$$

• ***Percentage error in experimental oxygen determination:***

$$\text{Percent error} = \left|\frac{\left[(\text{Experimental value}) - (\text{Theoretical value})\right]}{(\text{Theoretical value})}\right| \times 100\%$$

Procedure

CAUTION: Potassium chlorate is a strong oxidizing agent and may cause fires or explosions if mixed or heated with combustible materials. Observe the following safety precautions when working with potassium chlorate

1. Wear safety goggles.
2. Use clean crucibles.
3. Dispose of any excess or spilled potassium chlorate as directed by your instructor. Check before disposal.
4. Heat samples slowly and carefully to avoid spills and spattering.

NOTES

1. Make all weighings to the highest precision possible with the balance. Use the same balance for each weighing.
2. Make sure that you record your results AT THE BALANCE. Do not write on a separate sheet of paper only to transfer the data to your lab book later. This is one of the best ways to make errors.
3. After the initial heating it is best to handle your crucible with the crucible tongs.

Determining Percent Composition

1. Place a clean, dry crucible on a clay triangle and heat for 2-3 minutes at the maximum flame temperature.
2. Allow the crucible to cool, and weigh the crucible and cover.
3. Add between 1 and 1.5 grams of potassium chlorate, and weigh again.
4. Place the covered crucible on the clay triangle and heat gently for 8 minutes. This is best accomplished by moving the burner in and out from under the crucible.
5. Heat the crucible more vigorously for 10 minutes by placing the top of the bright, blue cone of the burner so that it touches the bottom of the crucible. The bottom of the crucible should be heated to a dull, red color.
6. Remove the crucible from the wire triangle using the crucible tongs. Let the crucible cool and weigh it.
7. After weighing, heat the sample for an additional 6 minutes at the maximum flame temperature.; Cool and reweigh.
8. The last two weighings should be in agreement. If they are not within 0.05 g of each other then repeat the heating and weighing until two agree within the specified tolerance (±0.05 g).

Qualitative Examination of Residue

This part of the experiment should be started as soon as the final heating and weighing of the first sample is completed.

1. Number and place three clean test tubes in a rack.

2. Place a pea-sized quantity of potassium chloride into tube no. 1, a pea-sized quantity of potassium chlorate in tube no. 2.

3. Add approximately 10 mL of distilled water to each of the test tubes.

4. Add 10 mL of distilled water to the crucible containing the KCl residue from your first sample.

5. Gently warm this sample for approximately 1 minute, and transfer 1 to 2 mL of the resulting solution from the crucible to test tube no. 3.

6. Add approximately 10 mL of distilled water to tube no. 3.

7. Test each of the tubes in the same manner: Add 5 drops of 6 M nitric acid (HNO_3), and 5 drops of 0.10 M silver nitrate solution ($AgNO_3$). Mix thoroughly. The formation of a white precipitate indicates the presence of the chloride ion. Your crucible residue should test positive for chloride ions as it is potassium chloride.

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Experiment 6

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Determining Percentage Composition

Make all calculations necessary to complete the following. Be sure to carry the proper number of significant figures in your computations. Be sure to show all steps and work in the space provided. Do Not Obliterate Errors.

	Sample #1	Sample #2	Sample #3 (if needed)
Mass of Crucible + Cover			
Mass of Crucible + Cover and Sample			
Mass of Crucible + Cover and Sample after First Heating			
Mass of Crucible + Cover and Sample after Second Heating			
Mass of Crucible + Cover and Sample after Third Heating (If Needed)			
Mass of Original Sample			
Mass Lost Upon Heating			
Mass of Residue			
Sample Calculations			

Experimental Composition of Potassium Chlorate

	Sample #1	Sample #2	Sample #3 (if needed)
Mass of Original Sample (From data table)			
Mass Lost Upon Heating (From data table)			
Mass of Residue (From data table)			
Percentage of Oxygen in $KClO_3$ **(Experimental Value)**			
Sample Calculations			
Percentage of K + Cl in $KClO_3$ **(Experimental Value)**			
Sample Calculations			
Percentage of Oxygen in $KClO_3$ from Formula **(Theoretical Value)**			
Sample Calculations			
Percentage of K + Cl in $KClO_3$ from Formula **(Theoretical Value)**			
Sample Calculations			

Name

Percentage Error in Experimental Percent Oxygen Determination

	Sample #1	Sample #2	Sample #3 (if needed)
Experimental Percentage of Oxygen in $KClO_3$ (From previous page.)			
Theoretical Percentage of Oxygen in $KClO_3$ (From previous page.)			
Percentage Error in Experimental Percent Oxygen Determination			
Sample Calculations			

Qualitative Examination of Residue

Record what you observed when silver nitrate was added to the following:

a. potassium chloride solution;

b. potassium chlorate solution;

c. residue solution.

1. What evidence did you observe that would lead you to believe that the residue that remained in crucible was potassium chloride?

2. Did the evidence obtained in the silver nitrate tests of the three solutions prove conclusively that the residue actually was potassium chloride? Explain.

Post-lab Questions & Problems

1. Calculate the percentage of Cl in $Ca(ClO_3)_2$.

2. Other metallic chlorates exhibit the same behavior as potassium chlorate when heated. Write the balanced chemical equation for the decomposition of calcium chlorate.

EXPERIMENT 7

PREPARATION AND PROPERTIES OF OXYGEN

Objective

To generate oxygen gas and investigate how oxygen influences a variety of chemical reactions.

Introduction

Oxygen is a colorless, odorless gas under ordinary conditions, and is found as a diatomic molecule in its natural or elemental state (O_2). Our atmosphere contains approximately 21% O_2. It is very slightly soluble in water (important for aquatic animals to get oxygen), but due to limited solubility we can collect oxygen gas under water as is done in this experiment.

There are many ways to generate oxygen gas including the decomposition of hydrogen peroxide, potassium chlorate, water, and mercury(II) oxide. **Decomposition reactions** involve one substance braking apart into two or more new substances. In 1774, Joseph Priestly (1733-1804) prepared and isolated oxygen gas by focusing the sun's rays on a sample of mercury(II) oxide. Chemists represent this reaction as follows. Notice that in the formula equation the coefficients are added to balance the equation. Chemical formula equations must show adherence to the law of conservation of mass.

WORD EQUATION Solid mercury(II) oxide decomposes to produce mercury metal and oxygen gas.

FORMULA EQUATION $2\ HgO\ (s) \longrightarrow O_2\ (g) + 2\ Hg\ (l)$

Combination reactions involve two substances reacting to form new substances. A special type of combination reaction is the **combustion reaction** in which oxygen gas reacts with another substance. Oxygen may combine with many other elements to form oxides.

An **oxide** is defined as a compound containing oxygen combined with one other element. If an element is burned in the presence of oxygen an oxide typically results. For example magnesium will burn in the presence of oxygen to form magnesium oxide.

WORD EQUATION Magnesium strips react with oxygen gas to form magnesium oxide powder.

FORMULA EQUATION $2\ Mg\ (s) + O_2\ (g) \longrightarrow 2\ MgO\ (s)$

Notice that in the formula equation the coefficients are added to balance the equation. Chemical formula equations must show adherence to the law of conservation of mass.

Experimental Procedure

Generation of oxygen gas

Hydrogen peroxide decomposes spontaneously to form oxygen gas and water.

WORD EQUATION Hydrogen peroxide decomposes to form water and oxygen gas.

FORMULA EQUATION $2\ H_2O_2\ (aq) \longrightarrow 2\ H_2O\ (l) + O_2\ (g)$

Hydrogen peroxide normally decomposes very slowly, but the catalyst manganese dioxide greatly increases the rate of this reaction. A **catalyst** is a substance which increases the rate of a chemical reaction.

Be sure to wear safety glasses for this experiment. If any hydrogen peroxide gets on your skin, wash it off with water. The hydrogen peroxide will sometimes temporarily bleach your skin.

1. Assemble the apparatus shown in Figure 7.1

Figure 7.1 Gas Collection Apparatus

2. Fill the trough up to the shelf with tap water. Fill six wide-mouthed bottles with water, cover them with a glass plate, and invert them into the trough of water. Remove the glass plates.

3. Add a pea-sized scoop of manganese dioxide to the Erlenmeyer flask. (If you do not add enough MnO_2 to the flask the reaction will proceed very slowly.) Stopper the flask and adjust enough distilled water to the flask through the thistle tube to cover the bottom of the tube. Be sure that the bottom of the thistle tube is covered with water and the glass connecting tube is above the level of the water.

4. Pour approximately 50 mL of 9% hydrogen peroxide into a small beaker.

5. When you are ready to begin generating oxygen gas, place one of the inverted glass bottles over the hole in the bottom of the trough and add 5 to 10 mL of hydrogen peroxide to the Erlenmeyer flask through the thistle tube. If the hydrogen peroxide does not drain into the flask it is due to pack pressure from the oxygen gas. Gently loosen the stopper to release pressure and reseal. Be careful not to spill the hydrogen peroxide at this point.

6. You should begin to see oxygen gas bubbling into the bottle at this point. If nothing happens, check the seals of rubber tubing for air leaks, check for ample MnO_2, and add more H_2O_2 if necessary.

7. When the bottle is filled with gas, cover it with the glass plate and remove it from the trough. Store the bottle (still covered with the glass plate) upright on the bench top. Because oxygen gas is more dense than air it will escape slowly.

Testing properties of Oxygen

Combustion Reactions

In this portion of the experiment you will observe the reaction (combustion) of a variety of materials in both a 20% oxygen atmosphere (room air) and a 90%+ oxygen atmosphere (in the bottles) to determine how the concentration of oxygen affects reactivity.

Combustion of carbon in wood. Light a wood splint and blow out the flame. Thrust the glowing wood splint into the bottle of oxygen. Record your observations. Repeat with a bottle of air. How will you get a bottle filled with room air? Is the bottle of oxygen you just finished satisfactory?

Combustion of sulfur. Go to the hood and fill a deflagration spoon with a small amount of sulfur. Using the burner in the hood, ignite the sulfur. Plunge the burning sulfur into a bottle of air. Record your observations. Repeat with a bottle of oxygen, covering the mouth of the bottle as much as possible with a glass plate. When the reaction is over, add distilled water, replaced the plate and shake. Test the water in the bottle with litmus paper, and record your observations. Red litmus paper turns blue in the presence of a base, but stays red in the presence of an acid. Blue litmus paper turns red in the presence of an acid, but stays blue in the presence of a base.

Combustion of iron. Fill one of the bottles with approximately 25 mL of water and replace the cover, being careful not to lose the oxygen. Take a small piece of steel, holding the steel wool with you crucible tongs, hold it in the flame of you Bunsen burner until it begins to glow. (This only takes a couple of seconds!) Immediately thrust the glowing steel wool into the bottle of oxygen and observe. When the reaction is complete replace the glass plate, shake, and test the water with litmus paper. Record your observations. Repeat with a bottle of air.

Combustion of a candle. In this next section you will compare the length of time a candle will burn in the room air compared to the nearly pure oxygen. Stand a small candle on a glass plate and light it. Cover the candle with one of the bottles of oxygen and record the number of seconds the candle burns in the oxygen. Record and repeat with a bottle of air.

Relative Densities

As a last experiment you will make some predictions regarding the relative densities of air and oxygen gas.

1. Set a bottle of oxygen on the bench top right side up (mouth up) and remove the glass plate. Wait 60 seconds and insert a glowing splint into the bottle. Do the same with a bottle of air. Record your observations. Does the oxygen escape?

2. Set a bottle of oxygen upside down on top of a bottle of air and remove the glass plate. Wait 60 seconds and cover both bottles with a glass plate. Insert a glowing splint into each bottle and compare their behavior. Where is the oxygen gas?

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Generation of Oxygen

1. Write the word and formula equation for the preparation of oxygen gas from hydrogen peroxide. For the word equation be sure to use complete sentences.

   WORD EQUATION

   FORMULA EQUATION

2. This is not a good method because it is so slow. What are some methods that can be used to speed up its decomposition?

3. What physical/chemical properties must oxygen gas possess to allow it to be collected by water displacement? (Remember that ALL gases are less dense than water.)

Properties of Oxygen

4. What is the symbol of the element oxygen? ____________

5. What is the formula for oxygen gas? ____________

Combustion of carbon in wood.

Observations	
Carbon in oxygen:	Carbon in air:

a. Write a word equation for the combustion of carbon in oxygen. The oxide of carbon that is formed is carbon dioxide. Use complete sentences.

b. Write a formula equation for the combustion of carbon in oxygen. The formula of carbon dioxide is CO_2.

Combustion of sulfur.

Observations	
Sulfur in oxygen:	Sulfur in air:

a. Write a word equation for the combustion of sulfur in oxygen. The oxide of sulfur that is formed is sulfur dioxide. Use complete sentences.

b. Write a formula equation for the combustion of sulfur in oxygen. The formula of sulfur dioxide is SO_2.

Name

c. What color did the “sulfur water” turn the red litmus paper? Blue litmus paper?

Combustion of iron.

Observations	
Iron in oxygen:	Iron in air:

Make sure that there is about 25 mL of water in the bottle.

a. Write a word equation for the combustion of iron in oxygen. Call the oxide of iron that is formed iron oxide. Use complete sentences.

b. Write a formula equation for the combustion of iron in oxygen. The formula of the iron oxide formed is Fe_3O_4.

c. What color did the “iron water” turn the red litmus paper? Blue litmus paper?

Combustion of a Candle.

Observations	
Time (seconds) candle burned in a bottle of oxygen.	Time (seconds) candle burned in a bottle of air.

a. Explain this difference in combustion time.

b. Is it scientifically sound to conclude that all the oxygen in the bottle was reacted when the candle stopped burning? Explain.

c. What is your conclusion about the rate or speed of a chemical reaction with respect to the concentration of the reactants? For example, a combustion in a high concentration of oxygen (pure oxygen) compared to a combustion in a low concentration of oxygen.

d. What evidence did you observe in the burning of sulfur to confirm your conclusion?

