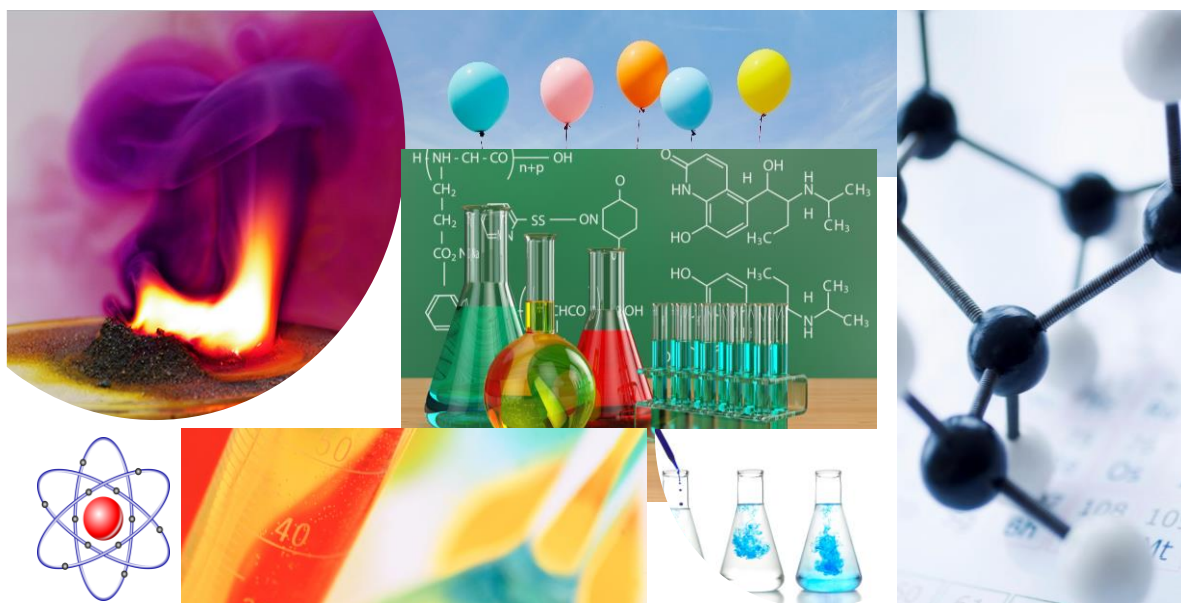


Introduction to Chemistry Lab Manual

CHM 103



Created and written by Dr. Sophia Koziatek and the St. Charles Community College
Chemistry Department

Periodic Table of the Elements

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	IA																	IIIA
1	hydrogen 1 H 1.00794																	helium 2 He 4.002602
2	lithium 3 Li 6.941	beryllium 4 Be 9.012182																neon 10 Ne 20.1797
3	sodium 11 Na 22.98977	magnesium 12 Mg 24.3050																argon 18 Ar 39.948
4	potassium 19 K 39.0983	calcium 20 Ca 40.078	scandium 21 Sc 44.95591	titanium 22 Ti 47.867	vanadium 23 V 50.9415	chromium 24 Cr 51.9961	manganese 25 Mn 54.93805	iron 26 Fe 55.845	cobalt 27 Co 58.9332	nickel 28 Ni 58.6934	copper 29 Cu 63.546	zinc 30 Zn 65.382	gallium 31 Ga 69.723	germanium 32 Ge 72.64	arsenic 33 As 74.9216	selenium 34 Se 78.96	bromine 35 Br 79.904	krypton 36 Kr 83.798
5	rubidium 37 Rb 85.4678	strontium 38 Sr 87.62	yttrium 39 Y 88.90585	zirconium 40 Zr 91.225	niobium 41 Nb 92.90638	molybdenum 42 Mo 95.94	technetium 43 Tc [98]	ruthenium 44 Ru 101.07	rhodium 45 Rh 102.9055	palladium 46 Pd 106.42	silver 47 Ag 107.8682	cadmium 48 Cd 112.411	indium 49 In 114.818	tin 50 Sn 118.710	antimony 51 Sb 121.760	tellurium 52 Te 127.60	iodine 53 I 126.9045	xenon 54 Xe 131.293
6	cesium 55 Cs 132.90545	barium 56 Ba 137.327	lutetium 71 Lu 174.967	hafnium 72 Hf 178.49	tantalum 73 Ta 180.9479	tungsten 74 W 183.84	rhenium 75 Re 186.207	osmium 76 Os 190.23	iridium 77 Ir 192.217	platinum 78 Pt 195.078	gold 79 Au 196.96655	mercury 80 Hg 200.59	thallium 81 Tl 204.3833	lead 82 Pb 207.2	bismuth 83 Bi 208.980	polonium 84 Po [209]	astatine 85 At [210]	radon 86 Rn [222]
7	francium 87 Fr [223]	radium 88 Ra [226]	lawrencium 103 Lr [262]	rutherfordium 104 Rf [261]	dubnium 105 Db [262]	seaborgium 106 Sg [263]	bohrium 107 Bh [262]	hassium 108 Hs [265]	meitnerium 109 Mt [266]	damstadtium 110 Ds [281]	roentgenium 111 Rg [280]	copernicium 112 Cn [285]	nihonium 113 Nh [284]	flerovium 114 Fl [289]	moscovium 115 Mc [288]	livermorium 116 Lv [293]	tennessine 117 Ts [294]	oganesson 118 Og [294]

lanthanum 57 La 138.9055	cerium 58 Ce 140.116	praseodymium 59 Pr 140.90765	neodymium 60 Nd 144.24	promethium 61 Pm [145]	samarium 62 Sm 150.36	europium 63 Eu 151.964	gadolinium 64 Gd 157.25	terbium 65 Tb 158.9253	dysprosium 66 Dy 162.50	holmium 67 Ho 164.930	erbium 68 Er 167.259	thulium 69 Tm 168.934	ytterbium 70 Yb 173.04
actinium 89 Ac [227]	thorium 90 Th 232.038	protactinium 91 Pa 231.0359	uranium 92 U 238.0289	neptunium 93 Np [237]	plutonium 94 Pu [244]	americium 95 Am [243]	curium 96 Cm [247]	berkelium 97 Bk [247]	californium 98 Cf [251]	einsteinium 99 Es [252]	fermium 100 Fm [257]	mendelevium 101 Md [258]	nobelium 102 No [259]

Table of Contents

Lab Safety Contract.....	2
Lab 1: Significant Figures Activity	3
Lab 2: Scavenger Hunt, Measurement and Unit Conversions Lab.....	17
Lab 3: Density and Graphing Lab	37
Lab 4: Periodic Properties Lab	55
Lab 5: Electron Configurations and Light Lab	75
Lab 6: Nomenclature Activity	99
Lab 7: Writing and Balancing Equations Activity	123
Lab 8: Copper Transformation Lab	145
Lab 9: Stoichiometry Lab	161
Lab 10: Molecular Models Activity	179
Lab 11: Gas Laws Lab	199
Lab 12: Solutions and Electrolytes Lab	219
Lab 13: Acid Base Lab	237
Appendix I.....	255
Appendix II.....	256
Appendix III.....	257
Appendix IV	258
Appendix V	259
Appendix VI	260



Some sections of this text were taken and modified from Wet Lab Experiments, [Chem LibreTexts](https://chem.libretexts.org/), Santa Monica College, under a Creative Commons Attribution-Noncommercial-ShareAlike 3.0 License. To view this license, visit: <https://creativecommons.org/licenses/by-nc-sa/4.0/legalcode>

Lab Safety Contract

On the first day of lab students will watch the Lab Safety Video, take a Safety Quiz and sign a Safety Contract. More information about the video and quiz will be provided by the lab instructor. Information about the safety contract is found below.

Please read the following rules of the laboratory that appear in the safety contract.

- I have received instructions in laboratory safety and have seen the safety film.
- I understand that these safety regulations are for my own protection and for the protection of others in the laboratory and pledge to abide by them.
- I know the location of the eye wash, fire extinguisher, first aid kit & emergency phone.
- **I agree to purchase and bring my own safety glasses. I agree to wear safety glasses in the laboratory at all times.**
- I agree to wear long-pants with no holes that cover all of my legs and ankles, closed toe and closed heel shoes with rubber soles that cover my entire foot including the top of the foot, and wear shirts that cover the entire front and back of my torso with sleeves that cover (at a minimum) my shoulders. *I will be covered from my shoulders to the floor with no skin showing.*
- It is at the instructor's discretion if clothing/goggles do or do not meet the criteria above.
- I agree that if I arrive at least 20 minutes late to lab I will be unable to perform the lab and will receive a zero for the lab report.
- I agree to handle all chemicals appropriately.
- I agree to dispose of all chemical waste appropriately.
- I will read the entire lab procedure before coming to class.
- I will wipe down my workbench before and after each lab with water and paper towels, cleaning any liquid or solid spills. This includes not only my lab bench but any other common area used during the lab period, like balances and fume hoods.
- I will not leave any containers sitting out when I am finished with the lab, especially unmarked containers containing chemicals.
- I will report any accident, spill, glass or other equipment breakage to the instructor immediately.
- I will not bring food or drink into the lab.
- If I am unable to operate a motor vehicle due to medication I cannot participate in lab.
- If I have health considerations or impairments (such as pregnancy, a compromised immune system, etc.) that may affect my ability to safely participate in lab, I understand that it my responsibility to discuss those concerns with my physician.

Any deviations from these rules will result in being required to leave the lab or deduction of points from the lab report. If I am required to leave the lab, I will earn a zero (0) for that lab report grade. If I blatantly disregard safety rules, I can be required to undergo more safety training and only allowed back in the lab after demonstrating outstanding safety consciousness based on testing by my instructor.

Lab 1: Significant Figures Activity

Lab Objectives

- Identify the number of significant figures, digits or decimal places in a value
- Round numbers to a specific number of significant figures
- Interpret or write numbers in scientific notation to the correct number of significant figures
- Determine the number of significant figures or decimal places in a calculation involving addition, subtraction, multiplication and/or division
- Report an answer to the correct number of significant figures when performing a calculation

Background Information

Keeping track of numbers is of utmost importance in chemistry. Numbers are a valuable tool used to quickly provide information about an experiment or a process. They convey the magnitude (how big or small), precision (consistency in measurement) or accuracy (proximity to a given value) of a measurement. This is done by keeping track of the number of significant figures or significant decimal places in a value or calculation. The term significant digit can refer to significant figures (SFs), decimal places (DP) or place values (ones, tens, etc.). Common numerical place value names are shown in **Figure 1**. Adept knowledge of rounding, scientific notation, significant figures, is necessary to do well and that is the focus of this lab.

Figure 1: Place Value Names

1	,	2	3	4	,	5	6	7	.	8	9	0	1
Millions		Hundred Thousands	Ten Thousands	Thousands		Hundreds	Tens	Ones	Decimal Point	Tenths	Hundredths	Thousandths	Ten-Thousandths

However, sometimes numbers are either too small or too large to write conveniently. For example, the mass of the earth is about 5,972,000,000,000,000,000,000 g and the average length of an O-H bond is 0.0000000009572 m. Both of those numbers may be difficult to conceptualize and are tedious to write. Therefore, they are more easily expressed in scientific notation.

Numbers in scientific notation have two parts: a decimal component and an exponent. The decimal portion of the number must be a value between one and ten and may or may not contain a decimal point. It is located on the left of the “ $\times 10$ ” part of the expression. The exponent is on the right of the “ $\times 10$ ” expression and written as a superscript. The exponent will be a whole number that can be a positive number, negative number or zero.

Large numbers, like the mass of the earth, will have a positive exponent when written in scientific notation. In general, large numbers are values greater than one. To change a large number from standard form to scientific notation use the following steps.

1. Locate the decimal point. If there is no decimal point, put one temporarily at the end of the number.
2. Move the decimal point enough places the left to wind up with a number between one and ten. This often involves moving the decimal point between the first two numbers of the value (e.g. between the ones and tens place).
3. The number of places the decimal point moved is the exponent.
4. Remember that the exponent will be *positive* for this number.
5. Always maintain the correct number of significant figures in the final value.

Example

Converting a large number from standard form to scientific notation

Number in Standard Form	Number in Scientific Notation
5,972,000,000,000,000,000,000,000 g	5.972×10^{27} g

Small numbers, like the length of an O-H bond, will have a negative exponent when written in scientific notation. Small numbers are values that are smaller than one. To change a small number from standard form to scientific notation use the following steps.

1. Locate the decimal point.
2. Move the decimal point enough places the right to wind up with a number between one and ten. This often involves moving the decimal point between the first two numbers that are not themselves both zero.
3. The number of places a decimal point moved will be the exponent for that number.
4. Remember that the exponent will be *negative* for this number.
5. When possible, try to maintain the correct number of significant figures in the value.

Example

Converting a small number from standard form to scientific notation

Number in Standard Form	Number in Scientific Notation
0.0000000009572 m	9.572×10^{-11} m

If a number is already a number between one and ten and needs to be written in scientific notation the decimal point should not move and the exponent will be zero.

Example

Converting a number already between one and ten to scientific notation

Number in Standard Form	Number in Scientific Notation
2.453 s	2.453×10^0 s

When converting from a number in scientific notation to standard form the rules be used vice versa, or the other way around. However the number is being expressed, the number in scientific notation or standard form must always maintain the same number of significant figures. Significant figures are used to report a value, measured or calculated, to the correct number of decimal places or digits that will reflect the precision of the value.

Rules for Determining the Number of Significant Figures (SF) in a Value

When determining the number of significant figures in a value, keep the following in mind.

1. All digits 1 – 9 are always significant. In a number that does not contain zeros, all numbers are considered significant.
2. Zeroes that come between other numbers that are not zero are significant.
3. When a value is less than one and begins “zero point something”, the zeroes at the beginning of that number are NOT significant. However, once the first nonzero number is reached then every number that follows is significant (even if it is a zero).
4. When a number is one or greater than one and contains a decimal point, then all values in front and behind the decimal point are significant (even if they are zeros)
5. Numbers that end in zero but do NOT contain a decimal point are called ambiguous. Only the nonzero numbers and any interior zeros will be significant.
6. In numbers written in scientific notation only the numbers to the left of the “ $\times 10^n$ ” are significant.

Example

Determine the number of significant figures (SF) in each value.

32	7503	0.000460	15.000	908,000	6.022×10^{23}
2 SF	4 SF	3 SF	5 SF	Ambiguous but has 3 SF	4 SF
Rule 1	Rule 2	Rule 3	Rule 4	Rule 5	Rule 6

Notice the difference between using Rule 3, Rule 4 and Rule 5 when determining significant figures in a value. Both the numbers 0.000460 and 15.000 contain a decimal point and end in zero, however, only in 15.000 do all of the zeros count as significant. In the value 0.000460 the zeros at the beginning of the number that are underlined are called placeholding zeros. Placeholding zeros are never considered significant (Rule 3), only occur at the beginning of a number and are only used to indicate how small the value is. The number 15.000 does not begin with zero but rather with 15 which is followed by a decimal point and zeros. This is an example of Rule 4 where all numbers before and after the decimal point are significant. The number 908,000 does not contain decimal point and has zeros at the end. The zeros that are underlined in 908,000 are called ambiguous zeros. Ambiguous zeros only occur at the end of numbers that do not contain a decimal point and are also not considered significant. If a value ends with ambiguous zeros the entire number is considered an ambiguous number (Rule 5). However, the nonzero numbers (or any interior zeros) are significant. So although the number 908,000 is ambiguous it can also be said to contain three significant figures (908).

The Significant Figures Flowchart in **Figure 2** can also help to determine the number of significant figures in a value. To use the flowchart, start in the upper left corner with the box containing the question, *Is the number written in scientific notation?* Then follow the appropriate *yes* or *no* response to determine the correct number of significant figures for a number. Underlining the last significant digit can help to keep track of values.

Some numbers are considered exact numbers and have an infinite number of significant figures. Therefore, they are never used to determine the number of significant figures in the final answer of a calculation. A number is considered an exact number if it falls into one of these three categories.

1. Counted numbers are exact. The number of students in the lab is an exact number.
2. Definitions or equivalent statements are exact. There are exactly 2.54 centimeters in one inch. This equivalent statement is written as $2.54 \text{ cm} = 1 \text{ in.}$
3. Integers in an equation or formula are exact. In the formula, $\text{radius} = \frac{\text{diameter}}{2}$, the number two in the denominator is an exact number.

When performing a calculation, the number of significant figures or decimal places in each number will affect how the final answer should be reported. The type of mathematical operation (addition, subtraction, multiplication or division) is what determines how the final answer or number is reported.

Decimal Place (DP) Rules for Addition and Subtraction

If a calculation only involves addition or subtraction as the operation, then decimal place rules should be followed. This involves keeping track of the numbers after the decimal point in a number.

1. Round the sum/difference to match the value with the fewest decimal places after the decimal point.
2. If a number does not contain a decimal point, then the answer is limited to lowest place value (ones place, tens place, hundreds place, etc.)

Example

$$\begin{aligned} 3.\underline{2} + 12.\underline{15} \\ = 15.\underline{35} \\ = \mathbf{15.4} \end{aligned}$$

The number 3.2 has 1 DP and the number 12.15 has 2 DP. The final answer, 15.4, can only have 1 DP.

Example

$$\begin{aligned} 45.\underline{00} + 6.\underline{1} + \underline{12} \\ = \underline{63.1} \\ = \mathbf{63} \end{aligned}$$

The number 45.00 has 2 DP, 6.1 has 1 DP, and 12 has no decimal point and thus is reported to the ones place. The final answer is limited to the same place value that does not contain a decimal point (the ones place).

Significant Figure (SF) Rules for Multiplication and Division

If a calculation only involves multiplication or division as the operation, then significant figures rules should be followed. This involves keeping track of the number of significant figures in a value.

1. Determine the number of significant figures in each number.
2. Round the product/quotient to match the value the fewest number of significant figures.

Example

$$\begin{aligned} 25.\underline{21} \times 3.\underline{0} \\ = \underline{75.63} \\ = \mathbf{76} \end{aligned}$$

The number 25.21 has 4 SF and the number 3.0 has 2 SF, so the final answer, 76, can only have 2 SF.

Example

$$\begin{aligned} 0.968 \div 0.9\underline{5} \\ = \underline{1.018947368} \\ = \mathbf{1.0} \end{aligned}$$

The number 0.968 has 3 SF and the number 0.95 has 2 SF, so the final answer, 1.0, can only have 2 SF.

Rules for Problems with Both Sets of Operations

When problems have mixed operations (e.g. addition and division) always perform each operation in PEMDAS (Parentheses, Exponents, Multiplication/Division, Addition/Subtraction) order. If there are multiple steps in the calculation, use unrounded numbers in all the steps and only round as the very last step of the problem. Remember, underlining in each step helps.

Example

$$\begin{aligned} & 3.252 \times (0.125 + 1.30) \\ & \qquad \qquad \qquad \mathbf{3 \text{ DP} \quad 2 \text{ DP}} \\ & = 3.252 \times 1.425 \\ & \qquad \mathbf{4 \text{ SF} \quad 2 \text{ DP and } 3 \text{ SF}} \\ & = 4.6341 \\ & \qquad \mathbf{3 \text{ SF}} \\ & = \mathbf{4.63} \end{aligned}$$

Start with DP rules (the first step involves addition), then follow SF rules (the second step involves multiplication). The number 1.425 has 2 significant DP which gives it 3 SF. The final answer must therefore have three significant figures.

Example

$$\begin{aligned} & (2.31 \times 0.014) + 6.8579 \\ & \qquad \mathbf{3 \text{ SF} \quad 2 \text{ SF}} \\ & = 0.03234 + 6.8579 \\ & \mathbf{2 \text{ SF and } 3 \text{ DP} \quad 4 \text{ DP}} \\ & = 6.88934 \\ & = \mathbf{6.889} \end{aligned}$$

Start with SF rules (the first step involves multiplication), then follow DP rules (the second step involves addition). The number 0.03234 has 2 SF which gives it 3 significant DP. The final answer must therefore have three decimal places.

Example

$$\begin{aligned} & = \frac{36.45 - 1.450}{5.00} \\ & = \frac{35.00}{5.00} \\ & = 7 \\ & = \mathbf{7.00} \end{aligned}$$

Start with DP rules, then follow SF rules. The number 36.45 has 2 DP and 1.450 has 3 DP. Once subtracted the difference is limited to 2 DP, which gives it 4 SF. That value is divided by 5.00 (3SF). The final answer is limited 3 SF and two zeros therefore must be added to give the correct SF but not change the numerical value.

Simplified Rounding Rules

When rounding the final answer, if the number to the right of the last number to keep is five or above, round up. If the last number to the right of the last digit kept is four or below, the number will round down or is said to stay the same.

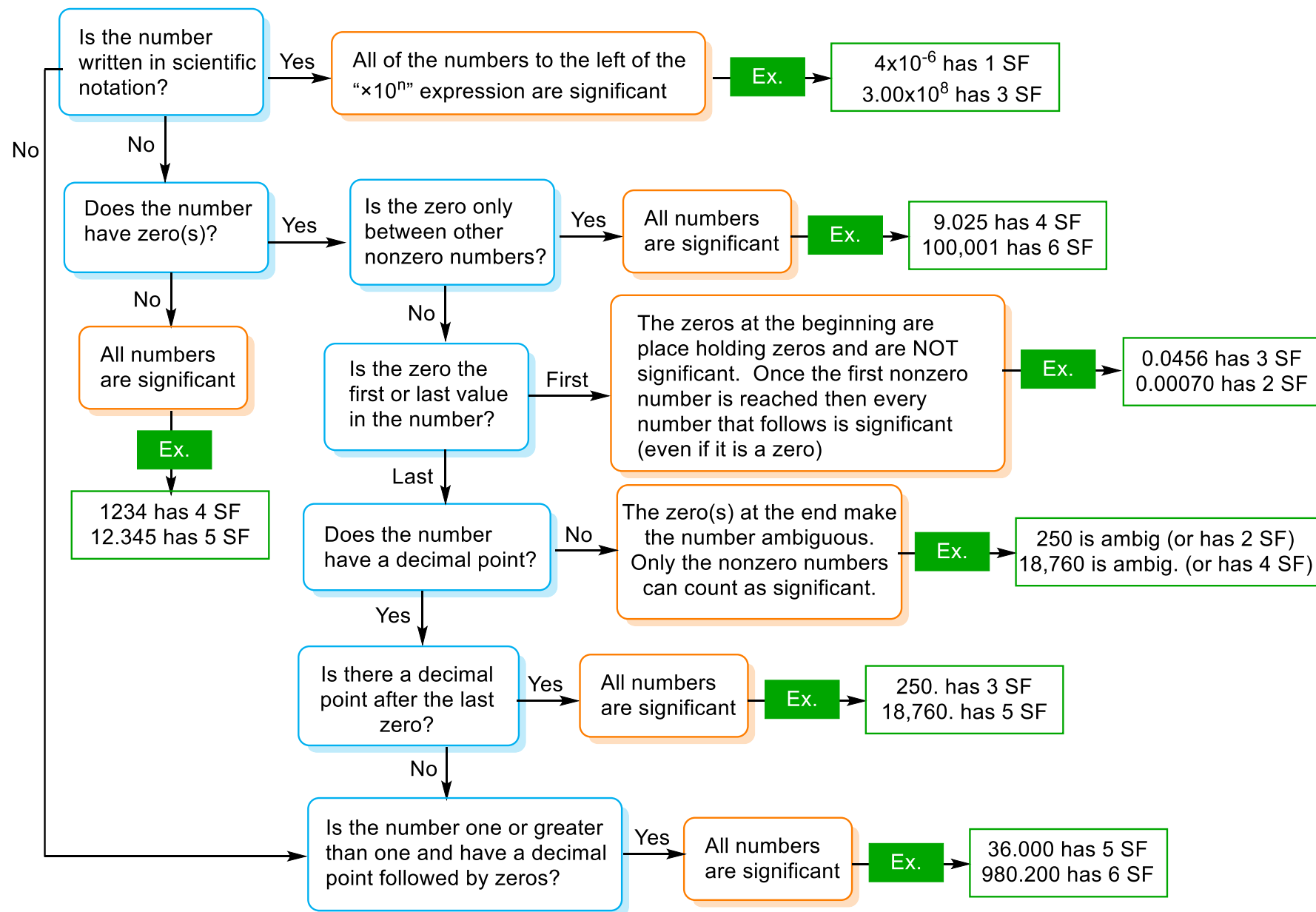
Example

Round 16.9 mL to two significant figures.
16.9 mL Rounds to **17 mL**

Example

Round 16.3 mL to two significant figures.
16.3 mL Rounds to **16 mL**

Figure 2: Significant Figures Flowchart



Materials and Reagents Needed for Lab

This activity is considered a dry lab thus no chemicals will be used for this lab. Only a writing utensil and a calculator are needed for this lab.

Safety Considerations

Although this is a dry lab, students must wear appropriate PPE when in the Intro Chem Lab. No goggles or safety glasses are needed for this lab, however, students must still wear closed-toe shoes, long pants, shirts with a front/back and some type of sleeve. Read the board to determine if gloves must be worn for this lab.

Procedure

Follow the directions in the manual and given by the lab instructor.

Before Leaving Lab

- Check the board for any additional instructions
- Wipe down the benchtops
- Push in stools
- Check with your instructor that your area has been sufficiently cleaned and reorganized

Lab 1: Significant Figures Activity Pre-Lab

Name _____

Section _____

Date _____

Pre-lab Questions

Read through all the background information and the procedure of the lab before answering these pre-lab questions.

1. When should a number be written in scientific notation?

2. Determine the word or phrase that correctly completes the sentences below.

A) Numbers with a _____ exponent represent large numbers.

B) The number 0.002350 has _____ significant figures and is written as _____ in scientific notation.

3. Write each number in scientific notation.

A) 602,200,000,000,000,000,000,000 _____

B) 0.0005485799 _____

4. Write each number in standard form.

A) 1.41×10^8 _____

B) 2.119×10^{-4} _____

5. Determine the number of significant figures in each value.

A) 831,725 _____

E) 0.0019 _____

B) 3.00×10^8 _____

F) 91.00 _____

6. In the value below, identify the number in the indicated place value.

29,615.04738

Ones place _____

Tenths place _____

Hundredths place _____

Hundreds place _____

7. Explain the difference between a placeholding zero and an ambiguous zero.

8. Briefly explain whether the decimal point is necessary to indicate the number of significant figures in the value below.

567,123.

9. When doing calculations, what set of rules should be followed when only addition or subtraction is performed?

10. When doing calculations, what set of rules should be followed when only multiplication or division is performed?

Lab 1: Significant Figures Activity – Data/Calculations

Name _____

Section _____

Drawer _____

Balance _____

Date _____

Partner(s) _____

Part A: Significant Figures

Directions: Determine the number of significant figures in each value.

1. 1106 _____

5. 0.38 _____

2. 902.040 _____

6. 5.74×10^9 _____

3. 0.000194 _____

7. 31,700. _____

4. 3060 _____

8. 3×10^{-5} _____

Part B: Scientific Notation

Directions: Write each number in either scientific notation or standard form.

1. 0.80 _____

6. 7.12×10^2 _____

2. 4230. _____

7. 8.040×10^{-4} _____

3. 5.8 _____

8. 9.156×10^1 _____

4. 0.000327 _____

9. 3.7×10^5 _____

5. 36,040 _____

10. 5×10^{-3} _____

Part C: Rounding

Directions: Round each number to four significant figures. Final answers must be unambiguous.

1. 564.491 _____

3. 10.0968 _____

2. 291,945 _____

4. 0.00021713 _____

Part D: Calculations

Directions: Complete each calculation and round the final answer to the appropriate number of significant figures or decimal places.

Problem		Unrounded Answer	Rounded Answer
1.	$805.913 - 60.04$		
2.	0.00089×0.2211		
3.	$82.86 \div 3.10$		
4.	$7.691 - 1.346589$		
5.	$5.67 + 56.33$		
6.	$(25.335 \div 735) \times 8.302$		
7.	$680 \div 20.5$		
8.	$335.70 + 96.0$		
9.	$\frac{7600}{0.53}$		
10.	$0.5 \times 26 \times 0.081$		

Part E: Calculations with Mixed Operations

Directions: Complete each calculation and round the final answer to the correct appropriate number of significant figures or decimal places. The work shown box must include the unrounded answer.

Problem		Work Shown and Unrounded Answer	Rounded Answer
1.	$9.732 + (9.14 \times 0.048)$		
2.	$(9.732 + 9.14) \times 0.048$		
3.	$\frac{6.744 \times 10^{-2}}{3.70 \times 10^3}$		
4.	$\frac{1079.8}{(33.8 - 155.88)}$		
5.	$\frac{1.028 + 18.2}{0.004460}$		
6.	$\frac{1.039}{0.00557} + 29.3$		
7.	$\frac{3.0 - 0.00742}{2.406 + 1.0972}$		

Problem		Work Shown	Rounded Answer
8.	$\frac{4.6 + 7.551}{10.891 - 0.20}$		
9.	$\frac{27.136 + 8.45}{0.7024}$		
10.	$2.34 \times 10^3 \div 1.67 \times 10^2$		
11.	$(218.8 \times 4.0251) + (35.91 \times 24.3466)$		
12.	$(14.224 \times 0.890) \div 0.62$		
13.	$8.58 + 201 + 0.8$		
14.	$\frac{6.842 \times 10^{32}}{6.51 \times 10^{32}}$		

Lab 2: Scavenger Hunt, Measurement and Unit Conversions Lab

Lab Objectives

- Locate and identify laboratory safety equipment and sketch the location of equipment
- Identify lab equipment by name and determine its purpose and how to use it
- Use units to determine or indicate types of measurement
- Convert units and report answers to the correct number of significant figures
- Indicate the place value a measurement should be reported to based on the type of lab equipment used
- Identify the type of measurement (mass, volume, temperature, energy, etc.) obtained by a specific piece of lab equipment

Background Information

Chemistry is the study of matter and the changes it undergoes. Our understanding of chemical processes thus depends on the ability to acquire accurate information about matter. Often, this information is quantitative and in the form of measurements. All quantitative measurements must contain both a number and a unit. In this lab, students will be introduced to some common measuring devices and will learn how to use them to obtain correct measurements.

Types of Measurements

Time

Time can be described as a change in physical quantity, like the number on a clock, or a measurable period during which an action exists (like the position of the sun in the sky). The interval over which the change occurs can be measured in seconds (s), minutes (min), or hours (hr).

Length

Length is an indication of distance and is often measured using a ruler. Length measurements will often be recorded in meters or derivatives of meters, like centimeters (cm), during lab.

Temperature

Temperature is a measure of how hot a substance is and is measured using a thermometer. There are three common temperature scales: degrees Celsius, degrees Fahrenheit and Kelvin. For this lab course, temperature measurements should always be recorded in degrees Celsius ($^{\circ}\text{C}$).

Mass

Mass is defined as the quantity of matter an object contains and is commonly measured using a balance. Digital balances are often very sensitive, even to air flow, so it is important to close all three doors on the balance when taking measurements. When using a digital scale, always record all numbers displayed. For this lab course mass measurements will be recorded in grams.

Volume

Volume is temperature dependent and is defined as the space matter occupies. It can be measured using a variety of different devices like graduated cylinders, volumetric pipets, and burets. The scale on the measuring device will determine how a volume should be recorded. The volume measurements will be recorded in liters or derivatives of liters, most commonly in milliliters (mL).

Density

Density is a derived unit meaning it requires more than one measurement to express it. To find the density of an object the mass of the object must be divided by its volume. Units of density will always be expressed with units of “mass over volume” or $\frac{\text{mass}}{\text{volume}}$, for example $\frac{\text{grams}}{\text{milliliters}}$.

When making a measurement using a non-digital readout (ruler, thermometer, graduated cylinder, etc.) it is important to use the following steps.

- 1. Always read between the lines.**

The last digit that is measured is an estimate of how far between the smallest two lines the quantity lies. The next place value down should be “estimated”. If an instrument gives readings to the ones place, then the measurement must be reported to the tenths place.

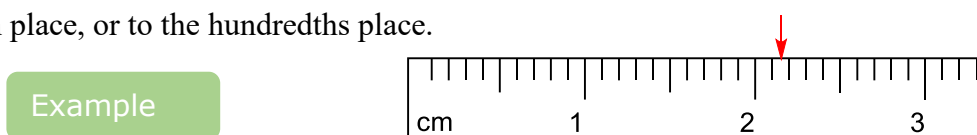
- 2. When measuring a liquid volume, always read from the bottom of the meniscus.**

The meniscus is the curved surface of the liquid. The steepness or shallowness of the meniscus will depend on the liquid and the container it is in.

All measuring devices are subject to error, making it impossible to obtain exact measurements. Measurements contain both certain and uncertain digits. When making a measurement record all the digits based on the markings of the device, those are the certain digits. The very last digit in the measurement must be estimated and is determined to be one place value over from what is shown on the scale. This is the uncertain digit. The *uncertain digit* is a best guess estimate using the smallest unit of measurement given and estimating between two of these values. These digits are collectively referred to as significant figures.

When making measurements, it is important to be as accurate and precise as possible. Accuracy is a measure of how close an experimental measurement is to the true, accepted value. Precision refers to how close repeated measurements (using the same device) are to each other.

On the ruler below, the large numbers are changing by 1 cm and the small markings indicate one-tenths of a centimeter (0.1 cm). The last digit would have to be estimated in the next place over from the tenth place, or to the hundredths place.



For example, the arrow is pointing to a position between 2.1 cm and 2.2 cm, those digits are known with certainty. The actual measurement, however, must be a number between those two values and must be estimated. The measurement for this position might be written as 2.14 cm or 2.16 cm. In this case, the four or six is the estimated or uncertain value.

In this course measurements will be recorded using metric units (grams, liters, meters, degrees Celsius) rather with English units (pounds, gallons, feet, degrees Fahrenheit). With the metric systems the base units (grams, liters, meters, etc.) can often be combined with a prefix (see Metric Prefix Table below) to describe large or small numbers. Any of the metric prefixes (e.g. centi, milli, deci, kilo, etc.) may be combined with any base unit. For example, combining the prefix *kilo* with the base *grams*, results in the unit *kilogram*. This means that in one kilogram there are 1×10^3 grams and the equivalent statement that relates them would be $1 \text{ kg} = 1 \times 10^3 \text{ g}$.

Metric Prefix Table

Prefix	Abbreviation or Symbol	Meaning	Decimal Equivalent
kilo	k	1×10^3	1000
hecto	h	1×10^2	100
deca	da	1×10^1	10
deci	d	1×10^{-1}	0.1
centi	c	1×10^{-2}	0.01
milli	m	1×10^{-3}	0.001

Unit Conversions and Conversion Factors

When converting between units a conversion factor is often used. Conversion factors are ratios used to express a measured quantity in different units and are based on an equivalence statement. An equivalence statement relates different types of units to each other. For example, 1 inch is equivalent to 2.54 centimeters. This may be written as the equivalent statement $1 \text{ in} = 2.54 \text{ cm}$. This allows for a convenient reference to convert from one unit to the other. A few examples of metric to English equivalent statements may be found below and more are in Appendix I.

Common English-Metric Equivalencies

Length		
1 kilometer (km)	=	0.6214 miles (mi)
1 inch (in)	=	2.54 centimeters (cm)
Mass		
1 pound (lb)	=	453.59 grams (g)
1 ounce (oz)	=	28.35 grams (g)
Volume		
1 liter (L)	=	1.057 quarts (qt)
1 gallon (gal)	=	3.785 liters (L)

For any one equivalence statement, two conversion factors may always be written. The conversion factors are reciprocals of each other. In the example below, the equivalence statement between kilometers (km) and miles (mi) has been used to write two different conversion factors.

Equivalence statement: $1 \text{ km} = 0.6214 \text{ mi}$ **Conversion factors:** $\frac{1 \text{ km}}{0.6214 \text{ mi}}$ and $\frac{0.6214 \text{ mi}}{1 \text{ km}}$

The “correct” conversion factor to use in a problem is the one that allows for correct cancelation of units to arrive at the final desired unit. In solving a problem sometimes one or more conversion factors may be used. The final answer should be rounded to the proper number of significant figures or digits. Remember that an equivalent statement is a definition; therefore, it should never be used to determine the number of significant figures in the final answer of a calculation.

Example

Convert 455 km to mi.

$$455 \text{ km} \times \frac{0.6214 \text{ mi}}{1 \text{ km}} = 282.737 \text{ mi} = \mathbf{283 \text{ mi}}$$

Equivalent statements may also be used to relate a metric prefix (e.g. centi-, mili-, deci-, kilo-) to its base (grams, liter, meters, etc.) in the International System of Units (commonly called the SI system or the metric system of measurement). Metric prefixes are listed in Appendix II.

Equivalence statement: $1 \text{ km} = 1 \times 10^3 \text{ m}$ **Conversion factors:** $\frac{1 \text{ km}}{1 \times 10^3 \text{ m}}$ and $\frac{1 \times 10^3 \text{ m}}{1 \text{ km}}$

Example

Convert 7.82 mi to m using conversion factors with units.

$$7.82 \text{ mi} \times \frac{1 \text{ km}}{0.6214 \text{ mi}} \times \frac{1 \times 10^3 \text{ m}}{1 \text{ km}} = 12,584.48664 \text{ m} = 12,600 \text{ m} = \mathbf{1.26 \times 10^4 \text{ m}}$$

Materials and Reagents Needed for Lab

Lab Equipment

10 mL graduated cylinder	Balance	Scoopula
150 mL beaker	Glass stir rod	Small test tube
Evaporating dish	Large test tube	Thermometer
100 mL graduated cylinder	Rubber stopper	Weigh boat
25 mL graduated cylinder	Ruler	250 mL beaker

Reagents

Water	Ice	Sodium chloride
-------	-----	-----------------

Safety Considerations

Be careful when handling glass containers as they can break. Balances should not be moved or shifted to different positions. If balances are moved too often, they become unlevelled and will not record the mass correctly. Pipets should never be placed in wash bottles, including the deionized water bottle. Water should always be poured into a beaker and pipetted from the beaker. Gloves must be worn for this lab.

Procedure

Part A: Lab Safety Sketch

1. Walk around the lab to find the following safety equipment. The numbers in parentheses represent the number of each type of equipment in the room.
2. Indicate on the Lab Safety Sketch (under Part A of the Data/Calculations Section) where each item is located. If an item is present in more than one place indicate all locations on the map.

Chemical waste container/ waste container area (2)

Emergency gas shut off valve (1)

Broken glass container (1)

Fire blanket (2)

Community sink (1)

Eyewash station (1)

Fire extinguishers (2)

Fume hood (1)

Safety shower (1)

Lab Phone (1)

First aid kit (1)

Deionized water (2)

Spill kits (4)

Coat racks/book bag shelf (2)

Part B: Lab Scavenger Hunt

1. Match the written description (below) with the picture of each item in **Table 1: Common Laboratory Equipment** in Part B of the Data/Calculation Section. Write the appropriate letter in the Purpose column next to each image.
2. Starting at your lab bench, find each item and then record its location in Table 1 under the location heading. Be specific in the describing where the item can be found (1st lab drawer, 2nd lab drawer, 3rd lab drawer, 4th lab drawer, bench top, fume hood, etc.)

Purpose of Lab Items

- A. Heat resistant mitt used to handle hot objects.
- B. A ceramic dish with a spout; used to heat a solution to dryness.
- C. Used to create a spark to light a Bunsen burner
- D. A circular concave piece of glass used to cover a beaker or hold solids.
- E. A clear glass container used to hold or transfer liquids.
- F. A plastic or glass cylindrical container used for measuring specific amounts of liquids.
- G. A plastic tube used to transfer small amounts of liquid; meant to be thrown away when finished after its use.
- H. A ceramic or plastic plate with multiple wells used to hold very small amounts of chemicals
- I. Device used to measure the mass of solids and liquids.
- J. Soap solution used to wash all glassware in lab.
- K. A small adjustable gas burner
- L. Measures the temperature of a liquid.
- M. Plastic bottle used to dispense water.
- N. A metal tool used to hold test tubes.
- O. An electrical heating unit; usually with a ceramic top and dial for adjusting temperature.
- P. Used to stir a liquid solution.
- Q. Plastic disposable dish used to temporarily hold solids.
- R. Metal tool used to transfer solids from one container to another.

- S. A narrow brush used to clean the inside of test tubes or narrow necked containers.
- T. A large container used to safely store and transport hazardous waste

Part C: Measurements

Mass Measurements

1. Locate a 250 mL beaker and a balance to use for this lab. Use the same balance for the remainder of the lab.
2. Make sure all the doors are closed on the balance and press the "Tare" button. The balance should now read 0.000 g. If it does not read zero, wait a moment ($\sim 15 - 20$ s) and press the tare button again. Repeat if necessary. This is called "taring" or "zeroing" the balance.
3. Open one of the balance doors and place the beaker on the balance. Close the doors and wait for the numbers to stabilize.
4. Record the mass of the beaker (with the correct unit) in **Table 2: Mass Measurements**.
5. Remove the beaker from the balance.
6. Repeat steps 2 – 5 for the 10 mL graduated cylinder and spatula.
7. Place a weigh boat on the balance and tare the balance with the weigh boat. The balance should now read 0.000 g with the weigh boat on it.
8. Using the scoopula add between 1.9 – 2.0 g of sodium chloride to the weigh boat. Any excess sodium chloride can go in the trash can.
9. Record the actual mass of sodium chloride in the same table.
10. Remove the weigh boat with the sodium chloride and set it aside as it will be used in the next section.

Temperature Measurements

1. Fill a 150 mL beaker more than halfway full with deionized water.
2. Submerge the bulb of the thermometer in the water. Be sure the bulb is fully submerged below the level of the water and not touching the bottom or the walls of the beaker. This means the thermometer should be held by hand and should never rest on the bottom or against the sides of the beaker when recording the temperature.

3. Record the temperature of the water in **Table 3: Temperature Measurements**.
4. Pour the water in the beaker down the drain.
5. Fill the 150 mL beaker with ice. Add enough deionized water to cover all the ice in the beaker.
6. Using a glass stir rod, stir the water in the beaker for 30 – 45 seconds.
7. Record the temperature of the ice-water mixture.
8. Use the salt (from the previous section) and add it to the ice-water mixture. Stir the mixture for about one minute with a glass stir rod.
9. Record the temperature.
10. Pour the solution down the drain with running water and set the beaker aside.

Volume Measurements

1. Fill a large test tube with deionized water until completely full.
2. Carefully, without spilling any water, pour the contents of the large test tube in a 100 mL graduated cylinder.
3. Record the volume of the liquid in **Table 4: Volume Measurements for Trial 1**. Pour the water out of the graduated cylinder into the sink.
4. Repeat this procedure two more times for Trials 2 and 3, respectively.
5. Completely fill a small test tube with deionized water.
6. Pour the contents of the small test tube in the 10 mL graduated cylinder.
7. Record the volume of the liquid in **Table 4**. Pour out the water in the graduated cylinder.
8. Repeat this procedure two more times for Trials 2 and 3, respectively.
9. Determine the volume for each trial and the average volume of both the large and small test tube.

Length Measurements

1. Use a ruler to measure the length of both lines in centimeters.
2. Record the length data in **Table 5: Length Measurements**.
3. Convert each of these lengths to meters and to inches. Record in **Table 5**.

Short Line



Long Line



First clean up, then complete all the remaining conversions and calculations.

Clean Up and Waste Disposal: All solutions can be poured down the drain with running water. Used weigh boats, gloves and disposable pipets or any excess sodium chloride should be placed in the regular trash. Wash all glassware with Alconox (soap) solution and then rinse with tap water. Dry off any wet glassware. Return all other items to their original locations.

Before Leaving Lab

- Make sure all items have been cleaned and put away.
- Check that drawers have been organized similar to the pictures in the lab drawers.
- Check the board for any additional clean-up or waste instructions.
- Wipe down the benchtops.
- Push in stools.
- Check with your instructor that your area has been sufficiently cleaned and reorganized.

Lab 2: Scavenger Hunt and Measurement – Pre-Lab

Name _____

Section _____

Date _____

Pre-lab Questions

Read through all the background information and the procedure of the lab before answering these questions. For calculations, work must be shown for full credit and a box should be placed around the final answer.

1. List two pieces of safety equipment in the lab.
2. Describe the procedure for how glassware should be cleaned at the end of lab.
3. What does it mean to “tare” a balance.
4. What should be done with the sodium chloride after its mass has been measured?
5. Describe how the thermometer bulb should be correctly used.

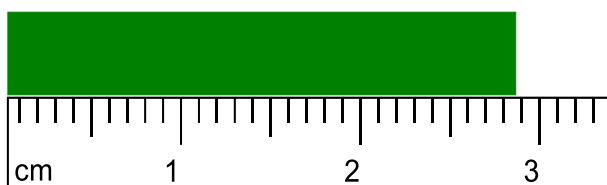
6. For each, indicate the certain and the uncertain digit(s) in each measurement.

24.1356 g Certain: _____ Uncertain: _____

7. In the table, indicate what (volume, length, mass, time, temperature, or density) is being measured based on the unit. Then write the abbreviation for each unit. Then The first one has been done as an example.

Unit	Type of Measurement	Abbreviation
Degrees Celsius	Temperature	°C
$\frac{\text{grams}}{\text{liter}}$		
Centimeter cubed		
Decisecond		
Millimeter		
Kilogram		

8. Record the measurement shown in the image.



Measurement: _____

9. Convert 8.50 inches to centimeters. Place a box around the final answer.

10. Convert 4678.9 g to kilograms. Place a box around the final answer.

Lab 2: Scavenger Hunt, Measurement and Unit Conversions – Data/Calculations

Name _____

Section _____

Drawer _____

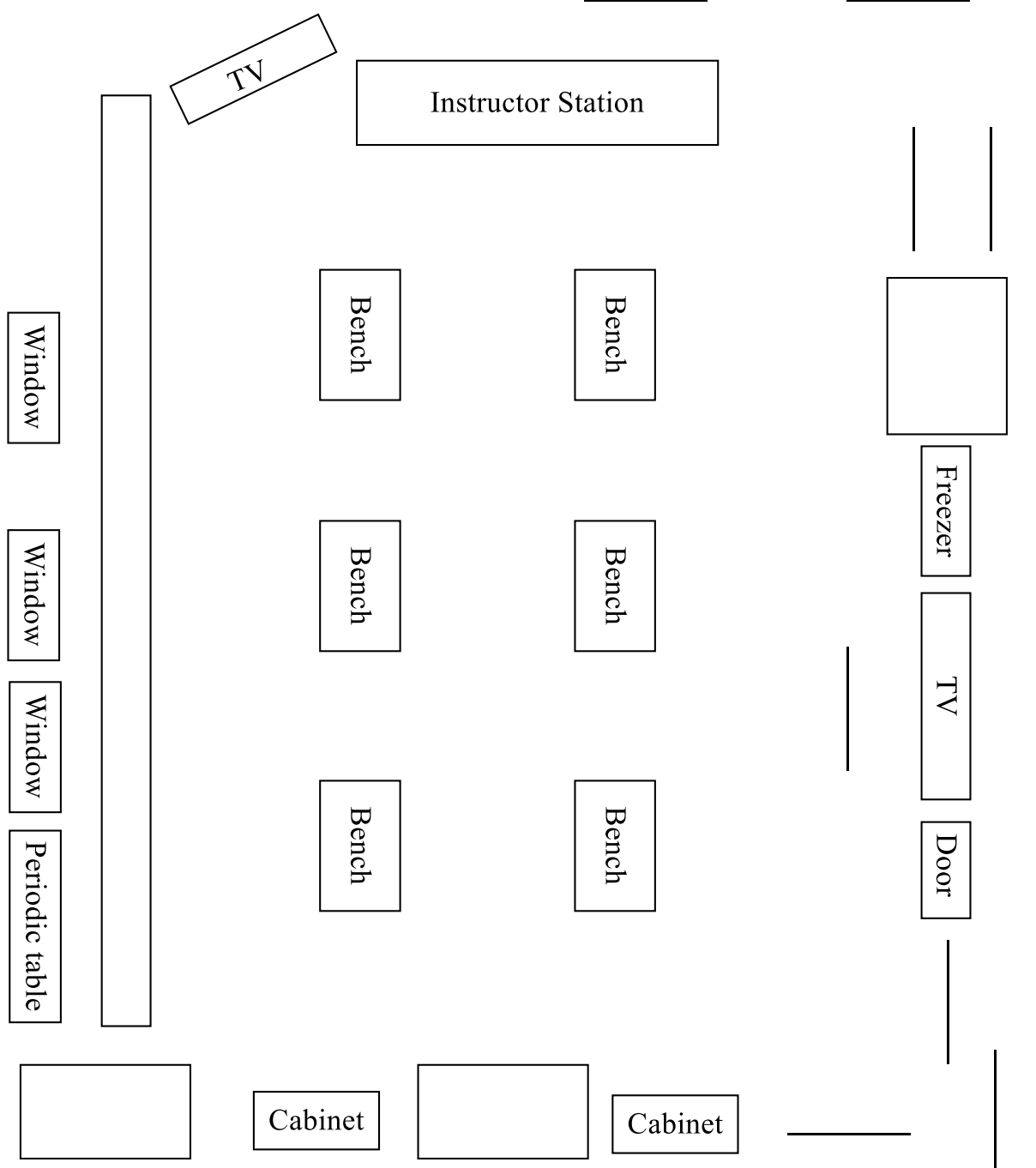
Balance _____

Date _____

Partner(s) _____

Part A: Lab Safety Sketch

Lab Floor Plan



Part B: Lab Scavenger Hunt

Table 1: Common Laboratory Equipment

Item	Image	Location	Purpose	Item	Image	Location	Purpose
Balance				Scoopula			
Beakers				Stir rod			
Bunsen Burner				Test Tube clamp			
Striker				Thermometer			
Ceramic spot plate				Chemical Waste Container			

Item	Image	Location	Purpose	Item	Image	Purpose	Description
Disposable Pipet				Hot hands			
Evaporating Dish				Test tube brush			
Graduated Cylinders				Wash Bottle			
Hot plate				Watch glass			
Alconox wash bottle				Weigh boat			

Part C: Measurements

Table 2: Mass Measurements

	Mass	Mass Data
A.	250 mL beaker	
B.	10 mL graduated cylinder	
C.	Spatula	
D.	Sodium chloride	

	Mass	Mass Calculations
E.	Convert the mass of sodium chloride to kilograms	
F.	Convert the mass of sodium chloride to pounds	

Table 3: Temperature Measurements

	Temperature	Temperature Data
A.	Room temperature water	
B.	Ice water	
C.	Salt, water, and ice mixture	

Table 4: Volume Measurements

	Volume	Volume Data		
		Trial 1	Trial 2	Trial 3
A.	Water in large test tube			
B.	Water small test tube			

	Volume	Volume Calculations
C.	Average volume of large test tube	
D.	Convert volume to liters	
E.	Average volume of small test tube	
F.	Convert volume to liters	

Table 5: Length Measurements

	Length	Length Data
A.	Length of short line	
B.	Length of long line	

	Length	Length Calculations
C.	Length of short line in meters	
D.	Length of short line in inches	
E.	Length of long line in meters	
F.	Length of long line in inches	

Lab 2: Scavenger Hunt and Measurement – Post-Lab Questions

Name _____

Section _____

Date _____

Post-lab Questions

1. Where is the eyewash station located in the lab? Why might this be the least used piece of safety equipment?
2. List two different types of lab safety equipment that can be used to put out fires.
3. For the equipment used in this lab, indicate to what place value the final measurement should be recorded.
A) Thermometer
B) Digital balance
C) 10 mL graduated cylinder
D) 100 mL graduated cylinder
4. The mass of a scoopula was measured to be 2.34 mL using a balance in the lab. Why is this an incorrect measurement?
5. Why should a pipet never be put into a wash bottle?

6. Identify each lab item and briefly describe its purpose.

Lab Equipment			
Name			
Purpose			

7. Briefly explain what effect adding salt to the ice water mixture had.

8. Measurements are used outside of lab in our everyday lives. Give two examples of measurements you have encountered this week.

9. Convert 40.62 gal to mL. Put box around final answer.

10. Convert 107 km to miles. Put box around final answer.

Lab 3: Density and Graphing Lab

Lab Objectives

- Determine the density of different items
- Use the density equation to solve for other variables
- Determine the slope of a line based on a graphical data
- Identify a value from a graph
- Identify the independent or dependent variable, the units or type of measurement from a set of data or a graph

Background Information

Matter can be described in terms of its physical or chemical properties. The chemical properties of a substance cannot be measured without changing the identity of a substance. For example, to determine if something is flammable, it must be burned, and a new substance is formed. Other examples of chemical properties include combustibility, radioactivity, or acidity. Physical properties can be measured without changing the identity of a substance. Determining the color of an object does not alter its identity. Melting point, boiling point, mass and volume are just a few other examples of physical properties.

Density is also a physical property and is expressed the ratio of mass to volume for a substance. Because molecules pack differently at different temperatures, the density of a substance also depends on the temperature of the sample. The unit for density must always involve a mass unit over a volume unit, for example $\frac{\text{grams}}{\text{milliliter}}$ or $\frac{\text{pounds}}{\text{gallons}}$.

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{m}{v}$$

Although mass is commonly measured using a digital balance, the volume may be measured in a variety of ways. Volume can be measured using an accurate measuring device like a graduated cylinder or a volumetric pipet. It can also be measured by volume difference or a change in volume (ΔV). This requires taking an initial and a final volume measurement. The difference between the two measurements gives the volume of the object or liquid.

$$\Delta V_{\text{object}} = V_{\text{final}} - V_{\text{initial}}$$

Volume can also be measured using the volume formula specific to an object.

volume of a block

$$V = l \times w \times h$$

volume of a cylinder

$$V = \pi \times r^2 \times h$$

volume of a sphere

$$V = \frac{4}{3} \times \pi \times r^3$$

Since pure substances have unique density values, measuring the density of a substance can help identify that substance. The densities (at 20°C) of several metals are given below in the density table.

Density Table

Metals	Density, $\frac{g}{cm^3}$
aluminum	2.702
boron	2.393
copper	8.960
gold	19.32
lead	11.34
silver	10.49
zinc	7.140

When working with volume units, keep in mind that 1 mL = 1 cm³.

Example

When a 25.752 g metal object is placed in 12.8 mL of water, the volume of water changes to 18.6 mL. Calculate the density of the object.

$$d = \frac{\text{mass}}{\text{volume}} \quad \text{volume} = V_f - V_i = 18.6 \text{ mL} - 12.8 \text{ mL} = 5.8 \text{ mL}$$

$$d = \frac{21.752 \text{ g}}{5.8 \text{ mL}} = 3.750345 \frac{\text{g}}{\text{mL}} = \mathbf{3.8 \frac{g}{mL}}$$

Graphing

A graph is a pictorial method used to quickly describe the relationship between variables. Line graphs describe the linear relationship between two variables, for example, mass and volume. When the mass and volume of a metal are graphed, the density of the metal can be determined. The equation for a line is found from the equation,

$$y = mx + b$$

Where y is the value found on the y-axis, m is the slope, x is value found on the x-axis and b is the y-intercept. Line graphs are also useful because once the equation of the best fit line is known, data can

be extrapolated (determine a value outside of the data range) or interpolated (determine a value that lies within the data range). When graphing volume on the y-axis and the mass on the x-axis, slope of the line will give the density of the sample.

All line graphs have the following features in common:

Title

The title of a graph is a brief description what has been plotted.

Axes and Labels

The horizontal axis on a line graph is the x-axis and refers to the independent variable, and the vertical axis on a line graph is the y-axis and refers to the dependent variable. Both axes should always be labeled with what is being measured and a unit. For example, the “Volume in mL” or “Volume (mL)” would be an appropriate way to indicate what the x-axis represents in a density graph.

Scale

A scale for each axis should be chosen such that the line of the graph is spread out over the majority of the graph and not clustered in just one small area of the graph. The scale should also be written so that numbers increase or decrease in a systematic way, 5 mL, 15 mL, 25 mL, etc. Often the range of data values may help to choose appropriate scales for a graph.

Lines and data points

Data points always have two coordinates, are plotted in the order (x, y) and should be clearly visible on the graph. In the case of a linear graph, a best fit line should be used. The intent of the best fit line is it “fits” all the data and is drawn as close as possible to all data points without all of the points falling too far above or below the line.

How quickly the data are increasing or decreasing will depend on the slope of line. The slope for a line is found from the following equation, commonly referred to as “rise over run”, or the change in y over the change in x.

$$\text{Slope} = m = \frac{\Delta y}{\Delta x} = \frac{y_2 - y_1}{x_2 - x_1}$$

When picking points to use in the slope equation, they must be points on the best fit line. This means that the points used to calculate the slope may or may not be data points. If a data point does not fall directly on the best fit line it cannot be used to calculate the slope of the line. Pick points that are far apart on the best fit line rather than very close together.

Once the slope is known, points on the best fit line can be used to determine the value of b and then to ultimately write the equation that describes the best fit line.

Materials and Reagents Needed for Lab

Lab Equipment

Ruler	Balance	10 mL graduated cylinder
100 mL graduated cylinder	150 mL beaker	50 mL beaker
Disposable pipet	Copper block	Aluminum cylinder

Reagents

Water	Unknown A
-------	-----------

Safety Considerations

Use caution when adding a metal cylinder or block to a glass or plastic container as they may break or crack the container. Balances should not be moved or shifted to different positions. Moving the balance can change the calibration or damage the spring or load cell which will cause errors in mass. Gloves must be worn for this lab.

Procedure

Part A: Determine the Density of Liquids

Density of Water

1. Tare the balance. Place an empty 10 mL graduated cylinder on the balance.
2. Record the mass of the empty graduated cylinder in **Table 1** for Trial 1.
3. Remove the cylinder from the balance.
4. Fill a 150 mL beaker about halfway full of deionized water. Use this water to add exactly 10 mL of water to the graduated cylinder using a disposable pipet.
5. If the meniscus of water is above or below the 10 mL mark, use the disposable pipet to adjust the water level until the bottom of the meniscus is on the top of the line and it measures 10.00 mL.
6. Record the volume of the liquid.
7. Tare the balance. Place the graduated cylinder back on the balance.
8. Measure and record the new combined mass.

9. Pour the water down the drain.
10. Repeat steps 1 – 9 two more times and record data in Trial 2 and 3, respectively.
11. Determine the density of water for each trial. Then calculate the average density of water.

Density of Unknown Liquid

1. Fill a 50 mL beaker about halfway full of Unknown A. Set it aside.
2. Tare a 10 mL graduated cylinder on the balance. This means the balance should read 0.000 g while the graduated cylinder is still on the balance. Leave the graduated cylinder on the balance.
3. Using a disposable pipet, add Unknown A dropwise to the graduated cylinder until approximately 5 g has been measured.
4. Record the exact mass of Unknown A in **Table 2**.
5. Remove the graduated cylinder from the balance.
6. Measure the volume of Unknown A in the graduated cylinder. Record the volume in the same table.
7. Pour Unknown A into a waste container.
8. Repeat steps 2 – 7 two more times for Trials 2 and 3 respectively.
9. Determine the density of Unknown A for each trial. Then calculate the average density of Unknown A. Record in table.

Part B: Determine the Density of Metal Objects

Density of Copper Block

1. Tare the balance.
2. Measure the mass of a copper block by placing it on the tared balance. Record the data in **Table 3**.
3. Use a ruler to measure the length, width and height of the block. Record the data for each.
4. Calculate and record the volume of the copper block using the length, width and height measurements.
5. Calculate and record the density of copper from this data.

Density of Aluminum Cylinder

1. Tare the balance.
2. Add an aluminum cylinder to the balance and record its mass in **Table 4**.
3. Use a 100 mL graduated cylinder and fill it at least halfway full with deionized water. Record the initial volume of water.
4. Add the aluminum cylinder to the graduated cylinder by tilting the graduated cylinder slightly on its side and sliding the aluminum down the side of the cylinder. Record the new volume of water.
5. Pour out the water and dry off the aluminum cylinder.
6. Repeat steps 1 – 5 two more times for Trials 2 and 3 respectively.
7. Determine the change in volume and density of the aluminum cylinder for each trial. Then calculate the average density of aluminum. Record in table.

Part C: Determining Density from a Graph

1. Plot the points from the table below onto the provided graph paper. For the graph, volume plotted should be on the x-axis and mass should be plotted on the y-axis.

Table 5: Mass and Volume Data
for Unknown B

Mass (g)	Volume (mL)
0.00	0.00
1.05	2.00
2.07	4.00
3.10	6.00
4.14	8.00
5.18	10.00

2. Pick an appropriate scale that allows the data points to spread out as much as possible. Make sure the data points cover at least 75% of the graph paper.
3. Include the point (0, 0) on the graph.

4. Use a ruler or straight edge to draw the best fit line of the graph.
5. Be sure to include all parts of the graph.
6. When finished, calculate the slope of your line. This slope represents the density of the unknown.

Clean Up and Waste Disposal: Return all objects (rulers, cylinders, blocks, etc.) to their original location. Any leftover unknown A not in the reagent bottle must go in the labeled waste container. Any water waste may be poured down the drain. Remove any sharpie from glassware. Clean all glassware with alconox/soap solution and rinse thoroughly with tap water. After drying the glassware, return items to their original location.

Before Leaving Lab

- Make sure all items have been cleaned and put away.
- Check the board for any additional clean-up or waste instructions.
- Wipe down the benchtops.
- Refill the water bottles.
- Push in stools.
- Check with your instructor that your area has been sufficiently cleaned and reorganized.

Lab 3: Density and Graphing Lab– Pre-Lab

Name _____

Section _____

Date _____

Pre-lab Questions

Read through all the background information and the procedure of the lab before answering these questions. For calculations, all work must be shown. All work shown must include the formula, the unrounded answer, and the correctly rounded final answer. Answers without work may not receive full credit.

1. Use the website <http://chemicool.com> to find the densities of the following elements.

Iron _____ Iodine _____

Mercury _____ Phosphorous _____

Nickel _____ Carbon _____

2. Is it OK to move the balances in lab? Briefly explain.
3. Briefly explain whether a temperature change would be considered a chemical or physical property.
4. What measuring device will be used to measure the volume of water in Part A?
5. What formula should be used to calculate the volume of copper in Part B?

6. What should be done with any leftover Unknown A at the end of the lab?
7. Describe the procedure for how to clean glassware at the end of lab.
8. What is the relationship between the unit milliliter (mL) and the unit centimeters cubed (cm³)?
9. What volume, in cm³, would 35.917 g of silver occupy? Use the information in the density table to answer.
10. What mass, in g, of boron would occupy a volume of 2.60 cm³? Use the information in the density table to answer.

Lab 3: Density and Graphing Lab – Data/Calculations

Name _____

Section _____

Drawer and Balance Number _____

Date _____

Partner(s) _____

Be sure to show all steps in calculations. Include formulas, units, unrounded answers and final rounded answers (with appropriate units) in the calculation table. Remember to include the appropriate unit(s) for all measurements. *Only the calculation work for Trial One must be shown.

Part A: Density of Liquids

Table 1: Water Density Data

Data		Water Density Data		
		Trial 1	Trial 2	Trial 3
A.	Mass of empty 10 mL graduated cylinder			
B.	Volume of water			
C.	Mass of 10 mL graduated cylinder and water			

Calculations		Water Density Calculations		
		Trial 1	Trial 2	Trial 3
D.	Mass of water*			
E.	Density of water*			
F.	Average density of water			

Table 2: Unknown A Density Data

Data		Unknown A Density Data		
		Trial 1	Trial 2	Trial 3
A.	Mass of Unknown A			
B.	Volume of Unknown A			

Calculations		Unknown A Density Calculations		
		Trial 1	Trial 2	Trial 3
C.	Density of Unknown A*			
D.	Average Density of Unknown A			

Part B: Density of Metal Objects**Table 3: Density Copper Block Data**

Data		Copper Density Data
A.	Mass of Copper Block	
B.	Length of Copper Block	
C.	Width of Copper Block	
D.	Height of Copper Block	

Calculations		Copper Density Calculations
E.	Volume of Copper Block	

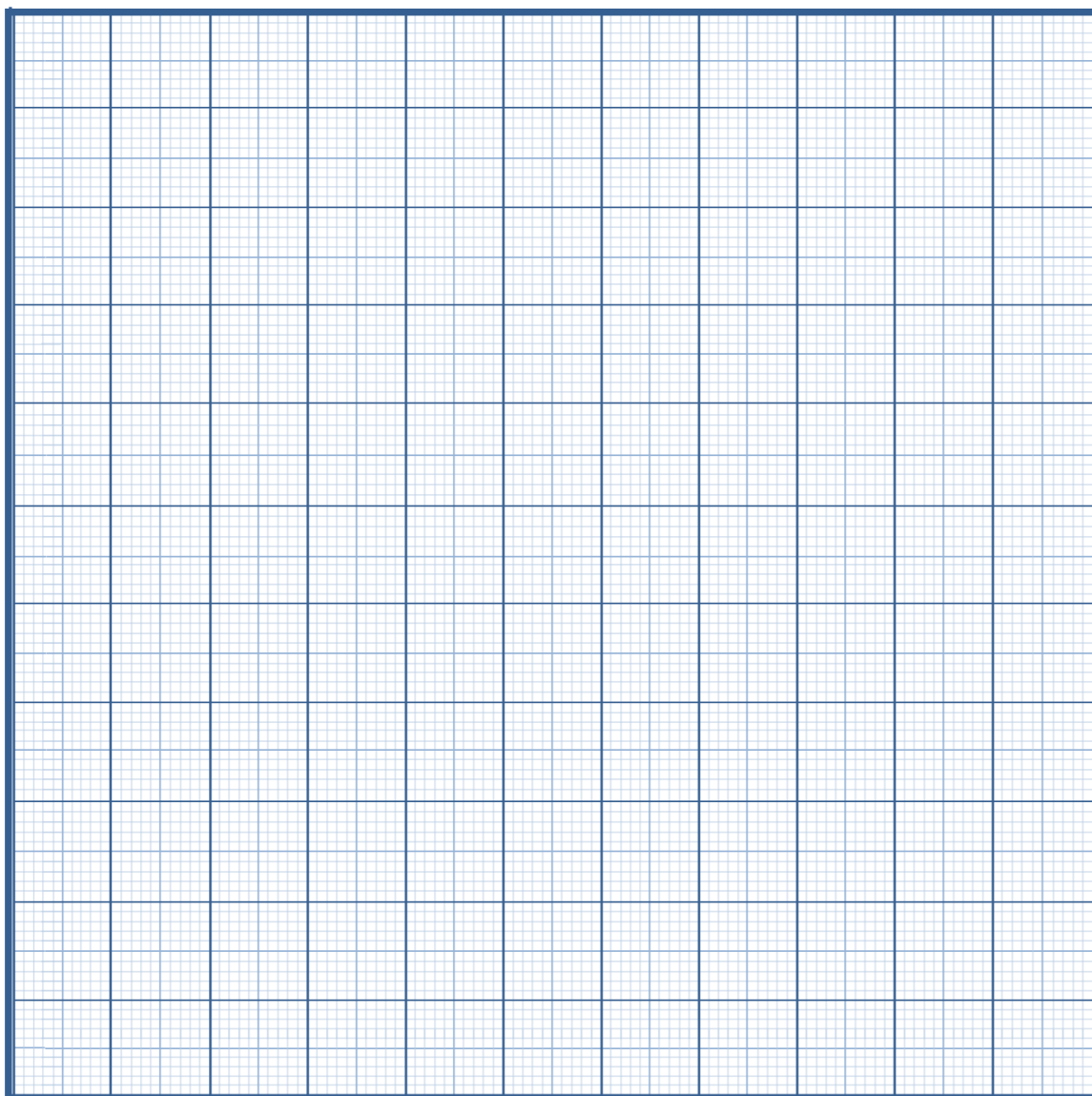
F.	Density of Copper Block	

Table 4: Density Aluminum Cylinder Data

Data		Aluminum Density Data		
		Trial 1	Trial 2	Trial 3
A.	Mass of Aluminum Cylinder			
B.	Initial Volume of Water			
C.	Final Volume of Water			

Calculations		Aluminum Density Calculations		
		Trial 1	Trial 2	Trial 3
D.	Change in Volume*			
E.	Density of Aluminum Cylinder*			
F.	Average Density of Aluminum			

Part C: Density from a Graph



1. Which variable is on the independent axis? _____
2. Which variable is on the dependent axis? _____
3. On the best fit line of the graph, circle two points that can be used to calculate the slope of the line and write their coordinates (x, y) below.

Point 1: _____

Point 2: _____

4. Calculate the slope (as a decimal not a fraction) of the line using the points above. Give the slope to two significant figures. Show work.

Lab 3: Density and Graphing Lab – Post-Lab Questions

Name _____

Section _____

Date _____

Post-lab Questions

For calculations, all work must be shown. Work shown must include the formula, the unrounded answer and the correctly rounded final answer. Answers without work may not receive full credit.

1. The true density of water at 25°C is 0.9970 $\frac{g}{mL}$. Use the average density you determined in this lab as the experimental value to calculate the percent error. Give the absolute value of the final answer.

$$\text{Percent error} = \frac{|\text{true value} - \text{experimental value}|}{\text{true value}} \times 100$$

2. Does your density match the true value? Explain why this may or may not be the case.

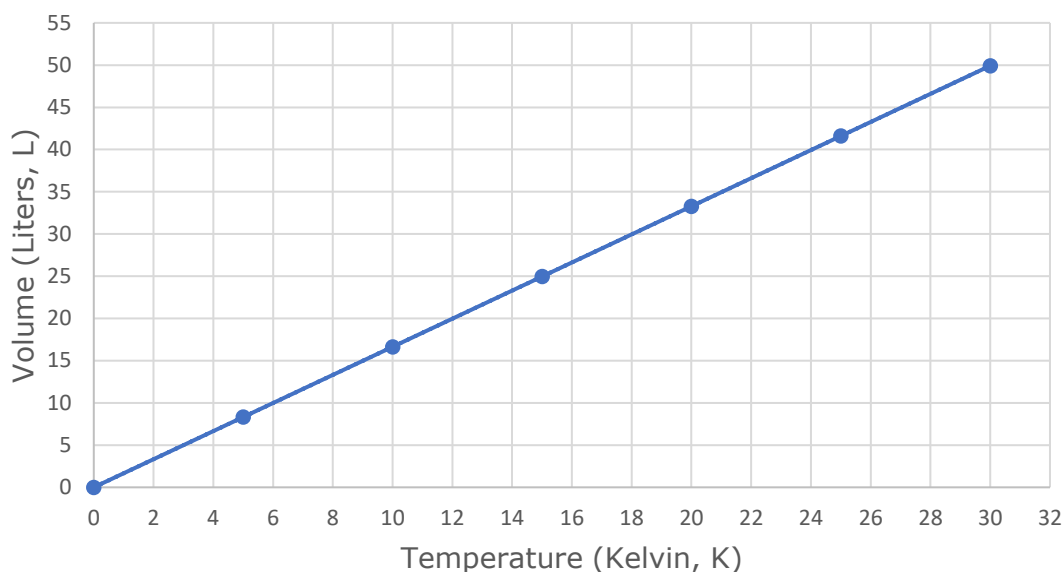
3. What was the density for Unknown A. _____

4. Based on your graph from Part C for Unknown B, solve for the y-intercept (to the hundredths place) and then determine the linear equation for the line. Show work.

5. Using the best fit line of your graph from Part C, what would be the volume, mL, of Unknown B if the mass is 12.8 g? Give answer to three significant figures. Show work.
6. Using the best fit line of your graph from Part C, what would be the mass, of Unknown B in g, if the volume is 36.2 mL? Give answer to three significant figures. Show work.
7. A certain liquid has a density of $3.00 \frac{\text{g}}{\text{mL}}$. Briefly explain what would happen to the copper block and the aluminum block when added to the liquid (float or sink). Assume neither reacts with the liquid.

8. Use the graph below to answer the following questions. Remember to show work.

Relationship of Gas #1



- A) Which variable is on the independent axis? _____
- B) Which variable is on the dependent axis? _____
- C) Place an "x" at the points (12, 20) and (24, 40) on the graph. Using those points, determine the slope (m) of the line to two significant figures.
- D) Find the y-intercept (b) and then write the equation for the line.
- E) Using the linear equation, calculate the volume, in L, when the temperature was 75 K. Give answer to the ones place.
- F) Using the linear equation, calculate the temperature, in K, when the volume was 84 L. Give answer to the ones place.

9. A 0.6908 g piece of foil was measured to be 4.00 cm wide, 3.12 cm long and 0.0205 cm tall. Determine the density of the foil.

10. An object has a mass of 0.02536 kg and a volume of 0.1578 L. Calculate the density, in $\frac{g}{mL}$, of this object.

Lab 4: Periodic Properties Lab

Lab Objectives

- Determine periodic trends for certain elements
- Determine properties of metals, nonmetals and semimetals
- Obtain information from the Periodic Table

Background Information

Atoms and Subatomic Particles

Trying to define matter goes as far back as at least 400 B.C. with the Greek philosopher Democritus. He suggested that matter could be broken down to the point where it eventually became indivisible or “atomos”. Hence the first written account for the basis of the word atom used in a chemical context. Atoms, however, are not quite indivisible; they can be broken down to even smaller pieces called subatomic particles. There are three fundamental subatomic particles: protons, electrons, and neutrons. Protons are positively charged particles that are found in the nucleus (the tiny dense center of an atom) and have a mass of one atomic mass unit (amu). Neutrons are also located in the nucleus and have a mass of one amu, but they do not have any charge. Electrons are negatively charged particles located outside the nucleus and are so small that they are considered to be “massless” when compared to protons and neutrons. All atoms are electrically neutral which means they have the same number of protons and electrons.

Every atom has a specific number of protons. There are 118 known elements and each element is unique with its own characteristics and properties. Hydrogen has just one proton and is assigned atomic number one. Helium has two protons and has atomic number two; and so forth for the remaining elements. The number of protons or the atomic number is therefore used to define the identity of an element. Each element consists of a one or two letter combination to represent the atomic symbol (or abbreviation) for that element.

Example

Atomic Symbols (Abbreviations) for Elements

Element:	Carbon	Bromine
Atomic Symbol:	C	Br

Some elements exist as isotopes. Isotopes are atoms of the same type of element with different numbers of neutrons. An isotope is identified by its mass number (the number of protons plus neutrons). The mass number is often written in the upper left-hand corner of the isotope. Carbon has three isotopes: ^{12}C , ^{13}C and ^{14}C . Even though the three carbon isotopes have different mass numbers, because they are all carbon they all have only six protons.

	mass number → 12 atomic symbol → C	¹³ C	¹⁴ C
Example	6 protons	6 protons	6 protons
	6 neutrons	7 neutrons	8 neutrons
	6 electrons	6 electrons	6 electrons

Atoms are electrically neutral which means they have the same number of protons and electrons. If the number of electrons differs from the number of protons, then it is no longer a neutral atom instead is called an ion. Ions are formed when an atom loses one or more electrons. Only electrons can be lost. If the number of protons changes, the atom becomes a different element not an ion. When electrons are lost, the atom becomes a positively charged ion and is called a cation. When electrons are gained, the atom becomes a negatively charged ion and is called an anion.

	Li	Li ⁺	O	O ²⁻
ions	atom	cation	atom	anion
	3 protons	3 protons	8 protons	8 protons
	3 electrons	2 electrons	8 electrons	10 electrons

When the lithium atom, Li, (three protons, three electrons) loses one electron it becomes the lithium ion and has the formula Li^+ (three protons, two electrons). Because there are more protons than electrons the ion is positively charged (a cation). When the oxygen atom, O, (eight protons, eight electrons) gains two electrons it becomes negatively charged and has the formula O^{2-} (eight protons, ten electrons). Because there are now fewer protons than electrons the ion is negative (an anion).

The Periodic Table

The Periodic Table is a useful tool for organizing the 118 elements that are currently known. The current version of the Periodic Table is based on the model originally proposed by Dimitri Mendeleev in the 1870s. Although Mendeleev is often called the “Father of the Periodic Table”, he was not the first person who attempted to organize the elements. In 1789, Lavoisier organized a list of 33 “simple substances” which were arranged by element type (gases, metals, nonmetals, and earths). In 1862 De Chancourtois proposed a descending spiral of elements according to their atomic weights. The elements were then placed in vertical groups with properties like the ones above and below it. In 1865 Newlands grouped elements with similar properties in eight columns. Similar organizational techniques were used by Mendeleev. He also arranged elements by increasing atomic mass and put elements with similar properties in the same group. The reason Mendeleev gets the credit for the periodic table, however, is because he could use his model of the table to predict the properties of elements not yet discovered. In his model he left gaps, or holes, in the columns for information that seemed to be “missing”. At the time Mendeleev created his table the elements silicon and tin were known, but not the element germanium. He proposed that there was an element that was like silicon, which he called eka-silicon (like silicon), that was yet to be discovered. He predicted that this eka-silicon (E) element would be a dark grey metal, have a density of 5.5 g/mL and a high melting point. When eka-silicon reacted with oxygen he predicted it would form a compound with the formula EO_2 . Almost a decade later germanium was discovered by Clemens Winkler and was found to match very closely with the properties predicted by Mendeleev. Germanium was indeed found to be a grey-white metal, have a density of 5.32 g/mL and a melting point of more than 900°C. When germanium reacts with oxygen it forms the compound germanium dioxide, or GeO_2 all of which were consistent with Mendeleev’s predictions. It is due to the predictive nature of Mendeleev’s table that he receives the credit for the design of the modern periodic table.

Today the elements in the periodic table are arranged by increasing atomic number (which indicates the number of protons in an atom). The elements are also arranged into 18 groups (vertical columns) and 7 periods (horizontal rows). Elements with similar physical and chemical properties are found in the same column called a group (or a family). For example, the metals in Group 1 (1A) are called Alkali Metals. The elements in that group all form 1+ cations (Li^+ , Na^+ , etc) and are very reactive with water. Although hydrogen is in Group 1, it is not a metal, but rather a nonmetal. It is the only nonmetal on the left of the periodic table. The elements in Group 2 (2A) are the Alkaline Earth Metals and they all form a 2+ cation (Mg^{2+} , Ca^{2+} , etc.) but are less reactive with water. The elements in Group 16 (6A) are the Chalcogens and form 2- ions, those in Group 17 (7A) are the Halogens and form 1- ions and the elements in Group 18 (8A) are the Noble Gases which do not form any ions. Not every column in the periodic table has a specific name, and for those groups they are often referred to as a “family”. The elements in Group 14 (4A) are said to be in the Carbon Family, while those in Group 15 (5A) are in the Nitrogen Family. Different regions of the periodic table also have specific names. The metals in Groups 3 – 12 are all called the Transition Metals and elements with atomic numbers 57 – 70 and 89 – 102 are all called Inner Transition Metals.

Elements exist in different physical states: solids, liquids, and gases. There are only two elements that are liquid at room temperature, mercury, and bromine. All the elements in group 18 (8A), hydrogen, nitrogen, oxygen, fluorine, and chlorine are all gases. The rest of the elements are solids or are made in a laboratory setting. Elements with atomic numbers 93 – 118 are made synthetically in a laboratory rather than occur naturally. Those elements often exist for only fractions of a second when made and therefore are not really classified as a solid, liquid or gas.

Finally, elements may also be classified as either a metal, nonmetal, or semimetal (or metalloids). Metals comprise most of the elements in the periodic table. Regardless of their identity, metals tend to share similar properties. Metals are usually solid and tend to have high melting and boiling points, are shiny and good conductors of heat and electricity. Some, but not all, metals are magnetic. Nonmetals on the other hand appear in a variety of physical states, are brittle or dull and are poor conductors. Their melting and boiling points are significantly lower than that of metals. There are no magnetic nonmetals. Semimetals have some metallic and some non-metallic properties and will be unique to the semimetal. Germanium is a semimetal that is shiny and conductive like a metal but is brittle and breaks easily like a nonmetal.

Properties of Metals, Nonmetals and Semimetals

Metals	Nonmetals
Most exist as solid	Exist as solid, liquid or gas
High melting points	Low melting points
Shiny	Dull
Good conductors of heat and electricity	Poor conductors of heat and electricity
Malleable (shapes) and ductile (thin wires)	Brittle
Can form cations	Can form anions
Metalloids (or semimetals) have properties of both, and each is unique	

Semimetals can be identified based on their placement in the periodic table. They are the elements, not including aluminum, that appear directly above and below the heavily outlined stair step that is

often seen in the periodic table. Metals are all the elements to the left of this semimetal staircase, and nonmetals are all to the right of the staircase.

An element's placement in the periodic table can provide valuable information that can be used to predict their chemical, physical, and atomic properties. Elements in similar columns or rows exhibit similar patterns in terms of their size (measured through the radius of an atom or ion), types of ions formed, density, melting points, etc. Understanding the structure of atoms and some of the periodic trends that exist in the Periodic Table are the focus of this lab.

Materials and Reagents Needed for Lab

Lab Equipment

Magnet	Periodic table	Colored pencils (or markers/pens)
--------	----------------	--------------------------------------

Lab Kit Reagents

Aluminum	Bismuth	Carbon
Cobalt	Copper	Gallium
Iodine	Iron	Magnesium
Nickel	Selenium	Silicon
Sulfur	Tin	Zinc

Safety Considerations

None of the vials should be opened for any reason. Observation of each element should be made by looking through the glass vial. If a vial breaks, stop what you are doing and immediately notify the instructor right away. Iodine is a severe skin/eye irritant and must stay in the hood during the lab. Bismuth can cause damage to the central nervous system upon prolonged exposure. Selenium powder is very hazardous if ingested. Carbon powder should not be inhaled or go down drains. Iron, sulfur, aluminum, cobalt, magnesium, tin, copper, silicon, nickel and zinc are all eye and skin irritants. Gallium can cause severe eye damage and can quickly react with other metals (including jewelry) causing irreversible damage. Gloves must be worn when handling all reagents.

Procedure

Part A: Organization of the Periodic Table

On the blank periodic table in Part A of the Data/Calculation section label and color code each section as follows. Then answer the questions that pertain to the periodic table.

1. Label each of the groups 1 – 18 on the periodic table.
2. Label each of the periods 1 – 7 on the periodic table.
3. Indicate the following areas of the periodic table as follows.
 - a. Leave hydrogen white but color the Alkali Metals red.
 - b. Color the Alkaline Earth Metals orange.

- c. Color the semimetals black.
- d. Color the nonmetals in the Chalcogen group yellow.
- e. Color the Halogen group green.
- f. Color the Noble Gas Metal group blue.
- g. Color the Transition Metal group purple.
- h. Color the Inner Transition Metal group pink.

Part B: Atoms and Subatomic Particles

1. On each image of an atom, indicate the location and number for the given subatomic particles for the given isotope. For protons use the symbol "+", for electrons use the symbol "-" and for a neutron use the symbol "0". Then write the appropriate isotope in the form of ${}^AX^Y$.
2. Use the indicated isotope and periodic table to determine the number of subatomic particles for each element in **Table 1** of the Data/Calculation section.
3. Then indicate the mass number and the atomic number for each isotope.

Part C: Properties of Metals, Nonmetals and Semimetals

1. Obtain a sample kit and write the symbol for each element in **Table 2**.
2. **Do not open any of the vials.** Observe and record the physical appearance of each element. Observations should be made only by looking through the walls of the vial. Gallium is located at the instructor station and iodine is in the fume hood. Describe the physical state (solid, liquid, or gas), color (white, yellow, black, etc.), luster (shiny, glittery, dark, dull, etc.), and appearance (crystalline, powder, sheet, smooth, flakes, etc.).
3. **Do not remove the samples from the vial.** Test the magnetic ability of each element by running a magnet along the outside of the glass vial. If the sample is attracted by the magnet, circle Yes, if not, circle No.