Oxygen Density.

a. What did you observe when you put the glowing splint into the bottle of oxygen which had been sitting open on the bench for 60 seconds?

b. How does this compare to a bottle of room air?

Name

c. Describe your observations for the experiment where you place a bottle of oxygen over a bottle of air and then put a glowing splint into each bottle.

Bottle on top:	Bottle on bottom:

d. What can you conclude regarding the relative densities of oxygen and air based on your observations?

Collection of Oxygen

1. What gas was in the apparatus before you started oxygen generation?

2. Where did it go?

3. What is different about the composition of the first bottle of gas collected compared to the other five?

Questions

1. Which of the following formulas represent oxides? (circle)
   MgO, $KClO_3$, SO_2, MnO_2, O_2, $NaOH$, PbO_2, O_3, Na_2O

2. Compare the litmus paper results of sulfur and iron. Suggest a reason why the results may have been the same or different.

3. If the same experiment were done with other elements you get the results in the table below. Are there some trends you observed from this data?

Element	Red Litmus Paper Color
N	red
Na	blue
K	blue
P	red
Cl	red
Mg	blue
C	red

EXPERIMENT 8

PREPARATION AND PROPERTIES OF HYDROGEN

Objective

To generate molecular hydrogen gas and investigate how hydrogen reacts in a variety of conditions.

Introduction

Hydrogen is found as a diatomic molecule in its natural or elemental state. This means that two hydrogen atoms are joined together to form a stable hydrogen molecule with the formula, H_2. Our atmosphere contains very little hydrogen gas because it has such a low density. Hydrogen is a very important component of many of the compounds we use everyday and composes about 0.9 % (by mass) of the earth's crust; in fact, it is the 9th most abundant element in the earth's crust.

Hydrogen, H_2, is a colorless odorless gas under ordinary conditions, and is very slightly soluble in water. Hydrogen is very reactive, and such, is generally found combined in nature. For example, when reacted with oxygen gas, O_2, it readily forms water.

WORD EQUATION	Hydrogen gas reacts with oxygen gas to form water vapor.
FORMULA EQUATION	$2\ H_2\ (g) + O_2\ (g) \longrightarrow 2\ H_2O\ (g)$

When hydrogen burns in the presence of oxygen it reacts explosively which results in a characteristic pop as the reaction occurs. Note that although hydrogen burns readily in the presence of oxygen, that is it is combustible, it does not support combustion.

Hydrogen is very important on the sun. There, it is involved in the nuclear reactions that generate energy.

Preparation of Hydrogen

Hydrogen may be prepared by the reaction of active metals with water or dilute acids. The alkali metals and some of the alkaline earth metals will react readily with water as shown in the following reactions:

WORD EQUATION — Cesium reacts with water to form cesium hydroxide and hydrogen gas.

FORMULA EQUATION — $2\ Cs\ (s) + 2\ H_2O\ (l) \longrightarrow 2\ CsOH\ (aq) + H_2\ (g)$

WORD EQUATION — Calcium reacts with water to form calcium hydroxide and hydrogen gas.

FORMULA EQUATION — $Ca\ (s) + 2\ H_2O\ (l) \longrightarrow Ca(OH)_2\ (aq) + H_2\ (g)$

Other active metals may require the presence of a dilute acid solution in order to react.

WORD EQUATION — Magnesium reacts with hydrochloric acid to form magnesium chloride and hydrogen gas.

FORMULA EQUATION — $Mg\ (s) + 2\ HCl\ (aq) \longrightarrow MgCl_2\ (aq) + H_2\ (g)$

WORD EQUATION — Zinc reacts with sulfuric acid to form zinc sulfate and hydrogen gas.

FORMULA EQUATION — $Zn\ (s) + H_2SO_4\ (aq) \longrightarrow ZnSO_4\ (aq) + H_2\ (g)$

Experimental Procedure

Preparation of Hydrogen Gas from Water + Sodium

1. Fill a test tube 1⁄2 full of water and place it behind the safety shield in the hood.

2. Sodium metal is stored under a non-reactive solvent such as hexane. Take a small piece of sodium (<4 mm/side) from the stock bottle and place it on a piece of filter paper, being careful not to touch it with your hands. Gently fold the filter paper over the sodium and blot it dry. As you dry the sodium, apply some pressure to it. Is it malleable?

3. Light a splint, then using tongs or tweezers, pick up the dry piece of sodium and drop it into the test tube filled with water. What happens?

4. Take a burning splint and bring it to the mouth of the test tube as the sodium is reacting. What happens?

5. Phenolphthalein is an acid/base indicator which is colorless in an acid solution (contains H^+ ions) and bright pink in a basic solution (contains OH^- ions). Take a few drops of phenolphthalein and place them into the test tube. What color is the resulting solution?

Preparation of Hydrogen from Hydrochloric Acid and Various Metals

In this portion of the experiment we wish to compare the relative activities of several different metals with respect to hydrochloric acid. The activity of a metal is defined as its propensity to react with the acid. In other words the more reactive, the more vigorous the reaction.

1. Set up four test tubes in a rack and label them zinc (Zn), copper (Cu), iron (Fe), and magnesium (Mg).

2. Obtain small samples of each of the above metals and place them into the appropriate test tubes. You will be using steel wool for the iron sample.

3. Quickly add enough dilute (6 M) HCl to just cover each of the samples and observe. Is gas evolved from the test tubes? How rapidly? Don't wait long before moving to the next step.

4. Immediately bring a burning splint to the mouth of each test tube. What happens? (Remember, a pop means that hydrogen is present.)

Preparation of Hydrogen from Zinc and Various Acids

In this portion of the experiment we wish to compare the relative reactivities of several different acids with respect to the same metal.

1. Set up four test tubes in a rack and label them hydrochloric acid, acetic acid, sulfuric acid, and phosphoric acid.

2. Dispense approximately 10 mL of each of the acids into the appropriate test tubes. You will be using dilute acids which means the hydrochloric acid and the acetic acid will both be 6 M and the sulfuric and phosphoric acids will both be 3 M.

3. Drop a small strip of zinc into each of the test tubes and observe. Is gas evolved from the test tubes? How rapidly?

4. Immediately bring a burning splint to the mouth of each test tube. What happens?

Collecting Hydrogen Gas to Examine its Properties

In this section of the experiment you will generate hydrogen gas by reacting zinc metal with dilute sulfuric acid. You will collect five bottles of hydrogen to use in the next section of the experiment.

1. Assemble the apparatus as shown in Figure 7.1, "Gas Collection Apparatus" on page 54.

2. Fill the trough up to the shelf with tap water. Fill five wide-mouthed bottles with water, cover them with a glass plate, and invert them into the trough of water. Remove the glass plates.

3. Add approximately 10 grams of mossy zinc to the Erlenmeyer flask.

4. Add enough distilled water to cover the bottom of the thistle tube when it is inserted into the flask. Use a minimum amount of water.

5. Seal the flask with the stopper and thistle tube.

6. Make sure the bottom of the thistle tube is covered with liquid. If not either lower the tube or add more water

7. Place one of the inverted glass bottles over the hole in the bottom of the trough and add about 50 mL of 6 M hydrochloric acid to the Erlenmeyer flask through the thistle tube. You should begin to see hydrogen gas bubbling into the bottle at this point. If nothing happens, check the seals of rubber tubing for air leaks, make sure there is plenty of zinc available to react and ask your instructor for some concentrated acid if necessary. When the bottle is filled with gas, cover it with the glass plate and remove it from the trough. Store the bottle (still covered with the glass plate) with the mouth down on the bench top.

   Note: If the hydrochloric acid does not drain into the flask it is due to back pressure from the hydrogen gas. Gently loosen the stopper to release pressure and reseal. Be careful not to spill the hydrochloric acid at this point.

Reactions of Hydrogen

1. Raise the first bottle of gas up from the bench top and bring a burning splint to the mouth of the bottle. What happens?

2. Raise the second bottle of gas up from the bench top and quickly insert a burning splint into the bottle and then slowly withdraw it. Be sure to do this slowly and observe the splint as you approach, when the splint is at the base of the bottle, and as it reaches the mouth of the bottle.

   Note: Unless otherwise specified, keep the bottles with the mouth down to prevent loss of hydrogen gas.

Relative Densities

As a last experiment you will make some predictions regarding the relative densities of air and oxygen gas.

1. Set a bottle of hydrogen on the bench top right side up (mouth up) and remove the glass plate. Wait 60 seconds and insert a burning splint into the bottle. Do the same with a bottle of air. Record your observations. Does the hydrogen escape?

2. Set a bottle of hydrogen upside down on top of a bottle of air and remove the glass plate. Wait 60 seconds and cover both bottles with a glass plate. Insert a burning splint into each bottle and compare their behavior. Where is the hydrogen gas?

3. Set a bottle of air upside down on top of a bottle of hydrogen and remove the glass plate. Wait 60 seconds and cover both bottles with a glass plate. Insert a burning splint into each bottle and compare their behavior. Where is the hydrogen gas?

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Preparation of Hydrogen Gas from Water + Sodium

1. What happened when you squeezed the sodium in the filter paper? Did it bend?

2. Describe you observations when you dropped the chunk of sodium into the water?

3. What happened when you brought a burning splint to the mouth of the test tube?

4. What color was the solution after you added phenolphthalein?

5. Was the solution acidic or basic? Explain.

6. Complete and balance the following equations

WORD EQUATION Sodium metal reacts water to form ____________________

____

FORMULA EQUATION ______ Na *(s)* + ______ H_2O *(l)* ⟶

7. Which of the products of the reaction above is responsible for the reaction with phenolphthalein?

Preparing Hydrogen Gas from Hydrochloric Acid and Various Metals

Fill in the data chart below.

Note: Be sure to use descriptive words, i.e. vigorous reaction or slow reaction, any color change, any temperature change, intensity of any pops heard.

Metal	Reaction w/ 6 M HCl	Reaction w/ Burning Splint
Zinc		
Copper		
Magnesium		
Iron		

Which is the most active metal with respect to reaction with hydrochloric acid?

Which is the least active metal with respect to reaction with hydrochloric acid?

Complete and balance the following equations

WORD EQUATION Zinc metal reacts with hydrochloric acid to form ________

____

FORMULA EQUATION ______ Zn *(s)* + ______ HCl *(aq)* $\longrightarrow$

Name

Preparing Hydrogen Gas from Zinc and Various Acids

Fill in the data chart below: (Remember your descriptive words.)

Acid	Reaction w/ Zinc	Reaction w/ Burning Splint
6 M HCl (hydrochloric acid)		
6 M $HC_2H_3O_2$ (acetic acid)		
3 M H_2SO_4 (sulfuric acid)		
3 M H_3PO_4 (phosphoric acid)		

1. Which is the most reactive acid with respect to reaction with zinc?

2. Which is the least reactive acid with respect to reaction with zinc?

3. Complete and balance the following equations

WORD EQUATION Zinc metal reacts with sulfuric acid to form ____________

____

FORMULA EQUATION ______ $Zn\ (s)$ + ______ $H_2SO_4\ (aq) \longrightarrow$

Reactions of Hydrogen

1. What happened when the burning splint was brought to the mouth of the first bottle?

2. Describe fully the results of testing the second bottle of gas.

3. Is hydrogen combustible? What is your evidence?

4. Does hydrogen support combustion? What is your evidence?

5. What compound was formed when you brought the burning splint to the mouth of the bottles of hydrogen?

Name

Hydrogen Density

Experiment 1. Bottle upright on bench for 60 seconds without a cover.

What happened when a burning splint was brought to the mouth of this bottle?

Experiment 2. Bottle of hydrogen on top of bottle of air.

Describe your observations of what happened when a burning splint was brought to the mouth of the top and bottom bottles?

Bottle on top:	Bottle on bottom:

Experiment 3. Bottle of air on top of bottle of hydrogen

Describe your observations of what happened when a burning splint was brought to the mouth of the top and bottom bottles?

Bottle on top:	Bottle on bottom:

What conclusions can you make regarding the relative densities of air and hydrogen from these experiments? Be sure to use your experimental data to justify your claims.

EXPERIMENT 9

OXIDATION REDUCTION REACTIONS

Introduction

Oxidation reduction or redox reactions are probably the most common and most important reactions observed in chemistry. All combustion reactions, as well as many important biochemical reactions are redox reactions. By definition, a redox reaction is a reaction in which electrons are transferred from one element to another. The element which gains electrons is said to be reduced and the element which loses electrons is said to be oxidized. In this experiment you will see two types of redox reaction, combustion reactions and single replacement reactions.

In **combustion reactions** an element or compound is reacted with oxygen gas and donates one or more electrons to the oxygen atom. An example of this type of reaction is shown below:

$$2\ Na\ _{(s)} + {}^{1}/_{2}\, O_2\ _{(g)} \longrightarrow Na_2O\ _{(s)}$$

In this reaction the sodium atoms give up electrons to the oxygen atoms. This can be represented by writing ionic equations showing the transfer of electrons.

$$2\ Na \longrightarrow 2\ Na^+ + 2\ e^-$$

[Ne] $3s^1$ [Ne] Note that sodium goes from a charge of zero to a charge of plus one. As the electron configurations show electrons are lost by sodium. This is **oxidation**.

$$O + 2\ e^- \longrightarrow O^{2-}$$

[He] $2s^2\ 2p^4$ [He] $2s^2\ 2p^6$ = [Ne] Note that oxygen goes from a charge of zero to a charge of minus two. As the electron configurations show electrons are gained by oxygen. This is **reduction**.

In **single replacement reactions** an electron is transferred between an atom in an element and an atom in a compound. The general reaction for this is represented by the chemical equation below:

$$A + BC \rightarrow B + AC$$

Element A is the more active element and replaces element B from the compound BC. If element B is more active than element A; no reaction will occur. Let us consider two specific examples using copper and chromium.