Part D: Graphing Periodic Trends

Trends within a Period

1. Use a periodic table to find the atomic number for each element in **Table 3**, then record the value.

2. Choose an appropriate scale to plot the relationship between atomic number (x-axis) and atomic radius (y-axis) on Graph # 1. Be sure to include all the other appropriate parts (title, axes, labels, scales, etc.) of the graph. The axes do not necessarily need to start at (0, 0).
3. Use a pencil to connect the points on the graph. The line should not go through (0,0) and will not be a straight line.

Trends Down within a Group

1. Use the same procedure as above to find atomic number and graph the relationship between the atomic number and size for each element in **Table 4**.
2. Use a pencil to connect the points on the graph. The line should not go through (0,0) and will not be a straight line. The axes do not necessarily need to start at (0, 0).
3. Use the same procedure as above to find atomic number and graph the relationship between the atomic number and density for each element in **Table 5**.
4. Use a pencil to connect the points on the graph. The line will go through (0,0) and should be fairly close to a straight line.

Clean Up and Waste Disposal: Return all items to their appropriate locations.

Before Leaving Lab

- Make sure all items have been cleaned and put away.
- Check the board for any additional clean-up or waste instructions.
- Wipe down the benchtops.
- Push in stools.
- Check with your instructor that your area has been sufficiently cleaned and reorganized.

Lab 4: Periodic Properties Lab – Pre-Lab

Name _____

Section _____

Date _____

Pre-lab Questions

Read through all the background information and the procedure of the lab before answering these questions. For any calculations, all work must be shown. Work shown must include the formula, the unrounded answer, and the correctly rounded final answer. Answers without work may not receive full credit.

1. Describe the procedure that will be used to determine if an element is magnetic in this lab.
2. Describe what should be done if one of the vials breaks.
3. Which vial(s) can be opened during the lab? _____
4. What are the three subatomic particles?
5. For each element write the atomic number.
Hydrogen _____ Uranium _____ Chromium _____
6. List one element that has properties similar to helium. _____
7. What element is in Group 13 and Period 4? _____

8. Describe how elements are arranged in the modern version of the periodic table.

9. Write the atomic symbol (abbreviation) for each element.

Potassium _____

Tin _____

Iron _____

10. Use the periodic table to classify each element as a metal, nonmetal or a semimetal.

Boron _____

Lithium _____

Nitrogen _____

Sulfur _____

Lab 4: Periodic Properties Lab – Data/Calculations

Name _____

Section _____

Drawer _____

Balance _____

Date _____

Partner(s) _____

Be sure to show all steps in calculations when necessary. Include formulas, units, unrounded answers and final rounded answers (with appropriate units) in the calculation table.

Part A: Organization of the Periodic Table

[illegible]

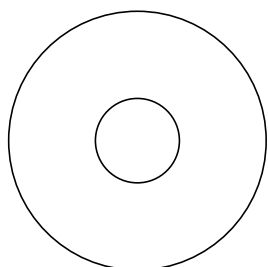
1. Which group(s) will only form 2+ cations? _____
2. Circle the element(s) that will only form a 3- ion. Na Ga P Cl He U
3. What is the name of the transition metal in period 5 and group 11?
4. Explain the relationship between the number of protons and electrons in each.
 - A) An atom
 - B) A cation
 - C) An anion

Part B: Atoms and Subatomic Particles

Examples of Isotopes

Isotope 1

6 protons, 6 electrons, 6 neutrons

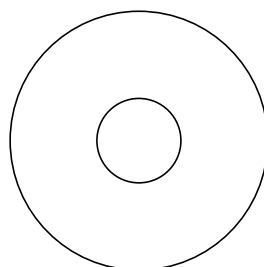


isotope:

atom anion cation

Isotope 2

3 protons, 2 electrons, 4 neutrons

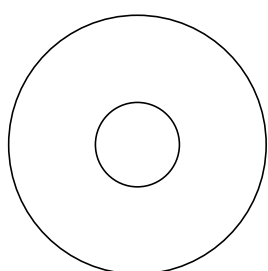


isotope:

atom anion cation

Isotope 3

8 protons, 10 electrons, 9 neutrons

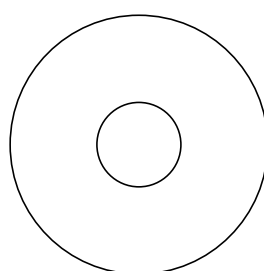


isotope:

atom anion cation

Isotope 4

5 protons, 5 electrons, 7 neutrons



isotope:

atom anion cation

Table 1: Isotopes of Different Elements

Element	³⁷ Cl	³⁰ Si	¹⁴ N	³⁹ K ⁺	³³ S ²⁻
Protons					
Electrons					
Neutrons					
Mass number					
Atomic number					

Part C: Properties of Metals, Nonmetals and Semimetals

Table 2: Elemental Properties

	Atomic Symbol	Physical state	Color	Luster	Appearance	Magnetic (circle one)
Aluminum						Yes No
Bismuth						Yes No
Carbon						Yes No
Cobalt						Yes No
Copper						Yes No
Gallium						Yes No
Iron						Yes No
Iodine						Yes No

	Atomic Symbol	Physical state	Color	Luster	Appearance	Magnetic (circle one)
Magnesium						Yes No
Nickel						Yes No
Selenium						Yes No
Silicon						Yes No
Sulfur						Yes No
Tin						Yes No
Zinc						Yes No

Part D: Graphing Periodic Trends

Table 3: Atomic Radii Trends Across a Row

Element	Atomic Number	Radius (pm)
Lithium		128
Beryllium		96
Boron		84
Carbon		76
Nitrogen		71
Oxygen		66
Fluorine		57

Graph #1

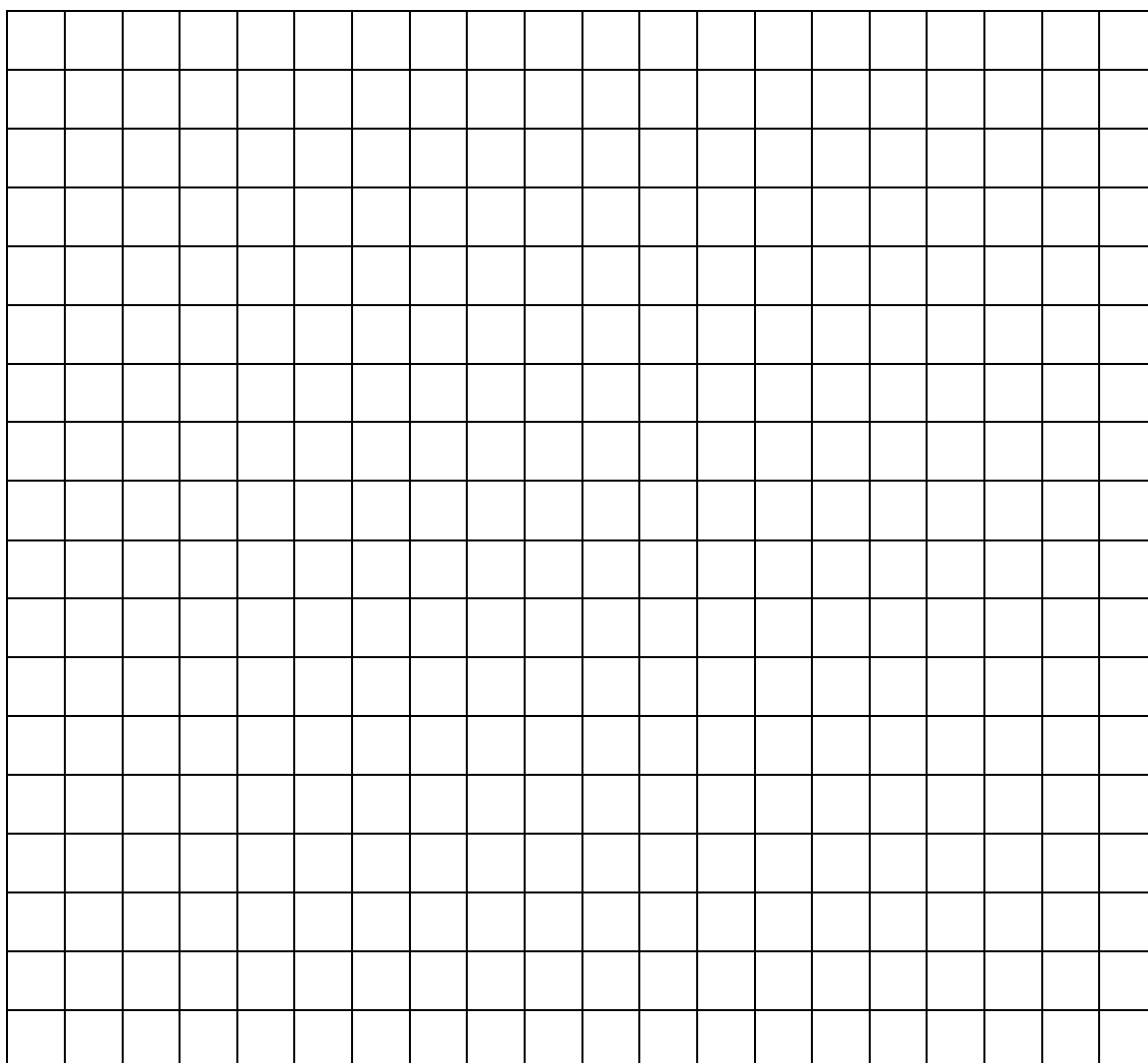


Table 4: Atomic Radii Trends Down a Column

Element	Atomic Number	Radius (pm)
Beryllium		96
Magnesium		141
Calcium		176
Strontium		195

Graph #2

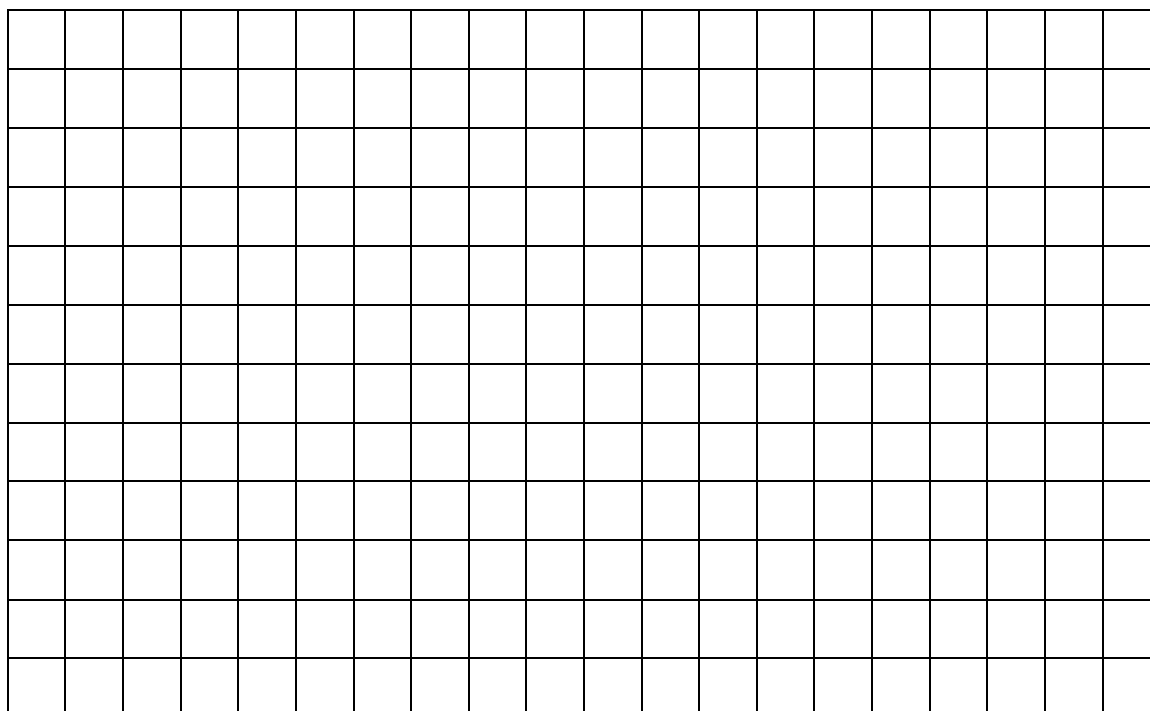
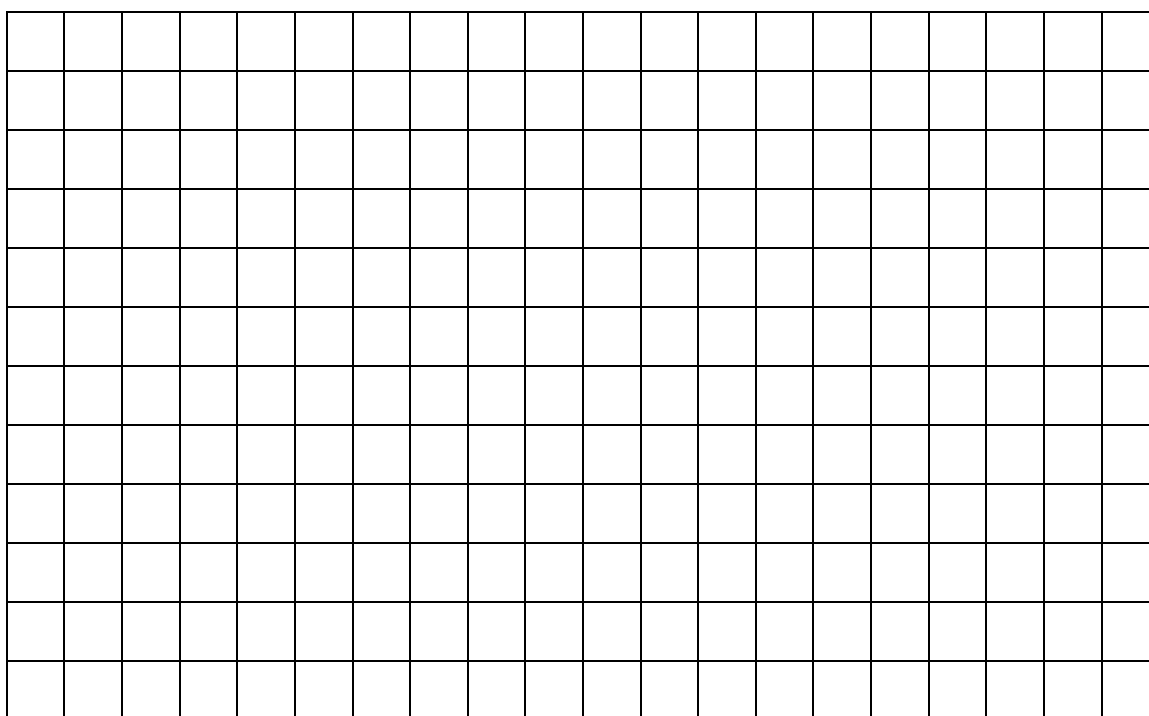


Table 5: Density Trends Down a Column

Element	Atomic Number	Density ($\frac{g}{cm^3}$)
Helium		1.79
Neon		9.00
Argon		17.8
Krypton		37.0
Xenon		58.5

Graph #3



Lab 4: Periodic Properties Lab – Post-Lab Questions

Name _____

Section _____

Date _____

Post-lab Questions

For calculations, all work must be shown. Work shown must include the formula, the unrounded answer the correctly rounded final answer. Answers without work may not receive full credit.

1. Write the formula of the isotope that has the following subatomic particles. Be sure to include the mass number or any charge for each isotope.

A) 19 protons, 19 electrons, 22 neutrons

B) 8 protons, 10 electrons, 10 neutrons

C) 56 protons, 54 electrons, 30 neutrons

2. How many electrons will each of the following have?

A) Ar

B) Rb^+

C) P^{3-}

3. Which metals were magnetic in Part C? _____

4. Give the name of each periodic group.

Group 1: _____

Group 2: _____

Group 17: _____

Group 18: _____

5. What is the most common physical state of an element? _____
6. Which elements in Part C were consistent as having metallic properties. Briefly explain.
7. Which elements in Part C were consistent as having nonmetal properties. Briefly explain.
8. Based on the line drawn for Graph #1, what trend was observed regarding the relationship between atomic size and atomic number?
9. Based on the line drawn for Graph #2, what trend was observed regarding the relationship between atomic size and atomic number?
10. Answer the following questions about the elements: As Br Ga
- ii. Are they in the same group or period?
- iii. Based on the trends observed in the periodic table, rank them in atomic size from smallest to largest atomic radius.
- _____ < _____ < _____
- Answer the following questions about the elements: Br Cl F
- iv. Are they in the same group or period?
- v. Based on the trends observed in the periodic table, rank them in atomic size from smallest to largest atomic radius.
- _____ < _____ < _____

Lab 5: Electron Configurations and Light Lab

Lab Objectives

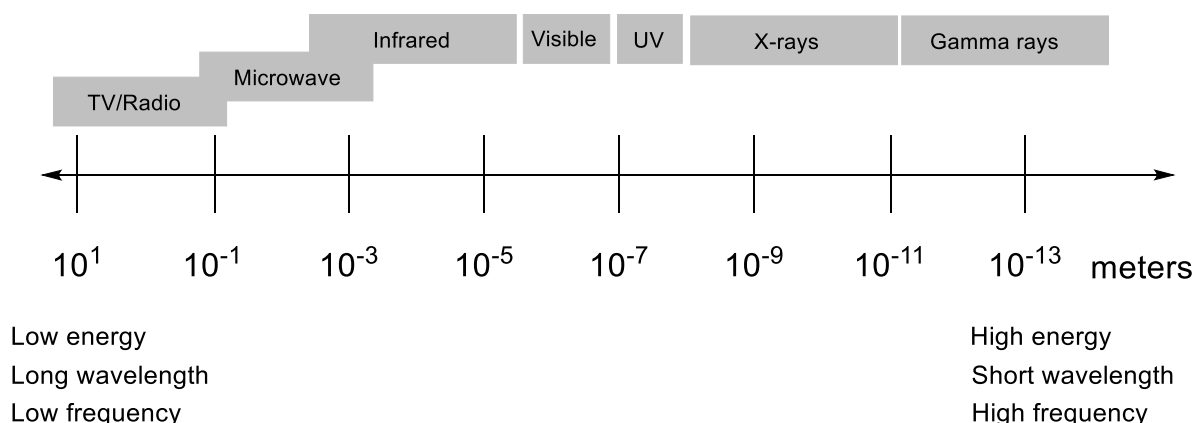
- Observe and predict the colors of metal cations in a flame test
- Calculate frequency and wavelength of different types of light
- Write or recognize the electron configuration (expanded and noble gas) for an atom or ion
- Identify core and valence electrons
- Compare and calculate wavelength, frequency and energy of light

Background Information

Light and Energy

Light is a complicated topic that cannot easily be described in a single sentence. One way to describe light is as energy in the form of electromagnetic radiation. Electromagnetic radiation varies from low energy radiation (Radio waves) to very high energy radiation (Gamma rays) and is displayed on the Electromagnetic Radiation (EMR) Spectrum (see Figure 1). The EMR spectrum shows the different types of energy ranked from the lowest energy (left side) to highest energy (right side).

Figure 1: Electromagnetic Radiation (EMR) Spectrum



Light has both wave-like features (e.g. wavelength, frequency) and particle-like features (e.g. photons, energy). In a vacuum, all light travels at the same constant speed, the speed of light, and is equal to $3.00 \times 10^8 \frac{m}{s}$. Although the speed of light (c) is constant, light can vary in terms of its wavelength (λ , pronounced like “lamb-duh”) and frequency (ν , pronounced like “new”).

speed of light equation

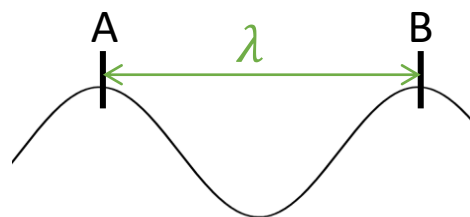
$$c = \lambda \times \nu$$

The wavelength (λ) of light is the distance between two consecutive peaks on adjacent waves. In the example below, the wavelength is the distance measured between Position A and Position B on the

wave. Because wavelength is a distance measurement, its units are commonly given in meters or nanometers (nm). When doing calculations that involve the speed of light, the unit of wavelength must be in meters (m).

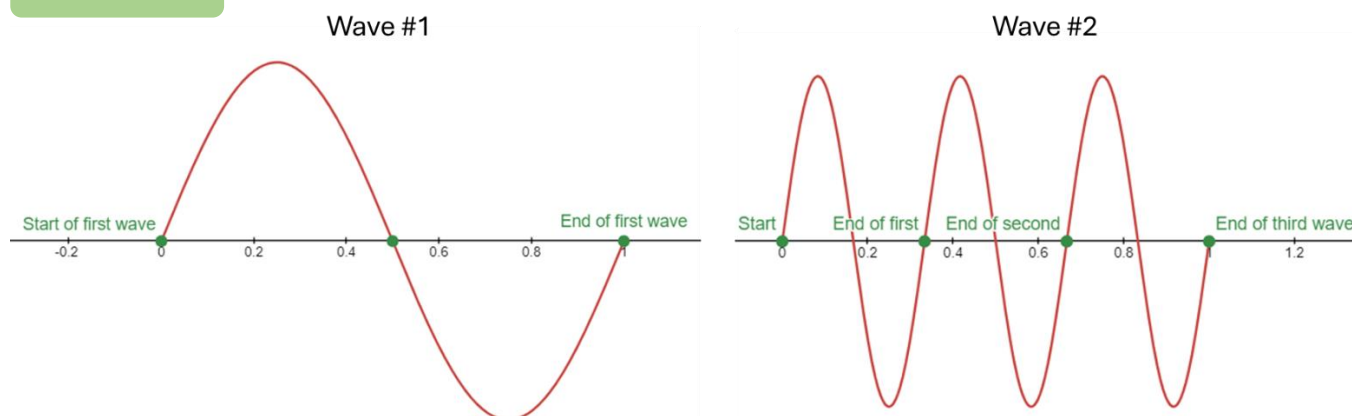
Example

The wavelength is the distance measured from position A to position B on the wave.



The frequency (ν) of a wave indicates the number of complete cycles that occur in one second. The units of frequency are given in Hertz (Hz), reciprocal second (s^{-1}), or per second, $\frac{1}{\text{s}}$. If one cycle is completed in one second, the wave will have a frequency of one hertz (1 Hz or 1 s^{-1} or $1 \frac{1}{\text{s}}$). If 100 cycles occur in one second, then the frequency is 100 hertz (100 Hz or 100 s^{-1} or $100 \frac{1}{\text{s}}$). A wave that completes many cycles per second is said to have a high frequency and a low frequency when only a few waves are completed.

Example



Wave #1 shows a complete wave cycle from the beginning of the wave until the end. On Wave #2, each circle marks the start of a new wave. In this time frame, 3 cycles have been completed. This represents a wave with a frequency of 3 Hz.

Wavelength and frequency are inversely related to each other. This means as one becomes larger, the other becomes smaller. So, waves with a long (large) wavelength have a low (small) frequency, but waves with a short (small) wavelength have a high (large) frequency. If either frequency or wavelength is known, the other variable can always be calculated since the speed of light is constant.

Example

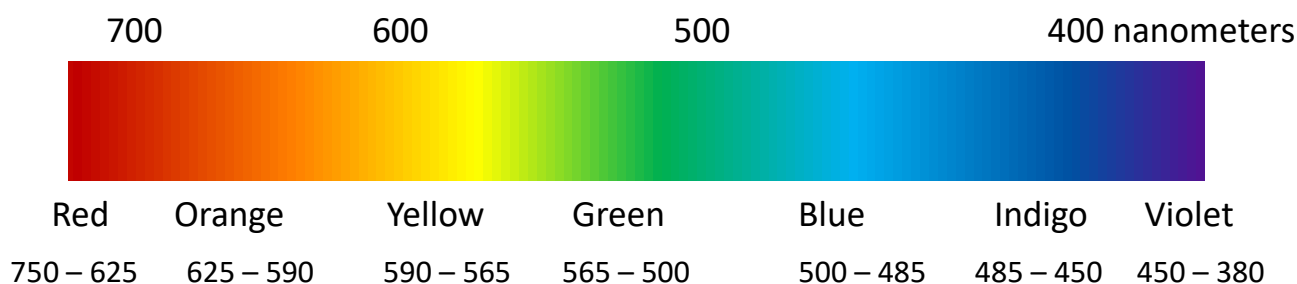
Calculate the frequency, in Hz, of light that has a wavelength of $2.6 \times 10^{-2} \text{ m}$.

$$c = \lambda \times \nu \quad \text{rearranges to} \quad \nu = \frac{c}{\lambda}$$

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \frac{\text{m}}{\text{s}}}{2.6 \times 10^{-2} \text{ m}} = 1.15385 \times 10^{10} \frac{1}{\text{s}} = \mathbf{1.2 \times 10^{10} \text{ Hz}}$$

Light visible to the human eye falls in a small region of the EMR spectrum between infrared (IR) and ultraviolet (UV). The different colors that are seen in a rainbow, for example, are the result of light having different wavelengths. The wavelength for the visible spectrum ranges from 750 nm to roughly 380 nm. A colored version of this spectrum may be found in Appendix VI.

Figure 2: Visible Spectrum



When atoms or ions temporarily absorb energy (through heat for example), their electrons can move from lower energy levels to higher energy levels. Eventually, the electrons release this energy and move back down to the lower energy level. When the energy that is released occurs within the visible region, it can be seen in the form of a color. Many metals when heated often produce a distinctive flame color that can then be used to identify them. For example, cesium metal produces a light purple or lilac flame.

Electron Configurations

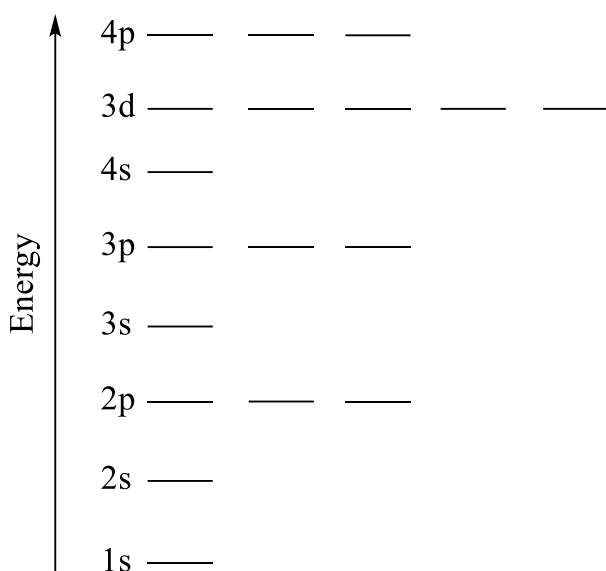
The electrons in each atom are arranged in a very specific manner and are described by their electron configuration. Electron configurations show the arrangement of electrons in orbitals (three-dimensional probability regions of space) for an atom or ion.

Electrons can occupy different energy levels in an atom. These different energy levels are called principle quantum numbers. The principle quantum numbers (n) indicate sequentially different energy levels that get larger with each new number and represent a farther distance from the nucleus. The lowest energy level, which is closest to the nucleus, is designated as $n = 1$. The next level is $n = 2$, which is larger and higher in energy (than $n = 1$). The next level is $n = 3$, which is even larger and higher in energy and so on with each new level.

Each energy level can contain one or more sublevels. The sublevel, represented by a letter, designates the overall shape of the three-dimensional region of space where the electron can be found. An s-sublevel is characterized by a spherical shape and has only one type of orbital (single line in Figure 3). The p-sublevel is commonly said to have a dumbbell shape and has three types of orbitals (three lines in Figure 3). The d-sublevel has a characteristic cloverleaf shape and has five types of orbitals. There is an additional shape associated with the d-sublevel, but for this course, only the cloverleaf shape will be emphasized. The f-sublevel is more complicated to describe but has seven types of orbitals and is not further discussed in this lab.

Each orbital can hold up to two electrons. Since each orbital can hold up to two electrons, an s-sublevel can have a maximum of two electrons (one orbital), a p-sublevel can up to six electrons (three orbitals), and a d-sublevels can hold up to ten electrons (five orbitals).

Figure 3: Orbital Filling Diagram



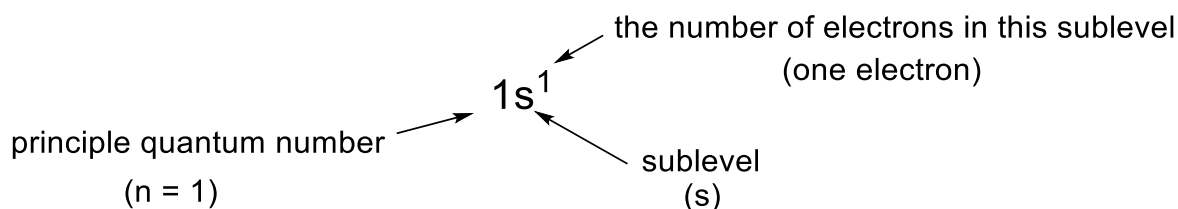
The 1s orbital is the lowest energy level. This is the first level that electrons can fill. Once the 1s level is full, electrons must fill the next lowest energy level, 2s. This process repeats until all electrons have been used.

NOTE: This is only a partial orbital filling diagram. There are additional orbitals beyond the 4p level.

When all this information is combined (principle quantum number levels, sublevels and electrons) an electron configuration can be written. The principle quantum number is written first, followed by the sublevel and the electrons in that sublevel are written as superscripts to the right of the sublevel.

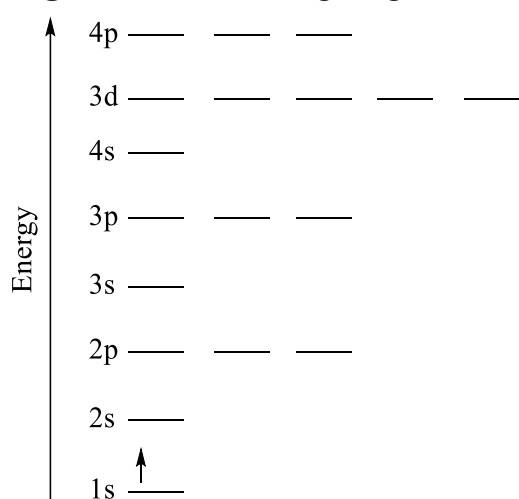
Example

The electron configuration of a hydrogen atom



A hydrogen atom has only one electron. This one electron must go in the lowest energy level (1s) first. The electron configuration for hydrogen is therefore $1s^1$.

Figure 4: Orbital Filling Diagram for H

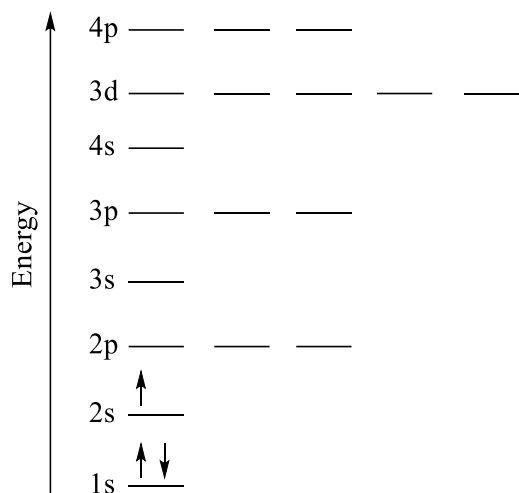


The one electron in hydrogen (represented by an arrow) must go in the lowest energy level first.

The electron configuration for a hydrogen atom is $1s^1$

When completing an orbital diagram, electrons are represented by arrows. The first arrow will be written pointing upward (spin up) and if a second arrow is present it will be written pointing downward (spin down). Keep in mind that each orbital (line) can only hold two electrons. Once the orbital is full, the next electron must go into the next lowest energy orbital.

Figure 5: Orbital Filling Diagram for Li

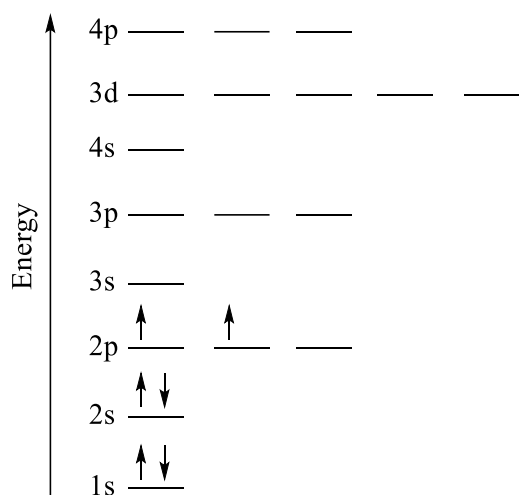


The first two electrons in lithium must go in the lowest energy level first, the third electron goes into the next lowest energy level (2s)

The electron configuration for a lithium atom is $1s^2 2s^1$

Once electrons enter a p-sublevel or a d-sublevel, they must spread out into an empty orbital pointing up before they will pair up.

Figure 6: Orbital Filling Diagram for C



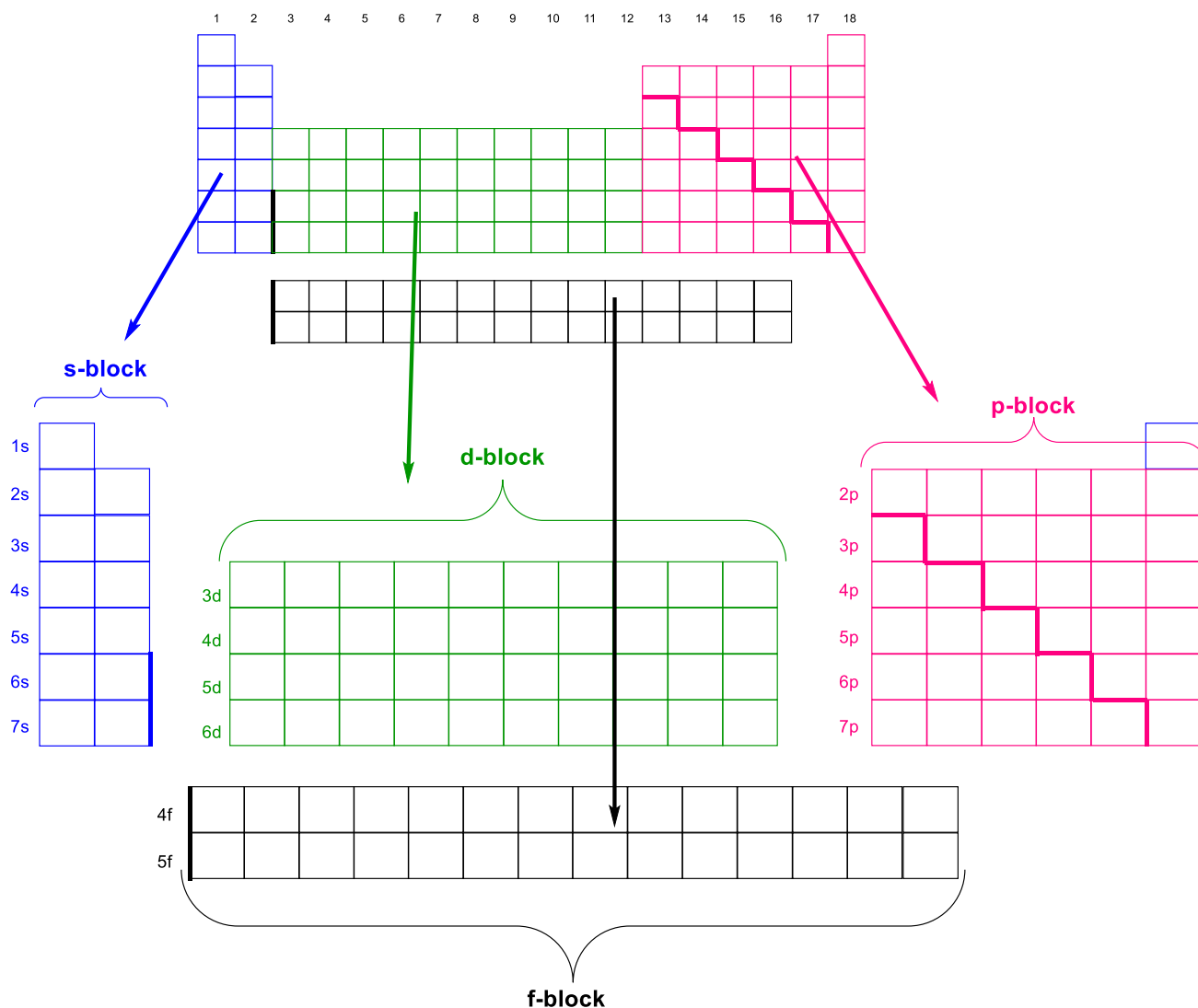
The first two electrons in carbon must go in the lowest energy level first, the third and fourth electrons go into the next lowest energy level, the fifth electron starts the 2p level and the sixth electron will spread out to fill one of the other empty 2p sublevels.

The electron configuration for a carbon atom is $1s^2 2s^2 2p^2$

An atom's placement in the periodic table can also help to determine its electron configuration. Figure 7 shows the periodic table broken into blocks that represent the principle quantum levels and sublevels.

The s-block numbering starts at 1s and goes to 7s, the p-block numbering starts at 2p and goes 7p, and the d-block numbering starts at 3d and goes to 6d. When using the periodic table to write an electron configuration, always start in the upper left corner (1s) and move from left to right across a row until all electrons have been used. Once the end of a row is reached, move to the next row if there are additional electrons present.

Figure 7: Orbital Blocks of the Periodic Table



Example

Write an electron configuration for a carbon atom.

1. Locate the atom in periodic table.
2. Determine the number of electrons it has. Carbon has six electrons.
3. Start in upper left corner and move from left to right until all electrons have been accounted for.

The first two electrons complete the 1s level: $1s^2$

4. Once a level has been completed, move to the next row.

The next two electrons complete the 2s level: $1s^2 2s^2$
 Electrons five and six are in the 2p level: $1s^2 2s^2 2p^2$

Notice this gives the same electron configuration that is obtained by the orbital filling diagram.

Noble Gas Electron Configuration

A shorthand way of writing electron configurations is to use the noble gas configuration (NGC). For example, a sodium atom have 11 electrons and its electron configuration is $1s^2 2s^2 2p^6 3s^1$. The first ten electrons, $1s^2 2s^2 2p^6$, are isoelectronic (have the same electron configuration) with neon. Therefore, the noble gas electron configuration for sodium is $[\text{Ne}]3s^1$. The ten electrons in sodium that are isoelectronic with neon are represented by $[\text{Ne}]$ in the square bracket, the remaining $3s^1$ electron is written outside of the brackets. Only the noble gas that comes directly before the element can be used in the NGC and only an element in group 18, a noble gas, can be used in this notation.

When writing a condensed noble gas configuration for atoms use the following steps.

1. Locate the atom on the period table and determine the number of electrons it has.
2. Find the last noble gas (in Group 18) that comes before it and place it in square brackets [Ex].
3. Continue writing out the remaining electron configuration outside of the square brackets until all electrons have been used.

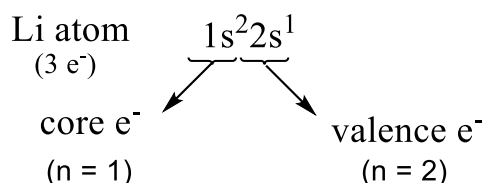
Example

Write a noble gas electron configuration for a carbon atom.

1. Locate the atom on the period table and determine the number of electrons it has.
Carbon has six electrons.
2. Find the last noble gas (in Group 18) that comes before it and place it in square brackets [Ex].
For the noble gas that comes before carbon is helium.
 $[\text{He}]$
3. Continue writing out the remaining electron configuration outside of the square brackets.
 $[\text{He}]2s^2 2p^2$

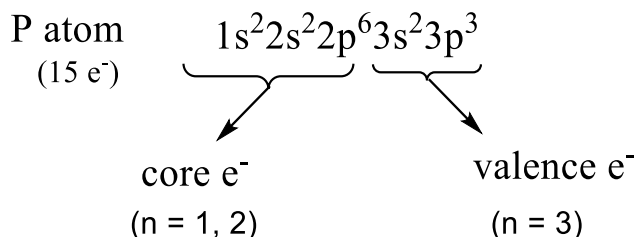
Valence and Core Electrons

Not all electrons in an atom behave in the same manner. The electron(s) in the highest principle quantum level are called valence electrons. These electrons are responsible for the reactivity of an atom. All other electrons are called core electrons and they contribute to the stability, rather than the reactivity, of an atom. For a lithium atom there is just one valence electron in the 2s orbital and the two electrons in the 1s orbital are its core electrons.



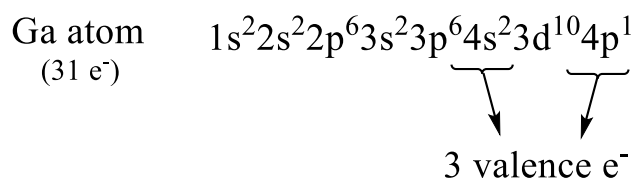
Example

Determine the number of core and valence electrons in a phosphorous atom.



Phosphorous has 10 core electrons and 5 valence electrons.

When determining the number of valence electrons, it will not always be the electrons at the end of the electron configuration. The valence electrons are all of the electrons in the highest principle quantum level. For gallium the highest principle quantum number is $n = 4$. So, all of the electrons in the $n = 4$ level are the valence electrons. Therefore, gallium has three valence electrons, two in the 4s level and one in the 4p level, and the remaining 28 electrons are core electrons.



Materials and Reagents Needed for Lab

Lab Equipment

7 long wooden splints	7 large test tubes	Test tube rack
400 mL beaker	Bunsen burner	Striker/Spark gun
Sharpie		

Reagents

1 M copper (I) chloride	1 M iron (II) nitrate	1 M potassium chloride
1 M copper (II) chloride	1 M lithium chloride	1 M calcium chloride
1 M barium chloride	1 M sodium chloride	1 M Unknown
1 M strontium chloride	1 M lead (II) nitrate	Water

Safety Considerations

A Bunsen burner and a live flame will be used in this lab. Long hair should be tied back in a bun, ponytail or braid. Keep the open flame away from hand sanitizer. Close the gas nozzle when not in use. Copper (II) chloride is toxic and should not come into contact with eyes or skin. Lithium chloride is a body tissue irritant. Rinse or soak the wooden splints in water for at least 30 – 60s before discarding them in the trash to avoid trashcan fires. Solutions of copper, barium, strontium and lead cannot go down the drain. Students will not dispose of the Cu, Ba, Sr and Pb solutions, they will be disposed of by the instructor. Gloves must be worn for this lab.

Procedure

Part A: Instructor Flame Test Demonstration

1. The instructor will demonstrate how to set up and use the Bunsen burner and perform the flame test for several metal ion containing solutions.
2. While watching the demonstration, record the charge of the cation and the flame color in **Table 1** under Data/Calculations section.

Part B: Flame Test of Metals

1. Use a Sharpie to label one large test tube with calcium chloride.
2. Fill the test tube about half full with the calcium chloride solution.

3. Place the test tube in the test tube rack.
4. Place one wooden splint into the test tube. Let the splint soak in the test tube for at least 60 seconds.
5. While the wooden splint is soaking, repeat steps 1 – 4 for the remaining solutions but labeling them with their correct names.
6. Fill a 400 mL beaker about three-fourths full of deionized water. Set the beaker aside, but close by. This is the waste beaker.
7. Set up the Bunsen burner and light it.
8. Remove the splint from the calcium chloride test tube. Gently shake any excess solution off the splint into the test tube.
9. Place the wet end of the splint directly into the flame.
10. Observe and record the flame color in **Table 2** under Data/Calculations.
11. Once the color has been observed, place the wet/burned end of the wooden splint in the waste beaker
12. Repeat steps 8 – 11 for the remaining solutions.
13. When finished, turn off the Bunsen burner.
14. Leave the wooden splints soaking in the beaker. The same waste beaker will be used in the next section.

Part C: Flame Test of Unknowns

1. Obtain one new large test tube and an unknown.
2. Record the unknown number or letter in **Table 3**.
3. Fill the test tube about half full of the appropriate unknown solution and place it in the test tube rack.
4. Place one wooden splint into the test tube and let it soak in the test tube for at least 60 seconds.
5. Use the same procedure from Part B Steps 7 -11 to observe the flame color of the unknown.
6. When finished, turn off the Bunsen burner.
7. Based on flame color, determine the identity of the metal present.

Part D: Wavelength and Frequency of Light

1. Write the wavelength range (in nm) in **Table 4** for the color listed from the information given in the introduction.

2. Determine the middlemost value (to 3 significant figures) of the above range and use it to calculate the frequency (in $\frac{1}{s}$) and energy (in J) of the light. Remember to convert the wavelength to meters before calculating the frequency.
3. For light with a frequency of 5.71×10^{14} Hz determine the wavelength in nm, the color and the energy of the light.

Part E: Electron Configurations of Atoms

1. Write the full electron configuration for each atom and ion in **Table 5**.
2. Write the condensed noble gas electron configuration for each atom in **Table 6**.

Clean Up and Waste Disposal: The used wooden splints may be placed in the trash can after they have soaked in water for at least 30 – 60s. The solutions in each test tube and the waste beaker may be poured down the drain with running water. Use soap and water and a scrub brush to clean and remove the sharpie from the test tubes. After rinsing the test tubes with tap water and again with deionized water, the cleaned test tubes should be placed in the test tube bucket on the counter. Clean all other glassware withalconox/soap solution, then rinse with tap water and again with deionized water. After drying the glassware, return items to their original location.

Before Leaving Lab

- Make sure all items have been cleaned and put away.
- Check the board for any additional clean-up or waste instructions.
- Wipe down the benchtops.
- Refill the water bottles.
- Push in stools.
- Check with your instructor that your area has been sufficiently cleaned and reorganized.

Lab 5: Electron Configurations and Light Lab – Pre-Lab

Name _____

Section _____

Date _____

Pre-lab Questions

Read through all the background information and the procedure of the lab before answering these questions. For any calculations, all work must be shown. The work shown must include the formula, the unrounded answer, and the correctly rounded final answer. Answers without work may not receive full credit.

1. Explain how the wooden splints should be extinguished after each flame test.
2. Explain what should be done with each solution in the test tubes at the end of lab?

3. Write the speed of light. _____

4. Identify each variable.

λ _____ ν _____ c _____

5. Rank the following types of radiation in order of longest to shortest wavelength: infrared, radio, ultraviolet.

_____ > _____ > _____
longest shortest

6. Use the website <http://chemicool.com> to find the flame test color of the following elements.

Copper _____

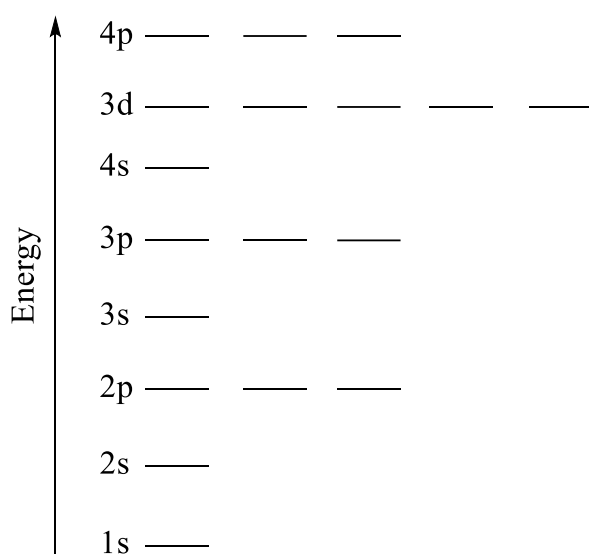
Potassium _____

Strontium _____

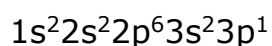
Sodium _____

7. Calculate the wavelength, in m, for light that has a frequency of 2.14×10^{14} Hz. Write answer in scientific notation.

8. Complete the orbital diagram for a sulfur atom and write its full electron configuration.



9. Use the electron configuration of the below atom to determine the number of valence and core electrons.



Valence electrons: _____ Core electrons: _____

10. Identify each atom from the electron configuration.

A) $1s^2 2s^2 2p^6 3s^2 3p^5$ _____ B) $[\text{He}] 2s^2 2p^2$ _____

Lab 5: Electron Configurations and Light Lab – Data/Calculations

Name _____

Section _____

Drawer _____

Balance _____

Date _____

Partner(s) _____

Be sure to show all steps in calculations when necessary. Include formulas, units, unrounded answers and final rounded answers (with appropriate units) in the calculation table.

Table 1: Instructor Flame Test of Metal Cations Demonstration

Solution	Metal Ion	Color of Flame
Copper (I) chloride		
Copper (II) chloride		
Iron (II) nitrate		
Lead (II) nitrate		
Barium chloride		
Strontium chloride		

Table 2: Flame Test of Metal Cations

Solution	Metal Ion	Color of Flame
Calcium chloride		
Potassium chloride		
Lithium chloride		
Sodium chloride		

Table 3: Flame Test of Unknown

Solution	Label	Color of Flame
Unknown		

Table 4: Wavelength and Frequency of Light**Red Light Wavelength Data**

Light Color	Wavelength Range (λ)	Middle Value
Red		

Red Light Frequency Calculation

Frequency (ν)	
Calculation of frequency	

Red Light Energy Calculation

Energy (J)	
Calculation of Energy	

Frequency Data

Frequency (ν)	Color
5.71×10^{14} Hz	

Wavelength Calculation

Wavelength (λ)	
Calculation of wavelength	

Energy Calculation

Energy (J)	
Calculation of Energy	

Table 5: Electron Configurations of Atoms and Ions

Atom or Ion	Full Electron Configuration
Ca	
Ca ²⁺	
S	
S ²⁻	
Li	
Li ⁺	
P	
P ³⁻	
F	
F ⁻	
Al	
Al ³⁺	

Table 6: Noble Gas Configurations of Atoms

Atom	Noble Gas Configuration
Si	
Li	
Se	
B	
C	
K	
As	
Mg	
Cl	

Lab 5: Electron Configurations and Light Lab – Post-Lab Questions

Name _____

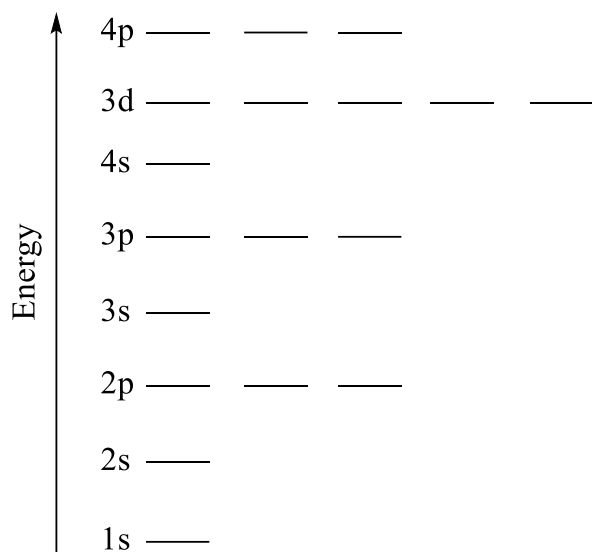
Section _____

Date _____

Post-lab Questions

For calculations, all work must be shown. Work shown must include the formula, the unrounded answer the correctly rounded final answer. Answers without work may not receive full credit.

1. Complete the orbital diagram for a selenium atom then answer the questions that follow.



A) Which level(s) contain the valence electron(s)? _____

B) How many valence electrons does Se have? _____

C) Write the full electron configuration for this atom.

D) Write the condensed noble gas configuration.

E) Will Se will form a cation or anion. Briefly explain your answer.

2. Identify the number of core and valence electrons in a iodine atom.

Core electrons _____ Valence electrons _____

3. Write the full electron configuration for each.

Al atom

Al³⁺ ion

Cl atom

Cl⁻ ion

4. Write the condensed noble gas electron configuration for each.

Rb atom

P atom

5. List one **anion** that has the configuration: $1s^2 2s^2 2p^6$ _____

List one **cation** that has the configuration: $1s^2 2s^2 2p^6 3s^2 3p^6$ _____

6. From the instructor demonstration, which metal cation was observed to emit the longest wavelength? Explain.

7. What was the unknown label used by your group? _____

Briefly explain what metal ion might be present in your unknown.

8. A firework produced a purple flame. List one possible metal that could be present.
9. Calculate the wavelength, in nm, of light that has a frequency of 4.40×10^{14} Hz.

What color light would this be? _____

10. A St. Louis Radio station broadcasts a radio signal that has wavelength of 3.308 meters. What is the frequency of the signal in MHz? Report to three significant figures.

Lab 6: Nomenclature Activity

Lab Objectives

- Write the name or formula of an acid, ionic compound, covalent compound, or polyatomic ion
- Differentiate or classify compounds as ionic, molecular or acidic
- Predict the chemical formula with a given cation and anion and vice versa
- Classify, name and write the formula of a simple or polyatomic ion
- Categorize bonding as either ionic or covalent

Background Information

Naming a compound in chemistry is fundamental to discussing chemical concepts. In this lab, simple ions, polyatomic ions, ionic compounds, covalent compounds and acids will be discussed.

Naming Simple Ions

Metallic elements can form positively charged ions called cations. Simple cations occur when a metal atom loses one or more electrons. Ions are written with their charge indicated in the upper right-hand corner as a superscripted number followed by a positive or negative sign. The number of electrons lost by the atom is consistent with the charge on the cation. If one electron is lost, the cation will have a plus one charge; if two electrons are lost the cation will have a plus two charge and so on.

Metallic cations are named depending on whether they have one specific charge (monatomic cations or Type I ions) or can form cations with different positive charges (variably charged cations or Type II ions). The monatomic cations are listed in Table I below. These ions are named by adding the word “ion” to the end of the name given in the periodic table regardless of their charge. For example, when the sodium atom, Na, loses one electron it forms positively charged ion, Na^+ , called sodium ion.

Table I: Monatomic (Type I) Metal Cations without Roman Numerals

Metals that only form 1^+ cations	Metals that only form 2^+ cations	Metals that only form 3^+ cations
Alkali metals	Alkaline Earth metals	Aluminum, Gallium, Indium
Silver	Zinc, Cadmium	

Example

Name each monatomic cation.

Cation
Ion Name

Na^+
Sodium ion

Mg^{2+}
Magnesium ion

Al^{3+}
Aluminum ion

Any other metal not listed in Table I uses a Roman Numeral in parentheses to indicate the charge on the cation. Table II shows the Roman Numerals for the values 1 – 10.

Table II: Roman Numeral 1 – 10

Number	Roman Numeral	Number	Roman Numeral
1	I	6	VI
2	II	7	VII
3	III	8	VIII
4	IV	9	IX
5	V	10	X

Example

Name each variably charged cation.

Cation
Ion Name

Fe^+
Iron (I) ion

Fe^{2+}
Iron (II) ion

Fe^{3+}
Iron (III) ion

Nonmetal elements can form negatively charged ions called anions. Simple anions occur when a nonmetal gains one or more electrons. Their charge is written in the same place as a cation but is followed by a negative sign. They are named by adding “ide” to the end to their elemental name. The nonmetal elements that form simple anions are found in Table III.

Table III: Monatomic Nonmetal Anions

Nonmetals that form 1 ⁻ anions	Nonmetals that form 2 ⁻ anions	Nonmetals that form 3 ⁻ anions
Fluorine, Chlorine, Bromine, Iodine	Oxygen, Sulfur, Selenium	Nitrogen, Phosphorous

Example

Name each simple anion.

Anion
Ion Name

F^-
Fluoride ion

O^{2-}
Oxide ion

N^{3-}
Nitride ion

Notice that when the charge on an ion is 1⁺ (as in Fe^+) or 1⁻ (as in F^-) the number one is not written but is understood to be present.

Hydrogen is the only element that can form either a cation (H^+ , hydrogen ion) or an anion (H^- , hydride). All other elements will only form cations (metals), anions (nonmetals) or do not form ions (semimetals and noble gases).

Naming Polyatomic Ions

Polyatomic ions are made of two different covalently bonded elements that have an overall charge. The names of polyatomic ions may be recognized because they often, but not always, end in “-ate” or “-ite”. The names and formulas of some polyatomic ions are listed in Table IV and Appendix III. However, because the naming

convention of polyatomic ions is not as straightforward as simple ions, their names must often be memorized. For example, nitrate (NO_3^-), nitrite (NO_2^-) and hydroxide (OH^-) are all polyatomic ions.

Table IV: Polyatomic Ions

Polyatomics with 1 ⁻ Charge		Polyatomics with 2 ⁻ Charge		Polyatomics with 3 ⁻ Charge	
Formula	Name	Formula	Name	Formula	Name
$\text{C}_2\text{H}_3\text{O}_2^-$	Acetate	CO_3^{2-}	Carbonate	PO_4^{3-}	Phosphate
HCO_3^-	Bicarbonate	SO_4^{2-}	Sulfate	PO_3^{3-}	Phosphite
NO_3^-	Nitrate	SO_3^{2-}	Sulfite	SiO_3^{3-}	Silicate
NO_2^-	Nitrite	$\text{C}_2\text{O}_4^{2-}$	Oxalate		
OH^-	Hydroxide	CrO_4^{2-}	Chromate		
ClO_4^-	Perchlorate	O_2^{2-}	Peroxide		
ClO_3^-	Chlorate				
ClO_2^-	Chlorite				
ClO^-	Hypochlorite				
MnO_4^-	Permanganate				

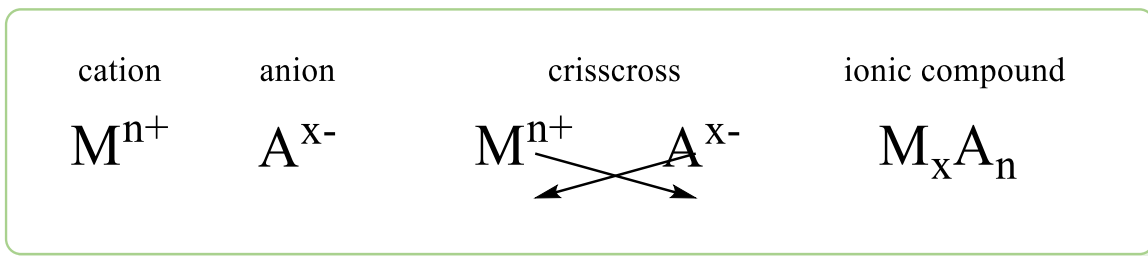
Polyatomics with 1 ⁺ Charge	
Formula	Name
NH_4^+	Ammonium

Writing Formulas of Ionic Compounds

Like their name suggests, ionic compounds are comprised of positively charged ions (cations) attracted to negatively charged ions (anions). When the ions combine, a neutral ionic compound is formed. Ionic compounds will always have both a cation and anion. All ionic compounds contain ionic bonds. **Ionic bonds** are the attractive forces between oppositely charged ions. Only ions of opposite charge can combine to create ionic compounds. Enough cations and anions will be attracted to each other to create a compound that has an equal amount of positive and negative charges. Thus, ionic compounds always have an overall charge of zero. The result is an ionic compound that represents the lowest ratio of cations and anions necessary to neutralize charge – this is the empirical formula.

The “crisscross” method is often used to help write the formula of an ionic compound. Using this method, the charge of each ion becomes the subscript on the other ion.

Figure 1: Example of the Crisscross Method



Example

Write the formula of the ionic compound formed when the aluminum ion combines with the chloride ion.

Formula	Al^{3+}	Cl^-	AlCl_3
	cation	anion	Neutral ionic compound

NOTE: Using the crisscross method, the superscripted three on aluminum becomes the subscript on chloride. And vice versa for the superscripted one on the chloride becoming the subscript on aluminum. In other words, since chloride has a negative one charge, there are three chloride ions needed to balance the plus three charge on the aluminum ion.

Example

Write the formula of the ionic compound formed when the lithium ion combines with the bromide ion.

Formula	Li^+	Br^-	LiBr
	cation	anion	Neutral ionic compound

NOTE: Using the crisscross method, the superscripted one on lithium becomes the subscript on bromide. And vice versa for the superscripted one on the bromide. This time only one bromide is needed to balance the plus one charge on the lithium ion.

Parenthesis is used when more than one polyatomic ion is needed to make the ionic compound neutral. When the calcium ion (Ca^{2+}) combines with nitrate (NO_3^-), parentheses are used around the nitrate anion to show two nitrates are needed to balance the charge on the calcium ion and the formula is written as $\text{Ca}(\text{NO}_3)_2$. However, when the sodium ion (Na^+) combines with nitrate (NO_3^-), the formula is NaNO_3 and no parenthesis is needed.

Example

Write the formula of the ionic compound formed when the calcium ion combines with the nitrate ion.

Formula	Ca^{2+}	NO_3^-	$\text{Ca}(\text{NO}_3)_2$
	cation	anion	Neutral ionic compound

NOTE: Using the crisscross method, the superscripted two on calcium becomes the subscript on nitrate. However, since nitrate is a polyatomic ion, the entire ion is placed in parenthesis with the two written outside of the parenthesis. In other words, two nitrate ions are needed to balance the plus two charge on calcium.

Example

Write the formula of the ionic compound formed when the titanium (I) ion combines with the nitrate ion.

Formula	Ti^+	NO_3^-	TiNO_3
	cation	anion	Neutral ionic compound

NOTE: Using the crisscross method, the superscripted one on titanium (I) becomes the subscript on nitrate. However, since it is not necessary to write the

number one as a subscript, it is understood to be present, and parentheses are not needed.

When writing the formulas of ionic compounds the formula must always be reduced to the lowest set of possible subscripts. Using the crisscross method, Cu^{2+} and SO_4^{2-} would initially form $\text{Cu}_2(\text{SO}_4)_2$ but must be reduced to the empirical formula of CuSO_4 . Paying close attention to the combination of ions is important when using the crisscross method. Lead (IV) sulfide is a good example of this. When the cation Pb^{4+} is combined with S^{2-} , the crisscross method initially gives the formula Pb_2S_4 ; this is not, however, the empirical formula. Both the two and four can be reduced by dividing by two, the correct formula for lead (IV) sulfide, therefore, is written as PbS_2 .

Example

Write the formula of the ionic compound formed when lead (IV) ion combines with the sulfide ion.

Formula	Pb^{4+}	S^{2-}	Pb_2S_4	PbS_2
	cation	anion	Needs to reduce	Empirical formula

There are some ionic compounds where it appears the formula has not been reduced to the empirical formula. But remember it is only the number in superscripted charge that can be reduced once the crisscross method is used. For example, sodium peroxide has the formula Na_2O_2 which at first glance looks like it should be written as NaO . But Na_2O_2 is formed from the sodium ion (Na^+) and peroxide (O_2^{2-}). When those two ions are combined using the crisscross method the formula becomes Na_2O_2 . In this case, the subscripted two on the oxygen is not from the charge of the sodium ion but is from the two oxygens that are present in peroxide. Another way to view it, is that there is two Na^+ ions needed to balance the charge on one O_2^{2-} ion.

Naming Ionic Compounds

When naming any ionic compound, the cation is named first followed by anion name. Once the name of the cation and anion are known they are written together to give the name of the ionic compound. The metals with monatomic ionic charges do not use Roman Numerals in their names. Roman Numerals are only used for metal cations that can have a variable charge. Doing the crisscross method in reverse can help to determine the charges on the ions in the compound.

Figure 2: Reverse Crisscross Method

ionic compound	crisscross	cation	anion
M_xA_n		M^{n+}	A^{x-}

Example

Write the name of the ionic compound Fe_3P_2 .

Formula	Fe^{2+}	P^{3-}	Fe_3P_2
	Iron (II) ion	Phosphide	Iron (II) phosphide

NOTE: Using the reverse crisscross method, the cation present is Fe^{2+} and the anion is P^{3-} therefore the ionic compound is named iron (II) phosphide.

Naming Binary Molecular Compounds

These compounds contain only nonmetal and/or metalloids (also called semimetals) in their chemical formulas. When only nonmetals and/or metalloids are present in a compound, electrons are not transferred, instead they are shared between the atoms and form **covalent bonds**. These compounds therefore are also called covalent compounds. The crisscross method is not used with binary molecular compounds because no ions are present. When writing their formulas, they should never be reduced to their empirical formulas.

Example

The compound N_2H_4 is not reduced to NH_2 .

To name binary molecular compounds, the subscripts on each of the two elements must be specified using the Greek prefixes shown in Table V. The first element retains its name from the periodic table and the second element's name is changed to end in "ide". There are two rules associated with their use. First, if there is only one of the first elements shown, the prefix "mono-" is not included, it is assumed. Second, if the prefix ends with an "a", such as "penta-", and the element that comes next in the name begins with a vowel, such as oxygen or iodine, the "a" is dropped. For instance, N_2O_5 is called dinitrogen pentoxide (not pentaoxide).

Table V: Prefixes for Binary Molecular Compounds

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

Example

Name each covalent compound.

Formula
Name

NO
Nitrogen monoxide

N_2O_4
Dinitrogen tetroxide

Example

Write the formula of each covalent compound.

Name
Formula

Carbon dioxide
 CO_2

Triboron pentabromide
 B_3Br_5

Example

What type of bonding occurs in each compound?

Name
Bonding Type

NH_3
Covalent bonding

AlCl_3
Ionic bonding

Naming Acids

Only the names and formulas of Arrhenius acids will be used this semester. Those acids produce a hydrogen ion, or H^+ , when placed in water. The cation for all acids therefore is always the hydrogen ion, H^+ . They are easily recognized because their formula begins with hydrogen as the first element and often contains the aqueous physical state (aq) at the end of their formulas. There are two types of Arrhenius acids: binary acids and oxyacids.

Binary acids contain hydrogen and only one other element that is not oxygen. The names of binary acids always begin with “hydro” and end in “-ic acid”. When writing their formula, there are enough hydrogen ions (H^+) to balance the charge on the simple anion. For example, only one H^+ is needed to balance the charge on Cl^- to create the acid HCl(aq) named hydrochloric acid.

Oxyacids are formed from polyatomic ions that contain one or more oxygen atoms, many of which are given in Table IV. The name of the acid depends on which polyatomic ion is present. Polyatomic ions ending in “-ate” become “-ic” acids, while polyatomic ions ending in “-ite” become “-ous” acids (see Table VI).

Table VI: Examples of Oxyacids

Polyatomic ion Formula	Polyatomic ion Name	Acid Formula	Acid name
BrO_3^-	bromate	$\text{HBrO}_3(\text{aq})$	Bromic acid
BrO_2^-	bromite	$\text{HBrO}_2(\text{aq})$	Bromous acid

Materials and Reagents Needed for this Activity

This activity is considered a dry lab thus no chemicals will be used for this lab. Only a writing utensil and a periodic table are needed for this lab.

Safety Considerations

Although this is a dry lab, students must wear appropriate PPE when in the Intro Chem Lab. No goggles or safety glasses are needed for this lab, however, students must still wear closed-toe shoes, long pants, shirts with a front/back and some type of sleeve.

Procedure

Follow the directions in the manual given by the lab instructor.

Part A: Simple and Polyatomic ions

Answer the questions about ions.

Part B: Ionic Compounds

1. In Table 1, write the name for each ionic compound that contains a monatomic metal cation.
2. In Table 2, write the formula for each ionic compound that contains a monatomic metal cation.
3. In Table 3, write the name for each ionic compound that contains a variably charged metal cation.
4. In Table 4, write the formula for each ionic compound that contains a variably charged metal cation.
5. In Table 5, write the formula for each cation, anion and ionic compound in each row.

Part C: Binary Molecular Compounds

1. In Table 6, write the name for each molecular compound.
2. In Table 7, write the formula for each molecular compound.

Part D: Acids

1. In Table 8, first classify each acid as a binary acid or an oxyacid. Then write the name for each acid.
2. In Table 9, write the formula for each acid.

Part E: Name and Formulas of Compounds and Ions

In Table 10, write either the name or the formula for each.

Before Leaving Lab

- Check the board for any additional instructions.
- Wipe down the benchtops
- Push in stools
- Check with your instructor that your area has been sufficiently cleaned and reorganized

Lab 6: Nomenclature Activity – Pre-Lab

Name _____

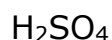
Section _____

Date _____

Pre-lab Questions

Read through all the background information and the procedure of the lab before answering these pre-lab questions.

1. Ionic compounds must contain a cation and an anion. True or False
2. The names of covalent compounds have Roman Numerals. True or False
3. Classify each as either an ionic compound, a molecular compound or a binary acid or an oxyacid.



4. Give the name for NaCl. _____

5. Write the name of CO₂. _____

6. Write the name for each ion.

Rb⁺ _____

OH⁻ _____

S²⁻ _____

Cr⁶⁺ _____

7. Write the formula of the ionic compound that will form when Zn²⁺ and N³⁻ combine.

8. Write the formula of the ionic compound that will form from each set of ions.

copper (II) ion and nitrate

gallium ion and oxide

9. Briefly explain the type of bonding (ionic or covalent) present in NH_3 .

10. A certain compound contains magnesium and fluorine. Explain whether the compound would be ionic or covalent.

Lab 6: Nomenclature Activity – Data/Calculations

Name _____

Section _____

Drawer _____

Balance _____

Date _____

Partner(s) _____

Date _____

Part A. Simple and Polyatomic Ions

1. Circle the element(s) that will form a cation: Ne Cr Si Cl Ca H

2. Circle the element(s) that will form an anion: N Na Se S Ar Co

3. Circle the element(s) that do not form an ion: Li Mg As Ne Fe C

4. Write the name of each ion.

NH_4^+ _____

O^{2-} _____

Cl^- _____

Al^{3+} _____

Cu^{4+} _____

NO_2^- _____

5. Write the formula of each ion.

Sulfate _____

Calcium ion _____

Titanium (I) ion _____

6. Explain how a simple ion is different than a polyatomic ion.

7. Define monatomic cation.

8. What type of ions use Roman Numerals in their name?

Part B. Ionic Compounds

Table 1: Naming Ionic Compounds with Monatomic Cations

Compound Formula		Compound Name
	NaCl	Sodium chloride
a.	AgNO ₃	
b.	Li ₃ N	
c.	Ga ₂ (SO ₄) ₃	
d.	Mg(OH) ₂	
e.	ZnS	
f.	(NH ₄) ₂ C ₂ O ₄	

Table 2: Writing Formulas of Ionic Compounds with Monatomic Cations

Compound Name		Compound Formula
	Sodium chloride	NaCl
a.	Potassium selenide	
b.	Calcium phosphide	
c.	Aluminum phosphite	
d.	Sodium acetate	
e.	Cadmium bicarbonate	
f.	Sodium carbonate	

Table 3: Naming Ionic Compounds with Variably Charged Cations

Compound Formula		Compound Name
	Fe_3P_2	Iron (II) phosphide
a.	SnBr_4	
b.	SnBr_2	
c.	$\text{Ni}_3(\text{PO}_4)_5$	
d.	$\text{Fe}_2\text{C}_2\text{O}_4$	
e.	$\text{V}_2(\text{CO}_3)_3$	
f.	CrI_6	

Table 4: Writing Formulas for Ionic Compounds with Variably Charged Cations

Compound Name		Compound Formula
	Iron (II) phosphide	Fe_3P_2
a.	Bismuth (III) bromide	
b.	Platinum (III) sulfite	
c.	Lead (IV) oxide	
d.	Rhodium (I) carbonate	
e.	Cobalt (II) fluoride	
f.	Manganese (VI) nitrite	

Table 5: Ionic Compound Summary

	Compound Name	Cation Formula	Anion Formula	Compound Formula
	Iron (II) phosphide	Fe^{2+}	P^{3-}	Fe_3P_2
a.	Titanium (III) nitrate			
b.	Aluminum bromide			
c.	Nickel (II) phosphate			
d.	Silver selenide			
e.	Ammonium acetate			
f.	Zinc phosphide			

Part C. Binary Molecular Compounds

Table 6: Naming Binary Molecular Compounds

	Compound Formula	Compound Name
	NO	nitrogen monoxide
a.	Si_4H_{10}	
b.	B_7F_9	
c.	SeF_2	
d.	S_3O	
e.	I_2F_6	
f.	N_5Cl_8	

Table 7: Writing Formulas of Binary Molecular Compounds

Compound's Name		Compound Formula
	Dinitrogen tetroxide	N_2O_4
a.	Bromine monofluoride	
b.	Boron dioxide	
c.	Trinitrogen pentaphosphide	
d.	Tetrasulfur heptabromide	
e.	Phosphorus trihydride	
f.	Xenon tetrafluoride	

Part D. Acids

Table 8: Classifying and Naming Acids

Acid Formula		Type of Acid	Acid Name
	HCl(aq)	Binary acid	Hydrochloric acid
a.	$\text{HC}_2\text{H}_3\text{O}_2\text{(aq)}$		
b.	HBr(aq)		
c.	$\text{H}_2\text{CO}_3\text{(aq)}$		
d.	HI(aq)		
e.	$\text{H}_2\text{SO}_3\text{(aq)}$		
f.	$\text{H}_2\text{Se(aq)}$		

Table 9: Writing the Formulas of Acids

	Acid Name	Acid Formula
	Hydrochloric acid	HCl(aq)
a.	Nitrous acid	
b.	Hydrosulfuric acid	
c.	Chloric acid	
d.	Hypochlorous acid	
e.	Hydrofluoric acid	
f.	Phosphoric acid	

Part E. Names and Formulas of Compounds and Ions

Table 10: Writing Names and Formulas

	Name	Formulas
a.	Zinc phosphite	
b.	Sulfur hexafluoride	
c.		NH ₄ Cl
d.		Co(ClO ₃) ₃
e.	Iron (II) oxide	
f.		CrO ₄ ²⁻
g.		HClO ₄ (aq)
h.	Calcium acetate	
i.		HCO ₃ ⁻

j.	Iodide	
k.	Sulfide	
l.		N_3O_5
m.		PbS_2
n.	Copper (II) sulfate	

Lab 6: Nomenclature Activity – Post-Lab Questions

Name _____

Section _____

Date _____

Post-lab Questions

1. Some compounds are referred to by their common (or household) names, several of which are shown below. Write their name using the rules learned in this lab.

Bleach: NaClO _____

Rust: Fe_2O_3 _____

Baking Soda: NaHCO_3 _____

2. Write the name of each.

A) CO

B) AgBr

C) $\text{Ni}_2(\text{SO}_4)_3$

3. The compounds below are named incorrectly (the formulas are correct as written). Write the correct name for each.

SiBr_4

PbS

$\text{H}_3\text{PO}_3(\text{aq})$

Silicon bromide

Lead (I) sulfide

Phosphoric acid

4. Write the formula for each ion.

A) Sulfite

B) Nitride

C) Cadmium ion

5. For the two compounds named below, only one should use Roman Numerals. Briefly explain which compound is named incorrectly then write its correct name.

Compound 1: Copper (II) chloride **Compound 2:** Barium (II) chloride

6. Briefly explain the type of bonding (ionic or covalent) present in $\text{Cr}(\text{OH})_2$.

Name the compound: _____

7. Explain how to recognize the difference between an ionic compound and a covalent compound.
8. Explain the difference between an oxyacid and a binary acid.

9. When should the prefix "hydro" be used when naming acids? Briefly explain. Then give an example (formula and name) of this type of acid.

10. Complete the following table by writing the correct name or formula for each.

	Name	Formula
a.		Se_5F_4
b.	Zinc iodide	
c.		HCO_3^-
d.	Copper (III) carbonate	
e.		N_6S_7
f.	Nitric acid	
g.		K_2SO_4
h.	Hydroiodic acid	
i.		$\text{W}_2(\text{C}_2\text{O}_4)_5$
j.	Lithium phosphite	
k.		P^{3-}

Lab 7: Writing and Balancing Chemical Equations Activity

Lab Objectives

- Interpreting information about a reaction or balanced chemical equation
- Classify reactions as combustion, decomposition, synthesis, single displacement or double displacement
- Balance an equation or determine the coefficients for a balanced equation
- Predict the products of nonmetal/metal synthesis, hydrocarbon combustion reactions, precipitation, acid-base neutralization reactions
- Apply the solubility rules for nitrates, chlorides, iodides, bromides, sulfates, carbonates, phosphates, hydroxides and sulfides

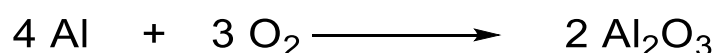
Background Information

Chemical reactions occur every day all round us. Although reactions may be different in many ways, they all have several features in common. In all chemical reactions all the atoms present at the beginning of the reaction must still be present at the end; however, the atoms are reorganized to form new substances. The substances present at the start of the reaction are called the reactant(s) and the new species that are formed will be the product(s). Each reactant or product can exist in one of several physical states. The physical state of an atom, molecule or compound will appear in parenthesis at the end of the chemical form. Examples include solid (s), liquid (l), gas (g) or aqueous (aq). The term “aqueous” comes from the Latin word *aqua* meaning water. Therefore, when an ionic compound is dissolved in water it will have an aqueous physical state (aq).

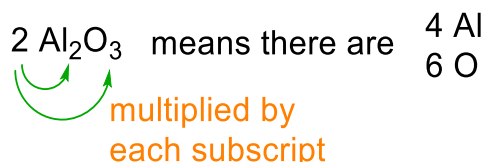
The new substances that are formed have physical and chemical properties that are separate and different than the reactant(s) from which they came. Even though atoms are rearranged to form new bonds, the same number and type of atoms must be present before and after the change. When a chemical reaction occurs there are some physical changes that may be observed. Color or odor changes, appearance of bubbles, temperature change, production or disappearance of a solid, generation of light, the formation or loss of heat are all common indications that a chemical reaction has taken place.

Interpreting Chemical Reactions

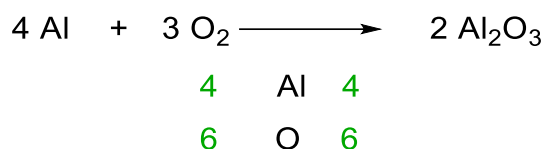
All chemical reactions are written in the same way. The reactant(s) will be on the left of the reaction arrow and the product(s) will be on the right. In the reaction below, both the Al and O₂ are on the reactant side of the arrow and Al₂O₃ is on the product side. When chemical reactions are balanced correctly the same number and type of atoms are present on both sides of the equation.



Reactions are balanced by writing coefficients in front of each reactant and product. Coefficients are always whole numbers and apply to all the atoms in the formula that immediately follow it. A coefficient of one (1) is not written and instead is “understood” to represent the number one. In the reaction above, there is a coefficient of four (4) in front of Al, a coefficient of three (3) in front of O₂ and a coefficient of two (2) in front of Al₂O₃. When counting the number of atoms, the coefficient (if present) is multiplied by the subscript of the formula. For example, on the product side of the above reaction the two in front of Al₂O₃ is multiplied by the subscripted two on Al for a total of 4 aluminums and multiplied by the subscripted three on the oxygen for a total of six oxygen.



This reaction is balanced because there are an equal number of each type of atom on both sides of the equation.



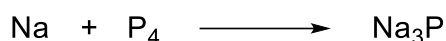
Balancing Chemical Reactions

When trying to balance a chemical equation, keep the following information in mind.

1. A balanced chemical equation has the same number and type of atom on both sides of the reaction arrow.
 - a. Identify the number of each atom or group of atoms on both sides of arrow.
 - b. Count how many there are on both sides of the reaction. If polyatomic ions are present, count them as a unit rather than as individual atoms.
 - c. If the number of atoms is all same the reaction is already balanced; if they are different, the reaction must still be balanced.
2. Pick one atom that is unbalanced and try to balance it by placing a coefficient in front of the formula. It helps to balance the atoms that appear in more than one compound on both sides of the reaction. Balance any free elements last.
 - a. Remember, only the coefficients may be adjusted in a reaction (never the subscripts in a formula).
 - b. Coefficients must be integers (2, 3, 4, 5, etc.); a coefficient of one (1) is usually not written.
 - c. The coefficients apply to all atoms that immediately follow it.
 - d. If the formula contains a polyatomic ion in parentheses, the coefficient is still multiplied by all of the atoms inside of the parentheses.
3. Once that atom is balanced, check to see if the coefficient affects other atoms in the formula, then balance the corresponding atom(s) on the other side of the equation.
4. Continue ping ponging back and forth until all atoms are balanced.

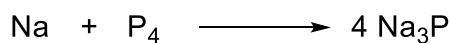
Example

Balancing when no polyatomic ions are present



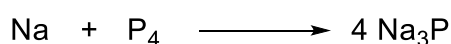
1	Na	3
4	P	1

Step 1: Reaction is not balanced (there are different numbers of each type of elements on both sides)



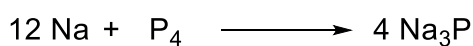
1	Na	3
4	P	1 4

Step 2: Try balancing the phosphorous by placing a 4 in front of the product.



1	Na	3 12
4	P	1 4

Step 3: The 4 in front of the product also changes the number of sodiums on the product side to 12.



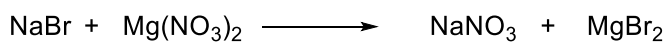
12 1	Na	3 12
4	P	1 4

Step 4: A coefficient of 12 is now needed on the reactant side of the equation to balance the sodium

The reaction is now balanced

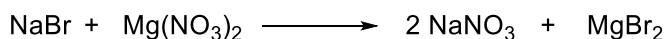
Example

Balancing when polyatomic ions are present



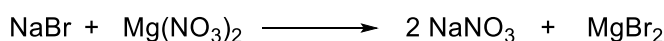
1	Na	1
1	Br	2
1	Mg	1
2	NO ₃	1

Step 1: Reaction is not balanced (there are different numbers of each type of elements on both sides)



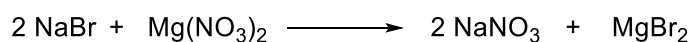
1	Na	1
1	Br	2
1	Mg	1
2	NO ₃	1 2

Step 2: The magnesium is already balanced. Try balancing the nitrate group first by placing a 2 in front of sodium nitrate



1	Na	1 2
1	Br	2
1	Mg	1
2	NO ₃	1 2

Step 3: The 2 in front of sodium nitrate also changes the number of sodiums on the product side to 2.



2 1	Na	1 2
2 1	Br	2
1	Mg	1
2	NO ₃	1 2

Step 4: A coefficient of 2 is now needed on the reactant side of the equation to balance the sodium. This also balances the bromine.

The reaction is now balanced

Classifying Reactions

There are several types of reactions that will be learned in this course.

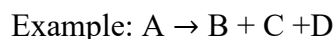
1. Synthesis (also called Combination) Reactions

- Recognized by two or more reactants forming a single type of product.
- These often contain a metal and a nonmetal forming an ionic compound.



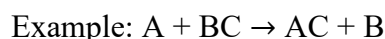
2. Decomposition Reactions

- Recognized by a single reactant forming two or more products.
- Predicting the product can often be more difficult than for other reactions but there will always be at least two different products formed.



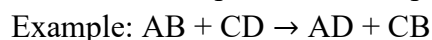
3. Single Displacement (or Replacement) Reactions

- Recognized by a free element reacting with a compound to form a new compound and a new element.
- The free element replaces the cation of the compound to form a new compound. The original cation becomes the new element.



4. Double Displacement (or Replacement) Reactions

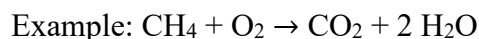
- Recognized by two different compounds forming two new compounds.
- The cation of each compound will swap the anion to which it is originally paired.