A strip of metallic copper is immersed in a solution of chromium(II) nitrate, $Cr(NO_3)_2$.

$$Cu\ (s) + Cr(NO_3)_2\ (aq) \rightarrow Cr\ (s) + Cu(NO_3)_2\ (aq)$$

Notice that the charge on the chromium changed from +2 to 0, and the charge on the copper changed from 0 to +2. This shows the transfer of electrons from the copper to the chromium. So, copper is oxidized and chromium is reduced.

If a metal of variable charge is active enough to replace the aqueous cation already in the solution, then it usually obtains a charge of +2 as a product in the aqueous phase.

A strip of metallic chromium is immersed in a solution of copper(II) nitrate, $Cu(NO_3)_2$.

$$Cr\ (s) + Cu(NO_3)_2\ (aq) \longrightarrow \text{No Reaction}$$

No change is observed even after the solution has been standing for a prolonged time, and we conclude that there is no reaction, and Cu is more active than Cr.

From this set of experiments it is clear that copper will lose its electrons to chromium, but chromium will not lose its electrons to copper. This suggests that copper will more easily lose electrons than chromium and thus it is considered to be more reactive or more active. Thus, copper is more active than chromium because it is more easily oxidized. In this experiment you will perform a series of single replacement reactions to determine the relative activity of a variety of elements.

Procedure

Part I – Instructor Demo

Describe the reaction which takes place when:

a. A piece of sodium metal is immersed into water.

b. A piece of potassium metal is immersed into water.

Part II – Combustion Reactions

Hold a small piece of copper foil over a Bunsen burner for at least 30 seconds and observe the change in the copper (note that copper becomes Cu^{2+}). Repeat with a small piece of magensium strip. Do not look directly at the glowing magnesium; it emits UV radiation, which can be damaging to your eyes.

Part III – Determination of an Activity Series

With a few of these reactions, the evidence for reaction is not immediately apparent. Before making a decision be sure to let the reaction mixture stand for approximately 10 minutes.

The evidence for reaction can include:

• The evolution of a gas (bubbles)

• The appearance of a metallic deposit on the surface of another metal. These metals are often black and do not often look as you would expect.

Assume that the metals of variable charge (copper & lead) become +2 ions in solution.

1. Obtain 3 pieces of zinc, 2 pieces of copper, and 1 piece of lead.
2. Clean the metal pieces with sandpaper to expose a fresh surface.
3. Place 6 test tubes in a rack and label each of them appropriately and add the following reactants.
    a. Tube 1: Copper strip and approximately 4 mL silver nitrate;
    b. Tube 2: Lead strip and about 4 mL copper(II) nitrate;
    c. Tube 3: Zinc strip and about 4 mL lead(II) nitrate;
    d. Tube 4: Zinc strip and about 4 mL magnesium sulfate;
    e. Tube 5: Copper strip and about 4 mL dilute (3 M) sulfuric acid;
    f. Tube 6: Zinc strip and about 4 mL dilute (3 M) sulfuric acid.
4. Observe the contents of each of the test tubes for evidence of reaction.

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Part I – Instructor Demo

Evidence of Reaction (Bubbles, Precipitate, Heat, Color Change, Etc.)	Equation	More Active Element	Less Active Element
1	$Na_{(s)} + H_2O_{(l)} \longrightarrow$		
2	$K_{(s)} + H_2O_{(l)} \longrightarrow$		

Most Active		Least Active

Which elements lost electron(s)?

Were these element(s) oxidized or reduced?

Which element(s) gained electrons?

Were these element(s) oxidized or reduced?

Part II – Combustion Reactions

Evidence of Reaction	Equation
	$Cu\,(s) + O_2\,(g) \longrightarrow$
	$Mg\,(s) + O_2\,(g) \longrightarrow$

Which elements lost electron(s)?

Were these element(s) oxidized or reduced?

Which element(s) gained electrons?

Were these element(s) oxidized or reduced?

Name

Part III - Determination of an Activity Series

Evidence of Reaction (Bubbles, Precipitate, Heat, Color Change, Etc.)	Equation	More Active Element	Less Active Element
1	$Cu_{(s)} + AgNO_{3(aq)} \longrightarrow$		
2	$Pb_{(s)} + Cu(NO_3)_{2(aq)} \longrightarrow$		
3	$Zn_{(s)} + Pb(NO_3)_{2(aq)} \longrightarrow$		
4	$Zn_{(s)} + MgSO_{4(aq)} \longrightarrow$		
5	$Cu_{(s)} + H_2SO_{4(aq)} \longrightarrow$		
6	$Zn_{(s)} + H_2SO_{4(aq)} \longrightarrow$		

Post-Lab Questions

1. Arrange Pb, Mg, and Zn in order of their reactivities, listing the most active first. Explain how you determined this order using experimental data.

Most Active		Least Active

2. Arrange Cu, Ag, and Zn in order of their activities, listing the most active first. Explain how you determined this order using experimental data.

Most Active		Least Active

3. Arrange Mg, H, and Ag in order of their activities, listing the most active first. Explain how you determined this order using experimental data.

Most Active		Least Active

Name

4. Arrange all five of the metals (excluding hydrogen) in an activity series, listing the most active first. Explain how you determined this order.

Most Active				Least Active

5. On the basis of the reactions observed in the six tubes, explain why the position of the hydrogen cannot be fixed exactly with respect to all of the other elements listed in the activity series in Question 4.

6. What reagents would you use to establish the exact position of hydrogen in the activity series of the elements listed in Question 4? (i.e. What additional test(s) would be needed?)

7. On the basis of the evidence developed in this experiment, predict the results of the following experiments.

    a. Would silver react with dilute sulfuric acid? Why or why not?

    b. Would magnesium react with dilute sulfuric acid? Why or why not?

    c. Would lead react with zinc nitrate? Why or why not?

EXPERIMENT 10

DOUBLE DISPLACEMENT REACTIONS

Introduction

Double displacement reactions are very common in aqueous environments and thus it is important to recognize them easily. Double displacement reactions generally occur between two ionic compounds which exchange their anions. Occasionally acids and bases will also be involved in these reactions. The chemical equation for a double displacement reaction has several common characteristics which are identified below:

$$AB + CD \longrightarrow AD + CB$$

A and C represent **cations** or atoms (or groups of atoms) which have lost electrons and as a result carry a positive charge.

B and D represent **anions** or atoms (or groups of atoms) which have gained electrons and as a result carry a negative charge.

In double displacement reactions, the new substances formed must be composed of a cation combined with an anion. If two cations or two anions came together, their charges would repel and no new compound would form. This means that the only way to make a new compound is to combine each cation with a new anion.

Before looking at any specific reactions it is important to identify the indicators of a reaction. The following examples will show a variety of double displacement reactions and how to recognize them. There are three common types of double displacement reactions and each produces a characteristic observation. These types are:

1. Precipitation Reactions - Reactions which produce an insoluble substance which will precipitate out of solution giving either a crystalline solid which settles to the bottom of the container or a suspended precipitate which will result in a cloudy solution.

2. Gas Formation Reactions - Reactions which produce a gas which may result in the production of bubbles in the solution. Sometimes the gas is quite soluble, but it may be detected by smell if it is an oderific substance. **Remember to waft!**

3. Exothermic Reactions - Reactions which produce heat. These reactions generally produce a slightly ionized substance which results in a slight warming of the solution. For dilute solutions this change in temperature may be difficult to detect.

Example #1 - Precipitation Reaction

When a solution of copper(I) acetate is mixed with a solution of lithium hydroxide a white precipitate forms and settles to the bottom of the test tube. Describe this reaction using a balanced chemical equation.

In this case the reactants are $CuC_2H_3O_2$ (composed of Cu^+ and $C_2H_3O_2^-$ ions) and LiOH (composed of Li^+ and OH^- ions). When a double displacement reaction occurs, the Li^+ ion will combine with the $C_2H_3O_2^-$ ion forming $LiC_2H_3O_2$ and the Cu^+ ion will combine with the OH^- ion forming CuOH. This is shown in the chemical equation below:

reactants — predicted products

$$\underbrace{Cu|C_2H_3O_2 + Li|OH}_{\text{reactants}} \longrightarrow \underbrace{LiC_2H_3O_2 + CuOH}_{\text{predicted products}}$$

Note that lines have been drawn separating the anions and cations in the reactants to facilitate their identification.

Now we must determine which of the products is insoluble in water and has formed a precipitate. To help with this look at the solubility rules at the top of the next page. Looking at the solubility rules, lithium containing compounds are soluble, so $LiC_2H_3O_2$ must be soluble, but hydroxides are generally insoluble suggesting that the CuOH forms a precipitate. We will rewrite the equation to show the states of all reactants and products as shown below. This is the complete balanced equation:

$$CuC_2H_3O_2\,(aq) + LiOH\,(aq) \longrightarrow LiC_2H_3O_2\,(aq) + CuOH\,(s)$$

State Labels (aka Phase Labels)

The state of a substance is identified using state labels.

The common state labels are listed below:

solid	*(s)*	
liquid	*(l)*	
gas	*(g)*	
aqueous	*(aq)*	aqueous means dissolved in water

Solubility Rules

1. All common compounds of alkali metals (Li^+, Na^+, K^+...) and ammonium (NH_4^+) ions are soluble (i.e. aqueous, *(aq)*).
2. All nitrates (NO_3^-), acetates ($C_2H_3O_2^-$), and chlorate (ClO_3^-) are soluble.
3. All binary compounds of the halogens (Cl^-, Br^-, I^-) (other than fluoride (F^-)) with metals are soluble, except those of silver (Ag^+), mercury(I) (Hg_2^{2+}), and lead(II) (Pb^{2+}) (Pb halides are soluble in hot water.).
4. All sulfates (SO_4^{2-}) are soluble, except those of barium (Ba^{2+}), strontium (Sr^{2+}), calcium (Ca^{2+}), lead(II) (Pb^{2+}), silver (Ag^+), and mercury(I) (Hg_2^{2+}). The latter three are slightly soluble.
5. Except for rule 1, carbonates (CO_3^{2-}), hydroxides (OH^-), oxides (O^{2-}), silicates (SiO_4^{2-}), and phosphates (PO_4^{3-}) are insoluble (i.e. solid, *(s)*).
6. Sulfides (S^{2-}) are insoluble except for sulfides of calcium (Ca^{2+}), barium (Ba^{2+}), strontium (Sr^{2+}), magnesium (Mg^{2+}), sodium (Na^+), potassium (K^+), and ammonium (NH_4^+).

Example #2 - Precipitation Reaction with unequal charges

When a solution of lead(II) nitrate is mixed with a solution of potassium chromate a bright yellow precipitate forms and settles to the bottom of the test tube. Describe this reaction using a balanced chemical equation.

In this case the reactants are $Pb(NO_3)_2$ (composed of Pb^{2+} and NO_3^- ions) and K_2CrO_4 (composed of K^+ and CrO_4^{2-} ions). When a double displacement reaction occurs, the Pb^{2+} ion will combine with the CrO_4^{2-} ion forming $PbCrO_4$ and the K^+ ion will combine with the NO_3^- ion forming KNO_3. This is shown in the chemical equation below:

$$Pb|(NO_3)_2 + K_2|CrO_4 \longrightarrow PbCrO_4 + 2\ KNO_3$$

Note that the ions in the products are combined in the smallest ratios that will give a neutral compound.

Looking to the solubility rules at the top of the page, the lead(II) chromate is insoluble, so the complete and balanced chemical equation for this process is:

$$Pb(NO_3)_2\,_{(aq)} + K_2CrO_4\,_{(aq)} \longrightarrow PbCrO_4\,_{(s)} + 2\ KNO_3\,_{(aq)}$$

Example #3 - Gas Formation Reactions

When a solution of potassium sulfite is mixed with a solution of hydrochloric acid a colorless solution results. Upon closer inspection, it is observed that small bubbles are beginning to form and an acrid odor is detected indicating the formation of a gas. Describe this reaction using a balanced chemical equation.

In this case the reactants are K_2SO_3 (composed of K^+ and SO_3^{2-} ions) and HCl (composed of H^+ and Cl^- ions). When a double displacement reaction occurs, the K^+ ion will combine with the Cl^- ion forming KCl and the H^+ ion will combine with the SO_3^{2-} ion forming H_2SO_3. This is shown in the chemical equation below:

$$2\,H|Cl \;+\; K_2|SO_3 \;\longrightarrow\; 2\,KCl \;+\; H_2SO_3$$

Note again that the ions in the products are combined in the smallest ratios that will give a neutral compound.

Looking at the list of compounds which decompose to form gases below, it is clear that the H_2SO_3 will decompose to form water and sulfur dioxide. This decomposition is reflected in the final balanced equation:

$$2\,HCl\,(aq) + K_2SO_3\,(aq) \longrightarrow 2\,KCl\,(aq) + H_2O\,(l) + SO_2\,(g)$$

Here H_2SO_3 has been replaced with H_2O and SO_2.

Slightly Ionized Substances		
Name	Formula	Decomposes to
water	H_2O *(l)*	
Acids		
acetic acid	$HC_2H_3O_2$ *(aq)*	
carbonic acid	H_2CO_3 *(aq)*	H_2O *(l)* & CO_2 *(g)*
hydrofluoric acid	HF *(aq)*	
oxalic acid	$H_2C_2O_4$ *(aq)*	
phosphoric acid	H_3PO_4 *(aq)*	
sulfurous acid	H_2SO_3 *(aq)*	H_2O *(l)* & SO_2 *(g)*
Bases		
hydroxide	NH_4OH *(aq)*	H_2O *(l)* & NH_3 *(g)*
methyl amine	CH_3NH_2 *(aq)*	

The reaction of an acid and a base is really just another form of a double replacement reaction. It is also called a **neutralization reaction**. All acid-base reactions result in the formation of a salt and water or another slightly ionized substance.