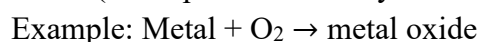
5. Combustion Reactions

There are two types of combustion reactions. Combustion of hydrocarbons and combustion of metals.

- Combustion of hydrocarbons are recognized by a hydrocarbon (a compound with only carbon and hydrogen) reacting with elemental oxygen (O_2) to form carbon dioxide and water.



- Combustion of metals is recognized by a metal reacting with elemental oxygen to form a metal oxide (a compound with only a metal and oxygen).

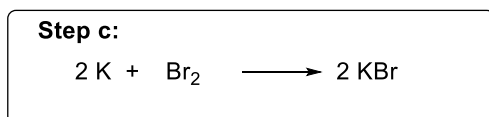
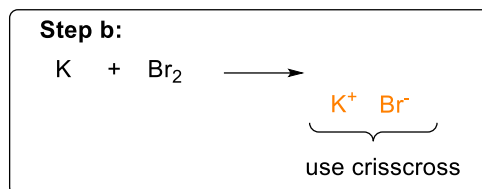
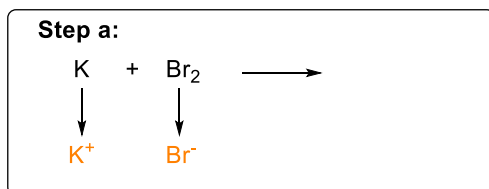


Predicting Products in a Reaction

When trying to predict the products in a reaction, being able to translate a name into a chemical formula is often required. It may be helpful to refer to the back to the Nomenclature Activity Handout to help with writing chemical formulas before beginning.

1. Synthesis/Combination) Reactions

- Determine the cation the metal forms and anion the nonmetal forms.
- Use the crisscross method to determine the formula of the new ionic compound (see Activity C Nomenclature for more information on the crisscross method).
- Balance as discussed above.



2. Decomposition Reactions

The products in a decomposition reaction depend on the type of reactant present. However, because the types of products that are formed are less straightforward than with other reactions, predicting the products for these reactions are not covered in this course.

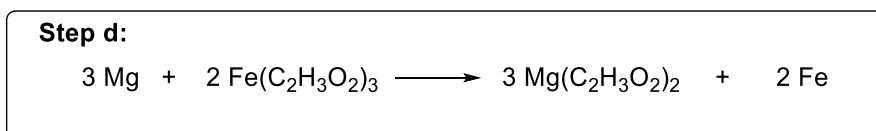
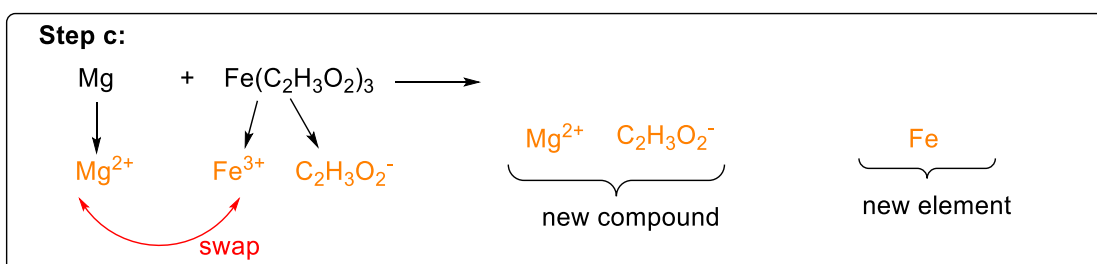
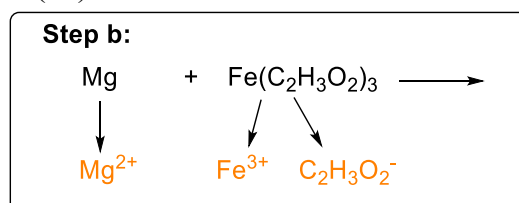
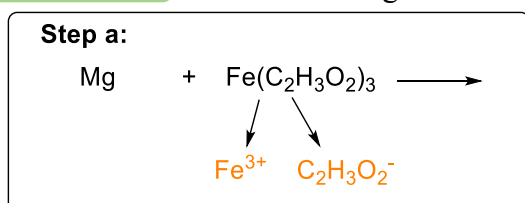
3. Single Displacement (or Replacement) Reactions

- Break apart the compound into the ions showing their respective charges. Use the reverse crisscross method if necessary.
- Determine the charge that the free element normally forms.
- Swap the cation that the anion is bonded to form the new compound. Use the crisscross method to determine the formula. The cation that now has nothing to bond to becomes the new element* (no charge is shown).
 - Most elements will be written as atomic elements (K, Cu, Pb, C, As, etc.)
 - Seven elements are diatomic and are written as X_2 and include H_2 , N_2 , O_2 , F_2 , Cl_2 , Br_2 , I_2
- Balance as discussed above.

*The wording in a reaction can indicate when an atom is in its elemental state. For example, descriptors like “solid copper” really means elemental copper or Cu, “liquid bromine” means elemental bromine or Br_2 , “gaseous nitrogen” means elemental nitrogen or N_2 , or “molecular oxygen” means elemental oxygen or O_2 .

Example

Predict the products and write a balanced equation for the single displacement reaction of magnesium and iron (III) acetate



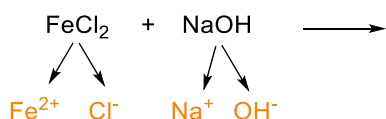
4. Double Displacement (or Replacement) Reactions

- Break apart compound into the ions showing their respective charges.
- The two cations will swap the anion to which they are bonded.
 - Keep in mind two cations cannot combine and two anions cannot combine.
 - Use the crisscross method to determine the formulas of the new compounds.
- Balance as discussed above.

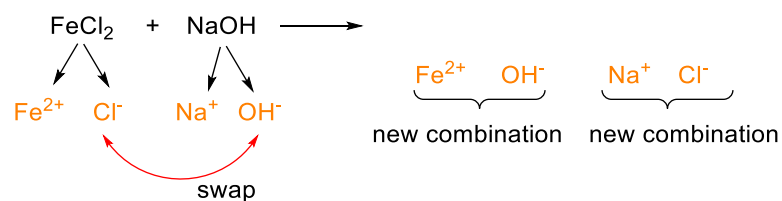
Example

Predict the products and write a balanced equation for the double displacement reaction of iron (II) chloride and sodium hydroxide

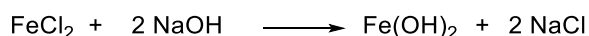
Step a:



Step b:



Step c:



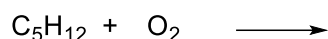
5. Combustion Reactions

- When balancing combustion of hydrocarbons remember O_2 must be added as a reactant.
- The only products will be CO_2 and H_2O .
- Try balancing carbon first, then hydrogen and finally oxygen.

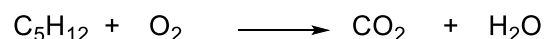
Example

Predict the products and write a balanced equation for the combustion of C_5H_{12} .

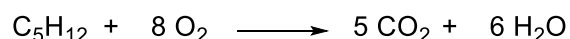
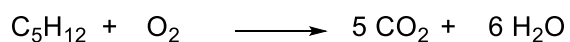
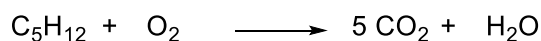
Step a:



Step b:



Step c:



There are now a total of 16 total oxygens on the product side (10 from the CO_2 and 6 from the H_2O). Therefore, a coefficient of 8 is needed on the reactant side to balance the oxygen.

- When balancing combustion of metals reactions, the formula of the metal oxide can be predicted by using the steps in balancing a synthesis reaction as described above. The charge of the metal cation will depend on if it is a monatomic or a variably charged metal. The charge of the monatomic metals is determined by their group number. If the metal is a variably charged metal, more information must be provided to know what cation is formed. The anion of the metal oxide will always be O^{2-} .

Predicting Physical States Using the Solubility Rules

The Solubility Rules (see Table 1 or Appendix IV in the lab manual) should be used to predict the physical state of an ionic compound in water. Ionic compounds with any of the ions listed under the **Soluble Ionic Compound** header are all soluble (have an aqueous physical state when) formed in water. Ionic compounds with any of the ions listed under the **Insoluble Ionic Compounds** header are all insoluble (have a solid physical state) when formed in water.

Table 1: Solubility Table

Top	Soluble Ionic Compounds	Important Exceptions
	Alkali metal ions	None
	Ammonium, NH_4^+	None
	Nitrate, NO_3^-	None
	Acetate, $\text{C}_2\text{H}_3\text{O}_2^-$	None
	Halides, Cl^- , Br^- , I^-	Compounds with Ag^+ , Hg_2^{2+} and Pb^{2+} are insoluble
	Sulfate, SO_4^{2-}	Compounds with Sr^{2+} , Ba^{2+} , Hg_2^{2+} and Pb^{2+} are insoluble
Bottom	Insoluble Ionic Compounds	Important Exceptions
	Carbonate, CO_3^{2-}	Compounds with ammonium and alkali metals are soluble
	Phosphate, PO_4^{3-}	Compounds with ammonium and alkali metals are soluble
	Chromate, CrO_4^{2-}	Compounds with ammonium and alkali metals are soluble
	Hydroxide, OH^-	Compounds with alkali metals and Ca^{2+} , Sr^{2+} and Ba^{2+} are soluble
	Sulfides, S^{2-}	Compounds with ammonium, the alkali metals and Ca^{2+} , Sr^{2+} and Ba^{2+} are soluble

Use the following steps to determine the solubility (or physical state) of an ionic compound in water.

- Determine which ions (cation or anion) are present in the ionic compound.
- Determine if solubility rules exist for either ion
- If the cation or anion is present in the top part (Soluble Ionic Compounds) of the table, the compound is soluble and will have the physical state “(aq)” for aqueous.
 - If an exception is listed, the compound will instead be insoluble and will have the physical state “(s)” for solid.
 - If neither ion is listed in the top part of the table, look in the bottom portion of the table (Insoluble Ionic Compounds)
- If the cation or anion is present in the bottom part (Insoluble Ionic Compounds) of the table, the compound is insoluble and will have the physical state “(s)” for solid.
 - If an exception is listed, the compound will instead be soluble and will have the physical state “(aq)” for solid.

ExampleDetermine the solubility $\text{Cu}(\text{NO}_3)_2$.**Step a: Determine ions present** $\text{Cu}(\text{NO}_3)_2$ means Cu^{2+} and NO_3^- **Step b: Determine if rule exist for ion(s)** Cu^{2+} not in table, can't use
to determine solubility NO_3^-

in solubility table

Step c: Assign Solubility using Top Soluble Table $\text{Cu}(\text{NO}_3)_2(\text{aq})$ nitrate has no exceptions so the
compound has the physical state "aq"**Step d: Assign Solubility using Bottom Insoluble Table**not needed; solubility has already
been assigned**Example**Determine the solubility FeOH .**Step a: Determine ions present** FeOH means Fe^{+} and OH^- **Step b: Determine if rule exist for ion(s)** Fe^{+} not in table, can't use
to determine solubility OH^-

in solubility table

Step c: Assign Solubility using Top Soluble Tablehydroxide not in the top table; must
use bottom table to decide solubility**Step d: Assign Solubility using Bottom Insoluble Table** $\text{FeOH}(\text{s})$ Fe^{+} is not one of the exceptions for hydroxide so
the compound will have the physical state "s"**Example**Determine the solubility AgI .**Step a: Determine ions present** AgI means Ag^{+} and I^- **Step b: Determine if rule exist for ion(s)** Ag^{+} not in table, can't use
to determine solubility I^-

in solubility table

Step c: Assign Solubility using Top Soluble Table $\text{AgI}(\text{s})$ Ag^{+} is an exceptions for iodide so the
compound has the physical state "s"**Step d: Assign Solubility using Bottom Insoluble Table**not needed; solubility has already
been assigned

Materials and Reagents Needed for Lab

This activity is considered a dry lab thus no chemicals will be used for this lab. Only a writing utensil and a calculator are needed for this lab.

Safety Considerations

Although this is a dry lab, students must wear appropriate PPE when in the Intro Chem Lab. No goggles or safety glasses are needed for this lab, however, students must still wear closed-toe shoes, long pants, shirts with a front/back and some type of sleeve.

Procedure

Follow the directions in the manual and given by the lab instructor

Part A: Classifying Reactions

Determine type of reaction that is shown. Use the descriptions in the Classifying Reactions section of the Background Information to help.

Part B: Balancing Equations

Balance each equation by writing the correct coefficients on the line in front of the chemical formula. Only coefficients greater than one (1) must be shown.

Part C: Determining Ions Present in a Compound

Break apart each ionic compound into its cation and anion. Using the reverse crisscross method can help with ion determination. Be sure to show both the formula and charge for each ion.

Part D: Predicting Products of Reactions

Predict the product(s) for each reaction. Use the descriptions in the Predicting Products in a Reaction section of the Background Information to help.

Part E: Using Solubility Rules

Use the Solubility Table in the Background Information to help the physical state of each compound in water.

Part F: Writing Balanced Equations

Use all of the skills described throughout the lab to predict the product(s) and write a balanced chemical equation for each reaction. Physical states must be included unless otherwise indicated.

Before Leaving Lab

- Make sure all items have been cleaned and put away.
- Check the board for any additional clean-up or waste instructions.
- Wipe down the benchtops.
- Refill the water bottles.
- Push in stools.
- Check with your instructor that your area has been sufficiently cleaned and reorganized.

Lab 7: Writing and Balancing Chemical Equations Activity – Pre-Lab

Name _____

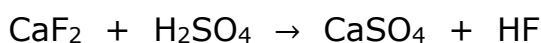
Section _____

Date _____

Pre-lab Questions

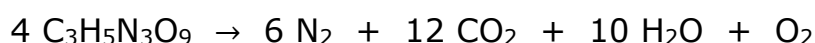
Read through all the background information and the procedure of the lab before answering these questions. For any calculations, all work must be shown. The work shown must include the formula, the unrounded answer, and the correctly rounded final answer. Answers without work may not receive full credit.

1. When writing a balanced equation, on what side of the arrow are products located?
2. Write the chemical formula(s) of all the reactants that are present in the following reaction.



Formula(s): _____

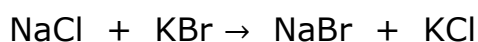
3. When balancing an equation, which can be changed or adjusted: subscripts or coefficients? Briefly explain.
4. Use the balanced equation below to determine the coefficients of each formula.



A) _____ $\text{C}_3\text{H}_5\text{N}_3\text{O}_9$

B) _____ CO_2

5. Explain why no coefficients are written for the reaction below even though it is a balanced chemical equation.

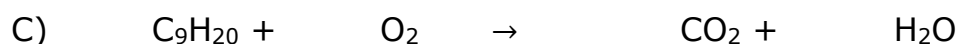
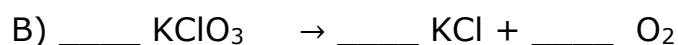
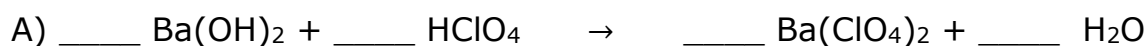


6. Classify each generic reaction as decomposition, single displacement, double displacement or synthesis.

A) $A + BC \rightarrow AC + B$ Reaction type: _____

B) $A \rightarrow B + C$ Reaction type: _____

7. Write out the coefficients needed to balance each reaction.



8. Use the Solubility Table to determine if each compound will be soluble or insoluble in water. Write the term **soluble** or **insoluble** next to each compound.

A) CaCO_3 _____ B) $\text{Ca}(\text{NO}_3)_2$ _____

9. What ions are present in solution when $\text{Zn}(\text{NO}_3)_2$ is dissolved in water.

10. Circle the correct answer.

A) When an ionic compound is soluble in water, what physical state will it have?

Aqueous Gas Liquid Solid

B) When an ionic compound is insoluble in water, what physical state will it have

Aqueous Gas Liquid Solid

Lab 7: Writing and Balancing Chemical Equations Activity – Data/Calculations

Name _____

Section _____

Drawer _____

Balance _____

Date _____

Partner(s) _____

Part A: Classifying Reactions

Directions: Classify each reaction as a combustion, decomposition, synthesis, single displacement or double displacement.

1. $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl}$ Reaction Type: _____

2. $\text{Mg} + 2 \text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$ Reaction Type: _____

3. $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ Reaction Type: _____

4. $2 \text{C}_4\text{H}_{10} + 13 \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$ Reaction Type: _____

5. $\text{Sr}(\text{OH})_2 + 2 \text{HClO}_3 \rightarrow \text{Sr}(\text{ClO}_3)_2 + 2 \text{H}_2\text{O}$

Reaction Type: _____

Part B: Balancing Equations

Directions: Write the coefficients needed to balance each reaction.

1. _____ $\text{Li} + \text{N}_2 \rightarrow \text{Li}_3\text{N}$

2. _____ $\text{Zn}(\text{C}_2\text{H}_3\text{O}_2)_2 + \text{Al} \rightarrow \text{Al}(\text{C}_2\text{H}_3\text{O}_2)_3 + \text{Zn}$

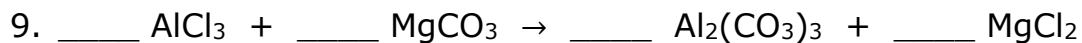
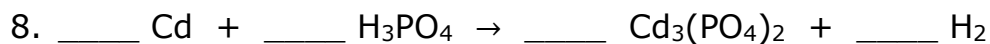
3. _____ $\text{N}_2 + \text{H}_2 \rightarrow \text{N}_2\text{H}_4$

4. _____ $\text{NaOH} + \text{MnSO}_4 \rightarrow \text{Na}_2\text{SO}_4 + \text{Mn}(\text{OH})_2$

5. _____ $\text{Fe}(\text{NO}_3)_3 + \text{H}_2\text{CO}_3 \rightarrow \text{Fe}_2(\text{CO}_3)_3 + \text{HNO}_3$

6. _____ $\text{PCl}_5 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{PO}_4 + \text{HCl}$

7. _____ $\text{C}_5\text{H}_{12} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$



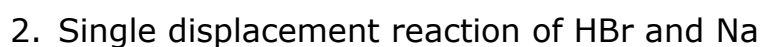
Part C: Determining Ions Present in a Compound

Directions: Break apart each compound into its ions. Be sure to show the charge on each ion.



Part D: Predicting Products of Reactions

Directions: Predict the products and write a balanced chemical equation for each reaction. Physical states do not need to be shown.



5. Single displacement reaction of Li and AlBr_3

6. Combustion reaction of C_7H_{16}

7. Double displacement reaction of Na_2SO_4 and $\text{Mg}(\text{NO}_3)_2$

Part E: Using Solubility Rules

Directions: Use the Solubility Table to determine if each compound will be soluble or insoluble in water. Write the word soluble or insoluble on the line.

- | | |
|---------------------------------------|---|
| 1. $\text{Cd}(\text{NO}_3)_2$ _____ | 9. Li_2CO_3 _____ |
| 2. AgCl _____ | 10. $\text{Ba}(\text{OH})_2$ _____ |
| 3. AlPO_4 _____ | 11. $\text{Hg}_2(\text{NO}_3)_2$ _____ |
| 4. $\text{Fe}_2(\text{SO}_4)_3$ _____ | 12. SrS _____ |
| 5. Ag_2S _____ | 13. $(\text{NH}_4)_2\text{SO}_4$ _____ |
| 6. $\text{Cu}(\text{OH})_4$ _____ | 14. BaSO_4 _____ |
| 7. CaCO_3 _____ | 15. $\text{KC}_2\text{H}_3\text{O}_2$ _____ |
| 8. KBr _____ | 16. $\text{Mg}_3(\text{PO}_4)_2$ _____ |

Part F: Writing Balanced Equations

Directions: Write a balanced equation for each reaction. For some reactions the product(s) will need to be predicted before balancing. Physical states must be indicated unless otherwise stated.

1. Gaseous carbon monoxide and gaseous oxygen forming gaseous carbon dioxide.

2. The combustion reaction of C_4H_8 . Ignore physical states.
3. Single displacement reaction of zinc and copper (II) sulfate. Ignore physical states.
4. Aqueous hydrobromic acid and gaseous elemental fluorine forming aqueous hydrofluoric acid and liquid elemental bromine.
5. Synthesis reaction of elemental chlorine with aluminum. Ignore physical states.
6. Soluble copper (II) chloride reacts with solid lead to form solid lead (II) chloride and copper metal.
7. Solid lead (II) nitrate decomposes to form solid lead (II) oxide, gaseous nitrogen monoxide and gaseous nitrogen dioxide and gaseous oxygen.

8. Aqueous solutions of acetic acid and calcium hydroxide forming liquid water and soluble calcium acetate.

9. Double displacement reaction of solid iron (III) hydroxide and aqueous sulfuric acid. Predict physical states of the products.

10. Ammonia (NH_3) reacting with hydrochloric acid to form ammonium chloride. Ignore physical states.

Lab 7: Writing and Balancing Chemical Equations Activity – Post-Lab Questions

Name _____

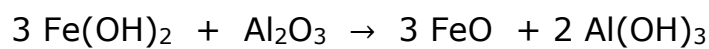
Section _____

Date _____

Post-lab Questions

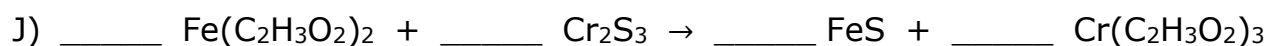
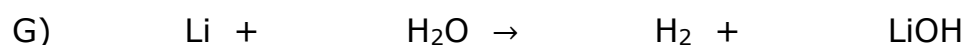
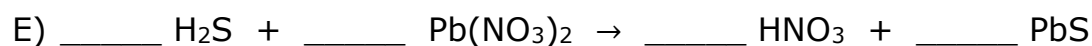
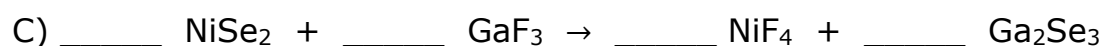
For calculations, all work must be shown. The work shown must include the formula, the unrounded answer and the correctly rounded final answer. Answers without work may not receive full credit.

1. When balancing an equation, what do the coefficients represent?
2. In the balanced equation below, how many hydroxide ions are present on the reactant side?



3. Classify each reaction as combustion, decomposition, single displacement, double displacement or synthesis.
 - A) The reaction of calcium metal reacting with fluorine gas to form solid calcium fluoride.
 - B) When methane (CH_4) is heated in the presence of oxygen it forms carbon dioxide and water.
 - C) The reaction of sodium chloride and silver nitrate forming sodium nitrate and silver chloride.

4. Balance each equation.



5. Predict the product(s) and write a balanced equation for each. Physical states may be ignored.

A) Solid lithium reacting with gaseous oxygen in a synthesis reaction.

B) The double displacement reaction between barium bromide and sulfuric acid.

C) The combustion of C₈H₁₆.

D) The neutralization reaction of acetic acid and potassium hydroxide.

6. What precipitate(s) will form when aqueous solutions of cadmium chloride and sodium sulfide are mixed?

7. Write the ions that are present for each aqueous solution.

A) $\text{H}_2\text{SO}_4(\text{aq})$ _____ C) $\text{Mg}(\text{NO}_3)_2(\text{aq})$ _____

B) $\text{K}_2\text{C}_2\text{O}_4(\text{aq})$ _____ D) $\text{AlCl}_3(\text{aq})$ _____

8. What does it mean for a compound to be insoluble in water?

9. Determine if the following are soluble or insoluble in water.

A) PbBr_2 _____ E) Fe_2SO_4 _____

B) NaI _____ F) $\text{Ca}(\text{OH})_2$ _____

C) SrSO_4 _____ G) HNO_3 _____

D) $(\text{NH}_4)_2\text{CrO}_4$ _____ H) $\text{Cu}(\text{OH})_2$ _____

10. Write a balanced equation for the precipitation reaction between aqueous solutions of aluminum nitrate and calcium hydroxide. Physical states must be included.

Lab 8: Copper Transformation Lab

Lab Objectives

- Perform a series of copper transforming reactions
- Write balanced equations for a series of copper reactions
- Identify physical evidence of a chemical reaction
- Classify reactions as decomposition, synthesis, single displacement or double displacement
- Apply the solubility rules for products formed in aqueous reactions

Background Information

In a chemical reaction one or more substances (the reactants) are changed into new and different substances (the products). The new products that are formed have physical properties and characteristic different from those of the reactants.

When a chemical reaction is performed often there are observable changes as evidence of the reaction. Common observations that a chemical change has taken place include color changes, formation of a gas, heat changes, production of light, or formation of a solid precipitate.

When a color change occurs the reaction mixture may become a different color (red to blue) or gain or lose color (colorless to pink). When a gas is produced faint or vigorous bubbling may occur. If a heat change occurs, the reaction flask may begin to feel warm or cold to the touch even when it is not actively being heated on a hot plate or cooled in an ice bath. Causing glowsticks to glow is an example of a reaction where a bright light is produced when specific reactants are mixed. A precipitate formation is evidenced by a cloudy solution, formation of a gel or jelly-like substance or a solid in solution. For some reactions one or more of these changes may be seen for the same reaction.

Common reaction types include decomposition, synthesis, single displacement or double displacement. **Decomposition** reactions involve a single reactant breaking down to form two or more products. **Synthesis** reactions are the opposite of decomposition reactions. In this type of reaction, several reactants combine to form a single type of product. In a **single displacement** reaction an element reacts with a compound to form a new element and a new compound. **Double displacement** reactions occur when two compounds that form ions in solution swap the anion to which they are bonded to form two new compounds.

Many of the above reactions take place in water (aqueous reactions) and it is important to remember that although water is present, it is often not a reactant. It can be formed as a product in a reaction as a liquid. But usually water serves as the solvent or the media that is used to increase the volume of a reaction or make it easier for the reaction to take place. The solubility (soluble or insoluble) of a substance in water can be determined using the solubility table (see Appendix IV).

In this lab a series of copper reactions will be performed and careful observation will be needed to determine what changes (physical and chemical) are taking place.

Materials and Reagents Needed for Lab

Lab Equipment

50 mL beaker (2)	Disposable pipet	10 mL graduated cylinder
250 mL beaker	100 mL graduated cylinder	Glass stir rod
Hot plate	Ceramic spot plate	Hot mitts
400 mL beaker	Graduated pipet	Weigh paper
Balance	Spatula	Small watch glass
Sharpie		

Reagents

Copper	Concentrated nitric acid	Copper (II) nitrate
Ice bath	Deionized water	3 M Sodium hydroxide
3 M Sulfuric acid	Zinc	

Safety Considerations

Gloves must be worn during the entire experiment. None of the blue solution from Part B can go down the drain or placed in the trash can. Nitric acid, sulfuric acid and sodium hydroxide can cause severe chemical burns and should never be handled with bare hands. Never put a pipet directly into a reagent bottle. Only deionized water should be used during the experiment. Turn off and unplug the hotplate when not in use. Never turn the hotplate on before it is needed. Since many of the solutions are all clear and colorless, label the beakers with a sharpie by writing directly on the side of the beaker (not the frosted glass). Keep hot glassware off of the lab mat or the mat will burn.

Procedure

Part A: Instructor Demo of Copper and Nitric Acid

1. The instructor will demonstrate the reaction between 1 g copper and 10 mL of concentrated nitric acid.
2. First record observations of each reactant before the reaction.
3. Watch the demonstration and record observations of any changes that occur.
4. Write a balanced chemical equation for the reaction.

Part B: Reaction of Copper (II) Nitrate and Sodium Hydroxide

1. Pour a small amount (less than half full) of the copper (II) nitrate stock solution into a 50 mL beaker. The stock solution is in a bottle labeled 0.5 g Cu in water. Record observations of the solution.
2. Use a disposable pipet to add 5 mL of the stock solution in the beaker to a 10 mL graduated cylinder.
3. Pour the solution into a 250 mL beaker. Rinse the graduated cylinder with deionized water into the same beaker until no blue solution remains in the graduated cylinder. Place the beaker in an ice bath. Leave beaker in the ice bath for the remainder of this section.
4. Measure out 100 mL deionized water into a 100 mL graduated cylinder. Pour the water into the same 250 mL beaker with the stock solution.
5. Measure out between 19 – 21 mL sodium hydroxide into a 100 mL graduated cylinder. Record observations.
6. **Slowly** pour the sodium hydroxide (with stirring using a glass stir rod) into the same 250 mL beaker with the water and stock solution. Continue stirring the solution for one minute. Record all observations.
7. Any excess copper (II) nitrate (blue stock) solution should be poured into the $\text{Cu}(\text{NO}_3)_2$ waste container in the hood. Rinse the disposable pipet with any blue solution also into the same waste container.

Part C: Reaction of Copper (II) Hydroxide

1. Remove the beaker from the ice bath and dry the sides and bottom of the beaker with a paper towel.
2. Place a hot plate (do not turn on yet), a ceramic spot plate and a hot mitt on the bench top. Never turn the hotplate on until the beaker is on top.
3. Place the beaker on the hot plate, turn on the hot plate and adjust the heat setting to between a setting of 6 – 7. While heating, **constantly stir** the solution with a glass stir rod. Be sure to stir the contents of the entire beaker, not just the top of the solution.
4. Continue stirring and heating the solution until the blue color is gone and only a black solid is dispersed throughout the solution. If any residual blue solid adheres to the walls of the beaker, use the deionized water bottle to

rinse the solid off the walls. Ask the instructor if it is okay to adjust the setting of the hotplate.

5. Once only a black solid in solution remains, turn off the hot plate. Use a hot mitt to remove the beaker from the hot plate and place it on the ceramic spot plate to cool. Do not place the hot beaker on the lab mat. Record all observations.
6. Once the solid has settled to the bottom of the beaker and the solution has cooled to room temperature (or just barely warm), slowly decant the clear, colorless (or slightly gray) solution into a 400 mL beaker leaving the black solid behind. If any black residue is seen in the 400 mL beaker, decant the solution again until no solid remains.
7. There should now only be black solid and a *minimal* amount of liquid in the 250 mL beaker. The 400 mL beaker should contain only a clear, colorless solution.

Part D: Reaction of Copper (II) Oxide and Sulfuric Acid

1. Pour a small amount (less than half full) sulfuric acid solution into a 50 mL beaker. Record any observations.
2. Using a graduated pipet, measure out 15 mL of sulfuric acid and add it to the contents in the 250 mL beaker with the black solid from Part C. Stir the mixture with a glass stir rod and record any observations. The reaction is finished once all the black solid is gone.

Part E: Reaction of Copper (II) Sulfate and Zinc

1. Place a folded (in half) piece of weigh paper on the balance and tare the balance. This means the balance should now read 0.000 g with the weigh paper on top.
2. Use a scoopula to measure out 1.9 – 2.0 g of zinc. Record any observations.
3. Add the zinc to the 250 mL beaker from Part D and cover with a small watch glass. Record any observations. Once all the bubbling has stopped and no additional color changes are seen the reaction is over.

Part F: Writing and Classifying Reactions

First clean up, then write the balanced equation that occurred in Parts A – E. Include physical states for all reactants and products.

Clean Up and Waste Disposal: Any waste from Part B, Part C or leftover sodium hydroxide should be added to Reaction 3 Waste Container. Any waste from Part D, Part E or leftover sulfuric acid should be added to Reaction 5 Waste Container. Excess zinc should go in the Excess Zinc Waste Container. Rinse the graduated pipet with water 2 – 3 times into an empty beaker. This acid waste can go into the Reaction 5 waste container. The graduated pipet can now be cleaned with soap and water. Remove the sharpie from any glassware withalconox/soap and water. Clean all glassware withalconox/soap solution, then rinse with tap water and again with deionized water. After drying the glassware, return items to their original location.

Before Leaving Lab

- Make sure all items have been cleaned and put away.
- Check the board for any additional clean-up or waste instructions.
- Wipe down the benchtops.
- Refill the water bottles.
- Push in stools.
- Check with your instructor that your area has been sufficiently cleaned and reorganized.

Lab 8: Copper Transformation Lab – Pre-Lab

Name _____

Section _____

Date _____

Pre-lab Questions

Read through all the background information and the procedure of the lab before answering these questions. For any calculations, all work must be shown. The work shown must include the formula, the unrounded answer, and the correctly rounded final answer. Answers without work may not receive full credit.

1. What are two safety considerations for this lab?
2. Use the website <https://www.chemicool.com/elements/> to find the following physical properties of elemental copper.
Atomic symbol: _____ Color _____
Physical state: _____ Density: _____
3. Give two physical changes that are indications that chemical reaction has occurred.
4. How are single displacement reactions different than double displacement reactions?
5. On what page number(s) in the lab manual can the solubility table be found?

6. What physical state (solid – s, liquid – l, gas – g, or aqueous – aq) should be written for each description?

Copper metal _____ Soluble zinc sulfate _____

7. What acid will be added to copper in the first step of the lab?
8. How much water should be added to the 250 mL beaker in Part B?
9. What mass of zinc will be used for this lab?
10. At the end of the lab, what should happen to any leftover sodium hydroxide?

Lab 8: Copper Transformation Lab – Data/Calculations

Name _____

Section _____

Drawer _____

Balance _____

Date _____

Partner(s) _____

Be sure to show all steps in calculations when necessary. Include formulas, units, unrounded answers and final rounded answers (with appropriate units) in the calculation table.

Table 1: Instructor Demo of Copper and Nitric Acid

Part A	Copper	Nitric acid	Nitrogen dioxide	Copper (II) nitrate	Water
Initial Observation					
Observations During and After Reaction					
Reaction	solid copper and aqueous nitric acid forms aqueous copper (II) nitrate and gaseous nitrogen dioxide and liquid water				
Balanced Equation					

Table 2: Copper (II) Nitrate and Sodium Hydroxide

Part B	Copper (II) nitrate	Sodium hydroxide	Copper (II) hydroxide	Sodium nitrate
Initial Observation				
Observations During and After Reaction				
Reaction	Aqueous copper (II) nitrate and aqueous sodium hydroxide form solid copper (II) hydroxide and aqueous sodium nitrate			
Balanced Equation				

Table 3: Copper (II) Hydroxide

Part C	Copper (II) hydroxide	Copper (II) oxide	Water
Initial Observation			
Observations During and After Reaction			
Reaction	Solid copper (II) hydroxide forms solid copper (II) oxide and liquid water		
Balanced Equation			

Table 4: Copper (II) Oxide and Sulfuric Acid

Part D	Copper (II) oxide	Sulfuric Acid	Copper (II) sulfate	Water
Initial Observation				
Observations During and After Reaction				
Reaction	Solid copper (II) oxide and aqueous sulfuric acid form aqueous copper (II) sulfate and liquid water			
Balanced Equation				

Table 5: Copper (II) Sulfate and Zinc

Part E	Copper (II) sulfate	Zinc	Copper	Zinc sulfate
Initial Observation				
Observations During and After Reaction				
Reaction	Aqueous copper (II) sulfate and solid zinc form aqueous zinc sulfate and solid copper			
Balanced Equation				

Lab 8: Copper Transformation Lab – Post-Lab Questions

Name _____

Section _____

Date _____

Post-lab Questions

For calculations, all work must be shown. The work shown must include the formula, the unrounded answer and the correctly rounded final answer. Answers without work may not receive full credit.

1. Give three examples of physical changes that occurred in this lab that were used to indicate a chemical reaction took place.
2. The solubility of copper (II) oxide is not explained using the solubility rules. Based on your observations that occurred during lab, what is the solubility of copper (II) oxide in water? Briefly explain.
3. Classify the reaction type (single displacement, double displacement, decomposition, synthesis, combustion) from each part of the lab.

Part B: _____

Part C: _____

Part D: _____

Part E: _____

4. Write the name and formula of copper product that was produced in each part of the lab.

Part A: Formula _____ Name _____

Part B: Formula _____ Name _____

Part C: Formula _____ Name _____

Part D: Formula _____ Name _____

Part E: Formula _____ Name _____

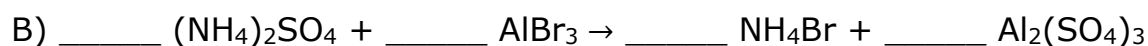
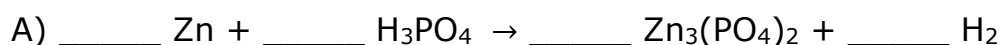
5. What physical state does water have when it is formed as a product? _____

6. Write the physical state solid (s) or aqueous (aq) in the parenthesis for each compound when it is formed in water using the Solubility Rules.

Na_2S () Fe_2S_3 () $(\text{NH}_4)_3\text{PO}_4$ () $\text{Cu}_3(\text{PO}_4)_2$ ()

$\text{Cr}(\text{OH})_6$ () $\text{Ca}(\text{OH})_2$ () PbBr_2 () TiBr_2 ()

7. Balance the following reactions. If a coefficient of one (1) is needed, leave the line blank.



8. Write a balanced chemical equation for each reaction. Must include physical states.

A) Soluble copper (II) chloride reacts with solid lead to give solid copper and lead (II) chloride.

B) Solid iron (III) hydroxide reacts with sulfuric acid to form soluble iron (III) sulfate and water. It may help to write water as HOH when balancing.

C) Aqueous solutions of acetic acid and calcium hydroxide form water and soluble calcium acetate. It may help to write water as HOH when balancing.

D) Fluorine gas reacts with soluble sodium bromide to form aqueous sodium fluoride and liquid bromine.

9. Write a balanced chemical equation that will occur in double displacement reaction between aqueous solutions of silver nitrate and lithium chloride. Include physical states.

10. Write a balanced chemical equation that will occur in single displacement reaction between magnesium metal and hydrochloric acid. Include physical states.

Lab 9: Stoichiometry Lab

Lab Objectives

- Describe or determine the mole ratio from a balanced chemical equation
- Use a balanced equation to solve for the mass of a substance that involves a mole-to-mole step
- Calculate the molar mass for elements and compounds using the periodic table
- Write a balanced chemical equation for a gas evolution reaction
- Perform a percent yield calculation after calculating the theoretical yield

Background Information

The driving force of some reactions is the formation or evolution of a gas. Certain compounds when formed as a result of a double displacement reaction will decompose to form a gas. For example, when carbonic acid, H_2CO_3 , is formed it will decompose to CO_2 and water. Other examples are in Table 1. When writing these reactions, it is helpful to balance it first as a double displacement reaction to help predict the formulas of the products. Then evaluate the products to see if any of them appear in Table 1. If a product appears in Table 1, then the reaction should be rewritten with the product(s) shown in the table. Nothing else about reaction should be changed. Notice in the table that when H_2S , hydrogen sulfide gas, forms it does not decompose but is itself a gas. To determine the physical state of the other product, the solubility table should be used. Remember when water is formed in a double displacement reaction, it will have a liquid (l) physical state.

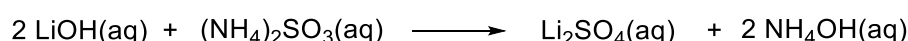
Table 1: Gas Evolution Products

Compound	Does it decompose?	Gas Evolution Products	
H_2S	No	$\text{H}_2\text{S(g)}$	
H_2CO_3	Yes	$\text{CO}_2\text{(g)}$	$\text{H}_2\text{O(l)}$
H_2SO_3	Yes	$\text{SO}_2\text{(g)}$	$\text{H}_2\text{O(l)}$
NH_4OH	Yes	$\text{NH}_3\text{(g)}$	$\text{H}_2\text{O(l)}$

Example

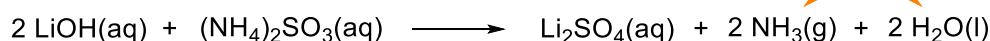
Write the gas evolution reaction for aqueous solutions of LiOH and $(\text{NH}_4)_2\text{SO}_3$

initial double displacement reaction:



This compound appears in the gas evolution table. It decomposes to form gaseous NH_3 and liquid water.

final gas evolution reaction:

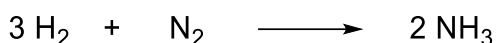


Reaction Stoichiometry

When a balanced chemical equation is written it can be used to relate the amount of one reactant or product to another. This is known as reaction stoichiometry. Since balanced equations show the molar relationships of all reactants and products the coefficients in the reactions represent the number of moles of each reactant and product. Mole ratios can compare any reactant to any product, any product to any reactant, any reactant to another reactant or any product to another product. In other words, any two parts of a reaction may be related as a mole ratio. For any ratio ($\frac{a}{b}$) that can be written, its reciprocal ($\frac{b}{a}$) will also be true.

Example

Write three different mole ratios for the reaction below.



Mole ratio 1:

$$\frac{3 \text{ mol H}_2}{1 \text{ mol N}_2}$$

Mole ratio 2:

$$\frac{1 \text{ mol N}_2}{2 \text{ mol NH}_3}$$

Mole ratio 3:

$$\frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2}$$

For each of the mole ratios written above their reciprocals will also be valid.

Mole ratio 1:

$$\frac{1 \text{ mol N}_2}{3 \text{ mol H}_2}$$

Mole ratio 2:

$$\frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2}$$

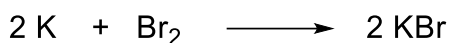
Mole ratio 3:

$$\frac{3 \text{ mol H}_2}{2 \text{ mol NH}_3}$$

These mole ratios can relate stoichiometric amounts in a calculation like the one below, acting as conversion factors.

Example

Calculate the number of moles of KBr that can form from 0.754 mol Br₂.



$$0.754 \text{ mol Br}_2 \times \frac{2 \text{ mol KBr}}{1 \text{ mol Br}_2} = 1.508 \text{ mol KBr} = \mathbf{1.51 \text{ mol KBr}}$$

Other types of measurements can also be determined using stoichiometry. For example, converting between mass and moles is a very common calculation used in stoichiometry problems. When performing this type of calculation, molar mass must be used. The molar mass of an element relates to the mass (in grams) of one mole of the element. The masses for each element can be found from the average atomic mass given in the periodic table. For example, 1 mole boron has a mass of 10.811 g. That means that the molar mass of boron can also be written as a conversion factor.

Example

Write the molar mass of boron as a conversion factor.

$$\frac{10.811 \text{ g B}}{1 \text{ mol B}} \quad \text{or} \quad \frac{1 \text{ mol B}}{10.811 \text{ g B}}$$

The molar mass of a compound relates the mass in grams for one mole of the compound. This can be determined by adding together all the masses for each element in the formula. For compounds, report the molar mass to four places after the decimal point.

Example

Write the molar mass of water as a conversion factor.

$$\frac{18.0153 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \quad \text{or} \quad \frac{1 \text{ mol H}_2\text{O}}{18.0153 \text{ g H}_2\text{O}}$$

$$2(\text{H}) + \text{O} = 2(1.00794) + 15.9994 = 18.01528$$

Since the formula of water (H₂O) has two hydrogens and one oxygen, the molar mass of water is calculated multiplying the mass of hydrogen from the periodic table by 2 and adding it the mass of oxygen.

When performing mole to mole, or mole to mass or mass to mole calculations, never start with the conversion factor. The first step in a calculation should always start with a measurement (just mass or just volume or just moles) *followed* by a conversion factor. Depending on the problem, one or more conversion factors may be necessary. To help keep track of the calculation be sure the units cancel correctly.

Example

Convert 0.8080 mol Mg to grams of Mg.

$$0.8080 \text{ mol Mg} \times \frac{24.3050 \text{ g Mg}}{1 \text{ mol Mg}} = 19.6384 \text{ g Mg} = \mathbf{19.64 \text{ g Mg}}$$

To convert from moles to mass, molar mass must be used. Since the calculation starts with moles, the conversion factor with moles in the denominator must be used to correctly cancel units.

Example

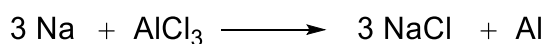
Convert 2.02 g CH₄ to moles of CH₄.

$$2.02 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.0425 \text{ g CH}_4} = 0.125916 \text{ mol CH}_4 = \mathbf{0.126 \text{ mol CH}_4}$$

To convert from mass to moles, again molar mass must be used. Since the calculation is starting with mass this time, the conversion factor with grams in the denominator must be used to correctly cancel units.

Example

Convert 6.4 mol AlCl₃ to moles of NaCl.



$$6.4 \text{ mol AlCl}_3 \times \frac{3 \text{ mol NaCl}}{1 \text{ mol AlCl}_3} = 19.2 \text{ mol NaCl} = \mathbf{19 \text{ mol NaCl}}$$

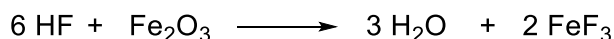
For this example, the amount of one compound needs to be converted to the amount of a different compound. To convert from moles of AlCl₃ to moles of NaCl the balanced chemical equation must be used to create a mole ratio.

Since the calculation starts with the moles of AlCl_3 , the mole ratio with AlCl_3 must be in the denominator to correctly cancel units.

Keep in mind that when doing calculations that involve molar mass and a balanced chemical equation, the molar mass is always based only on one mole (1 mol) of a substance regardless of the coefficients present in the balanced equation.

Example

Convert 0.370 mol Fe_2O_3 to grams of FeF_3 .



$$0.370 \text{ mol Fe}_2\text{O}_3 \times \frac{2 \text{ mol FeF}_3}{1 \text{ mol Fe}_2\text{O}_3} \times \frac{112.8402 \text{ g FeF}_3}{1 \text{ mol FeF}_3} = 83.50175 \text{ g FeF}_3 \\ = \mathbf{83.5 \text{ g FeF}_3}$$

For this example, two conversion factors are needed. The first step is to convert the moles of Fe_2O_3 to the moles of FeF_3 using the balanced chemical equation. Then the conversion factor for molar mass is needed to convert moles of FeF_3 to mass of FeF_3 .

Percent Yield

The mass that was calculated in the last example, is the theoretical yield of FeF_3 . The theoretical yield is the amount of a product that can form in a chemical reaction based on the balanced chemical equation and the amount of starting material present. This is the expected amount of product that can form and must be determined using the balanced equation. When the reaction is performed in the lab, and the mass of the product is determined experimentally, this is called the actual yield. The actual (or experimental yield) is the mass of a substance determined through experimentation. When the actual yield is compared to the theoretical yield and is expressed as a percentage, that is called the percent yield.

$$\% \text{Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

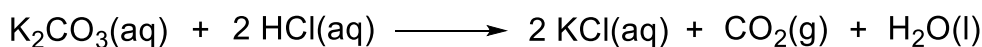
Example

Calculate the percent yield when 29.48 g is expected to form and 17.91 g are actually produced.

$$\% \text{Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{17.91 \text{ g}}{29.48 \text{ g}} \times 100 = 60.75305\% = \mathbf{60.75\%}$$

Example

A 0.562 g K_2CO_3 reacts with HCl . At the end of the reaction, 0.608 g KCl was obtained. Determine the experimental mole ratio of potassium chloride and potassium carbonate.



$$0.562 \text{ g K}_2\text{CO}_3 \times \frac{1 \text{ mol K}_2\text{CO}_3}{138.2055 \text{ g K}_2\text{CO}_3} = 0.0040664 \text{ mol K}_2\text{CO}_3$$

$$0.608 \text{ g KCl} \times \frac{1 \text{ mol KCl}}{74.5513 \text{ g KCl}} = 0.0081555 \text{ mol KCl}$$

$$\text{Experimental mole ratio} \quad \frac{\text{moles KCl}}{\text{mol K}_2\text{CO}_3} = \frac{0.0081555 \text{ mol}}{0.0040664 \text{ mol}} = 2.06 : 1.00$$

The experimental mole ratio (2:1) is consistent with the mole ratio of KCl and K₂CO₃ as defined by the balanced chemical equation.

Materials and Reagents Needed for Lab

Lab Equipment

150 mL beaker	Weigh paper	Balance
Spatula	10 mL graduated cylinder	Disposable pipet
Hot plate	Ceramic spot plate	Hot mitts or tongs
Glass stir rod	Casserole dish (2)	Small watch glass (2)
Large Test Tubes (2)	250 mL beaker	

Reagents

Sodium carbonate	Sodium bicarbonate	3 M hydrochloric acid
Deionized water		

Safety Considerations

Gloves must be worn during the entire experiment. Hydrochloric acid can cause severe chemical burns, eye damage and should never be handled with bare hands. Never place face directly over casserole dish when the acid is being heated. Do not turn on the hotplate until the casserole dish is on top. The hot casserole dish should not be placed on the lab mat. Turn off and unplug the hotplate when not in use. Never turn the hotplate on before it is needed. Label the test tubes with a sharpie.

Procedure

Part A: Reaction of Sodium Carbonate and Hydrochloric Acid

1. Obtain a casserole dish and small watch glass. Measure and record the mass of both together in **Table 1**. Set aside for later use. Keep the watch glass and the casserole dish together. Do not mix them up with the other set.
2. Place a piece of folded weigh paper on the balance and tare the balance.
3. Use a spatula to measure between 0.75 – 1.0 g sodium carbonate onto the weigh paper. Do not yet record this value (the actual mass will be recorded later).
4. Place a 250 mL beaker on the balance with one large test tube inside (to hold it upright) and tare the beaker/test tube.

5. Carefully pour all of the sodium carbonate solid into the test tube and now record the mass of the sodium carbonate in **Table 1**. Remove the beaker and test tube from the balance.
6. Measure between 5 - 7 mL of hydrochloric acid into a 10 mL graduated cylinder. Slowly pour the HCl into the test tube. Bubbling will occur so add the acid **slowly**. Once the bubbling has stopped gently stir the mixture with a glass stir rod.
7. If solid still exists, use the dropper vial to continue adding hydrochloric acid dropwise until all solid has dissolved. Remember to stir the solution after each addition. The reaction is finished when all bubbling has stopped, and the solution is clear and colorless with no visible solid.
8. Pour the solution from the test tube into the casserole dish.
9. Rinse the test tube with a minimal amount ($\sim 1 - 2$ mL) of deionized water and pour this mixture into the same casserole dish. Repeat once more.
10. Put the small watch glass on the casserole dish and place both on a hotplate. Adjust the hot plate to the maximum setting and turn it on.
11. Continue heating until all the liquid has been removed and only a white solid remains. There should be no liquid in the bottom of the casserole dish, and no condensation on the watch glass. Turn off the hotplate. At this point the casserole dish may be removed.
12. Use hot mitts or tongs to remove the casserole dish from the hotplate and place the casserole dish on a ceramic spot plate to cool. While waiting, Part B may be started.
13. Once the casserole dish is at room temperature, record together the mass of the casserole dish, watch glass and white solid in **Table 1**.
14. Clean up before working on the remaining calculations.
15. Write the balanced equation for the double displacement reaction between solid sodium carbonate and aqueous hydrochloric acid, including physical states, for this reaction in **Table 2**. Then write the gas evolution reaction for this same reaction.
16. Determine the actual mass of sodium chloride. Show work and report answer in **Table 3**.
17. Use the mass of sodium carbonate to calculate the theoretical mass, in grams, of sodium chloride.

18. Calculate the percent yield for this reaction.
19. Calculate the number of moles of sodium carbonate used in this experiment.
20. Calculate the number of moles of sodium chloride formed in this experiment.
21. Determine the experimental mole ratio of sodium chloride to sodium carbonate.

Part B: Reaction of Sodium Bicarbonate and Hydrochloric Acid

1. Obtain a casserole dish and small watch glass. Measure and record its mass in **Table 4**. Set aside for later use.
2. Place a piece of folded weigh paper on the balance and tare the balance.
3. Use a spatula to measure between 0.75 – 1.0 g sodium bicarbonate onto the weigh paper. Do not yet record this value (it will be recorded later).
4. Place a 250 mL beaker on the balance with one large test tube inside (to hold it upright) and tare the beaker/test tube.
5. Carefully pour all of the sodium bicarbonate solid into the test tube and now record the mass of the sodium bicarbonate in **Table 4**. Remove the beaker and test tube from the balance.
6. Measure between 5 - 7 mL of hydrochloric acid into a 10 mL graduated cylinder. Slowly pour the HCl into the test tube. Bubbling will occur so add the acid **slowly**. Once the bubbling has stopped gently stir the mixture with a glass stir rod.
7. If solid still exists, use the dropper vial to continue adding hydrochloric acid dropwise until all solid has dissolved. Remember to stir the solution after each addition. The reaction is finished when all bubbling has stopped, and the solution is clear and colorless with no visible solid.
8. Pour the solution from the test tube into the casserole dish.
9. Rinse the test tube with a minimal amount (~ 1 – 2 mL) of deionized water and pour this mixture into the same casserole dish. Repeat once more.
10. Put the small watch glass on the casserole dish and place both on a hotplate. Adjust the hot plate to the maximum setting and turn it on.
11. Continue heating until all the liquid has been removed and only a white solid remains. There should be no liquid in the bottom of the casserole dish, and

no condensation on the watch glass. Turn off the hotplate. At this point the casserole dish may be removed.

12. Use hot mitts or tongs to remove the casserole dish from the hotplate and place the casserole dish on a ceramic spot plate to cool.
13. Once the casserole dish is at room temperature, record together the mass of the casserole dish, watch glass and white solid **Table 4**.
14. Clean up before working on the remaining calculations.
15. Write the balanced equation for the double displacement reaction between solid sodium carbonate and aqueous hydrochloric acid, including physical states, for this reaction in **Table 5**. Then write the gas evolution reaction for this same reaction.
16. Determine the actual mass of sodium chloride. Show work and report answer in **Table 6**.
17. Use the mass of sodium bicarbonate to calculate the theoretical mass, in grams, of sodium chloride.
18. Calculate the percent yield for this reaction.
19. Calculate the number of moles of sodium carbonate used in this experiment.
20. Calculate the number of moles of sodium chloride formed in this experiment.
21. Determine the experimental mole ratio of sodium chloride to sodium carbonate.

Clean Up and Waste Disposal: Hot plates should be turned off, unplugged and left out to cool. Once the top of the hot plate is cool to the touch it may be put back in the lab drawer. Otherwise, leave the hot plate out on the bench top. Rinse the test tube, casserole dish and watch glass with the sodium chloride down the drain with running water. Any leftover acid should go in the acid waste container. Used weigh paper should go in the trash can. Any excess sodium carbonate or sodium bicarbonate can go in the trash can or down the drain with running water. Clean all glassware with alconox/soap solution, then rinse with tap water and again with deionized water. Cleaned test tubes should go in the white bucket by the window. After drying the remaining glassware, return items to their original location.

Before Leaving Lab

- Make sure all items have been cleaned and put away.
- Check the board for any additional clean-up or waste instructions.
- Wipe down the benchtops.
- Refill the water bottles.
- Push in stools.
- Check with your instructor that your area has been sufficiently cleaned and reorganized.

Lab 9: Stoichiometry Lab – Pre-Lab

Name _____

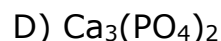
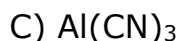
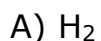
Section _____

Date _____

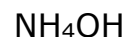
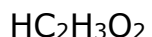
Pre-lab Questions

Read through all the background information and the procedure of the lab before answering these questions. For any calculations, all work must be shown. The work shown must include the formula, the unrounded answer, and the correctly rounded final answer. Answers without work may not receive full credit.

1. Write the molar mass for each formula. Give answer to four places after the decimal point for compounds.

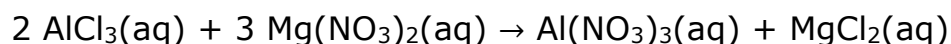


2. Circle the compound(s) that will decompose to form gas.

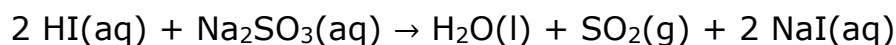


3. What gas is formed when carbonic acid decomposes? _____

4. Write the two different mole ratios for the relationship between $\text{Mg}(\text{NO}_3)_2$ and AlCl_3 using the balanced equation below.



5. Use the balanced equation below to determine the mass, in g, of NaI that can form from 0.564 g Na_2SO_3 .



6. In a reaction, the theoretical yield of the product is 3.421 g. Once the experiment is performed, if the 1.067 g are obtained, what is the percent yield?

7. After reading the procedure, indicate the total amount needed for this experiment.

A) Mass of Na_2CO_3 _____

B) Mass of NaHCO_3 _____

C) Volume of HCl _____

8. List two safety considerations for this lab.

9. When should the casserole dish be removed from the hotplate?

10. How should the sodium chloride be disposed at the end of the lab?

Lab 9: Stoichiometry Lab – Data/Calculations

Name _____ Section _____

Drawer _____ Balance _____ Date _____

Partner(s) _____ Date _____

Be sure to show all steps in calculations when necessary. Include formulas, units, unrounded answers and final rounded answers (with appropriate units) in the calculation table.

Part A: Reaction of Sodium Carbonate and Hydrochloric Acid

Table 1: Sodium Carbonate Data

	Measurement	Mass Data
A.	Mass of empty casserole dish and watch glass	
B.	Mass of sodium carbonate	
C.	Mass of casserole dish, watch glass and sodium chloride	

Table 2: Balanced Equation

Balanced Chemical Equation	
Double displacement	
Gas evolution	

Table 3: Sodium Carbonate Calculations

	Measurement	Calculation Table
D.	Actual mass of sodium chloride	
E.	Theoretical mass of sodium chloride	
F.	Percent Yield	
G.	Moles of sodium chloride	
H.	Moles of sodium carbonate	
I.	Experimental mole ratio sodium chloride to sodium carbonate	

Part B: Reaction of Sodium Bicarbonate and Hydrochloric Acid

Table 4: Sodium Bicarbonate Data

	Measurement	Mass Data
A.	Mass of empty casserole dish and watch glass	
B.	Mass of sodium bicarbonate	
C.	Mass of casserole dish, watch glass and sodium chloride	

Table 5: Balanced Equation

Balanced Chemical Equation	
Double displacement	
Gas evolution	

Table 6: Sodium Bicarbonate Calculations

	Measurement	Calculation Table
D.	Actual mass of sodium chloride	
E.	Theoretical mass of sodium chloride	

F.	Percent Yield	
G.	Moles of sodium chloride	
H.	Moles of sodium bicarbonate	
I.	Experimental mole ratio sodium chloride to sodium bicarbonate	

Lab 9: Stoichiometry Lab – Post-Lab Questions

Name _____

Section _____

Date _____

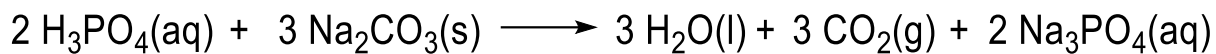
Post-lab Questions

For calculations, all work must be shown. All work shown must include the formula, the unrounded answer, and the correctly rounded final answer. Answers without work may not receive full credit.

1. What is the name of the white solid left in the beaker after the end of the experiment in both reactions.
2. Using the balanced equations for Parts A and B, write one mole ratio for the following.
A) Na_2CO_3 and HCl

B) NaHCO_3 and HCl
3. Did the experimental mole ratio you calculated match the mole ratio predicted by the balanced equation for this lab? Briefly explain.
4. How many grams of CO_2 are predicted to form from Part A of this experiment. (Hint: use your measured mass of Na_2CO_3)

5. Use the balanced equation to answer the following questions.



A) Write the two mole ratios relating phosphoric acid and sodium carbonate.

B) What mass, in g, Na_3PO_4 will theoretically form from 0.2468 g Na_2CO_3 ?

C) What is the percent yield of Na_3PO_4 if 0.1534 g are actually formed?

6. Write the gas evolution reaction for the reaction between aqueous ammonium chloride and aqueous calcium hydroxide. Include physical states.

Lab 10: Molecular Models Activity

Lab Objectives

- Determine the number of electrons in different bond types
- Explain electronegativity (EN) and rank elements based on their EN
- Draw or interpret Lewis-dot symbols for individual elements
- Explain the octet rule, its exceptions and when to use the octet rule
- Draw or interpret Lewis structures that contain up to four electron sets
- Identify the electronic geometry and molecular geometry of a molecule based on the VSEPR theory, including bond angles

Background Information

The structure of a molecule can have a big effect on both the physical and chemical reactivity of a molecule. One such effect is whether a covalent molecule is polar or nonpolar. Molecules that are polar mix well with other molecules that are also polar, but do not interact well with molecules that are nonpolar and vice versa for nonpolar compounds. For example, oil (a nonpolar substance) and water (a polar substance) do not mix well, but water (polar) and sugar (a polar substance) will mix. Whether a molecule is polar or nonpolar has a lot to do with the type and arrangement of atoms and bonds within the molecule.

Types of Bonds

In all covalent molecules electrons are shared between atoms to form a covalent bond. A covalent bond occurs when two atoms (nonmetals or semimetal) share two or more electrons. When two electrons are shared between them a single bond result. When two atoms share four electrons a double bond is formed, and a triple bond is the result of sharing six electrons (see Figure 1).

Figure 1: Types of Covalent Bonds

single bond	double bond	triple bond
X : Y	X :: Y	X ::: Y
X—Y	X=Y	X≡Y

However, sometimes the electrons are not “shared” equally. This can occur when one atom tends to pull the bonded electrons more toward itself, much like playing a game of tug of war with a body builder. This results in one end of the bond with more electron density (δ^-) and one end with less electron density (δ^+). When this happens, a bond has a dipole (two ends with different amounts of electron density), and a polar covalent bond is formed. To predict which atom is better at pulling electrons toward itself in a covalent bond, the electronegativity (EN) of the atom is needed. The electronegativity of an atom refers to how strongly an atom pulls bonded electrons toward itself. Atoms with large electronegativity values pull the electrons more strongly than atoms with small electronegativity values. In general, electronegativity

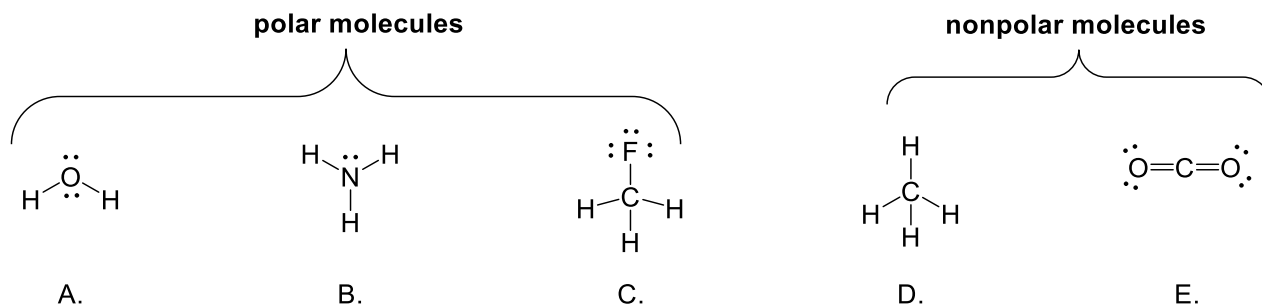
H 2.1																
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.7	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.1

180

Molecular Polarity

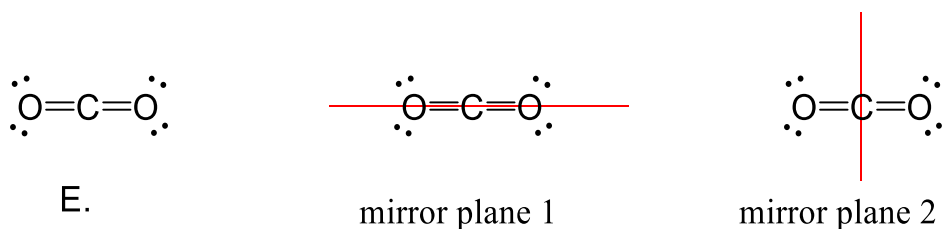
An entire molecule, not just a bond, can also be polar. Polar molecules always contain one or more polar bonds. They are also asymmetrical in how the electron density is distributed throughout the molecule. This often occurs if there are lone pair electron groups on the central atom or if the molecule lacks two perpendicular mirror planes (see Figure 2).

Figure 2: Polar and Nonpolar Molecules



Molecules A, B and C are all examples of polar molecules. In Molecule A, the O-H bonds are polar, and the central atom has two lone pair electron groups. Molecule B also has polar N-H bonds and has a lone pair electron group on the central atom. Molecule C does not have any lone pair electron groups on the central carbon. However, it is also polar due to the C-F bond being polar and the molecule lacking two reflective mirror planes that are 90° to each other. Molecules D and E are both nonpolar. For Molecule D none of the bonds are polar, that alone is enough to make the entire molecule nonpolar. Both of the C=O bonds in Molecule E are polar, but this molecule has two mirror planes that are perpendicular (90°) to each other so its electron distribution is symmetrical throughout the molecule (see Figure 3). Thus, it is a nonpolar molecule.

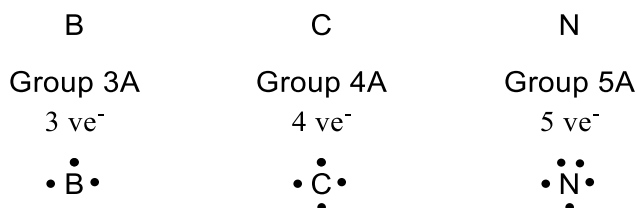
Figure 3: Two Reflective Perpendicular Mirror Planes



Molecules A – E are examples of Lewis structures. Lewis structures show the lone pair electrons (represented by dots) on an atom and the bonding electrons (represented by lines) used to connect an atom to another atom in a molecule or polyatomic ion. Lewis structures are usually only drawn for covalent compounds that contain Main Group elements (Groups 1A – 8A). The electrons used to draw Lewis structures are only the valence electrons on each atom. The group number that ends in A (1A, 2A, 3A, etc.) that an element is in will be the same as its number of electrons. For example, carbon is in Group 4A and thus has 4 valence electrons. The four valence electrons on carbon can be shown using a Lewis electron dot structure. To draw Lewis electron dot structure for an atom, imagine the atom in an imaginary box surrounded by four sides. Place one dot (the first electron) any one side of the atom, if additional electrons are present, continue placing the dots (one per side) until all four sides of the atom have a single dot. If additional electrons are present, the electrons can now be doubled until all electrons have been used.

Example

Draw a Lewis Dot Structure for a boron, carbon and nitrogen atom.



The Lewis electron dot structure can be expanded to more than just a single element. For molecules, compounds and polyatomic ions, Lewis structures are drawn to show their atom connectivity. To keep track of the number of valence electrons (ve⁻) in a molecule or polyatomic ion, add up all of the valence electrons based on the number and type of elements that are present.

Example

How many valence electrons are in C₃H₈?

$$\begin{aligned}
 \text{ve}^- &= 3(\text{C}) + 8(\text{H}) + \text{charge} \\
 &= 3(4) + 8(1) + 0 \\
 &= \mathbf{20 \text{ ve}^-}
 \end{aligned}$$

In the compound there are three carbon atoms (each with four valence electrons), eight hydrogen atoms (each has just one valence electron). Since the compound is neutral, no additional electrons must be added or removed. When the Lewis structure is drawn for C₃H₈, a total of 20 valence electrons must be shown in its Lewis structure.

Keep in mind that electrons are negatively charged. Therefore, when counting the number of valence electrons in a polyatomic ion, the charge of the ion must also be considered. For every negative charge that is present on a polyatomic ion, an additional electron must be added to the total count. For every positive charge on a polyatomic ion, an electron must be removed from the total count.

Example

How many valence electrons are in NH₄⁺?

$$\begin{aligned}
 \text{ve}^- &= 1(\text{N}) + 4(\text{H}) + \text{charge} \\
 &= 1(5) + 4(1) + -1 \\
 &= \mathbf{8 \text{ ve}^-}
 \end{aligned}$$

In this polyatomic ion there is one nitrogen atom (with five valence electrons), four hydrogen atoms (with one valence electron each) and an overall charge of 1+. Since this polyatomic ion has a *positive* overall charge, one valence electron must be *subtracted*. One electron is removed from the overall count of ve⁻. The Lewis structure for NH₄⁺ must have 8ve⁻.

Example

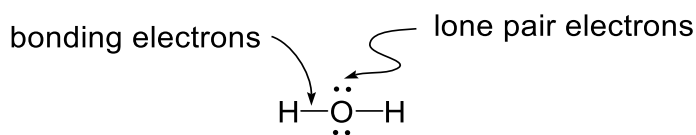
How many valence electrons are in SO₄²⁻?

$$\begin{aligned}
 \text{ve}^- &= 1(\text{S}) + 4(\text{O}) + \text{charge} \\
 &= 1(6) + 4(6) + 2 \\
 &= \mathbf{32 \text{ ve}^-}
 \end{aligned}$$

In this polyatomic ion there is one sulfur atom (with six ve⁻), four oxygen atoms (each with six ve⁻) and an overall charge of 2-. Since this polyatomic ion has a *negative* overall charge, two valence electrons must be *added* to the overall count of ve⁻. The Lewis structure for SO₄²⁻ must have 32ve⁻.

Lewis Structures

Once the number of valence electrons are known for a formula, a Lewis Structure may be drawn. Keep in mind that when drawing Lewis Structures, each line represents two electrons and any lone pair electrons must be shown.



In the Lewis structure for water, H_2O , oxygen has eight electrons around it. Two are in each of the single bonds and two are in each lone pair electron group. Remember bonding electrons are the electrons that are shared between atoms and lone pair electrons are the electrons that belong to only one atom.

Most atoms (except hydrogen) will also have an octet. Achieving an octet is one reason nonmetals (and semimetals) share electrons with other atoms. When an atom has an octet, it has an especially stable configuration of its electrons. This is so common for Main Group Elements, that it is called the Octet Rule. The Octet Rule states that an atom is stabilized by having its highest valence energy level filled with eight electrons (n^2p^6). This can help with predicting the number and types of bonds an atom will form with another atom. Looking at the Lewis Electron Dot structure for carbon, it is easy to see it would need four more electrons to fulfil its octet. Therefore, carbon will always bond in such a way that it has four bonds in its Lewis Structures.

For every rule, there are always exceptions. Hydrogen, for example, will NEVER have an octet. Since hydrogen is a small atom with only an s-orbital, it can only ever have two electrons around it. As a result, it should always be drawn with just a single bond. Boron also tends to defy the octet rule. Boron is usually drawn with only six valence electrons around it in its Lewis Structures. Both hydrogen and boron are electron deficient (have fewer than eight electrons in their structures) but are still stable. Sulfur and phosphorous are two examples that sometimes go in the other way. Those two atoms can occasionally be hypervalent (have more than eight valence electrons) in the Lewis Structures but are still stable. This does not mean that S and P will always be hypervalent, only that they can be. Other atoms can also be hypervalent or electron deficient, but never C, N, O or F.

Most of the compounds that will be drawn for this lab will be small molecules with 2 – 5 atoms. When drawing a Lewis Structure for a compound, use the following steps.

Step for Drawing Lewis Structures

1. Determine the total number of valence electrons in the compound.
2. Pick the central atom (the atom with the lowest group number that is not hydrogen) and then place the other atoms around it.
3. Use a single bond to connect the outer (terminal) atoms to the central atom.
4. If additional electrons are present, complete the octet of the outer atoms first by placing them in groups of two around the atom. (If needed as lone pair electron groups)
5. Complete the octet of the central atom. (If needed with any remaining lone pairs or multiple bonds as needed).
6. Use formal charge to evaluate the structure.

Keep in mind, these steps are *general steps* and all of them may not apply to all compounds. For example, some molecules may not have central atom. The compound HBr does not have a central atom but a Lewis structure may still be drawn for it. In some compounds once all the outer atoms are connected to the central atom all the electrons have already been used and steps 4 and 5 are not necessary. In the case of boron, although it is often the central atom, it will usually not have an octet. The best way to understand what to do and when is to practice drawing Lewis structures. Then keep practicing more.

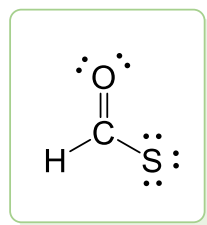
Once a Lewis Structure has been drawn it should be evaluated to see if it is a good structure. Calculating the formal charge (FC) on each atom can help with this. Formal charge is a way of identifying when an atom is not stable in a compound. To calculate formal charge, use the following equation.

Equation 1: Formal charge = valence electrons – covalent bonds – lone pair electrons.

When possible, draw Lewis Structures such that all atoms have a formal charge of zero. If an atom has a formal charge of zero, nothing needs to be indicated for that atom. However, if the formal charge is not zero for an atom, the formal charge must be shown by writing the charge near the atom. It sometimes helps to put the formal charge in a circle so that it is apparent to which atom it belongs. The sum of the formal charges on all atoms should be consistent with the overall charge of the molecule or polyatomic ion.

Example

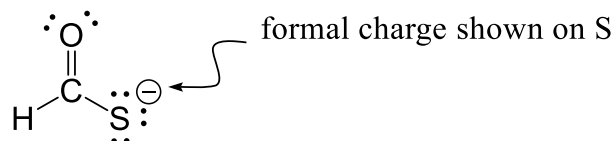
Show the formal charge on each atom when necessary for the CHOS⁻ structure below.



$$FC_H = 1 - 1 - 0 = 0$$

$$FC_O = 6 - 2 - 4 = 0$$

$$FC_S = 6 - 1 - 6 = -1$$



In this example, only sulfur has a formal charge that is not zero. Since it has a FC = -1, the negative one is written in small circle next to the sulfur atom. Since C, O and H all have a FC = 0, nothing needs to be indicated on those atoms.

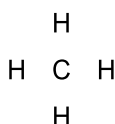
Example

Draw a Lewis Structure CH₄. Indicate formal charge when necessary.

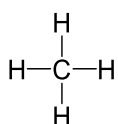
Step 1

$$ve^- = 8 \text{ } ve^-$$

Step 2



Step 3



Step 4, 5

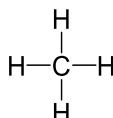
For this example these steps are unnecessary because all electrons have been used and the middle atom already has an octet.

Step 6

$$FC_C = 4 - 4 - 0 = 0$$

$$FC_H = 1 - 1 - 0 = 0$$

Final Structure

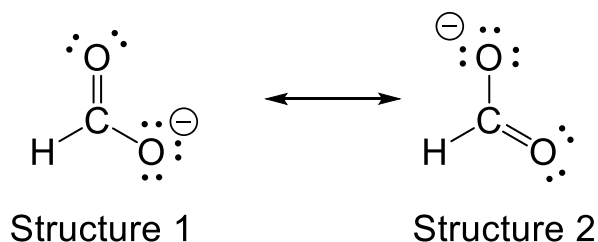


Since all atoms have a $FC = 0$, no formal charge needs to be written.

Sometimes more than one Lewis Structure can be drawn for the same formula, and it looks like a double bond, or a lone pair electron group has “moved”. These are usually examples of resonance structures. Resonance structures are used when a single Lewis structure cannot adequately show the bonding between the atoms.

Example

Draw a Lewis structure for CHO_2^- . Indicate formal charge when necessary.



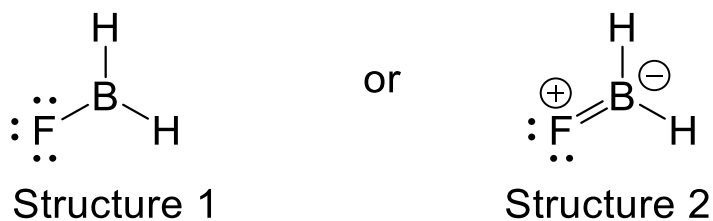
Both structures are valid. In both Lewis structures, carbon has a single bond to hydrogen, a single bond to one oxygen and a double bond to the other. The overall formal charge is the same for both structures, so they are equivalent. Since either oxygen can be drawn with the double bond and the same structure results, these are a pair of resonance structures.

When trying to evaluate if one Lewis Structure is better than another, or if the structure drawn is a “good” Lewis Structure, keep the following in mind. A good Lewis Structure will have all the following (in order of importance):

1. The correct number of valence electrons.
2. Atoms have a formal charge of zero when possible.
3. Appropriate atoms following the octet rule
 - a. FONC can never violate the octet rule
 - b. Some atoms can be electron deficient (H, B are common)
 - c. Some atoms can be hypervalent (S, P are common)

Example

Determine which is a better structure.



Both structures have the correct number of valence electrons. Structure 1 has a formal charge of zero for all atoms. In this structure, boron does not have an octet (it is electron deficient) but it still has a $FC_B = 0$. In Structure 2, although all atoms obey the octet rule, they do not all have a zero FC ($FC_F = +1$ and the $FC_B = -1$). Therefore, between these two resonance structures, Structure 1 is better Lewis structure.

Some Lewis structures can contain double or triple bonds. Knowing when to use a double or a triple bond is not always easy. One way to recognize it is when all the valence electrons have been used, but not all of the atoms have an octet.

Example

Draw a Lewis structure for C_2H_4 .

Step 1 $ve^- = 12 ve^-$

Step 2

```

  H   H
   C C
  H   H
  
```

Step 3

```

  H   H
 :C - C:
  H   H
  
```

All electrons have been use but both carbons do not have an octet.

Step 4, 5

The LP electron group has to be shared between the two carbon atoms. It is the only way to get rid of the formal charge on both carbons

```

  H   H
 :C - C:
  H   H
  
```

$FC_C = 4 - 3 - 2 = -1$ (left carbon)

$FC_C = 4 - 3 - 0 = +1$ (right carbon)

Step 6

$FC_C = 4 - 4 - 0 = 0$

$FC_H = 1 - 1 - 0 = 0$

Final Structure

```

  H   H
   C = C
  H   H
  
```

Electron Geometry and Valence Shell Electron Pair Repulsion

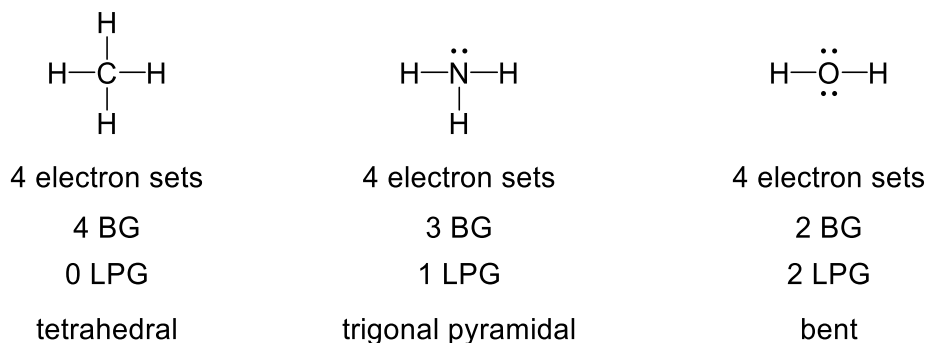
After drawing several Lewis structures a few trends may begin to emerge. Many elements tend to always have a certain number of bonds or lone pair groups when they have a formal charge of zero (see Table 2). Carbon, and other elements in group 14), tend to have four bonds and no lone pairs. The four bonds can be four single bonds, two double bonds, one double bond and two single bonds, or one single bond and one triple bond. However carbon is drawn, it always has four lines drawn to it in its Lewis structures.

Table 2: Common bonding Patterns for Elements with a Zero Formal Charge

Carbon and Group 14 Elements	Nitrogen and Group 15 Elements	Oxygen and Group 16 Elements	Fluorine and Group 17 Elements
4 Bonds	3 Bonds	2 Bonds	1 Bond
0 Lone Pair Groups	1 Lone Pair Group	2 Lone Pair Groups	3 Lone Pair Groups

Lewis structures are drawn as 2D images, but in reality, molecules are three dimensional and can adopt various shapes due to the arrangement of electrons present. The Valence Shell Electron Pair Repulsion (VSEPR) theory states that electron groups tend to spread out as much as possible in covalent compounds. The number of electron sets (or groups) around the central atom can be used to determine the three-dimensional electron and molecular geometry. The electron geometry indicates the arrangement of electrons around a central atom while the molecular geometry describes the arrangement of atoms within the molecule. When counting the number of groups around an atom, a single bond, a double bond, a triple bond and a lone pair of electrons all count as a single group of electrons.

For example, the compounds CH_4 , NH_3 and H_2O all have four electron sets (groups of electrons) around their respective central atoms, but yet they adopt different molecular geometries.



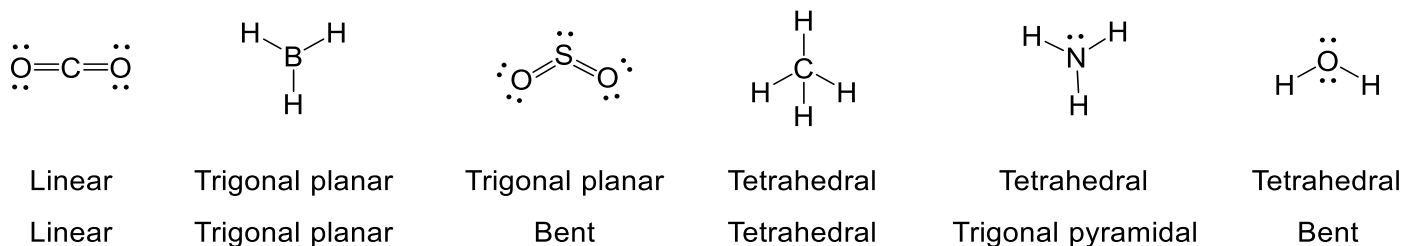
In the case of CH_4 , all four of the electron sets are bonding groups so the molecular geometry is the same as the electron geometry. Each H-C-H bond angle is 109.5° which is typical for all tetrahedral compounds. For the compound, NH_3 , of the four electrons sets, three are bonding groups and one is a lone pair electron group. The presence of the lone pair electron group causes a distortion of the tetrahedral geometry and causes the H-N-H bond angle to shrink down to 107° . This distorted shape is called trigonal pyramidal. Finally, for the compound H_2O there are two bonding groups and two lone pair electron groups. The second lone pair electron group causes an even greater distortion and compresses the H-O-H bond angle down to 104° . This molecular geometry is called a bent geometry.

Table 3 has the general bond angles and names of the electron and molecular geometries that will be used in this lab.

Table 3: Valence Shell Electron Pair Repulsion (VSEPR)

Electron Sets	Electron Geometry	Bond Angle	Bonding Groups	Lone Pair Groups	Molecular Geometry
2	Linear	180°	2	0	Linear
3	Trigonal planar	120°	3	0	Trigonal planar
			2	1	Bent
4	Tetrahedral	109.5°	4	0	Tetrahedral
			3	1	Trigonal pyramidal
			2	2	Bent

Each of the electron geometries and molecular shapes are shown in the examples below. The Electron Geometry is the top name and the Molecular Geometry is the bottom name.



Notice when counting electron sets, a single bond, double bond, triple bond and a lone pair electron group all count as one group. Knowing that molecule has a tetrahedral molecular geometry may be different than being able to visualize what that three-dimensional shape actually is, therefore, for this lab, ball and stick model kits will be used to make the visualization easier for this lab.

Materials and Reagents Needed for this Activity

This activity is considered a dry lab thus no chemicals will be used for this lab. Only a writing utensil and a model kit are needed for this lab.

Model Kit Includes: Element (color) – count

Carbon (black) – 6	Hydrogen (white) – 1
Nitrogen (blue, 3 holes) – 1	Nitrogen (blue, 4 holes) – 2
Oxygen (red, 2 holes) – 6	Oxygen (red, 4 holes) – 2
Halogens (green) – 6	Boron (beige, 5 holes) – 1
Phosphorous (beige, 4 holes) – 1	Phosphorous (beige, 5 holes) – 1
Sulfur (beige, 6 holes) – 1	Paddles (lone pair group) – 6
Single bonds (gray, short) – 20	Single bonds (purple, short) – 5
Double/triple bonds (gray, long) – 12	

When using the model kit, keep the following in mind.