Example #4 - Reaction forming a slightly ionized substance (often acid/base)

When a solution of phosphoric acid is mixed with a solution of sodium hydroxide a colorless solution forms. Upon closer inspection it is observed that the test tube has warmed slightly. Describe this reaction using a balanced chemical equation.

In this case the reactants are H_3PO_4 (composed of H^+ and PO_4^{3-} ions) and NaOH (composed of Na^+ and OH^- ions). When a double displacement reaction occurs, the

H^+ ion will combine with the OH^- ion forming H_2O and the Na^+ ion will combine with the PO_4^{3-} ion forming Na_3PO_4. This is shown in the chemical equation below:

$$H_3|PO_4 + 3\ Na|OH \longrightarrow Na_3PO_4 + 3\ H_2O$$

Note again that the ions in the products are combined in the smallest ratios that will give a neutral compound.

An energy change indicates the formation of a molecular or slightly ionized substance. Looking to the table on the previous page, water is a slightly ionized substance which will generate heat when formed. Thus, the complete and balanced chemical equation for this process is:

$$H_3PO_4\,(aq) + 3\ NaOH\,(aq) \longrightarrow Na_3PO_4\,(aq) + 3\ H_2O\,(l)$$

Note that many of the slightly ionized substances will also decompose to form a gas.

Example #5 - No Reaction

When a solution of magnesium acetate is mixed with a solution of chromium(III) nitrate a colorless solution results. There is no indication of any gas formation or any temperature change. Describe this reaction using a balanced chemical equation. Here is a balanced chemical equation showing reactants and possible products:

$$3\ Mg|(C_2H_3O_2)_2 + 2\ Cr|(NO_3)_3 \longrightarrow 3\ Mg(NO_3)_2 + 2\ Cr(C_2H_3O_2)_3$$

There is no experimental evidence of any reaction taking place in this example. Looking to the solubility table neither of the products are insoluble so it makes sense that there was no precipitate. Looking to the list of slightly ionized substances and gas producing compounds, none of these are formed either. This leaves the overall balanced equation as:

$$3\ Mg(C_2H_3O_2)_2\,(aq) + 2\ Cr(NO_3)_3\,(aq) \longrightarrow 3\ Mg(NO_3)_2\,(aq) + 2\ Cr(C_2H_3O_2)_3\,(aq)$$

Upon closer inspection, if a substance is soluble it means that it will break apart into ions in solution. Looking at the ions which will be present in solution at the beginning and end of the reaction we see:

$$3Mg^{2+}\,(aq) + 6C_2H_3O_2^-\,(aq) + 2Cr^{3+}\,(aq) + 6NO_3^-\,(aq) \longrightarrow 3Mg^{2+}\,(aq) + 6NO_3^-\,(aq) + 2Cr^{3+}\,(aq) + 6C_2H_3O_2^-\,(aq)$$

There is clearly no difference between the reactants and products so this would be better written as:

$$Mg(C_2H_3O_2)_2\,(aq) + Cr(NO_3)_3\,(aq) \longrightarrow \text{No Reaction}$$

Procedure

In this experiment equal volumes of each reagent will be mixed together in a test tube. Approximately 3 mL of each solution should be transferred to the test tube and mixed well. There is no need to measure this quantity in a graduated cylinder. The important part is to use equal volumes of solution. To determine where the 3 mL mark is on you test tube, fill your 10 mL graduated cylinder with 3 mL of water. Pour this water into the test tube and make a mental note where the fluid level is. This will serve as your 3 mL mark.

Look for evidence of chemical reaction. This may be the formation of a precipitate, the formation of a gas, or the evolution of heat. Make sure that you give the reaction mixtures sufficient time to react (at least 10 minutes).

Formation of a precipitate — Look for the formation of an insoluble compound.

Formation of a gas — Look for the formation of NH_4OH (NH_3), H_2CO_3, or H_2SO_3. These compounds decompose into gases.

Formation of a slightly-ionized substance — Heat usually accompanies the formation of H_2O, $HC_2H_3O_2$, or any other slightly-ionized compound.

In each instance where a reaction occurred, write the complete, balanced reaction including phase labels. Where there is no evidence of reaction write "No reaction".

Test Tube	Solutions to mix:
1	0.1 M NaCl and 0.1 M $AgNO_3$
2	0.1 M Na_2CO_3 and dilute (6 M) HCl
3	0.1 M NaCl and 0.1 M KNO_3
4	10% NaOH and dilute (6 M) HCl
5	0.1 M $BaCl_2$ and dilute (3 M) H_2SO_4
6	dilute (6 M) NH_4OH and dilute (3 M) H_2SO_4
7	0.1 M $CuSO_4$ and 0.1 M $Zn(NO_3)_2$
8	0.1 M Na_2CO_3 and 0.1 M $CaCl_2$
9	0.1 M $CuSO_4$ and 0.1 M NH_4Cl
10	10% NaOH and dilute (6 M) HNO_3
11	0.1 M $FeCl_3$ and dilute (6 M) NH_4OH
12	**Do this part under the hood**. Add 1 g of solid Na_2SO_3 to 3 mL of water and shake to dissolve. Add approximately 1 mL of concentrated (12 M) HCl solution, dropwise.

Name
Date
Section

Report

Experiment 10

Test	Reactants	Predicted Products	Evidence of Reaction
1	$NaCl_{(aq)}$ & $AgNO_{3\,(aq)}$		
	Complete Balanced Equation (Include phase labels)		
2	$Na_2CO_{3\,(aq)}$ & $HCl_{(aq)}$		
	Complete Balanced Equation (Include phase labels)		
3	$NaCl_{(aq)}$ & $KNO_{3\,(aq)}$		
	Complete Balanced Equation (Include phase labels)		
4	$NaOH_{(aq)}$ & $HCl_{(aq)}$		
	Complete Balanced Equation (Include phase labels)		
5	$BaCl_{2\,(aq)}$ & $H_2SO_{4\,(aq)}$		
	Complete Balanced Equation (Include phase labels)		
6	$NH_4OH_{(aq)}$ & $H_2SO_{4\,(aq)}$		
	Complete Balanced Equation (Include phase labels)		

Test	Reactants	Predicted Products	Evidence of Reaction
7	$CuSO_4$ *(aq)* & $Zn(NO_3)_2$ *(aq)*		
	Complete Balanced Equation (Include phase labels)		
8	Na_2CO_3 *(aq)* & $CaCl_2$ *(aq)*		
	Complete Balanced Equation (Include phase labels)		
9	$CuSO_4$ *(aq)* & NH_4Cl *(aq)*		
	Complete Balanced Equation (Include phase labels)		
10	$NaOH$ *(aq)* & HNO_3 *(aq)*		
	Complete Balanced Equation (Include phase labels)		
11	$FeCl_3$ *(aq)* & NH_4OH *(aq)*		
	Complete Balanced Equation (Include phase labels)		
12	Na_2SO_3 *(s)* & HCl *(aq)*		
	Complete Balanced Equation (Include phase labels)		

Name

Questions & Problems

1. Two aqueous solutions are mixed and the test tube begins to feel warm. What type of reaction has occurred?

2. Two aqueous solutions are mixed and small blue particles begin to float in the solution. What type of reaction has occurred?

3. Write the equation for the decomposition of sulfurous acid.

4. A tank of water is contaminated with mercury(II) ions. What reagent could be added to precipitate the mercury? Write a complete balanced equation to describe the reaction which would take place.

5. Using the three criteria for double displacement reactions and the solubility table, predict whether a double displacement reaction will occur for each of the following. If the reaction will occur, write the complete balanced reaction and include the proper phase labels. If no reaction will occur, write "no reaction".

$CaSO_4\,(aq)$ & $K_2CO_3\,(aq) \longrightarrow$
$VCl_3\,(aq)$ & $LiOH\,(aq) \longrightarrow$
$Na_2CO_3\,(aq)$ & $HNO_3\,(aq) \longrightarrow$
$Ca(ClO_3)_2\,(aq)$ & $NaBr\,(aq) \longrightarrow$
$NH_4OH\,(aq)$ & $H_2C_2O_4\,(aq) \longrightarrow$
$K_3PO_4\,(aq)$ & $MnCl_2\,(aq) \longrightarrow$
$KOH\,(aq)$ & $NH_4Cl\,(aq) \longrightarrow$
$SnCl_4\,(aq)$ & $Ti(NO_3)_3\,(aq) \longrightarrow$

EXPERIMENT 11

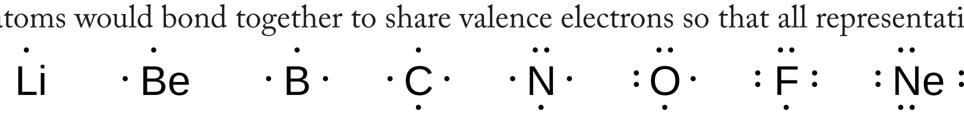
Molecular Models

Introduction

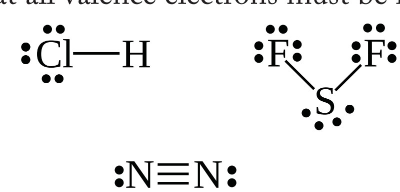
Models are often used by the chemist to visualize molecular structures. Molecular geometries can frequently influence chemical and physical properties, thus it is important to begin to recognize the way that atoms bond together and how they orient themselves in a molecule. In this experiment, you will build a variety of molecules using molecular models in order to become familiar with some of the more common geometries.

Lewis Electron Dot Structures

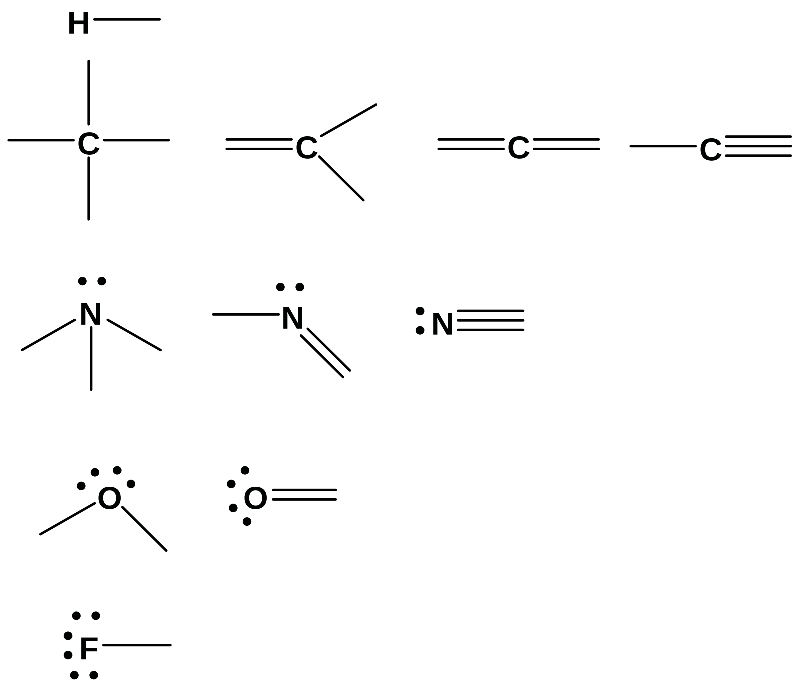
Bonding is based on the **valence electrons** in an atom. Valence electrons are found in the outermost shell of an atom. Often chemists will draw atoms surrounded by dots to represent the valence electrons. In 1916, G. N. Lewis developed a theory in which he proposed that atoms would bond together to share valence electrons so that all representative atoms



(with the exception of H and He) would be surrounded by 8 valence electrons. Chemists still use that model when drawing Lewis Electron Dot structures for molecules. Your textbook outlines the procedure for drawing correct electron dot structures for common compounds. A few examples of correctly drawn Lewis structures are shown in the figure to the right. Here lines or dashes represent two electron bonds and the dots represent non-bonding electrons or lone pairs. Remember that all valence electrons must be represented in a correct Lewis Structure!



To draw the Lewis Electron Dot structure for a molecule you must first determine the number of bonding electrons in the molecule. These electrons will then be distributed in such a way as to give all atoms a full outer shell, that is two electrons for hydrogen and helium and eight electrons for all other atoms. Common bonding patterns:

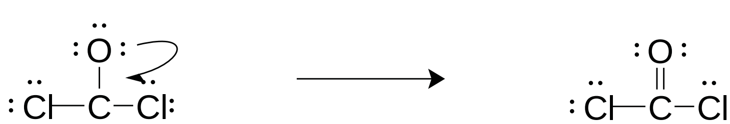


There are several ways to determine the correct electron dot structure. Two methods are as follows:

Method #1

1. Count the number of valence electrons available.
2. Using two electrons per bond, attach the atoms together using one bond for each connection. Generally the least electronegative atom will be the central atom.
3. Place the remaining electrons around the outer atoms to give them a complete octet (or duet for H). To give the central atoms an octet, move electrons from the outer atoms to form double and triple bonds.

$COCl_2$

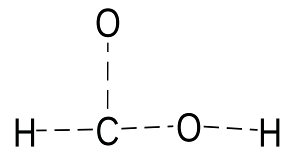


There are 24 valence electrons. Six electrons are used to attach the atoms and the remainder are put around outer atoms. Carbon does not have a full outer shell so 2 electrons from O are used to form a second bond giving C a full octet.

Method #2

1. Draw a skeleton structure, showing connectivity of atoms.

H_2CO_2



Skeleton Structure

2. Count the number of valence electrons available.

3. Count the number of electrons required to give all electrons a full outer shell. (This will be 2 electrons per H and 8 electrons for all other atoms.)