- For lone pair groups, use the paddles for the central atom only. (Remember a lone pair electron group (••) counts as a single group)
- For single bonds, the short stubby, bonds should be used.
- For double and triple bonds, the long skinny bonds should be used.
- If a double bond is needed, use two of the long bonds with both ends connecting both atoms. (Remember a double bond counts as one group. So, two double bonds would represent two groups).
- If a triple bond is needed, use three of the long bonds with both ends connecting both atoms. (Remember a triple bond counts as one group)

Safety Considerations

Although this is a dry lab, students must wear appropriate PPE when in the Intro Chem Lab. No goggles or safety glasses are needed for this activity; however, students must still wear closed-toe shoes, long pants, shirts with a front/back and some type of sleeve.

Procedure

Follow the directions in the manual, written on the board and given by your lab instructor.

Part A: Formal Charge

Show how the formal charge is calculated for the silicon and oxygen in the structures provided. Use the names of the polyatomic ions to write the chemical formula for each. Draw a Lewis structure. If more than one Lewis structure can be drawn only one Lewis structure needs to be drawn. Determine the number of valence electrons for each formula. Calculate the formal charge (FC) for each atom. Show the formal charge for all atoms where it is not zero.

Part B, C and D: Lewis Structures and Molecular Geometry

In Tables C, D and E, first determine the number of valence electrons (#ve-) for each structure. Then draw its Lewis structure. Use the model kit to build the ball-and-stick version of the Lewis structure. Sketch what the ball and stick model looks like. When appropriate, determine the number of bonding groups (#BG), the number of lone pair groups (#LPG) and the molecular geometry for the central atom.

Part E: Electronegativity and Bond Polarity

Use the electronegativity table to write the electronegativity (EN) of each element, determine the electronegativity difference (ΔEN) and type of bond that would be present between each pair of atoms. For the bonds that are polar, show the direction of the dipole using (δ^+ and δ^-). Then answer the remaining questions.

Clean Up: Disassemble all models and place back in the kit.

Before Leaving Lab

- Check the board for any additional instructions.
- Wipe down the benchtops
- Push in stools
- Check with your instructor that your area has been sufficiently cleaned and reorganized

Lab 10: Molecular Models Activity – Pre-Lab

Name _____

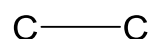
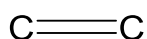
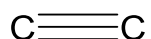
Section _____

Date _____

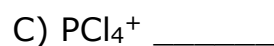
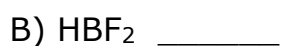
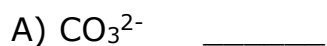
Pre-lab Questions

Read through all the background information and the procedure of the lab before answering these pre-lab questions.

1. Determine the number of electrons in each carbon carbon bond.



2. Calculate the total number of valence electrons in each formula.



3. Explain the two electronegativity trends found in the periodic table.

Trend 1: _____

Trend 2: _____

4. Draw a Lewis electron dot symbol for a Se atom. _____

5. Which element(s) must always follow the octet rule? _____

6. Answer the following for the structure shown in the box on the right.

A) Number of electron sets _____

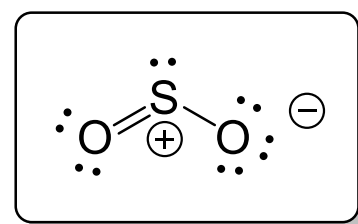
B) Number of bonding groups _____

C) Number of lone pair groups _____

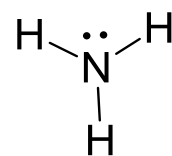
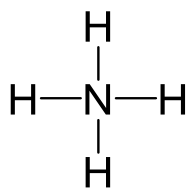
D) Electron geometry _____

E) Molecular geometry _____

F) Bond angle for O-S-O _____



7. Determine the formal charge on nitrogen in each structure below.



8. Draw a Lewis structure for CHCl₃.

9. Draw a Lewis structure for PF₃.

10. Draw a Lewis structure for SCl₂.

Lab 10: Molecular Models Activity – Data/Calculations

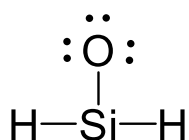
Name _____ Section _____

Drawer _____ Balance _____ Date _____

Partner(s) _____ Date _____

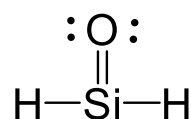
Part A. Formal Charge

- First calculate and show the formal charge (FC) on silicon and oxygen in each structure. Then explain which structure is better.



$$\text{FC}_{\text{Si}} =$$

$$\text{FC}_{\text{O}} =$$



$$\text{FC}_{\text{Si}} =$$

$$\text{FC}_{\text{O}} =$$

- Complete the following table.

Name	Valence Electrons	Chemical Formula	Lewis Structure (with FC)
Hydroxide			
Ammonium			
Carbonate			

Part B. Formulas with Single Bonds

Molecule	#Ve ⁻	Lewis Structure	Ball and Stick Model	Central Atom		Molecular Geometry
				# BG	#LPG	
H ₂						
Cl ₂						
HCl						
H ₂ O						
BCl ₃						
CH ₄						
CH ₂ Cl ₂						
H ₂ O ₂						
NH ₃						
N ₂ H ₄						
NH ₂ OH						

Part C. Formulas with at Least One Double Bond

Molecule	#Ve ⁻	Lewis Structure	Ball and Stick Model	Central Atom		Molecular Geometry
				# BG	#LPG	
O ₂						
C ₂ H ₄						
CO ₂						
HONO						
CH ₂ O						
HNO						

Part D. Formulas with at Least One Triple Bond

Molecule	#Ve ⁻	Lewis Structure	Ball and Stick Model	Central Atom		Molecular Geometry
				# BG	#LPG	
N ₂						
C ₂ H ₂						
HCN						

Part E. Electronegativity and Bond Polarity

Bond	Zn—H		F—C		C—S		N—O		Li—Cl	
EN										
Δ EN										
Bond Type										

1. Briefly explain which bond would be more polar: C-N or C-Br
2. The formulas HPS and HNO both have similar structures, but one molecule one is more polar than the other. Draw the Lewis structure for each. Then explain which molecule would be more polar and why.
3. Use the δ^+ and δ^- symbols to show the direction of the dipole.

S—C

S—O

S—F
4. The formulas HCl and LiH both involve hydrogen. Use electronegativity to explain why hydrogen is written in a different location in both formulas.

Lab 10: Molecular Models Activity – Post-Lab Questions

Name _____

Section _____

Date _____

Post-lab Questions

1. How many bonds does each atom typically form?

Carbon _____

Oxygen _____

Hydrogen _____

2. Which atom will never follow the octet rule? _____

3. Draw Lewis structure for Si_2H_2 . What type of bond is present between the silicon atoms?

4. Draw Lewis electron dot structures for each atom: Ge, I, H and Ar

5. Draw a Lewis structure for CF_2Cl_2 .

6. Draw a Lewis structure for H_2BI .

7. Draw a Lewis structure for PCl_3 .

8. Complete the following table.

Molecule	#Ve ⁻	Lewis Structure	Electron Geometry	Molecular Geometry
H ₃ O ⁺				
PF ₃				
NO ₂ ⁻				
CS ₂				
SiO ₃ ²⁻				
SO ₂				
HCN				
HBFCI				

Lab 11: Gas Laws Lab

Lab Objectives

- Identify and perform unit conversions for units of pressure
- Convert temperatures from degrees Celsius to Kelvin
- Calculate different variables (V, T, n or P) of gases using Boyle's Law, Charles's Law, Avogadro's Gas Law and the Ideal Gas Law Equation
- Perform stoichiometry for gases using the Ideal Gas Law
- Calculate using Dalton's Law of Partial Pressure

Background Information

Gases behave very differently than liquids and solids. For example, solids molecules are strongly attracted to each other due to electrostatic or intermolecular force attractions and are relatively stationary. These strong but stationary attractions give solids their definite shape and volume. The atoms and molecules in liquids are also attracted to each other but have more ability to move around compared to solids. The result is that liquids have a definite volume but no specific shape. If 25 mL of water is poured into a beaker it will take on the shape of that beaker, but if that same water is then poured on the table, it will not maintain the shape of the beaker. Instead, it will have an amorphous shape as nothing is there to determine what shape it should have.

Gas particles on the other hand are not strongly attracted to other gas particles, as a result there is a lot of empty space between them; therefore, gases have no definite shape and no definite volume. They can be compressed or expanded to fill any container they are in. Gases also have a very high kinetic energy (energy of movement) compared to liquids and solids. When these gas particles collide with the walls of a container, pressure is created.

The pressure of a gas can be measured in a variety of units including atmospheres (atm), millimeters of mercury (mm Hg), pounds per square inch (psi) or torr. The last unit is in honor of Evangelista Torricelli who invented the barometer. The relationship between the different units of pressure are given in Table 1. All of the values given in Table 1 are equivalent to 1 atm.

Table 1: Units of Pressure

Unit Name	Equivalent to 1 atm
Pascal	101,325 Pa
Inches of Hg	29.92 in Hg
Millimeter of Hg	760 mm Hg
Torr	760 torr
Pounds per square inch	14.70 psi

Example

Convert 0.645 atm to mm Hg.

$$0.642 \text{ atm} \times \frac{760 \text{ mm Hg}}{1 \text{ atm}} = 487.92 \text{ mm Hg} = \mathbf{488 \text{ mmHg}}$$

Individual Gas Laws

Pressure is not the only variable of a gas that can be measured. In addition to their pressures (P), gases can be described in terms of their volume (V), temperature (T) and amounts (n). When a change is made to one variable, the effect it has on another variable can be measured assuming the other variables are all held constant. For example, if a certain volume of a gas exerts specific pressure, and the volume changes the new pressure can be observed when temperature and amount are held constant.

The relationship that describes how the pressure and volume of a gas is related is called Boyle's Law. This law was named after Robert Boyle, a philosopher who studied the characteristics of air pressure and respiration. Boyle found that pressure and volume had an inverse relationship. As one goes up, the other goes down. Boyle's Law states that as the volume of a gas increases, the pressure will decrease and vice versa, as long as the amount and temperature of a gas are constant and do not change. This law is expressed by the relationship below.

Boyle's Law

$$P_1V_1 = P_2V_2$$

In Boyle's Law the subscript one and two refer to the respective initial and final conditions of the gas. As long as both pressure units or both volume units are in the same unit, they do not need to be converted when using this law. If, for example, the initial volume is in liters and the second volume is also in liters, no unit conversion need occur. But if the first volume is in liters and the second volume is in milliliters, then one of the units must be converted to the other.

Example

When a gas occupies 4.5 L it exerts a pressure of 2.7 atm. What pressure, in atm, will it exert if the volume is changed to 3.2 L?

$$P_2 = \frac{P_1V_1}{V_2} = \frac{2.7 \text{ atm} \times 4.5 \text{ L}}{3.2 \text{ L}} = 3.79688 \text{ atm} = \mathbf{3.8 \text{ atm}}$$

The formula must be rearranged to solve for the new pressure, P_2 . Notice the new volume went down therefore the pressure went up.

Jacques Charles studied the thermal expansion of gases and investigated the correlation between the temperature and volume of a gas. What he found was a direct relationship between temperature and volume. This means as one variable increase or decrease, the other does the same. Charles's Law states that at constant pressure, if the temperature of a gas increases, its volume will also increase. If the temperature decreases, the volume will also decrease.

Charles's Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

When using Charles's Law, the temperature must be in Kelvin. If another unit is given, it must first be converted to Kelvin before use. To convert a temperature in degrees Celsius to Kelvin 273.2 must be added to the temperature in degrees Celsius. When converting units of temperature, whatever decimal place or place value the original temperature is given in, the final temperature should be reported to the same place value or decimal place.

Example

When a gas is at 62.0°C, it occupies a volume of 9.80 mL? What volume, in mL, will it occupy at 88.4°C?

$$V_2 = \frac{V_1}{T_1} \times T_2 = \frac{9.80 \text{ mL}}{335.2 \text{ K}} \times 361.6 \text{ K} = 10.571838 \text{ mL} = \mathbf{10.6 \text{ mL}}$$

$$T_1 = 62.0^\circ\text{C} + 273.2 = 335.2 \text{ K} \quad T_2 = 88.4^\circ\text{C} + 273.2 = 361.6 \text{ K}$$

The formula must be rearranged to solve for the new volume, V_2 . Remember the temperatures must be converted to Kelvin before they can be used in the Charles's Law formula.

In this lab, the volume of a balloon will be calculated using the formula of a sphere. Although the balloons are not perfectly spherical that shape will be used as a close approximation for the volume of the balloon. The volume of a sphere is calculated from the below formula, where "r" is the radius of the sphere.

Sphere Volume

$$V = \frac{4}{3} \times \pi \times r^3$$

If the circumference of the sphere is known, the radius can be determined.

Circumference

$$C = 2\pi r$$

The volume and amount, in moles, of a gas are related through Avogadro's Law. Amedeo Avogadro studied gases and found the number of molecules of a gas was directly proportional to its volume at a constant temperature and pressure. As the number of molecules increases, the volume those molecules would occupy must also increase. This correlation between volume and amount is known as Avogadro's gas law.

Avogadro's Law

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

When using the Avogadro's Gas Law, the amount must always be in moles.

Example

If 206 g of CO_2 occupies a volume of 0.203 L, what amount, in moles, of gas will occupy 0.598 L?

$$n_2 = \frac{n_1}{V_1} \times V_2 = \frac{4.68081 \text{ mol CO}_2}{0.203 \text{ L}} \times 0.598 \text{ L} = 13.78879 \text{ mol CO}_2 = \mathbf{13.8 \text{ mol CO}_2}$$

$$n_1 = 206 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.0095 \text{ g CO}_2} = 4.68081 \text{ mol CO}_2$$

The formula must be rearranged to solve for the new amount, n_2 . The mass of a gas cannot be used in Avogadro's Gas Law. It must be converted to moles using the molar mass of the compound.

When trying to decide which gas should be used, it is helpful to look at what variables are changing versus what variables are staying the same. Remember that Boyle's Law involves changing conditions between pressure and volume, Charles's Law involves changing conditions between volume and temperature and Avogadro's gas Law involves the changing conditions between amount and volume.

There is another gas law, the Ideal Gas Law, which relates all of the variables, pressure, volume, temperature and amount. When using this gas law, the pressure must always be in atmospheres (atm), the volume is in liters (L), the amount in moles (mol) and the temperature in Kelvin (K).

Ideal Gas Law

$$PV = nRT$$

Notice that when using one of the other named gas laws, two of the units would also cancel so that only one final unit was obtained. With the Ideal Gas Law, none of the other units would cancel without the Ideal Gas Law Constant, R. This variable relates the other four variables with the specific units given above, has a value of $0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$ and is what allows the other units to cancel.

Example

When gaseous 6.0 g HCl is in a 1.02 L container at 298 K and what is the pressure, in atm?

$$P = \frac{nRT}{V} = \frac{0.16456 \text{ mol} \times 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \times 298 \text{ K}}{1.02 \text{ L}} = 3.94715 \text{ atm} = \mathbf{3.9 \text{ atm}}$$

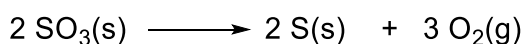
$$n = 6.0 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.4609 \text{ g HCl}} = 0.16456 \text{ mol} \quad T = 25^\circ\text{C} + 273.2 \text{ K} = 298 \text{ K}$$

The formula must be rearranged to solve for the Pressure, P. The mass must again be converted to moles using the molar mass and the temperature must be in Kelvin in order for the Ideal Gas Law Constant to correctly cancel units.

The Ideal Gas Law can also be used to convert the amount of a gas to a different substance that is related by a balanced chemical equation. Keep in mind the Ideal Gas Law can only be used for substances that are gases.

Example

When 7.932 g of SO₃ is decomposed, what volume, in mL, of O₂ will be produced if the below reaction takes place at 1.855 atm and 303.4 K?



$$V = \frac{nRT}{P} = \frac{0.14861 \text{ mol} \times 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \times 303.4 \text{ K}}{1.855 \text{ atm}} = 1.99555 \text{ L} \times \frac{1 \text{ mL}}{1 \times 10^{-3} \text{ L}} = 1995.55 \text{ mL} = \mathbf{1996 \text{ mL}}$$

$$n = 7.932 \text{ g SO}_3 \times \frac{1 \text{ mol SO}_3}{80.0632 \text{ g SO}_3} \times \frac{3 \text{ mol O}_2}{2 \text{ mol SO}_3} = 0.14861 \text{ mol O}_2$$

The formula must be rearranged to solve for the Pressure, P. The mass must again be converted to moles using the molar mass and the temperature must be in Kelvin in order for the Ideal Gas Law Constant to correctly cancel units.

Dalton's Gas Law and Vapor Pressure

When a mixture of gases is present in a container but not reacting, each individual gas will exert its own force, or pressure, on the walls of the container. If the pressure of each individual gas is then added together, a total pressure (P_{tot}) can be calculated for all the gases combined. This is known as Dalton's Law of Partial Pressures. The number of terms in the equation is dependent upon the number of different gases present.

Dalton's Law of Partial Pressure

$$P_{\text{tot}} = P_1 + P_2 + P_3 + \dots$$

Many reactions take place in water and as a result, water vapor (or gaseous water) may also be present. For example, if the SO_3 decomposition reaction above took place in water both $\text{O}_2(\text{g})$ and $\text{H}_2\text{O}(\text{g})$ would be present. If gases are in a closed container, the total pressure would come from the sum of the pressure from O_2 and the pressure of water. The vapor pressure of water that will be present is dependent upon the temperature of the reaction. This is because at higher temperatures, more liquid water can evaporate and form gaseous water, thus creating a higher pressure. Once the temperature is known, the vapor pressure of water can be determined using a vapor pressure table like the one in Table 2. A more complete vapor pressure table may be found in Appendix V.

Table 2: Vapor Pressure of Water

Temperature (°C)	Pressure (mm Hg)	Temperature (°C)	Pressure (mm Hg)	Temperature (°C)	Pressure (mm Hg)
21.0	18.5	21.4	18.9	21.8	19.3
21.1	18.6	21.5	19.0	21.9	19.4
21.2	18.7	21.6	19.1	22.0	19.5
21.3	18.8	21.7	19.2	22.1	19.6

When a reaction takes place in a container open to the atmosphere, the total pressure of the gaseous sample is the same as the pressure of the atmosphere. Since barometers can be used to measure the pressure of the atmosphere the total pressure is the barometric pressure.

Example

Calculate the barometric pressure, in mm Hg, if the partial pressure of hydrogen gas is 740.6 mm Hg and the reaction takes place over water at 21.6 °C?

$$P_{\text{Bar}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}} = 740.62 \text{ mm Hg} + 19.4 \text{ mm Hg} = 760.02 \text{ mm Hg} = \mathbf{760.0 \text{ mm Hg}}$$

The barometric pressure (P_{bar}) comes from using Dalton's Law of Partial pressure for hydrogen gas (P_{H_2}) and water vapor ($P_{\text{H}_2\text{O}}$). The partial pressure of water must come from a vapor pressure table.

Example

Calculate total pressure (atm, hundredths) when 0.348 g Zn reacts with HBr over water in a 0.105 L container at 22.0°C.

$$P_{\text{tot}} = P_{\text{H}_2\text{O}} + P_{\text{H}_2}$$



Step 1: $P_{\text{H}_2\text{O}} = 19.8 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.026053 \text{ atm}$

Step 2: $0.348 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.382 \text{ g Zn}} \times \frac{1 \text{ mol H}_2}{1 \text{ mol Zn}} = 0.005323 \text{ mol}$

Step 3:
$$P_{\text{H}_2} = \frac{nRT}{V} = \frac{(0.005323 \text{ mol H}_2) (0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}) (295.2 \text{ K})}{0.105 \text{ L}} = 1.2286 \text{ atm}$$

Step 4:
$$P_{\text{tot}} = 1.2286 \text{ atm} + 0.026053 \text{ atm} = 1.2547 \text{ atm} = \mathbf{1.25 \text{ atm}}$$

Dalton's Law of partial pressure must be used to find the total pressure. The vapor pressure of water comes from the vapor pressure table in Step 1. To find the partial pressure of H₂, first the number of moles of H₂ must be determined as in Step 2 then that be used in the Ideal Gas Law to find the pressure of H₂ as in Step 3. Finally, in Step 4 the two pressures can be added together to give the total pressure.

In this lab several gas laws will be investigated.

Materials and Reagents Needed for Lab

Lab Equipment

1 small balloon	1 medium balloon	60 mL syringe
Measuring tape	Thermometer	Tongs
Liquid nitrogen dewar	400 mL beaker	250 mL beaker
150 mL beaker	Sharpie	Buret clamp
Buret stand	Gas collection tube	Rubber stopper with copper loop
Ruler	Scissors	25 mL graduated cylinder
Disposable pipet		

Reagents

Deionized water	Magnesium	3 M hydrochloric acid
Liquid nitrogen		

Safety Considerations

The syringe cap must be held firmly in place, or it can become a projectile. HCl is corrosive to the eyes and skin. Be sure to wear gloves when handling hydrochloric acid. Liquid nitrogen may cause severe frost bite, cold burns or asphyxiation. Do not place face or bare hands near or in the liquid nitrogen vapor. Liquid nitrogen must be used in a well-ventilated room. Do not put the thermometers in the liquid nitrogen.

Procedure

Part A: Boyle's Law Experiment

1. Write the formula for Boyle's Law in **Table 1**.
2. Obtain and partially blow up a small balloon and tie it.
3. Remove the syringe tip and the plunger from a 60 mL syringe and insert the partially inflated balloon inside the syringe.
4. Replace the plunger into the syringe until the bottom of the syringe is close to the 30 mL mark.
5. Close the syringe tip with a syringe cap and hold with the palm of the hand.

6. Slowly push the plunger into the syringe and observe what happens. Record observations as Observations #1 in **Table 1**.
7. Slowly pull the plunger out of the syringe with the end still covered and observe what happens try not to let the balloon pop. Record observations as Observations #2 in **Table 1**.
8. Remove the balloon and replace the cap and the plunger on the syringe. Throw the balloon away.

Part B: Charles's Law Experiment

1. Write the formula for Charles's Law in **Table 2**.
2. Blow up a medium balloon to its maximum volume (without popping) and tie it.
3. Use a measuring tape to record the circumference (distance around the widest part) of the balloon. Record the circumference, in cm, of the balloon in **Table 2**.
4. Assume the balloon is a sphere. Calculate the radius, r , of the balloon in cm. Use the radius to find the volume, in cm^3 , of the balloon. Record as the initial volume of the balloon.
5. Measure the temperature of the room. Record this as the initial temperature of the gas in the balloon.
6. Take the balloon and a pair of tongs to the instructor station.
7. Place the balloon in the liquid nitrogen bath and use the tongs to push the balloon as far into the bath as possible.
8. Record your observations of what happens to the balloon once it is in the liquid nitrogen bath.
9. The temperature of the liquid nitrogen bath is -196.0°C . Record this as the final temperature of the gas in the balloon. Do not put the thermometer in the liquid nitrogen.
10. Using tongs, remove the balloon from the liquid nitrogen bath. Continue holding the balloon in the tongs and observe what happens. Record observations.
11. Take everything back to the lab bench.
12. Replace all labware and throw the balloon away.

13. After cleaning up, calculate the initial volume in cm^3 , the initial and final temperature in K and the final volume in cm^3 in **Table 3**.

Part C: Determination of the Ideal Gas Constant, R

1. Write the balanced equation for the single displacement reaction between magnesium and hydrochloric acid. Include physical states.
2. Write the formula for the Ideal Gas Law formula in **Table 4**.
3. Obtain a 400 mL beaker, a 250 mL beaker and a 150 mL beaker. Use a sharpie to label the 400 mL "reaction beaker", label the 250 mL beaker "waste beaker" and the 150 mL beaker "water".
4. Fill both the 400 mL beaker and the 150 mL more than half full with deionized water.
5. Obtain a buret clamp, gas collection tube and a rubber stopper with a copper loop.
6. Measure and cut 1 cm length of magnesium ribbon. Record its mass.
7. Gently fold the magnesium in half around the copper loop. Set the stopper with the magnesium aside.
8. Pour about 20 mL of HCl into a 25 mL graduated cylinder.
9. Slowly pour the HCl into the gas collection tube.
10. Use the water in the 150 mL beaker to pipet ~25 mL of water onto the acid layer. Use a circular motion to layer the water. At this point, pour the remaining water in the 150 mL beaker until the water layer is roughly halfway between the last marking on the tube and the top of the tube.
11. Firmly stopper the rubber stopper at the top of the gas collection tube.
12. Place finger on top of stopper for support and quickly invert the gas collection tube into the 400 mL beaker of water. Be sure the end of the gas collection tube is below the level of water in the beaker. Clamp the tube.
13. Once the air bubble has reached the top of the gas collection tube IMMEDIATELY record the volume of water. This is the initial volume.
14. The reaction is finished once the magnesium has been completely consumed.
15. Record the final volume from the gas collection tube and the temperature of the water in the 400 mL beaker.
16. Use temperature and the vapor pressure table (Appendix V) to record the vapor pressure of water.
17. Record the barometric pressure. This is the total pressure.

18. Remove the stopper and pour the solution in the gas collection tube into the waste beaker.
19. Clean up and complete any remaining calculations.

Clean Up and Waste Disposal: Return the syringe (cap and plunger) to the lab mat. The balloons may be placed in the trash can. Pour the contents of the Ideal Gas Law reaction beaker and the waste beaker into the acid Waste container. Any extra hydrochloric acid should also be placed in the acid waste container. Any extra pieces of magnesium should go in the large white weigh boat. Rinse the gas collection tube with deionized water only (not soap) and place the waste from the first rinse in the acid waste container. Rinse the gas collection tube twice more with deionized water and pour this waste down the drain both times. The rubber stopper should be replaced in the end of the gas collection tube and the gas collection tube should be put back in the clamp with the stoppered end toward the ceiling. Clean all remaining glassware with alconox/soap solution, then rinse with tap water and again with deionized water. After drying the glassware, return items to their original location.

Before Leaving Lab

- Make sure all items have been cleaned and put away.
- Check the board for any additional clean-up or waste instructions.
- Wipe down the benchtops.
- Refill the water bottles.
- Push in stools.
- Check with your instructor that your area has been sufficiently cleaned and reorganized.

Lab 11: Gas Laws Lab – Pre-Lab

Name _____

Section _____

Date _____

Pre-lab Questions

Read through all the background information and the procedure of the lab before answering these questions. For any calculations, all work must be shown. The work shown must include the formula, the unrounded answer, and the correctly rounded final answer. Answers without work may not receive full credit.

1. Write the formula for each gas law.

A) Boyle's Law _____

B) Charles's Law _____

C) Ideal Gas Law _____

2. Write the value of the Ideal Gas Law Constant. _____

3. Rearrange the Ideal Gas Law Equation to solve for R.

4. Convert the following units of pressure.

A) 0.805 atm to torr

B) 98,433 Pa to atm

C) 25.18 mm Hg to atm

5. Convert 22.8 °C to Kelvin.

6. Calculate the volume, cm^3 , of a sphere that has a circumference of 15.5 cm.

7. List two safety considerations.

8. What is the temperature of liquid nitrogen in degrees Celsius? _____

9. What sized beakers will be used for the Ideal Gas Law portion of the lab?
How should they be labeled?

10. Write a balanced equation for the single displacement reaction between
solid calcium and hydrobromic acid. Physical states may be ignored.

Lab 11: Gas Laws Lab – Data/Calculations

Name _____

Section _____

Drawer _____

Balance _____

Date _____

Partner(s) _____

Be sure to show all steps in calculations when necessary. Include formulas, units, unrounded answers and final rounded answers (with appropriate units) in the calculation table.

Table 1: Boyle's Law Data

	Measurement	Data
A.	Boyle's Law Formula	
B.	Observations #1	
C.	Observations #2	

Table 2: Charles's Law Data

	Measurement	Data
A.	Charles's Law Formula	
B.	Circumference of balloon, cm	
C.	Initial temperature, °C	
D.	Final temperature, °C	
E.	Observation of balloon in liquid nitrogen	

F.	Observation of balloon after removing from liquid nitrogen	
----	--	--

Table 3: Charles's Law Calculations

	Measurement	Calculations
G.	Radius of balloon, cm	
H.	Initial volume of balloon, cm ³	
I.	Initial temperature, K	
J.	Final temperature, K	
K.	Final volume of balloon, cm ³	

Table 4: Ideal Gas Law Data

Balanced Equation	

	Measurement	Data
A.	Ideal Gas Law Formula	
B.	Mass of Mg, g	
C.	Initial volume, mL	
D.	Final volume, mL	
E.	Temperature, °C	
F.	Vapor pressure of water, mmHg	
G.	Barometric pressure, mmHg	

Table 5: Ideal Gas Law Calculations

	Measurement	Calculations
H.	Moles of magnesium	

I.	Moles of hydrogen	
J.	Total Volume, in L	
K.	Temperature in Kelvin,	
L.	Partial Pressure of Hydrogen, mm Hg	
M.	Partial Pressure of Hydrogen, atm	
N.	Ideal Gas Law Constant, R	

Lab 11: Gas Laws Lab – Post-Lab Questions

Name _____

Section _____

Date _____

Post-lab Questions

For calculations, all work must be shown. The work shown must include the formula, the unrounded answer the correctly rounded final answer. Answers without work may not receive full credit.

1. Answer the following questions for Part A of the lab.

A) Which variables were changed in this experiment?

B) Which variables stayed the same?

C) Using an appropriate gas law, briefly explain your observation of the balloon when the plunger pushed/pulled.

2. Answer the following questions for Part B of the lab.

A) Which variables were changed in this experiment?

B) Which variables stayed the same?

C) Using an appropriate gas law, briefly explain your observation of the balloon before and after it was placed in the liquid nitrogen.

3. List two units of pressure used in this lab. _____
4. Describe an experimental set-up that could be used to test Avogadro's gas law. Answers from the internet or AI generated answers will be awarded zero points.
5. Explain what caused the volume to change inside the gas collection tube in Part C.
6. The true Ideal Gas Law Constant is $0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$. Use the value of R calculated in Part C of this lab as the experimental value to determine the percent error. Give the absolute value of the final answer to 3 significant figures.
- $$\text{Percent error} = \frac{|\text{true value} - \text{experimental value}|}{\text{true value}} \times 100$$
7. A gas occupied a pressure of 3.67 atm when it was at 0.48 L. For the same gas, what would be the final volume, in L, if the pressure was 2.16 atm.

8. A gas occupied a volume of 12.0 L at 343.2 K. At what temperature, in °C, will it occupy 18.2 L?
9. Calculate the pressure, in atm, for 9.00 g CH₄ when the gas is in a 3.5 L container at 28.2°C.
10. Calculate the pressure of CO₂ when 7.63 g Al₂(CO₃)₃ reacts with excess HCl in a 1.67 L container at 26°C.



Lab 12: Solutions and Electrolytes Lab

Lab Objectives

- Describe and classify a solution, solute, and solvent
- Describe and classify strong electrolytes, weak electrolytes and nonelectrolytes
- Define and calculate concentration using molarity
- Calculate the mass needed to make a given molarity of a solution
- Use the dilution equation to calculate molarity or volume of a new solution created from a standard
- Perform serial dilutions to determine the final molarity of a solution

Background Information

Many chemical reactions take place in water and are called aqueous solutions. However, the term “solution” is not limited to only liquids. There can also be solid solutions (called alloys) or gaseous solutions (the air we breathe). Regardless of the type of solution (solid, liquid or gas) all solutions are homogeneous mixtures that contain a solute and a solvent. A solute is any substance that dissolves in a solvent. The solute is usually present in a much smaller amount compared to the solvent. The solvent is the major component of the solution and is used to dissolve the solute. For example, when sodium chloride is placed in water it will dissolve. In this scenario, the solute is sodium chloride, and the solvent is water. After mixing the two together, the resulting mixture is a homogeneous aqueous solution. The solute and solvent can be in the same physical state or in different physical states, but the physical state of the solvent is what determines the type of solution. This lab will focus primarily on liquid solutions.

Table 1: Examples of Solutions

Example	Solution Type	Solute	Solvent
Brass	Solid	Zinc (33%)	Copper (67%)
Air	Gaseous	O ₂ (20.95%), Ar (0.93%), other gases (0.04%)	N ₂ (78.08%)
Saline	Liquid	NaCl(s)	H ₂ O(l)
Vinegar	Liquid	Acetic acid (5%)	H ₂ O (95%)
Soda	Liquid	CO ₂ (g)	H ₂ O(l)

Solubility and Electrolytes

When a solute dissolves in water and breaks apart into ions they can conduct an electrical current. These types of solution are called electrolytic solutions, and those ionic compounds are called electrolytes. An electrolyte is a substance that dissolves in water and has the ability to conduct an electrical current. Some electrolytes produce many ions in solution and are called strong electrolytes. Strong acids, strong bases and soluble ionic compounds are all examples of strong electrolytes. Strong acids and bases are listed in Table 2. Any other substance that is considered an acid or base but not in Table 2 is considered a weak acid or weak base. For example, ammonia, NH₃, is considered a base, but because it is not found in Table 2, it is considered a weak base. The solubility table (Appendix IV)

can be used to determine if an ionic compound is soluble in water. Ionic compounds that are soluble according to the solubility rules are also strong electrolytes.

Table 2: Strong Acid and Base Table

Strong Acids		Strong Bases	
HCl	Hydrochloric acid	LiOH	Lithium hydroxide
HBr	Hydrobromic acid	NaOH	Sodium hydroxide
HI	Hydroiodic acid	KOH	Potassium hydroxide
HNO ₃	Nitric acid	RbOH	Rubidium hydroxide
H ₂ SO ₄	Sulfuric acid	CsOH	Cesium hydroxide
HClO ₃	Chloric acid	Ca(OH) ₂	Calcium hydroxide
HClO ₄	Perchloric acid	Sr(OH) ₂	Strontium hydroxide
		Ba(OH) ₂	Barium hydroxide

When only a few ions are present in solution, then that solution is called a weak electrolyte. Weak acids and weak bases are examples of weak electrolytes. Compounds that dissolve in water but do not form ions are called nonelectrolytes. Nonelectrolytes are usually covalent compounds like sugar (C₆H₁₂O₆) or carbon dioxide (CO₂). An electrical conductivity meter can be used to indicate the strength of an electrical current in solution due to the presence of ions. The more ions that are in solution will lead to higher conductivity and fewer ions in solution will lead to lower conductivity. If no ions are in solution there will be no electrical current that can occur.

With the conductivity meter used for this lab, the scale below will help to identify the types of electrolytes that are in solution. When using the meter, look very closely at the lights as the red or green light can be very dim and may be difficult to see.

Table 3: Conductivity and Electrolyte Table

Scale	Red LED	Green LED	Conductivity	Electrolyte
0	Off	Off	Low or None	Nonelectrolyte
1	Dim	Off	Low	Nonelectrolyte
2	Medium or Bright	Off or Dim	Medium	Weak
3	Bright	Dim	High	Strong
4	Very bright	Medium	Very High	Strong

Example

Identify each substance as a strong electrolyte, weak electrolyte or nonelectrolyte: HC₂H₃O₂, NH₄Cl, C₂H₆O, HNO₃, NH₃.

Strong Electrolyte

NH₄Cl, HNO₃

Weak Electrolyte

HC₂H₃O₂, NH₃

Nonelectrolyte

C₂H₆O

NH₄Cl is soluble salt according to solubility rules and HNO₃ is a strong acid and can be found in Table 2. Therefore, both are also strong electrolytes. HC₂H₃O₂ is a weak acid and NH₃ is a weak base (neither are found in Table 2) and thus are weak electrolytes. The compound C₂H₆O is a covalent compound and does not produce ions when dissolved in water and as such is classified as a nonelectrolyte.

Molarity Formula

The amount of solute that is dissolved in solution is used to determine the concentration of a solution. Concentration is the amount of solute relative to the amount of solvent in a solution. Molarity (M) is one way to express concentration and is found from the number of moles of solute per liter of solution. The units of molarity are thus given as $\frac{\text{mol}}{\text{L}}$, or M. For example, a 1.00 M NaCl solution contains 1.00 mole of NaCl in one liter of solution and is written as $1.00 \frac{\text{mol}}{\text{L}}$ NaCl. When doing problems and trying to cancel units $\frac{\text{mol}}{\text{L}}$ is a more convenient unit to use as it helps to see how the units correctly (or incorrectly) cancel.

Molarity

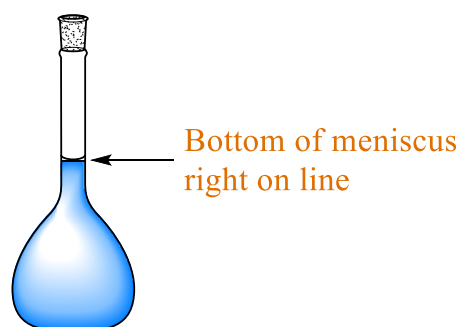
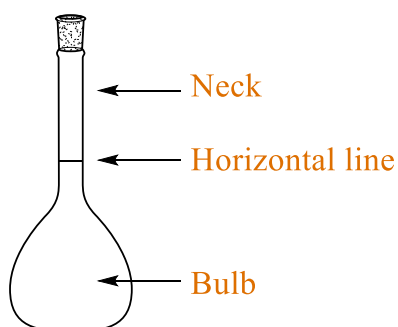
$$M = \frac{\text{moles of solute}}{\text{Liters of solution}}$$

When preparing a solution of a specific molarity a special flask, known as a volumetric flask, must be used. Volumetric flasks are used when very precise volumes are needed to make solution. This type of flask has a wide base and a very narrow neck. The neck has a horizontal mark that indicates the maximum amount of solution that should be added to the flask. Enough liquid should be added to the flask until the very bottom of the meniscus sits on the horizontal line. When making a solution in a volumetric flask, the total volume of the solution is the same as the size of the volumetric flask.

To make a solution in a volumetric flask use the following steps.

1. Add the solid or liquid solute.
2. Fill the bulb halfway full with solvent.
3. Swirl or mix the solution.
4. Add water to the flask until it just reaches the neck.
5. Use a disposable pipet to add the remaining solvent until the bottom of the meniscus is sitting on the horizontal line on the neck.
6. Cap the flask and invert it seven times.

Volumetric Flask



Example

How many moles of solute are in 3.25 L of a 0.258 M solution?

$$\text{moles of solute} = \text{Liters of solution} \times M = 3.25 \text{ L} \times 0.258 \frac{\text{mol}}{\text{L}} = 0.8385 \text{ mol} = \mathbf{0.839 \text{ mol}}$$

The formula must be rearranged to solve for moles.

Example

Calculate the mass (in g) of KBr needed to make 0.5240 L of a 0.1844 M solution.

$$\text{moles of solute} = \text{Liters of solution} \times M = 0.5240 \text{ L} \times 0.1844 \frac{\text{mol}}{\text{L}} \text{ KBr} = 0.096626 \text{ mol KBr}$$

$$\text{mass} = 0.096626 \text{ mol KBr} \times \frac{119.0023 \text{ g KBr}}{1 \text{ mol KBr}} = 11.49872 \text{ g KBr} = \mathbf{11.50 \text{ g KBr}}$$

Before the mass can be found the number of moles must be determined using the volume and molarity of the solution. Then using moles and molar mass the mass of KBr can be calculated.

Again, the wording is significant when calculating molarity. In the numerator, the amount of the solute must be in moles. Therefore, if a mass of solute is given, it must first be converted to moles. In the denominator, the volume is for the entire solution, not the solvent and must be in liters. If a different volume unit is used, it must be converted to liters before the molarity can be calculated.

Example

Calculate the molarity of solution when 5.042 g NaOH are dissolved in 800 mL of solution.

$$M = \frac{\text{moles of solute}}{\text{Liters of solution}} = \frac{0.126059 \text{ mol NaOH}}{0.800 \text{ L}} = 0.157574 \frac{\text{mol}}{\text{L}} \text{ NaOH} = \mathbf{0.158 \frac{\text{mol}}{\text{L}} \text{ NaOH}}$$

$$\text{mol} = 5.042 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{39.9971 \text{ g NaOH}} = 0.126059 \text{ mol NaOH}$$

$$V = 800 \text{ mL} \times \frac{1 \times 10^{-3} \text{ L}}{1 \text{ mL}} = 0.800 \text{ L}$$

In this example, the number of moles of NaOH is not directly given, but it can be found using the molar mass of NaOH. Similarly, the volume of solution was given in mL and needed to be converted to L before it could be used in the molarity formula.

Dilution Formula

Many chemistry labs store very concentrated solutions (large amount of solute in solution) and then dilute them to lower concentrations when needed. For example, concentrated nitric acid has a molarity of 15.7 M and that solution can be used to make a variety of diluted solutions (small amounts of solute in solution). The molarity of the concentrated solution (the stock solution) and the dilute solution are related through the dilution equation.

Dilution Formula

$$M_1 V_1 = M_2 V_2$$

In the formula, the M_1 refers to the molarity of the concentrated solution, V_1 refers to the volume of the concentrated solution, M_2 refers to the molarity of the diluted solution, and V_2 refers to the volume of the diluted solution. When using the dilution formula, as long as the volumes are both in the same unit, they can be in any unit (mL, L, μL , etc.). When making a diluted solution, a volumetric pipet (the V_1) is used to remove a very specific volume of the concentrated solution and add it to the volumetric flask (the V_2).

Example

When 10.00 mL of a 4.26 M solution is diluted to 500.0 mL, what is the new molarity?

$$M_2 = \frac{M_1 V_1}{V_2} = \frac{4.26 \text{ M} \times 10.00 \text{ mL}}{500.0 \text{ mL}} = 0.0852 \text{ M} = \mathbf{0.0852 \text{ M}}$$

Since a new molarity is to be solved for M_2 must be calculated. The new volume of the new solution is 500.0 mL, so that is V_2 and 10.00 mL of the concentrate solution was used so that is V_1 .

Example

What volume, in mL, of a HNO_3 solution is needed to prepare 250. mL of a 0.1885 M solution starting from a 15.7 M solution?

$$V_1 = \frac{M_2 V_2}{M_1} = \frac{0.1885 \text{ M} \times 250. \text{ mL}}{15.7 \text{ M}} = 3.00159 \text{ mL} = \mathbf{3.00 \text{ mL}}$$

When several sequential dilutions are performed from one stock solution, serial dilutions can be performed. Serial dilutions are stepwise solutions that are made with increasing lower concentrations.