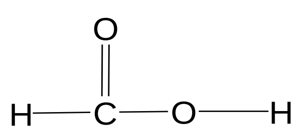
Atom	Electrons Needed	Electrons Available
C	1 C x 8 e^-/C = 8 e^-	1 C x 4 e^-/C = 4 e^-
2 O	2 O x 8 e^-/O = 16 $^{e-}$	2 O x 6 e^-/O = 12 e^-
2 H	2 H x 2 e^-/H = 4 e^-	2 H x 1 e^-/H = 2 e^-
Total	28 e^-	18 e^-

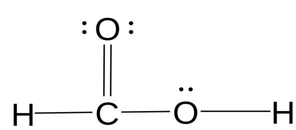
4. Subtract to determine how many electrons are needed to give all atoms a full outer shell (# electrons short). This number determines the number of bonds needed. One bond is needed for every 2 electrons short.

28 electrons needed
- 18 electrons available
----
10 electrons short

5 bonds needed

5. Put in the bonds needed, and arrange the remaining electrons to give all atoms a full outer shell.

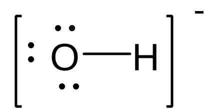




Five bonds are required. There are 10 electrons used to make the 5 bonds, leaving 8 electrons to be distributed as lone pairs. These electrons will go on the oxygen atoms to give them a complete octet.

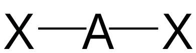
Ions

If you need to draw the Lewis Electron Dot structures for ions, add or subtract valence electrons based upon the charge on the molecule. For example OH^- would have 6 valence electrons from the oxygen, 1 valence electron from the hydrogen, and 1 extra valence electron due to the charge, giving it a total of 8 valence electrons:

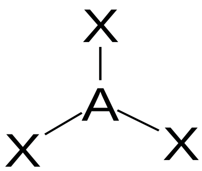


Molecular Geometries

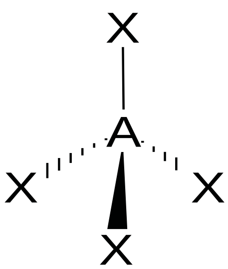
Valence **S**hell **E**lectron-**P**air **R**epulsion theory or **VSEPR** is used to predict molecular geometries. VSEPR theory proposes that the structure of a molecule is determined by the repulsive interaction of electron pairs in the valence shell of its central atom. In other words, the bonding pairs and the non-bonding (lone) pairs around a given atom are as far apart as possible. In order to determine the geometry you must first determine the number of electron clouds. Any bond (single, double, or triple) and any lone pair represents an electron cloud.



If an atom has two electron clouds surrounding it the geometry will be linear.



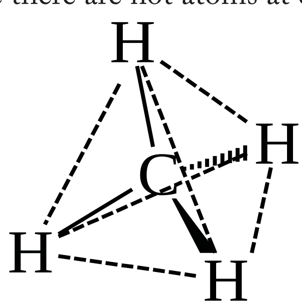
If an atom has three electron clouds the geometry will be trigonal planar.

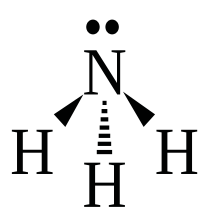


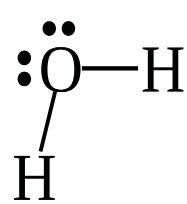
If an atom has four electron clouds, the geometry will be tetrahedral.

For example, methane CH_4, has four bonding electron pairs around the carbon atom. In order to maximize the distance between these bonding electrons (and minimize the repulsive forces) the hydrogen atoms will orient themselves forming a tetrahedron with 109.5° bond angles.

If pairs of non-bonding, lone pairs also surround an atom they will also require space. These lone pairs will distort the geometries of the molecule giving new geometries. An example of this is shown in both ammonia and water where the angles between the four sets of electron pairs are still approximately 109° but the molecules are not tetrahedral because there are not atoms at each bonding position.





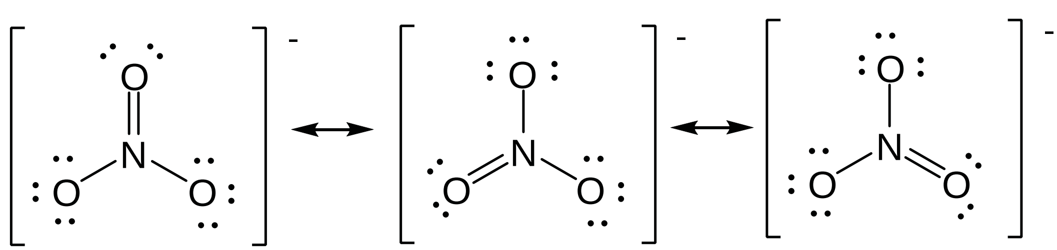


Ammonia becomes a trigonal pyramidal molecule and water is bend as seen in the figure. All of the common geometries are summarized in the table below:

Valence shell electron pairs	Bonding Electron clouds	Lone pairs	General Formula	Orbital geometry	Approximate bond angles	Molecular geometry
2	1	No central atom	AX	Linear	180°	Linear
2	2	0	AX_2	Linear	180°	Linear
3	3	0	AX_3	Trigonal Planar	120°	Trigonal Planar
3	2	1	$\ddot{A}X_2$	Trigonal Planar	<120°	Bent
4	4	0	AX_4	Tetrahedral	109.5°	Tetrahedral
4	3	1	$\ddot{A}X_3$	Tetrahedral	<109.5°	Trigonal Pyramidal
4	2	2	$:\ddot{A}X_2$	Tetrahedral	<109.5°	Bent

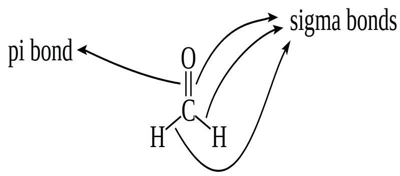
Resonance

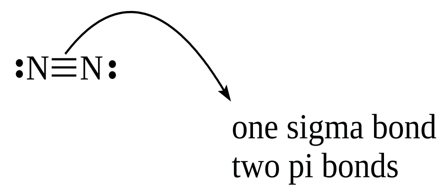
Another consideration is resonance. Molecules that have resonance structures have a double or triple bond that can exist in multiple positions without changing the geometry of the molecule. In other words, electrons are delocalized, they can exist in more than one location. The NO_3^- ion has three resonance structures as illustrated below: Note that ions structures are enclosed in brackets with the ion charge outside the brackets. Resonance structures are identified by the double headed arrows between the likely structures.



Sigma and Pi Bonds

Note that there is a difference between single, double, and triple bonds. Because all of the electrons cannot exist in the same place, each of the electron pairs must have a region in space to occupy. When a bond forms, the first pair of electrons occupies the region between the atoms and is called a sigma bond. As additional bonds are formed they begin to occupy the regions above and below the atoms and these bonds are called pi bonds. To recapitulate, the first bond to form between any two atoms is a sigma bond, and any additional bonds to form are pi bonds.





Molecular Polarity

The polarity of a molecule can be determined by looking at the shape of the molecule and the bonds within the molecule. A polar molecule occurs when there is a net dipole moment on the molecule as determined by electronegativity differences between atoms. A nonpolar molecule will not have a net dipole moment.

A molecule can usually be classified as polar if: there is a lone pair (or lone pairs) on the central atom, or there are different types of atoms (or groups of atoms) on the central atom.

Procedure

Build and draw the molecules and ions listed in the tables on the following pages. Be sure to answer the questions regarding these molecules. Remember, we're counting lone pairs and bonding electron electrons around the central atom.

Report

Molecular Models

Name	
Date	
Section	

Fill in the sheet below for the molecules you build.

Molecule	NH_3	CH_3Cl	CH_2O
Total # Valence Electrons			
Lewis Structure			
Number of Bonding Electrons Clouds			
Pairs of Nonbonding Electron Pairs on Central Atom			
Electronic or Orbital Geometry on Specified Atom	N	C	C
Molecular Geometry (Name)			
3-Dimensional Sketch			
Bond Angle			
Polarity of Molecule			

Molecule	H_2S	C_3H_8	CH_3COOH
Total # Valence Electrons			
Lewis Structure			
Number of Bonding Electrons Clouds		C	C (in CH_3) C (bonded to O)
Pairs of Nonbonding Electron Pairs on central atom.		C	C (in CH_3) C (bonded to O)
Electronic or Orbital Geometry on specified atom	S	C	C (in CH_3) C (bonded to O)
Molecular Geometry (Name)		C	C (in CH_3) C (bonded to O)
3-Dimensional Sketch			
Bond Angle			H-C-H C-C-O
Polarity of Molecule			

Name

Molecule	CO_2	C_2H_2	H_2O
Total # Valence Electrons			
Lewis Structure			
Number of Bonding Electrons Clouds			
Pairs of Nonbonding Electron Pairs on central atom			
Electronic or Orbital Geometry on specified atom	C	C	O
Molecular Geometry (Name)			
3-Dimensional Sketch			
Bond Angle		H-C-C	
Polarity of Molecule			

Molecule	NH_4^+	OH^-	CO_3^{2-}
Total # Valence Electrons			
Lewis Structure			Show all resonance structures.
Number of Bonding Electrons Clouds			
Pairs of Nonbonding Electron Pairs on central atom			
Electronic or Orbital Geometry of specified atom	N	O	C
Molecular Geometry (Name)			
3-Dimensional Sketch			
Bond Angle			

Name

Post Lab Questions

1. How many valence electrons does an atom of nitrogen contain?

2. How many valence electrons are in a molecule of $C_4H_6Br_2$?

3. How many valence electrons are in a $CH_3CO_2^-$ ion?

4. How many valence electrons are in a $NH_3C_2H_5^+$ ion?

5. What is the orbital or electronic geometry of a molecule with 2 nonbonding electron pairs and 2 bonding electron pairs?

6. What is the orbital or electronic geometry of a molecule with 0 nonbonding electron pairs and 2 bonding electron pairs?

7. What is the molecular geometry of a molecule with 1 nonbonding electron pair and 2 bonding electron pairs?

8. Why are the valence electrons of an atom the only electrons likely to be involved in bonding to other atoms?

9. Why do representative elements tend to form bonds giving them a total of 8 valence electrons?

10. How is the structure around a given atom related to repulsion between valence electron pairs on the atom involved?

EXPERIMENT 12

TITRATING A VINEGAR SOLUTION

Introduction

Titration is one of the methods used to determine the concentration of an unknown sample. The procedure involves a chemical reaction between two substances, usually an acid and a base. One of the substances is a standard solution, that is, the concentration of the standard solution is known with a high degree of precision. To the mixture is added a substance that changes color when the chemical reaction is complete. This substance is called an **indicator**. The indicator used in this experiment is phenolphthalein. This indicator is colorless in acid solutions and pink in basic solutions. When the indicator changes color, the end point of the titration has been reached.

Titration is defined as the measurement of a volume of standard solution that is needed to completely react with a known volume (or mass) of unknown sample. The unknown sample is often referred to as the **analyte**. Sodium hydroxide is the **standard solution** in this experiment. Vinegar, the analyte, is an aqueous solution of acetic acid whose concentration is unknown. The goal in this experiment is to determine the concentration of vinegar. Let us begin with a balanced chemical equation for the reaction:

$$CH_3COOH\ \textit{(aq)} + NaOH\ \textit{(aq)} \longrightarrow CH_3COO^- Na^+\ \textit{(aq)} + H_2O\ \textit{(l)}$$

Chemical formulas are often abbreviated; acetic acid (also written as $HC_2H_3O_2$) is abbreviated as AcOH (or HOAc) and acetate ion ($C_2H_3O_2^-$ or CH_3COO^-) is written as AcO^- (or OAc^-). Thus we can write the acid base reaction as:

$$AcOH\ \textit{(aq)} + NaOH\ \textit{(aq)} \longrightarrow NaOAc\ \textit{(aq)} + HOH\ \textit{(l)}$$

This reaction represents a neutralization reaction since an acid and base will react (will neutralize each other) to produce water and a salt. Notice in the reaction of acetic acid and NaOH that the stoichiometry between acid and base is 1:1. *The reaction stoichiometry is always part of the calculation in a titration*.

We start the titration by carefully measuring the volume of our analyte (vinegar) using a 25.00 mL volumetric pipet. The sample of vinegar is added to a clean Erlenmeyer flask and D.I. water is added to dilute the sample. Phenolphthalein is added to this solution and now we are ready to titrate the vinegar (acid) with base.

The base is added using a buret. The buret allows us to carefully measure the amount of base needed to reach the end point. The end point means that the neutralization reaction is complete; in other words, at the end point, the number of moles of acid and the number of moles of base (in the reaction) flask is equal. The end point is obvious since the solution turns pink. Once the **end point** is reached, we must perform some calculations to determine the concentration of our vinegar solution (the analyte).

The concentration of vinegar is obtained using the familiar equation for molarity:

Molarity = mole / liter

In order to calculate the concentration of vinegar, we need the volume and the number of moles in the 25.00 mL sample that we measured at the beginning of the experiment. Of course we have the volume of vinegar since we used the volumetric pipet to transfer the analyte. The number of moles of vinegar is calculated from the standard solution (base). Remember that since the stoichiometry is 1:1, the number of moles of vinegar (acid) is equal to the number of moles of base at the end point.

We obtain the number of moles of base by rearranging the molarity equation:

mole base = (Molarity base) × (liter base)

Since moles of base equal moles of acid at the end point:

mole base = mole acid

the concentration of acid can be calculated:

Molarity acid = mole acid / liter acid

Procedure

Your instructor will demonstrate the proper use of volumetric glassware. The main considerations are:

- glassware must be clean;
- glassware must be rinsed with the solution that will be measured;
- reading the volume requires an eye-level determination of the meniscus;
- buret volumes must be carefully controlled.

Practice transferring some water samples with the pipet until you are confident that you can do this step properly. Use the steps below for the actual titration once you have mastered the use of the pipet and buret. You can save much time if you first do a practice sample: pipet 25.00 mL of vinegar into a flask, dilute with D.I. water, add the indicator and titrate quickly. Open the stopcock of the buret and continue to add the base until the pink color persists. Most likely, you will have gone past the end point but you will have a rough estimation of the volume of base needed to titrate 25 mL of your unknown vinegar sample. For your actual samples, you can quickly add about 90% of this volume then carefully titrate drop-by drop to the end point.

1. From your instructor, obtain your unknown solution with your name on it. Record the name on the bottle of your vinegar sample on your data sheet.

2. Transfer 50 mL of your sample to a clean and dry beaker. You will pipet out of this beaker so as not to contaminate the original sample container. Rinse pipet with about 5 mL of vinegar two times. Using the pipet, transfer 25.00 mL of vinegar into a clean 250 mL Erlenmeyer flask. Label this as sample #1. Record this volume (25.00 mL) on your data sheet.

3. Dilute your 1st sample with about 40 mL of deionized water, then add 3 drops of phenolphthalein indicator to the sample. Swirl to mix contents. The sample is ready to titrate.

4. Now prepare the buret to deliver the base, NaOH. From your instructor, obtain approximately 150 mL of the standard NaOH solution in a clean, dry 250 mL beaker. Record the molarity of the base on your data sheet. Rinse the buret two times with small portions of the base. Be sure to wet the entire length of the buret and drain through the stopcock to ensure that all surfaces have been rinsed.

5. Fill the buret with base above the zero mark; open the stopcock and drain the solution until the bottom of the meniscus is at or just below the 0 mL mark of the buret. (It is not necessary to start at exactly zero mL!). Read the buret correctly and record this initial buret volume on your data sheet.

6. Place your 1st vinegar sample under the buret. You are now ready to titrate. Open the stopcock and add the NaOH to the vinegar, about 1-3 mL at a time. To mix the contents, swirl the flask gently between additions; do not splash any solution out of the flask. You will notice a pink color appear momentarily as the NaOH first encounters the acid (vinegar) in the flask. The pink color quickly fades as the base reacts with the acid. When you are close to the endpoint, the pink color will persist for longer periods of time. When you get close to the endpoint, add small amounts of base from the buret so that you do not go past the endpoint! When the addition of ½ of a drop results in the pink color persisting for 30 seconds, you have reached the endpoint. Stop the titration and record the final buret volume on your data sheet.

7. Refill the buret, pipet another 25.00 mL of vinegar into your flask (sample #2) and titrate sample #2 as before. If the volume of base used to reach the end point varies significantly between the two samples, then you must do a third trial until these volumes are within 0.05 mL of each other. In other words, you must have a minimum

of two trials in which the volume of base agrees to within 0.05 mL.

8. When finished with all trials rinse the buret and pipet with distilled water. Fill the buret with distilled water and replace in the container. Replace the pipet in the conatainer for the next class as well.

Sample Calculation

Using a 25.00 mL volumetric pipet, a sample of vinegar was placed into a clean Erlenmeyer flask to which was added 20 mL of D.I. water and 3 drops of phenolphthalein.

Using a buret, the solution was titrated to the end point; 34.45 mL of base was needed to reach the end point. The concentration of the standard solution (base) is 0.4589 *M*.

First, calculate the number of moles of NaOH added

$$\text{moles base} = \text{molarity of base} \times \text{liters of base}$$

$$\text{moles base} = 0.4589\ \text{mol/L} \times 0.03445\ \text{L}$$

$$\text{moles base} = 0.01581\ \text{mol}$$

Now, determine the number of moles of acid at the endpoint.

The balanced chemical equation for the reaction is

$$CH_3COOH\ (aq) + NaOH\ (aq) \longrightarrow CH_3COO^- Na^+\ (aq) + H_2O\ (l)$$

Notice that the stoichiometric ratio is 1:1, therefore the number of moles of acid at the endpoint is the same as the number of moles of base that was added.

$$\text{moles of base} = \text{moles of acid} = 0.01581\ \text{moles}$$

Third, we calculate the molarity of the acetic acid in the vinegar solution.

$$\text{molarity of acid} = \frac{\text{moles of acid at the endpoint}}{\text{L of acid}}$$

$$\text{molarity of acid} = \frac{0.01581\ \text{mol acid}}{0.02500\ \text{L}}$$

$$\text{molarity of acid} = 0.6324\ \text{mol/L}$$

Note: Remember that the volume of base is measured with the buret and the volume of acid is measured with the volumetric pipet. The molarity equation is defined as moles per liter; therefore, the buret volume and pipet volume must be converted from milliliters into liters.

Report

Experiment 12

Name	
Date	
Section	

Determining Vinegar Concentration

	Trial 1	Trial 2	Trial 3	Trial 4 (if needed)
Volume of analyte				
Molarity of standard NaOH solution				
Initial buret reading				
Final buret reading				
(1) Volume of titrant (base)				
(2) Number of moles of NaOH added				
(3) Number of mol acid				
(4) Molarity of vinegar				
(5) Average molarity of vinegar solution				

Show your calculations for each numbered row.

(1) Volume of titrant	(2) Number of moles NaOH added	(3) Number of moles of acid
(4) Molarity of vinegar solution	(5) Average Molarity of Vinegar Solution	

Post Lab Questions

1. Provide brief definitions for the following:

standard solution	indicator
analyte	end point

2. Consider the following and briefly explain what would be the effect on the calculated analyte concentration; i.e., why would the calculated analyte concentration be higher or lower (or unchanged) than the actual value if:

   a. the last drop of base was not rinsed from the buret into the reaction flask?

   b. the reaction flask contained traces of water in it before the acid was added with the volumetric pipet?

   c. the volumetric pipet used to transfer the acid solution contained traces of water?

3. Calculate the molarity of 25.00 mL of HCl if 15.78 mL of 0.5025 M NaOH is needed to reach the end point.

Name

EXPERIMENT 13

Molecular Models of Organic Molecules

Objectives

1. To become familiar with the three-dimensional aspects of molecular structure.

2. To examine the relationship between molecular models and structural formulas.

Introduction

Much of the chemical and biological behavior of molecules is determined by their molecular shape. Organic molecules are three-dimensional. Many people have difficulty visualizing two-dimensional diagrams on paper as three dimensional. The following experiments are designed to help the student of chemistry overcome these problems.

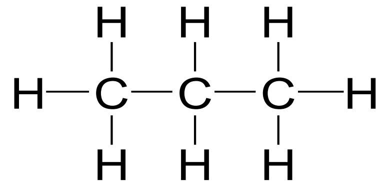
Molecular models are used by practicing chemists at all levels. From the shape of the molecules the chemist can often see important molecular features that might otherwise escape his or her attention. For example, propane (C_3H_8) looks rather ordinary as a structural formula; but, in three dimensions one can easily see a rich structure even in such a simple molecule. Although the three dimensional structure conveys a lot of information, it is difficult to draw. Chemists must rely on simpler ways to show structures, and structural formulas are often all that are used. We need to fully appreciate the structural details that are understood in the structural formulas. This laboratory exercise is designed to better acquaint you with the relationship of the two dimensional drawings with the three dimensional world of real molecules.

We use several types of drawings to illustrate the structure of molecules. They are as follows.

Chemical formula. A description of the number and types of atoms present in a molecule.

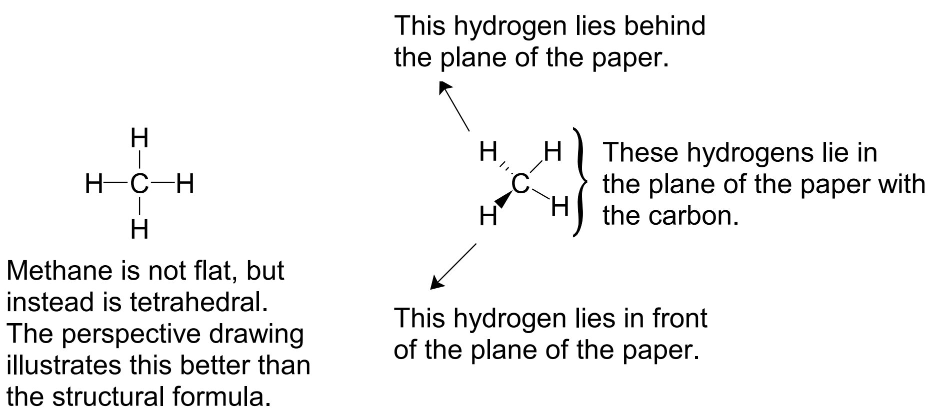
For example propane is C_3H_8.

Structural formula. A representation of a structure that emphasizes the bond connection between atoms.



Condensed formula. A condensed representation of a chemical structure that leaves out the vertical bonds and shows the most of the structure on one line. $CH_3CH_2CH_3$

Perspective formula (or 3-D formula). A representation of a chemical structure that conveys the three dimensional nature of a compound. Solid wedges represent bonds sticking out of the plane of the paper, dashed wedges represent bonds behind the plane of the paper, and straight lines represent bonds in the plane of the paper.



In this laboratory exercise we will build a number of different molecules and draw a variety of different types of formulas for them.

Alkanes

Alkanes are compounds containing only carbon and hydrogen atoms joined by single bonds. These alkanes are named according to the number of carbons in the molecule. A normal alkane (n-alkane) has all of the carbon atoms joined together in a straight line. Below is a list of the alkanes and their names.

Name

Name	Chemical Formula	Condensed Formula
methane	CH_4	CH_4
ethane	C_2H_6	CH_3CH_3
propane	C_3H_8	$CH_3CH_2CH_3$
butane	C_4H_{10}	$CH_3(CH_2)_2CH_3$
pentane	C_5H_{12}	$CH_3(CH_2)_3CH_3$
hexane	C_6H_{14}	$CH_3(CH_2)_4CH_3$
heptane	C_7H_{16}	$CH_3(CH_2)_5CH_3$
octane	C_8H_{18}	$CH_3(CH_2)_6CH_3$
nonane	C_9H_{20}	$CH_3(CH_2)_7CH_3$
decane	$C_{10}H_{22}$	$CH_3(CH_2)_8CH_3$

Alkyl Groups

The group of atoms that results when a hydrogen atom is removed from an alkane is called an alkyl group. The group is named by replacing the -ane suffix of the parent hydrocarbon with -yl. It is important to note that alkyl groups are not independent molecules. Rather, they exist as parts of molecules. Next we will look at the kinds of alkyl groups that can be made from some alkanes.

Make methane and ethane. How many ways can you remove one hydrogen from each of these molecules to make an alkyl group? Draw condensed formulas for the two alkyl groups that are formed. When you draw the alkyl groups a dash us used to show where the hydrogen was removed.

methane	methyl group $—CH_3$
ethane CH_3CH_3	ethyl group

Make propane. How many different ways can you remove a hydrogen from propane? ___

There are two possible propyl groups, they are normal propyl (n-propyl) and isopropyl (i-propyl). Look in your book to identify these two isomers and draw them.

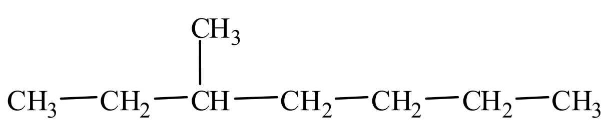
propane	n-propyl group
	isopropyl group

Nomenclature

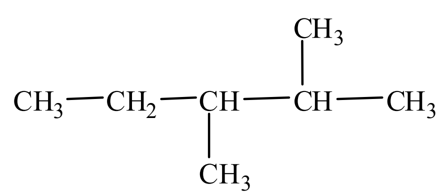
Organic compounds, even the most complex structures, are formally named from simpler pieces. The concept of the alkyl group is very useful in naming organic compounds.

Detailed rules are found in your text; however, they can be summarized as:

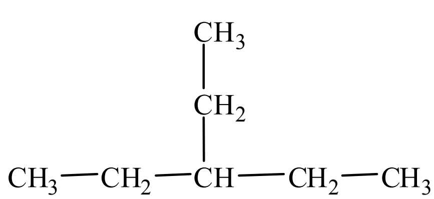
1. Find the longest continuous chain as your parent hydrocarbon.
2. Identify (with a position number and name) all of the groups attached to the main chain.
3. If two or more names seem possible, the correct name will be that which uses the longest chain for the parent hydrocarbon and whose alkyl groups total the lowest possible position numbers.



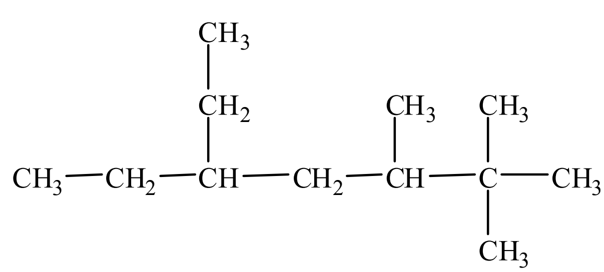
3-methylheptane



2,3-dimethylpentane



3-ethylpentane



5-ethyl-2,2,3-trimethylheptane

Name

Isomers

Isomers are compounds having the same molecular composition (numbers of atoms) but different arrangements of these atoms. Isomers are really different compounds. They have different solubility's, melting points, boiling points, etc. With your model of butane rearrange the atoms to make the other isomer of butane. Complete the table below:

Name	Chemical Formula	Structural Formula	Condensed Formula
Butane or n-butane			
Isobutane or 2-methylpropane			

Drawing Isomers

Make a model of n-hexane, C_6H_{14}, using the molecular models and write its condensed structural formula in the table below. Now rearrange the carbon atoms to make different isomers of hexane. Hexane has a total of 5 different isomers. Draw condensed structural formulas and name all 5 isomers.

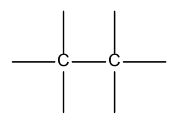
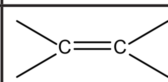
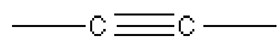
Substituted Alkanes

Make a model of butane. Remove one hydrogen and substitute a chlorine atom (green) to make C_4H_9Cl. (**All isomers will have a straight chain of carbons.**) Draw condensed formulas for each of these monochlorobutanes below:

How many different isomers of chlorobutane can you make? __________

Unsaturated Molecules

Alkenes and alkynes are hydrocarbons that have more than a single bond between carbon

Hydrocarbon Type	Condensed Formula	Structural Formula	Suffix
alkane	R_3CCR_3		-ane
alkene	R_2CCR_2		-ene
alkyne	RCCR		-yne

More nomenclature

1. For alkenes (alkynes) the parent hydrocarbon is the longest carbon chain containing the double (triple) bond.

2. The position of the double bond is given by the smaller of the two numbers containing the carbon atoms of the double bond.

$CH_2{=}CH{-}CH_3$

propene

$CH_2{=}CH{-}CH_2{-}CH_3$

1-butene

$CH_3{-}CH{=}CH{-}CH_3$

2-butene

$CH_3{-}CH{=}CH{-}CH_2{-}\overset{\displaystyle CH_3}{\overset{|}{C}}H{-}CH_3$

5-methyl-2-hexene

$CH_3{-}CH{=}CH{-}CH_2{-}CH{=}CH_2$

1,4-hexadiene

$CH_3{-}C{\equiv}C{-}CH_3$

2-butyne

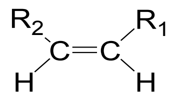
Name

Make models of ethane, ethene (ethylene), and ethyne (acetylene). (You will use the springs for multiple bonds.) For each of these molecules write the structural formula and draw a picture showing the geometry of the molecules.

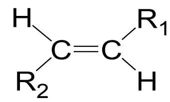
Molecule	C_2H_6	C_2H_4	C_2H_2
Total # Valence Electrons			
Lewis Structure			
Number of Bonding Electrons Clouds	C	C	C
Pairs of Nonbonding Electron Pairs	C	C	C
Electronic or Orbital Geometry	C	C	C
Molecular Geometry (Name)	C	C	C
3-Dimensional Sketch			
Bond Angle			
Hydrocarbon Type			
Saturated or Unsaturated			
Free Rotation around CC bond?			
Polarity of Molecule			

Geometric isomers: cis and trans

The *cis* isomers have the similar substituents on the same side of the alkene, while the *trans* isomers have these groups on opposite sides. As a result of restricted rotation, the *cis* and *trans* isomers do not interconvert.



cis



trans

Replace two of the hydrogen atoms of ethene with chlorine atoms. Draw them in the space below. Identify the isomers of dichloroethene below as cis, trans, or neither.

How many different structures can you make? ________

Hydrocarbon Questions

1. What is an alkyl group?

2. What is meant by a structural formula?

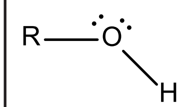
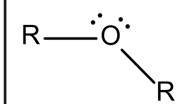
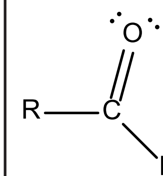
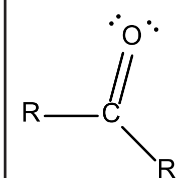
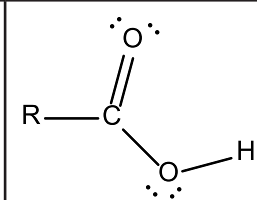
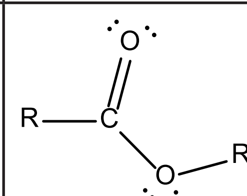
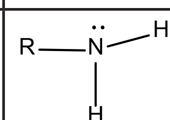
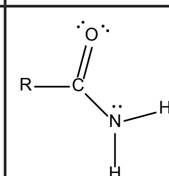
3. Why don't cis and trans isomers of alkanes exist?

4. Why do alkenes have geometric isomers and alkynes do not?

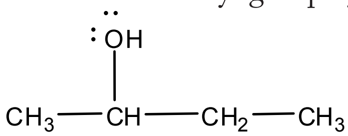
Name

Functional Groups

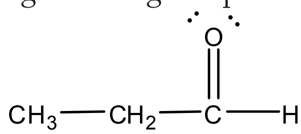
Organic molecules can be classified by their functional groups. A functional group is a characteristic group of atoms that are covalently bonded together, which give the molecule specific chemical and physical properties.

Functional Group	Condensed Formula	Structural Formula	Suffix
alcohol	R-OH		-ol
ether	R-O-R		-ether
aldehyde	R-CHO		-al
ketone	R-CO-R		-one
carboxylic acid	R-COOH or R-CO_2H		-oic acid
ester	R-COOR or R-CO_2-R		-oate
amine	R-NH_2; R_2-NH; R_3-N 1° 2° 3°		-amine
amide	R-CONH_2; R-CONHR; 1° 2° R-CONR_2 3°		-amide

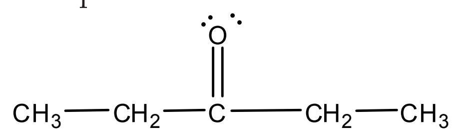
Organic compounds with functional groups are named in a similar fashion to alkanes, alkenes, and alkynes, but using different suffixes. From the alkane base remove the "e" and add the appropriate suffix. For alcohols and ketones indicate the position of the hydroxyl or carbonyl group by giving it the highest precedence. Examples are shown below:



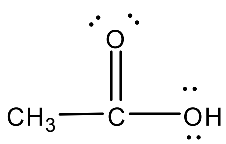
2-butanol



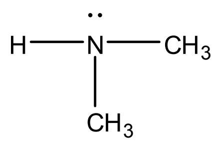
propanal



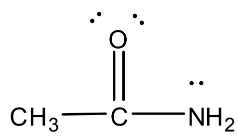
3-pentanone



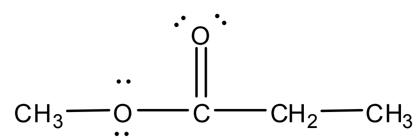
ethanoic acid
(aka acetic acid)



dimethylamine

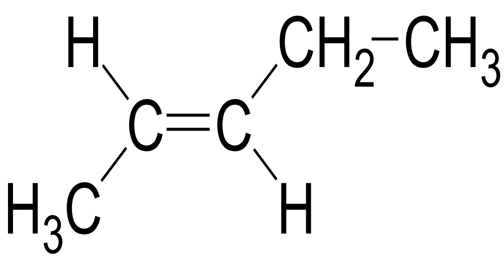


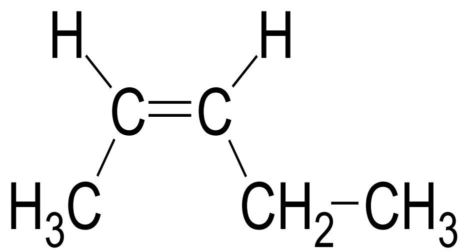
ethanamide



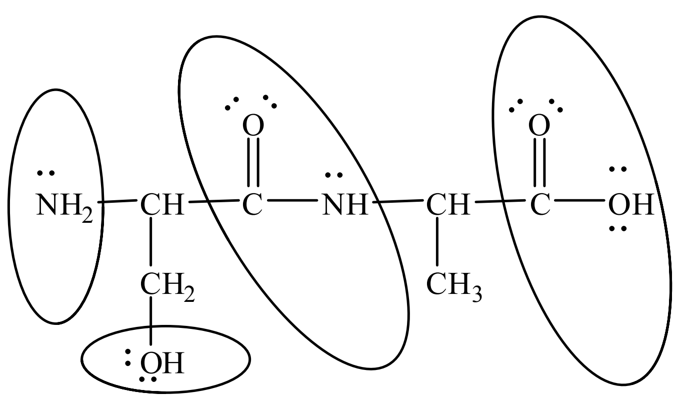
methylpropanoate

5. Name each of the following alkenes:





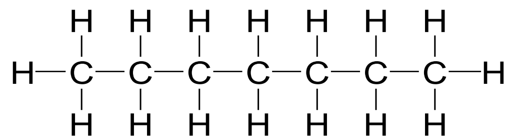
6. Identify the circled functional groups on the dipeptide amino acid, serine-alanine (Ser-Ala):



Name

7. Give the names for each of the following compounds:

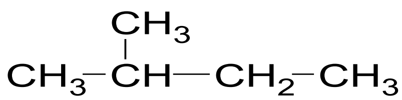
a.



b.

$CH_3(CH_2)_6CH_3$

c.



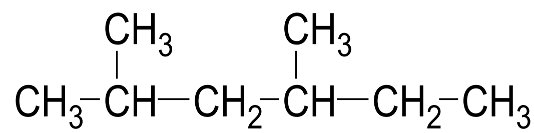
d.

$CH_3-CH_2-CH_2-\ddot{\underset{\cdot\cdot}{O}}H$

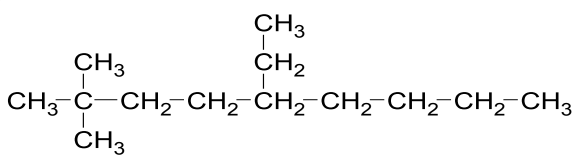
e.

$CH_3-CH=CH-CH_2-CH_2-CH_3$

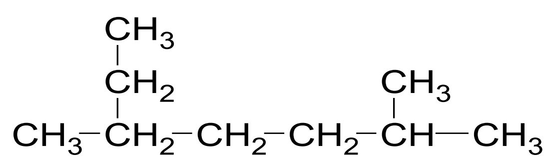
f.



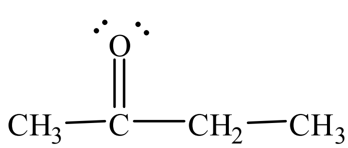
g.



h.



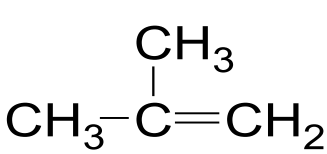
i.



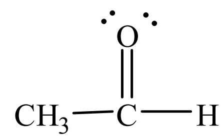
j.

$CH_3-CH=CH-CH_2-CH=CH-CH_2CH_3$

k.



l.



m.

$CH_3CH_2-C \equiv C-CH_2CH_2CH_2CH_2CH_2CH_3$

Name

EXPERIMENT 14

Nuclear Chemistry Worksheet

1. For each of the following atoms fill in the missing information:

Isotopic Elemental Symbol	Elemental Symbol	Number of protons	Number of neutrons
$^{115}_{48}Cd$			
	Fe		31
		73	105
$^{155}_{63}Eu$			
		81	115
	V		26

2. What is the correct atomic symbol for the following:

   a. An alpha particle (give 2)

   b. A beta particle (give 2)

3. Write a nuclear equation for the indicated decay of each of the following nuclides:

   a. U-234 by alpha emission

   b. Pb-214 by beta emission

   c. Th-230 by alpha emission

   d. Po-210 by alpha emission

   e. Tl-207 by beta emission

   f. Ac-227 by beta emission

   g. Tc-99m by gamma emission

Name

4. Fill in the missing particles in each of the following nuclear equations:

a. $\rightarrow {}^{217}_{85}At + {}^{4}_{2}He$

b. ${}^{75}_{34}Se + \qquad \rightarrow {}^{75}_{33}As$

c. ${}^{241}_{95}Am \rightarrow {}^{237}_{93}Np +$

d. ${}^{9}_{4}Be + \qquad \rightarrow {}^{6}_{3}Li + {}^{4}_{2}He$

e. ${}^{179}_{74}W + \qquad \rightarrow {}^{179}_{73}Ta$

f. $\rightarrow {}^{233}_{92}U + {}^{0}_{-1}e$

g. ${}^{241}_{94}Pu \rightarrow {}^{241}_{95}Am +$

h. ${}^{209}_{83}Bi + {}^{64}_{28}Ni \rightarrow {}^{272}_{111}Rg +$

5. A patient is given 0.050 mg of technetium-99m, a radioactive isotope with a half-life of about 6.0 hours. How much technetium will remain after 24 hours? After 48 hours?

6. A patient receives a 58-mg dose of I-131 to obtain an image of her thyroid. If the nuclide has a half-life of 8 days, how long will it take until less than 1 mg of I-131 remains in her body?

7. One of the nuclides in spent nuclear fuel is U-235, an alpha emitter with a half-life of 703 million years. How long would it take for the amount of U-235 to reach one-eighth of its initial amount?

8. Polonium-218 is an alpha emitter with a half-life of 3.0 minutes. Estimate how long it will take for a 312 g sample of polonium to decay to less than 1 gram?

Name

Worksheet 0 - Algebra Practice

1. What is $\frac{1}{2}+\frac{1}{3}$?

2. What is $\frac{1}{2}\times\frac{1}{3}$?

3. What is $\frac{1}{2}\div\frac{1}{3}$?

4. If $X \bullet Y = 3$, what does X equal?

5. D = M/V. Solve for V. V = ________

6. a/b = c/d. Solve for a. a = ________

7. Solve #6 for b. b = ________

8. Solve #6 for d. d = ________

9. If PV = nRT, what does R equal? R = ________

10. $\frac{1}{x} = \frac{1}{a} + \frac{1}{b}$. Solve for x. x = ________

11. $10^{-2} + 10^{-2}$ = ________

12. $10^{-2} + 10^{-3}$ = ________

13. $10^{2} \times 10^{3}$ = ________

14. $10^{-2} \times 10^{-3}$ = ________

15. $10^{-2} / 10^{-3}$ = ________

16. $y^{a} \bullet y^{b} =$ ________

17. Solve for A. 1.5A - 1.71 = -3.98 A = ________

18. (-1.10)(7.14 - 11.03) = ________

Name

Worksheet 1 - Dimensional Analysis I

1. Convert 61.4 g to mg.

2. Convert 0.693 cm to m.

3. Convert 6.08 mL to μL.

4. Convert 85.2 g to ounces.

5. Convert 6.83 miles to mm.

6. Convert 2.74 gallons to L.

7. Convert 78.7 ˚F to ˚C.

8. Convert 47.2 ˚C to ˚F.

Name

Worksheet 2 - Dimensional Analysis II

1. If a roll of nickel coins has a mass of 735 g, what is the mass of the roll of nickels in ounces?

2. Pistachio nuts cost $7.35/lb. How many grams of pistachios can be purchased for $5.00?

3. Convert 14.9 g/mL to lb/gallon.

4. Gold is a dense, shiny, ductile, malleable metal with a density of 19.30 g/mL. If a golden metal weighs 78.691 g and has a volume of 4.07 mL. Is the sample gold?

5. The mass of blood plasma in an adult is 7.0 pounds. Its density is 1.03 g/mL. Approximately how many liters of blood plasma are there in your body?

6. If the sun is 93,000,000 miles from earth, how long (minutes) does it take for sunlight to reach the earth? (Given: The velocity of light is 186,000 miles/second)

Name

Worksheet 3 - Inorganic Nomenclature

Directions

This is an example of the types of compounds that you will need to be able to name. The compounds will be a mixture of binary ionic, ternary ionic, binary covalent, binary acids, ternary acids, and hydrates.

1. Ag_2CO_3 ____________________
2. BrF_3 ____________________
3. HI *(aq)* ____________________
4. Na_3N ____________________
5. HCl *(aq)* ____________________
6. Cl_2O_7 ____________________
7. $KClO_3$ ____________________
8. $Ca(OH)_2$ ____________________
9. Au_2S_3 ____________________
10. H_3PO_4 *(aq)* ____________________
11. $V(NO_3)_5$ ____________________
12. Cu_2CO_3 ____________________
    or Latin ____________________
13. $ZnCl_2$ ____________________
14. AgCN ____________________
15. $Fe(ClO_4)_2 \cdot 7\ H_2O$ ____________________
    or Latin: ____________________
16. BaH_2 ____________________
17. H_2SO_3 *(aq)* ____________________
18. CCl_4 ____________________

1. oxalic acid
2. xenon tetrafluoride
3. potassium phosphide
4. magnesium cyanide
5. disilicon hexabromide
6. chromium(III) carbonate
7. ammonium chromate
8. silver nitride
9. sodium hypochlorite
10. chloric acid
11. sodium hydroxide
12. zinc acetate
13. lead(II) sulfite
14. tetrasulfur tetranitride
15. acetic acid
16. ammonium permanganate
17. iron(III) chloride or ferric chloride
18. magnesium nitrate
19. copper(II) bicarbonate

Name

Worksheet 4 - Moles I

1. How many chocolate bars are in 6.83 moles of chocolate bars?

2. If there are 20,600 students at Grossmont College, how many moles of students are there at Grossmont College?

3. How many atoms of gold are in 9.05 moles of gold?

4. If you have 2.85×10^{25} atoms of argon, how many moles of argon do you have?

5. What is the mass in grams of exactly one atom of antimony?

6. What is the mass in grams of 9.46 moles of molybdenum?

7. How many moles of nickel are there in 50.7 grams of nickel?

8. How many atoms of polonium are there in 3.86 grams of polonium?

9. What is the mass in grams of 2.07×10^{21} atoms of tungsten?

10. How many atoms of hydrogen are there in 652 molecules of oxalic acid, $H_2C_2O_4$?

11. How many moles of carbon are there in 7.93 moles of glucose, $C_6H_{12}O_6$?

Name

Worksheet 5 - Moles II

1. Calculate the mass of one atom of calcium in grams.

2. Salicylic acid is a natural pain reliever and fever reducer found in the bark of willow trees. Its chemical formula is $C_7H_6O_3$. What is the molar mass of salicylic acid?

3. How many molecules of salicylic acid are found in a 7.52 mole sample of salicylic acid?

4. What is the mass of 9.02×10^{24} molecules of salicylic acid?

5. How many atoms of carbon are in 470 molecules of salicylic acid?

6. How many moles of oxygen are in 4.07 moles of salicylic acid?

7. How many atoms of hydrogen are in 7.41 moles of salicylic acid?

8. How many grams of salicylic acid contain 64.2 moles of carbon?

9. What is the mass of hydrogen in a sample of salicylic acid with a mass of 6.84 grams?

10. If a sample of salicylic acid contains 28.5 grams of carbon, how many grams of hydrogen does it contain?

11. Determine the empirical formula of a compound that is composed of 2.82 g Na, 4.35 g Cl, and 7.83 g O.

12. The painkiller codeine is 72.22% C, 7.07% H, 4.68% N, and 16.03% O. Determine the empirical formula for codeine.

13. If a compound has and empirical formula of NH_2 and a molar mass of 64.1 g/mol, determine the molecular formula of the compound.

Name

Worksheet 6 - Stoichiometry

1. Balance the following chemical equation:

$$CoCl_2\,_{(aq)} + \quad Na_3PO_4\,_{(aq)} \longrightarrow \quad Co_3(PO_4)_2\,_{(s)} + \quad NaCl\,_{(aq)}$$

2. How many formula units of cobalt(II) phosphate can be produced by the reaction of 408 formula units of cobalt(II) chloride?

3. How many moles of sodium phosphate are required to react completely with 7.41 moles of cobalt(II) chloride?

4. How many grams of cobalt(II) phosphate can be produced by the reaction of 91.6 g of cobalt(II) chloride?

5. If 78.3 g of cobalt(II) phosphate are produced by the reaction in question 4, what is the percent yield?

6. How many grams of sodium phosphate are required to produce 3.77×10^{25} formula units of sodium chloride?

7. Balance the following chemical equation:

$$NH_{3\,(g)} + O_{2\,(g)} \longrightarrow NO_{(g)} + H_2O_{(l)}$$

8. How many grams of NO will be formed by the reaction of 8.55 moles of NH_3 with excess O_2?

9. How many moles of H_2O will be formed when 2.74 g of NH_3 react with excess O_2?

10. How many molecules of O_2 are required to react with 356 molecules of NH_3?

11. If 5.54 moles of NH_3 react with 6.25 moles of O_2, how many moles of NO will be produced?

12. If 78.2 g of NH_3 react with 83.4 g of O_2, how many grams of H_2O will be produced?

13. If 46.7 g of water are produced from #12, what is the percent yield?

Name

Worksheet 7 - Modern Atomic Theory and Periodic Trends

1. What is the maximum number of electrons that can be contained in the following:

   a. any orbital __________

   b. an s orbital __________

   c. an s sublevel __________

   d. a p orbital __________

   e. a p sublevel __________

   f. a d sublevel __________

   g. an f sublevel __________

2. Circle the best of the three choices:

a.	smallest atomic radius	O	S	F
b.	largest metallic character	Sn	I	Be
c.	largest ionic radius	Cl^-	Br^-	F^-
d.	largest first ionization energy	O	S	F
e.	smallest radius	Na^+	Na	I^-

3. Write the complete and condensed electron configurations for the following species:

   a. H

   b. O

   c. Fe

   d. Pb

   e. F^-

   f. Al^{3+}

4. Write the correct electron-dot formulas for the following:

   a. H

   b. K

   c. Mg

   d. Cl

   e. O^{2-}

Name

Worksheet 8 - Chemical Reactions

1. Single Replacement Reactions - Write the correctly balanced chemical equation for each and write which element is more active. Be sure to include state labels on your equations. Note that all metals will form a cation with a +2 charge if they react.
   a. If a piece of manganese metal is placed into a solution of zinc nitrate, $Zn(NO_3)_2$, the manganese begins to dissolve and a precipitate begins to form on the bottom of the test tube. Write a balanced chemical equation for the reaction that is taking place.

   b. Which element is more active, Zn or Mn?

   c. If a second piece of manganese metal is placed into a solution of barium nitrate, $Ba(NO_3)_2$, no visible reaction takes place. Which element is more active, Mn or Ba?

   Two test tubes were set up:
   Test tube #1 - Zinc metal in a solution of manganese(II) nitrate
   Test tube #2 - Barium metal in a solution of zinc nitrate

   d. Predict which test tube will show a reaction.

   e. Write a balanced chemical equation for this reaction.

2. Double Replacement Reactions - Write the correctly balanced chemical equation for each and indicate what kind of evidence of reaction you would be likely to observe. Be sure to include state labels on your equations.
   a. Na_3PO_4 *(aq)* + $FeCl_2$ *(aq)*

   b. Evidence of reaction:

c. $(NH_4)_2S$ *(aq)* + HBr *(aq)*

d. Evidence of reaction:

3. Combustion Reactions - If an inverted test tube is held over a sample of glucose, $C_6H_{12}O_6$, while it is combusted condensation is seen midway up the test tube. When a burning splint is placed into the test tube the flame is extinguished.
   a. What is the identity of the gas?

   b. Write a balanced equation for the reaction that took place. Be sure to include state labels on your equations.

4. Decomposition Reactions - A sample of white potassium perchlorate, $KClO_4$, solid was heated in a crucible with the lid slightly ajar. A sample of the gas that was evolved was collected in a test tube and when a glowing splint was inserted into the test tube the flame reignited. After the reaction a white solid was left in the crucible. The solid was found to be water soluble.
   a. What is the identity of the gas?

   b. Write a balanced equation for the reaction that took place. Be sure to include state labels on your equations.

Name

Worksheet 9 - Gas Laws

1. A sample of gas has a pressure of 515 mm Hg.
   a. What is the pressure of the gas in atmospheres?

   b. What is the pressure of the gas in kPa?

   c. What is the pressure of the gas in psi?

2. A sample of argon gas occupies a 4.0 L flask at a pressure of 2.6 atm. What volume will it occupy when the pressure is changed to 0.40 atm and the temperature remains constant?

3. A sample of methane gas, CH_4, is collected in a 3.55 L container with a pressure of 465 torr at 35.0 ˚C. What will the pressure be if the sample is transferred to a 10.0 L container and the temperature is raised to 95.0 ˚C?

4. A sample of nitrogen gas, N_2, is placed into a 30.0 L container at 26.0 atm pressure and 55.0 ˚C. To what temperature, in ˚C, must the nitrogen gas be reduced in order for the pressure to be reduced to 18.0 atm?

5. How many grams of nitrogen gas must be added to a container with 4.0 grams of nitrogen in order to increase the pressure of the nitrogen gas from 250 torr to 350 torr at a temperature of 18.0 ˚C?

6. A quantity of oxygen gas, O_2, occupies a volume of 18.6 L at 45.0 ˚C and 744 mm Hg. How many moles of oxygen are present?

7. If 41.7 grams of carbon dioxide, CO_2, are introduced into a 7.50 L container at 75.0 ˚C, what is the pressure of the carbon dioxide?

8. Calculate the density of sulfur dioxide, SO_2, at STP.

9. Calculate the molar mass of a gas with a density of 2.28 g/L at 25.0 ˚C and 1.32 atm pressure.

10. Butane gas is used in many lighters available. The butane, C_4H_{10}, evaporates and reacts with oxygen gas according to the unbalanced equation below:

$$C_4H_{10}(g) \; + \quad O_2(g) \longrightarrow CO_2(g) \; + \quad H_2O(g)$$

How many L of carbon dioxide would result from the reaction of 5.00 moles of butane with 10.0 moles of oxygen gas at a temperature of 65 ˚C and a pressure of 1.359 atm? (remember, this is a limiting reagent problem).

Name

Worksheet 10 - Solutions

1. A solution is prepared by dissolving 32.9 grams of sodium sulfate in 500 grams of water. Calculate the percent sodium sulfate in the solution.

2. A solution is prepared by dissolving 6.47 milligrams of aflatoxin in 2.53 kg of water. Calculate the concentration of aflatoxin in ppm.

3. A solution is prepared by dissolving 63.8 grams of copper(II) chloride in 500.0 grams of water. Calculate the mole fraction of copper(II) chloride in the solution.

4. A solution is prepared by dissolving 31.65 grams of potassium phosphate in enough water to make 1.500 L of solution. Calculate the molarity of the solution.

5. What mass of a 8.52% solution of ampicillin is required to supply 5.00 mg of ampicillin to a patient?

6. A solution is 0.675% acetaminophen by mass. How many grams of the solution are required to deliver 200.0 mg of acetaminophen?

7. What volume of 0.7526 M acetic acid ($HC_2H_3O_2$) is required to provide 3.000×10^{24} molecules of acetic acid?

8. How many grams of sodium sulfate are required to prepare 15.00 L of a 1.942 M solution of sodium sulfate?

9. If 35.72 mL of a 9.875 M solution of potassium chromate are diluted to a final volume of 850.0 mL, what is the final concentration of the solution?

10. What is the mass percent of potassium iodide in a 2.89 M solution of KI with a density of 1.27 g/mL?

Name

Worksheet 11
Solution Stoichiometry and Acid Base

1. How many grams of barium sulfate will precipitate if 418.6 mL of 0.353 M sodium sulfate are allowed to react with excess barium nitrate?

2. Consider the following reaction

$$Li_2S\ (aq) + Co(NO_3)_2\ (aq) \longrightarrow 2\ LiNO_3\ (aq) + CoS\ (s)$$

What volume of 0.275 M Li_2S solution is required to completely react with 125 mL of 0.150 M $Co(NO_3)_2$?

3. Determine the pH of a solution that is 2.7×10^{-5} M in HCl. (Hint: $[H_3O^+] = 2.7 \times 10^{-5}$ M)

4. Determine the pH of a solution that is 6.2×10^{-4} M in H_2SO_4. (Hint: $[H_3O^+] = 1.24 \times 10^{-3}$ M)

5. Determine the pOH of a solution that is 8.5×10^{-6} M in NaOH. (Hint: $[OH^-] = 8.5 \times 10^{-6}$ M)

6. What is the pH of a solution that is 8.62×10^{-4} M in OH^-?

7. What is the $[H_3O^+]$ in a solution with a pH = 8.63?

8. Fill in the chart below:

$[H_3O^+]$	$[OH^-]$	pH	pOH
7.37×10^{-5} M			
	4.61×10^{-3} M		
		10.742	
			5.291

9. A 35.00 mL sample of vinegar requires 51.74 mL of 0.4298 M sodium hydroxide to react with all of the acetic acid. What is the concentration of acetic acid (molarity) in the vinegar?

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