Example

Initially 5.00 mL of a 2.00 M solution are diluted to a new volume of 50.0 mL of solution. When 5.00 mL of the new solution are then again diluted to make a 50.0 mL solution, what will be the final concentration of the solution?

First dilution: $M_2 = \frac{M_1 V_1}{V_2} = \frac{2.00 \text{ M} \times 5.00 \text{ mL}}{50.0 \text{ mL}} = 0.2000 \text{ M}$

Second dilution: $M_2 = \frac{M_1 V_1}{V_2} = \frac{0.2000 \text{ M} \times 5.00 \text{ mL}}{50.0 \text{ mL}} = 0.02 \text{ M} = \mathbf{0.0200 \text{ M}}$

Materials and Reagents Needed for Lab

Lab Equipment

Sharpie	450 mL beaker	Ceramic spot plate
Flinn Conductivity probe	9-volt battery	Kimwipes
50.0 mL Volumetric flask (3)	Weigh paper	Balance
Spatula	Disposable pipets	10.00 mL volumetric pipet
150 mL beaker	Pipetter	

Reagents

Deionized water	Tap water	0.1 M sodium chloride
0.1 M acetic acid	0.1 M potassium iodide	0.1 M sucrose
0.1 M citric acid	0.1 M calcium acetate	Ethanol
Solid sodium chloride	Food coloring	

Safety Considerations

Do not allow the circuit board or battery to come in contact with water. Only the copper electrodes should be immersed in or rinsed with water. Wipe the copper electrodes with a Kimwipe to remove any residue before placing the electrodes in a solution. Leave the batteries in the probes at the end of lab. Set the conductivity meter on a dry paper towel when not in use during lab. Wear gloves when handling any of the reagents. Although in low concentration, acetic acid, citric acid, and potassium iodide may cause skin irritation; rinse any affected area with water. Some of the solutions may become discolored when the electrodes are placed in them. Sharpie should not be used on the ceramic spot plates.

Procedure

Part A: Conductivity of Standard Solutions

1. Obtain a 400 mL beaker and use a Sharpie to label it Waste Beaker. This waste beaker will be used for all parts of the lab.
2. Obtain a ceramic spot plate. Make sure it is clean and dry.

3. Obtain a conductivity meter and wipe the copper electrodes with a Kimwipe. The same Kimwipe may be used multiple times. The copper electrodes should be clean and dry. Turn the meter on (no lights will appear).
4. Carefully touch the end of the copper electrodes to the bottom of one of the empty spot plate wells. No lights will appear. This is an example of how the conductivity meter behaves when something does not conduct a current because no electrolytes are present. Keep the conductivity meter on but set aside on a dry paper towel.
5. Use the deionized water bottle to fill one of the spot plate wells.
6. Carefully place **ONLY** the end of the copper electrodes into the well with deionized water and observe what happens to both the red and green lights in terms of the intensity of each light. Use Conductivity and Electrolyte Table (in the Background Information) to determine the significance of the light intensity and electrolyte type. Record observation in **Table 1**.
7. Use a Kimwipe to wipe off the copper electrodes and set the meter aside.
8. Pour about 20 mL of tap water into the 50 mL beaker. Use a disposable pipet to add about 10 – 15 drops of tap water to one of the empty wells. The tap water can be discarded as it will no longer be needed for this lab.
9. Carefully place the copper electrodes in the tap water and observe the lights. Record in **Table 1**.
10. Use a Kimwipe to wipe off the copper electrodes and set the meter aside.
11. Add 10 – 15 drops of the 0.1 M sodium chloride solution into one of the empty wells. Again, carefully place the copper electrodes in the sodium chloride solution and observe the lights. Record in **Table 1**.
12. Use the deionized wash bottle to rinse the electrodes off over the waste beaker. Then use a Kimwipe to wipe off the copper electrodes and set the meter aside.

Part B: Conductivity of Sample Solutions

1. Obtain the following solutions: 0.1 M acetic acid, 0.1 M potassium iodide, 0.1 M sucrose, 0.1 M citric acid, 0.1 M calcium acetate and ethanol. Also obtain solid sodium chloride.
2. Place 10 – 15 drops of each solution in a different well of the spot plate. Keep track of where each solution is placed.

3. Use the spatula to measure out a small amount (just a few grains) of solid sodium chloride into an empty well. Be sure this well is dry beforehand and **DO NOT** add any water to it. This mass does not need to be recorded.
4. Measure the conductivity of each solution and the solid sodium chloride and record the results in **Table 2**. Be sure to rinse the electrodes into the waste beaker and dry with a Kimwipe between each trial.
5. After recording your observations of the conductivity of solid sodium chloride, add a few drops of deionized water until the salt dissolves. Stir the mixture, then replace the conductivity meter and record your observations.
6. When finished, rinse the probe one final time into the waste beaker and turn off. Leave the battery in the conductivity meter.
7. The solutions in the spot plate and waste beaker can be poured down the drain with running water.

Part C: Molarity and Serial Dilutions

1. Obtain three 50.0 mL volumetric flasks and use a sharpie to label one of them Stock Solution, then label the second one Dilution 1 and the last one Dilution 2.
2. Label a 150 mL beaker "water" and fill it more than halfway full with deionized water.
3. Tare a piece of folded weigh paper on the balance. Measure out between 0.2 – 0.3 g of sodium chloride onto the paper. Record the exact mass in **Table 3**.
4. Pour the solid into the 50.0 mL volumetric flask labeled Stock Solution. Add two drops of food coloring.
5. Carefully pour deionized water into the Stock Solution flask until the bulb is half filled. Then gently swirl the solution for about 15 seconds or until all the salt is dissolved. Then continue pouring water into the flask until the water just enters the neck of the flask.
6. Now use a disposable pipet to add water dropwise to the flask until the bottom of the meniscus sits on the line indicated on the neck of the flask.
7. Place a cap on the flask and invert the flask seven times.
8. This is the stock solution. Record visual observations of this solution.

9. Use a 10.00 mL volumetric pipet to remove 10.00 mL of the Stock solution directly from the flask and add it to the volumetric flask labeled Dilution 1. Again, add water to the new flask following a similar procedure in steps 5 - 7.
10. Record observations of this new solution.
11. Use 10.00 mL volumetric pipet to remove 10.00 mL of the Dilution 1 solution and add it to the flask labeled Dilution 2. Again, add water to the new flask following a similar procedure in steps 5 - 7.
12. Record observations of the final solution.
13. Clean up and then complete any remaining calculations.

Clean Up and Waste Disposal: Make sure the conductivity meter is turned off. To clean the 50 mL flasks addalconox solution and use a scrub brush to clean the neck and the inside of the bulb. Rinse the flask with tap water and again with deionized water. Dry off the outside with a paper towel. DO NOT attempt to dry the inside of the flask with a paper towel; allow it to air dry on the lab mat. To clean the volumetric pipet, put some soap solution into a beaker and pull the solution into the pipet, roll the pipet to coat all sides with the soap solution. Use tap water to rinse the inside and outside of the volumetric pipet and then rinse again with deionized water. Dry the outside only with a paper towel and allow to dry on lab mat. All solutions and liquid waste can go down the drain with running water. Remove sharpie from glassware using a scrub brush and Alconox. Clean the ceramic spot plate, the caps to the flasks and any remaining glassware withalconox/soap solution, then rinse with tap water and again with deionized water. After drying the glassware, return items to their original location.

Before Leaving Lab

- Make sure all items have been cleaned and put away.
- Check the board for any additional clean-up or waste instructions.
- Wipe down the benchtops.
- Refill the water bottles.
- Push in stools.

- Check with your instructor that your area has been sufficiently cleaned and reorganized.

Lab 12: Solutions and Electrolytes Lab – Pre-Lab

Name _____

Section _____

Date _____

Pre-lab Questions

Read through all the background information and the procedure of the lab before answering these questions. For any calculations, all work must be shown. The work shown must include the formula, the unrounded answer, and the correctly rounded final answer. Answers without work may not receive full credit.

1. What are the units of molarity? _____
2. What is the dilution equation? _____
3. Classify each as a strong, weak or nonelectrolyte. All are soluble in water.
A) $\text{H}_2\text{C}_2\text{O}_4$ _____ C) KOH _____
B) FeCl_3 _____ D) $\text{C}_3\text{H}_8\text{O}_3$ _____
4. A solution contains 25.75 g LiNO_3 dissolved in 46.26 g $\text{C}_3\text{H}_6\text{O}$.
A) Identify the solute _____
B) Identify the solvent _____
C) Briefly explain if this would be considered an aqueous solution.
5. List two safety considerations.

6. What is the purpose of adding the food coloring to the flask in Part C?
7. How many times should the volumetric flask be inverted? _____
8. What is the volume of the stock solution in Part C? _____
9. Calculate the molarity when 0.2546 g NaCl are dissolved in enough water to make 50.0 mL of solution.
10. When 8.00 mL of a concentrated solution are diluted to 25.00 mL, the resulting molarity is 0.304 M. What was the concentration of the original solution?

Lab 12: Solutions and Electrolytes Lab – Data/Calculations

Name _____

Section _____

Drawer _____

Balance _____

Date _____

Partner(s) _____

Date _____

Be sure to show all steps in calculations when necessary. Include formulas, units, unrounded answers and final rounded answers (with appropriate units) in the calculation table.

Table 1: Conductivity of Standard Solutions Data

	Measurement	Observations	Electrolyte
A.	Deionized water		
B.	Tap water		
C.	0.1 M NaCl		

Table 2: Conductivity of Sample Solutions Data

	Measurement	Observations	Electrolyte
A.	0.1 M acetic acid		
B.	0.1 M potassium iodide		
C.	0.1 M sucrose		
D.	0.1 M citric acid		

E.	0.1 M calcium acetate		
F.	Ethanol		
G.	Solid sodium chloride		
H.	Sodium chloride with water		

Table 3: Molarity and Serial Dilutions Data

	Measurement	Data
A.	Mass of NaCl	
B.	Observations of stock solution	
C.	Observations of Dilution 1 Solution	
D.	Observations of Dilution 2 Solution	

Table 4: Molarity and Serial Dilutions Calculations

	Measurement	Calculations
E.	Moles of NaCl	

F.	Molarity of Stock Solution	
G.	Molarity of Dilution 1 Solution	
H.	Molarity of Dilution 2 Solution	

Lab 12: Solutions and Electrolytes Lab – Post-Lab Questions

Name _____

Section _____

Date _____

Post-lab Questions

For calculations, all work must be shown. The work shown must include the formula, the unrounded answer, the correctly rounded final answer. Answers without work may not receive full credit.

1. What is an aqueous solution?
2. In this lab, what was the purpose of performing Part A? Briefly explain.
3. Using the data from Part B as a guide, what would be the predicted intensity of the red and green lights on a conductivity meter for each of the following. Briefly explain your answer.

Caffeine ($\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$)	Carbonic acid	Lithium iodide
--	----------------------	-----------------------
4. Why was it necessary to add water to the solid sodium chloride before the conductivity could be observed in Part B?
5. When making the stock solution for Part C, identify the solute and the solvent.

Identify the solute _____

Identify the solvent _____

Briefly explain if this would be considered an aqueous solution.

6. Explain what happened to the color of the stock solution after each dilution.

7. Citric acid and acetic acid are both considered weak acids. However, one of them is still a stronger acid than the other. Use your experimental results to explain this observation.

8. Calculate the mass (in g) of NaCl needed to make 35 mL of a 0.647 M solution.

9. When 1.5 mL of a 6.5 M solution are diluted to 3.5 M. What volume, in mL, of solution would be required?

10. Initially 2.00 mL of a 12.5 M solution are diluted to a new volume of 50.0 mL of solution. When 2.00 mL of the new solution are then again diluted to make a 50.0 mL solution, what will be the final concentration of the solution?

Lab 13: Acid Base Lab

Lab Objectives

- Define and identify Arrhenius and Bronsted-Lowry acids and bases
- Recognize the difference between strong acids/bases and weak acids/bases
- Define and identify equilibrium reactions and predict conjugate acids and bases
- Identify solutions as acidic or basic based on the pH scale
- Calculate the pH from the molarity or the molarity from the pH
- Determine the hydronium concentration or hydroxide concentration using the pH equation and the equilibrium expression of water
- Evaluate pH changes in buffer systems

Background Information

The definition of an acid or base can vary depending on the type of reaction. Some definitions are limited to how a species behaves in water (Arrhenius definition), or how a proton gets transferred (Brønsted-Lowry definition) or even whether or not a species will accept or donate a pair of electrons (Lewis definition). This lab will only focus on the Arrhenius and Brønsted-Lowry type of acids and bases.

Arrhenius Acids and Bases

Arrhenius acids and bases are classified based on their behavior in water. An Arrhenius acid is any substance that produces the hydrogen ion (H^+) when dissolved in water. The formulas of Arrhenius acids are easily recognized because they start with the element hydrogen and because they take place in water, their physical state is always aqueous (aq). These acids are often binary acids or oxyacids, like HCl or HNO_3 . Arrhenius bases will always produce the hydroxide ion (OH^-) when dissolved in water. These bases are usually a metal hydroxide like NaOH or $Ca(OH)_2$. They will also have an aqueous (aq) physical state.

Arrhenius acids and bases can also be further categorized as strong or weak acids/bases. When an acid is dissolved in water and it completely ionizes, it is considered a strong acid. Since 100% of the molecule dissociates in water there are many hydrogen ions present in solution.

Strong Acid

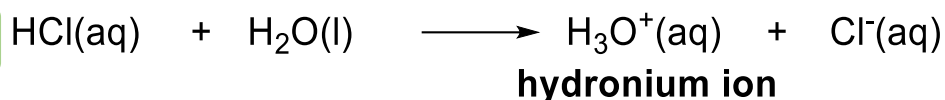


The hydrogen ion of a strong acid will never recombine with the anion to reform the acid once in solution. This is what makes an acid a strong acid with the Arrhenius definition. Because there are so many ions present in the solution, strong acids are also strong electrolytes. All the strong acids are given in Table 1. Any other acid that is not listed in Table 1 is considered a weak acid. Weak acids will be discussed later in this lab.

The hydrogen ion is very reactive and does not technically exist in water. When an Arrhenius acid is placed in water, it reacts with the water to form the hydronium ion, H_3O^+ . The hydronium ion is the true acidic component in water. As such, H^+ and H_3O^+ are often used interchangeably in their meaning.

Nonetheless, for the sake of convenience and clarity, often only the H^+ (hydrogen ion) is written. Both are used in throughout this lab.

Hydronium ion



Strong Arrhenius bases are usually metal hydroxides that completely dissociate in water and produce the hydroxide, OH^- , ion. Their physical states will also be aqueous due to the definition being limited to water. All the strong bases are also listed in Table 1.

Strong Base



The hydroxide ion of a strong base will never recombine with the cation to reform the base once in solution. This is what determines if base is a strong base using the Arrhenius definition. Strong bases are also strong electrolytes due to the large number of ions in solution. All the strong bases are given in Table 1. Any other aqueous base that is not listed in Table 1 is considered a weak base.

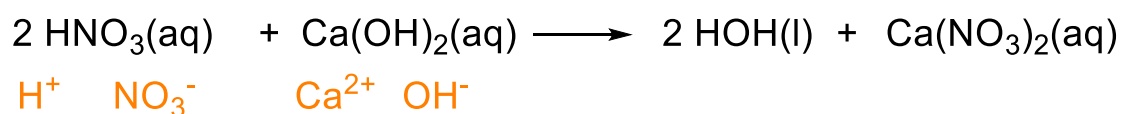
Table 1: Strong Acids and Bases

Strong Acid	Strong Acid Name	Strong Base	Name
HCl	Hydrochloric acid	LiOH	Lithium hydroxide
HBr	Hydrobromic acid	NaOH	Sodium hydroxide
HI	Hydroiodic acid	KOH	Potassium hydroxide
HNO_3	Nitric acid	RbOH	Rubidium hydroxide
H_2SO_4	Sulfuric acid	CsOH	Cesium hydroxide
HClO_3	Chloric acid	$\text{Ca}(\text{OH})_2$	Calcium hydroxide
HClO_4	Perchloric acid	$\text{Sr}(\text{OH})_2$	Strontium hydroxide
		$\text{Ba}(\text{OH})_2$	Barium hydroxide

When any Arrhenius acid reacts with any Arrhenius base in an acid-base reaction, water and a salt are always formed. Because water is produced as a product in these acid-base reactions, they are called neutralization reactions and are another example of a double displacement reaction. The hydrogen ion (H^+) of the acid combines with the hydroxide (OH^-) of the base to form liquid water, $\text{H}_2\text{O(l)}$. When balancing acid-base reactions, it is sometimes helpful to write the formula of water as HOH. Written in this way, it can be easier to see the “OH” on both sides of the equation when balancing. The formula of the salt can be determined from the cation of the base and the anion of the acid. The crisscross method can then be used to combine the ions to form the new ionic compound. The physical state of the ionic compound is determined by using the solubility table (see Appendix IV).

Example

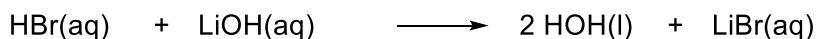
Write a neutralization reaction between aqueous solutions of HNO_3 and $\text{Ca}(\text{OH})_2$. Include physical states.



The solubility rules show that any ionic compound that contains nitrate will always be soluble.

Example

Calculate the molarity of HBr if 25.00 mL of HBr solution are needed to react with 0.03446 L of a 0.1295 M LiOH solution.



$$M_{\text{HBr}} = \frac{\text{mol HBr}}{V_{\text{HBr solution}}} = \frac{0.0044626 \text{ mol HBr}}{0.02500 \text{ L}} = 0.17850 \frac{\text{mol}}{\text{L}} \text{HBr} = \mathbf{0.1785 \frac{\text{mol}}{\text{L}} \text{HBr}}$$

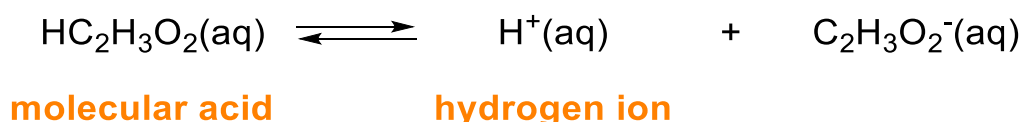
$$V_{\text{HBr solution}} = 25.00 \text{ mL} \times \frac{1 \times 10^{-3} \text{ L}}{1 \text{ mL}} = 0.02500 \text{ L}$$

$$\text{mol HBr} = 0.03446 \text{ L LiOH} \times \frac{0.1295 \text{ mol LiOH}}{1 \text{ L}} \times \frac{1 \text{ mol HBr}}{1 \text{ mol LiOH}} = 0.0044626 \text{ mol HBr}$$

To find the molarity of HBr, both the moles of HBr and the volume, in L, of HBr solution are required. The volume of HBr solution was given in the problem, but not the number of moles. Enough information about LiOH is provided to find the number of moles of LiOH and since the balanced chemical equation is known, the number of moles of LiOH can be directly related to the number of moles of HBr.

Weak acids also produce hydrogen ions in solution but to a much lesser extent. Weak acids only partially (1 – 5%) dissociate to ions when in water. Because of this, there is still a significant amount of intact molecular acid ($\text{HC}_2\text{H}_3\text{O}_2$) in solution and very little hydrogen ion (H^+) compared to strong acids. Thus when weak acids ionize, only a small amount of hydrogen ion is present in solution. The low amount of H^+ in solution is what designates an acid as a weak acid. Weak acids therefore are weak electrolytes.

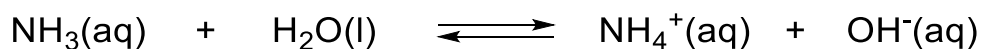
Weak Acid



The dissociation of weak acids are an example of an equilibrium reaction. With equilibrium reactions, a dual arrow ($\rightleftharpoons$) is used to indicate the rate of forward and reverse reaction are the same. In the dissociation of acetic acid, H^+ and $\text{C}_2\text{H}_3\text{O}_2^-$ are produced, and those two ions can then combine to reform the intact acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$.

Weak bases also only partially ionize produce only a small amount of hydroxide ion in solution. The low amount of OH^- in solution is what designates a base as a weak base. Weak bases are also weak electrolytes. Again, any other substance classified as a base but not in Table 1 is a weak base. For example, ammonia NH_3 , is classified as a base because it produces hydroxide in water. It only partially dissociates to form the ammonium cation (NH_4^+) and hydroxide (OH^-), thus it is a weak base. The ionization of ammonia is also an equilibrium process.

Weak Base



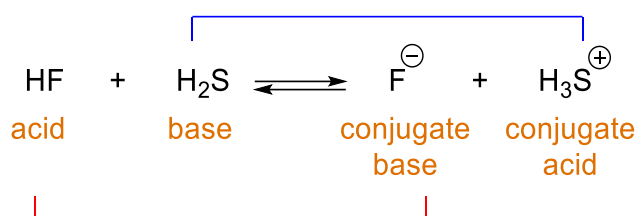
Keep in mind strong acids and bases are never written as equilibrium equations. This is because when a strong acid or strong base ionizes, the cation will never recombine with its anion to reform the intact

acid or base. Strong acids/bases only proceed in the forward direction, never in the reverse. This is a key difference between a strong acid/base and a weak acid/base.

Brønsted-Lowry Acids and Bases

While Arrhenius acids and bases are limited to water, Brønsted-Lowry (B.L.) acids and bases are not. This definition broadens the scope of what is considered an acid or a base. The B.L. acid definition is somewhat similar to the Arrhenius definition, but the base definition is very different. A Brønsted-Lowry acid is any compound that can donate a hydrogen ion, and a B.L. base is a compound that can accept a hydrogen ion. The transfer of H^+ can take place in water but can also occur in other solvents as well. In other words, Arrhenius acids are all Brønsted-Lowry acids but Brønsted-Lowry acids are not necessarily Arrhenius acids.

In the below reaction between HF and H_2S , both look like acids using the Arrhenius definition but only one is behaving as an acid using the Brønsted-Lowry definition. HF is considered a Brønsted-Lowry acid because it donates its hydrogen to H_2S , which therefore makes H_2S the Brønsted-Lowry base.



These reactions between Brønsted-Lowry acids and bases can also be reversible and exist as equilibrium reactions. When F^- and H_3S^+ react (on the right), they can reform HF and H_2S (on the left). In this case, H_3S^+ behaves as the new acid (the conjugate acid) because it donates the H^+ and F^- is the new base (the conjugate base) because it accepts the hydrogen ion. The relationship between the acid and the new base it forms is called an acid-conjugate base pair and the base and the new acid it forms are called base-conjugate acid pairs. If HF is the acid, then F^- is the conjugate base. The formula of the conjugate base differs from its acid by having one less hydrogen and one more negative charge. In a similar way, if H_2S is the base then H_3S^+ is its conjugate acid. The formula of the conjugate acid differs from its base by having one more hydrogen and one more positive charge.

Example

Write the conjugate base for CH_4 and for NH_4^+ .



acid



conjugate
base



acid



conjugate
base

Remember, the formula of the conjugate base differs from its acid by having one less hydrogen and one more negative charge.

Example

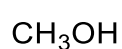
Write the conjugate acid for CO_3^{2-} and for CH_3OH .



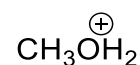
base



conjugate
acid



base

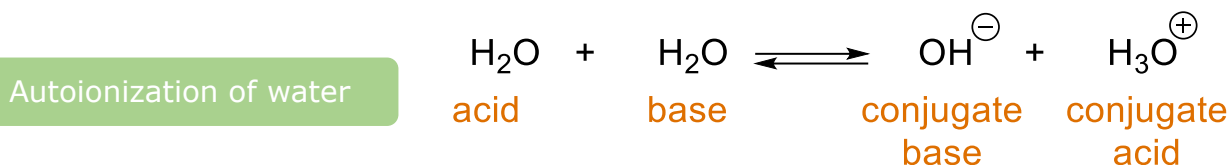


conjugate
acid

The formula of the conjugate acid differs from its base by having one more hydrogen and one more positive charge.

Concentration and pH

Water can in fact react with itself to behave as both an acid and a base and produce its respective conjugate base and conjugate acid. This is called the autoionization of water.



Notice there is always some amount of H^+ and OH^- present in solution. The amount of hydroxide and hydronium ion can be quantified by their concentration. It is common for the concentration (or molarity) to be written in square brackets. For example, $[\text{OH}^-]$ is read as “the concentration of hydroxide”, or $[\text{H}_3\text{O}^+] = 3.45 \times 10^{-2} \text{ M}$ is read as “the concentration of hydronium is $3.45 \times 10^{-2} \text{ M}$ ”. In neutral water, the concentration of hydronium ion and hydroxide are equal. In other words, $[\text{H}_3\text{O}^+] = [\text{OH}^-]$ and both are equal to $1.00 \times 10^{-7} \text{ M}$. When those values are multiplied the value of 1.00×10^{-14} is obtained, and this value is called the equilibrium constant for water and is expressed by the following equation.

Equilibrium Constant for Water

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14} \text{ M}^2$$

The $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ are in a delicate balance. If concentration of one goes up, the other will go down. When the concentrations of hydroxide and hydronium are different, the equilibrium equation can be rearranged to solve for one or the other.

Example Determine the molarity, in M, of H_3O^+ when the $[\text{OH}^-] = 1.00 \times 10^{-3} \text{ M}$.

$$[\text{H}_3\text{O}^+] = \frac{1.00 \times 10^{-14} \text{ M}^2}{[\text{OH}^-]} = \frac{1.00 \times 10^{-14} \text{ M}^2}{1.00 \times 10^{-3} \text{ M}} = 1 \times 10^{-11} \text{ M} = \mathbf{1.00 \times 10^{-11} \text{ M}}$$

Remember molarity and concentration are similar in meaning. Since the concentration of hydroxide is known, the equilibrium constant equation can be rearranged to solve for the concentration of hydronium ion.

The concentration of hydrogen ion in solution can be quickly represented by the pH of the solution. The pH of an aqueous solution is a unitless value and is found from the negative log of the acid concentration. The pH scale ranges from 0 to 14. Neutral solutions have a $\text{pH} = 7$ and any solution with a pH less than 7 is acidic, and any solution with pH greater than 7 is basic.

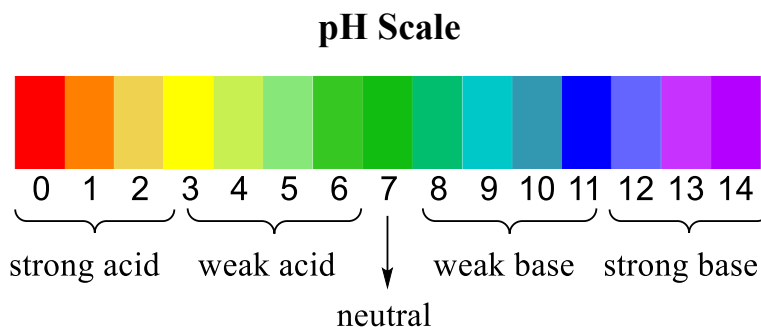
pH Formula

$$\text{pH} = -\log [\text{H}_3\text{O}^+] \text{ or } -\log [\text{H}^+]$$

The pH formula can only be used when the concentration of the hydronium ion, $[\text{H}_3\text{O}^+]$, or acid is known. The pH formula can never be used if only the concentration of hydroxide, $[\text{OH}^-]$, is given. Nonetheless the two, $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$, are related. When the $[\text{H}_3\text{O}^+]$ is high, the $[\text{OH}^-]$ is low and the pH is closer to 0. When the $[\text{OH}^-]$ is high, the $[\text{H}_3\text{O}^+]$ is low and the pH is closer to 14. A solution with a high concentration of $[\text{H}_3\text{O}^+]$ has a molarity higher than 1.00×10^{-7} and a solution with a low concentration of $[\text{H}_3\text{O}^+]$ has a molarity smaller than 1.00×10^{-7} .

Often an acid-base indicator can be added to a solution to quickly identify if a solution is acidic or basic. Litmus paper is red in acidic solutions and blue in basic solutions. The indicator

phenolphthalein is colorless in acidic solutions and pink in basic solutions. Universal indicator has a wider color range that varies with different pH values.



A colored version of the pH scale may be found in Appendix VI.

Example

If a solution has $[\text{H}_3\text{O}^+] = 2.1 \times 10^{-9} \text{ M}$, what is the pH of the solution. Would it be considered an acidic or basic solution?

$$\text{pH} = -\log [2.1 \times 10^{-9}] = 8.67751 = \mathbf{8.68 \text{ basic}}$$

Notice how the final numerical answer was reported in the previous example. The pH is reported to the same number of places after the decimal point as there are significant figures in the concentration of the acid. Since the concentration of $[\text{H}_3\text{O}^+]$ had two significant figures, the pH had to be reported to two places after the decimal point. The opposite would also be true. If the pH has a certain number of places after the decimal point, the concentration of the acid would be reported that number of significant figures. To calculate the concentration of an acid from its pH, use the following antilog formula.

Example

Find the pH of a nitric acid solution with $[\text{HNO}_3] = 9.72 \times 10^{-4} \text{ M}$.

$$\text{pH} = -\log [9.72 \times 10^{-4}] = 3.012334 = \mathbf{3.012}$$

The concentration of this acid is $9.72 \times 10^{-4} \text{ M}$. Since there is 1 mole of H^+ for every 1 mole HNO_3 , the concentration of the acid is directly related to the concentration of the hydrogen ion.

If the pH of a solution is known, the concentration of that solution can also be found. To find the concentration of an acid from its pH value, the antilog (10^{-x}) of the pH must be used. When it comes to acid-bases, the x is equal to pH.

Antilog Formula

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

The antilog is used to calculate the acid concentration by raising ten to the negative pH value. However, many places after the decimal point in the pH value there are, the acid concentration will be reported to the same number of significant figures.

Example

What is the concentration of a solution with a $\text{pH} = 5.423$?

$$[\text{H}_3\text{O}^+] = 10^{-5.423} = 0.000003776 \text{ M} = \mathbf{3.78 \times 10^{-6} \text{ M}}$$

The final answer is reported to 3 SF because there are 3 places after the decimal point of the pH value.

The concentration hydroxide in solution can also be calculated if the pH or $[\text{H}_3\text{O}^+]$ are known. There are similar pOH calculations that can be performed but are not discussed in this lab. Instead, the concentration of hydroxide, $[\text{OH}^-]$, can always be found using the equilibrium constant formula for water.

Example

Determine the $[\text{OH}^-]$ when the pH of a solution is 11.02. Write answer in scientific notation.

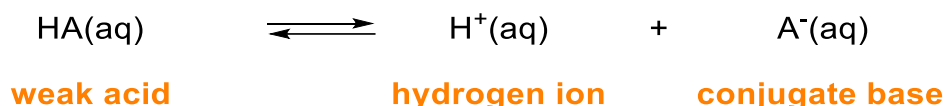
$$[\text{OH}^-] = \frac{1.00 \times 10^{-14} \text{ M}^2}{[\text{H}_3\text{O}^+]} = \frac{1.00 \times 10^{-14} \text{ M}^2}{9.54993 \times 10^{-12} \text{ M}} = 0.0010471 \text{ M} = \mathbf{1.0 \times 10^{-3} \text{ M}}$$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-11.02} = 9.54993 \times 10^{-12} \text{ M}$$

The equilibrium equation can be rearranged to solve for the concentration of hydroxide. However, the concentration of hydronium must first be found from the pH of the solution.

Buffer Solutions

Buffers are solutions that can resist drastic changes in pH values. In a buffer solution, the dissociation of a weak acid (HA) in water yields the hydrogen ion (H^+) ion and the conjugate base (A^-) of the acid.



Buffer solutions usually consist of large amounts of a weak acid and its conjugate base, or a weak base and its conjugate acid. When a weak acid (HA) is dissolved in a solution containing its conjugate base (A^-), the result is a mixture that is highly resistant to changes in acid concentration upon addition of either strong acid or strong base. For example, when sodium hydroxide (a strong base) is added to a buffer system like the one above, the added hydroxide ions will react with the weak acid and force the acid to dissociate and shift the reaction to the right. On the other hand, if hydrochloric acid (a strong acid) is added to a buffer system the additional H^+ ions react with the conjugate base to reform the intact acid and causes the reaction to shift back to the left producing more of the weak acid (HA). This type of reaction is also called an equilibrium reaction because the rate of the forward reaction and reverse reaction take place at the same rate, so the concentrations of the reactants and products do not change very significantly. Therefore, there is very little change in the overall concentration of H^+ in solution which causes a very minor change in the pH of the solution.

When no buffer is present, addition of a strong base like NaOH will produce a large amount of hydroxide to be present (and very little H^+) and the solution becomes more basic (pH goes closer to 14). Similarly, when a strong acid, like HCl, is added to a solution that does not contain a buffer, a large amount of hydrogen ion will be present, and the solution becomes more acidic (pH goes closer to 0). Investigating the pH of acids, bases and buffers is the purpose of this lab.

Materials and Reagents Needed for Lab

Lab Equipment

Sharpie	400 mL beaker	pH meter
25 mL graduated cylinder	50 mL beaker	Kimwipes
10 mL graduated cylinder	Ceramic spot plate	Wooden splints

Reagents

Calibration solution 1	Calibration solution 2	Calibration solution 3
0.1 M HCl	0.1 M KCl	0.1 M KOH
0.1 M $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$	0.1 NaHCO_3	0.2 M NaOH
Deionized water	Universal Indicator	Acetic acid buffer solution

Safety Considerations

Pipet directly from the bottle/jar with the dedicated plastic disposable pipet. Do not throw away the disposable pipets attached to the bottles/jars. Sodium hydroxide can cause eye damage and severe skin burns as it is corrosive. If it comes in contact with the skin, rinse with water for 15 minutes. Although in low concentration, hydrochloric acid and acetic acid may still cause skin irritation; rinse any affected area with water. Wear gloves when handling any of the reagents. The end of the pH probe is very fragile. To prevent breakage, do not press the end of the probe against the ceramic spot plate. Angle the end of the probe so that it is fully immersed in the liquid but not touching the sides or bottom of the spot plate.

Procedure

Part A: Calibrating the pH Meter

1. Obtain a 400 mL beaker and use a Sharpie to label it Waste Beaker. All waste from this lab should go into this beaker. If at any time the waste beaker becomes too full, the solution can be poured down the drain with running water.
2. Obtain a ceramic spot plate and the Calibration Solutions 1, 2 and 3.

3. Fill one of the wells (approximately 18 – 20 drops) of the spot plate with Calibration Solution 1 (neutral). The well should be almost full without going over.
4. Repeat with the remaining two calibration solutions in the next two wells of the spot plate. Record the color of each solution.
5. Remove the cover of pH meter rinse the tip with deionized water into the waste beaker. Gently dry the end of the meter with a Kimwipe.
6. Place the pH meter in the well for Calibration Solution 1 and gently swirl the solution for 4 – 5 seconds. The pH meter must be held at an angle so that it does not touch the bottom or sides of the spot plate. While the tip is still in the solution, press and hold the CAL (calibration) button for 5 seconds and then release. The screen should flash CAL three times and display a pH value. As soon as the flashing stops, immediately record the number displayed on the pH meter in **Table 1**.
7. Clean the end of the pH meter by rinsing the end of it with deionized water over the waste beaker and gently dry it with a Kimwipe.
8. Repeat steps 6 and 7 for Calibration Solution 2 (acidic) and Calibration Solution 3 (basic).
9. Clean the end of the pH meter with deionized water and dry it. Set aside on a clean and dry paper towel.
10. Show the calibration values to your instructor.
11. Once the values have been approved by the instructor, pour the contents of the spot plate into the waste beaker. Then rinse the spot plate with deionized water over the waste beaker.
12. Thoroughly dry the spot plate with a paper towel and set aside. It will be used again in the next section.

Part B: Measuring pH and Using Universal Indicator

1. Take three wooden splints and break them each apart into 3 – 4 pieces. Set them all aside on a clean and dry paper towel. Use a different splint to mix each well once the universal indicator is added.
2. Locate the following solutions: KCl, HCl, NaHCO₃, KOH, H₃C₆H₅O₇ (citric acid), deionized water and tap water (from the sink).

3. In one of the empty spot plate wells, fill it (approximately 18 – 20 drops) with KCl solution. Gently stir the solution with the pH meter and watch the reading. When the pH reading has stabilized (the numbers have stopped changing), record the pH of the KCl solution in **Table 2**.
4. Rinse the end of the pH meter with deionized water over the waste beaker and gently dry it with a kimwipe.
5. Add 1 – 2 drops of universal indicator to the same well. Use the wooden splint to gently mix the solution and record the color of the resulting solution.
6. Repeat steps 3 – 5 for the remaining solutions using different wells of the spot plate for the different solutions.
7. When finished, pour the contents of the spot plate into the waste beaker. Then rinse the spot plate with deionized water over the waste beaker.

Part C: Buffer Solutions

1. Use a 25 mL graduated cylinder to measure out 20 mL of deionized water. Pour the water into a 50 mL beaker.
2. Insert the pH meter into the same 50 mL beaker of water and gently swirl it in the water. Once the pH readings have stabilized record the pH of water in **Table 3**.
3. Add a drop of universal indicator to the solution, mix, then record the color.
4. Gently dry the end of the pH meter with a kimwipe and set it aside on a clean, dry paper towel.
5. Add 5 drops of the 0.1 M HCl solution to the beaker of water. Gently stir the solution with the pH meter. When the pH reading has stabilized, record the pH. Record the color of the solution.
6. Pour this solution into the waste beaker. Completely rinse the 50 mL beaker four times with deionized water into the waste beaker. Rinse the end of the pH meter with deionized water over the waste beaker and gently dry it with a kimwipe. Set the pH meter aside on a clean paper towel.
7. Use a 25 mL graduated cylinder to measure out 20 mL of deionized water. Pour the water into the 50 mL beaker. Add 5 drops of the 0.2 M NaOH solution to the same 50 mL beaker. Gently stir the solution with the pH meter and when the pH reading has stabilized, record the pH.
8. Add a drop of universal indicator, mix the solution, then record the color.

9. Pour this solution into the waste beaker. Rinse the 50 mL beaker four times with deionized water into the waste beaker. Dry the inside of the 50 mL beaker with a paper towel. Rinse the end of the pH meter with deionized water over the waste beaker and gently dry it with a kimwipe. Set pH meter aside on a clean paper towel.
10. Use the 25 mL graduated cylinder to measure out 20 mL of the buffer solution then pour the buffer solution into the 50 mL beaker.
11. Insert the pH meter into the buffer solution and gently swirl the water. Once the pH readings have stabilized record the pH of the buffer also in **Table 3**.
12. Add a drop of universal indicator to the solution, mix, then record the color.
13. Add 5 drops of the 0.1 M HCl solution to the buffer solution. Gently stir the solution with the pH meter. When the pH reading has stabilized, record the pH and the color of the solution.
14. Pour this solution into the waste beaker, rinse the 50 mL beaker four times with deionized water into the waste beaker. Rinse the end of the pH meter with deionized water over the waste beaker and gently dry it with a kimwipe. Set meter on a clean paper towel.
15. Use a 25 mL graduated cylinder to measure out 20 mL of the buffer solution then pour the solution into the 50 mL beaker and add 5 drops of the 0.2 M NaOH solution to the buffer solution. Gently stir the solution and when the reading has stabilized, record the pH.
16. Add a drop of universal indicator to the solution, mix, then record the color.
17. Rinse the end of the pH meter with deionized water over the waste beaker and gently dry it with a kimwipe. Turn off and replace the cap of the pH meter. The solution in the waste container can now be poured down the drain with running water.

Part D: Conjugate Acids and Bases

First clean up and then answer the questions about conjugate acids and bases.

Clean Up and Waste Disposal: Make sure the pH meter has been rinsed with deionized water, turned off, the cap replaced and returned it to its original

location. The wooden splints may go in the regular trash. All liquid waste can go down the drain with running water. Remove any sharpie from glassware using alconox and a scrub brush. Clean all glassware with alconox/soap solution, then rinse with tap water and again with deionized water. After drying the glassware, return items to their original location.

Before Leaving Lab

- Make sure all items have been cleaned and put away.
- Check the board for any additional clean-up or waste instructions.
- Wipe down the benchtops.
- Refill the water bottles.
- Push in stools.
- Check with your instructor that your area has been sufficiently cleaned and reorganized.

Lab 13: Acid Base Lab – Pre-Lab

Name _____

Section _____

Date _____

Pre-lab Questions

Read through all the background information and the procedure of the lab before answering these questions. For any calculations, all work must be shown. The work shown must include the formula, the unrounded answer, and the correctly rounded final answer. Numerical answers without work may not receive full credit.

1. Circle the strong acids from the examples below.

CH₄ H₂SO₄ HBr H₂CO₃ NaOH HClO₄ HF

2. Write the conjugate base of PH₃. _____

3. Write the conjugate acid of PH₃. _____

4. A solution has a pH = 6.5. Circle the type of solution this would be.

Strong acid Weak acid Neutral Weak Base Strong Base

5. Explain the difference between an Arrhenius Base and a Bronsted-Lowry base.

6. If a solution has a $[\text{H}_3\text{O}^+] = 1.7 \times 10^{-4}$, to how many decimal places (DP) should the pH be reported? Circle the correct answer.

1 DP 2 DP 3 DP 4 DP 5 DP

7. Calculate the pH of a solution with $[\text{H}_3\text{O}^+] = 8.1 \times 10^{-9}$.

8. List two safety considerations for this lab.

9. What is the concentration of NaOH used in this lab? _____

10. Explain why it is necessary to perform the calibration of the pH meter in Part A.

Lab 13: Acid Base Lab – Data/Calculations

Name _____

Section _____

Drawer and Balance Number _____

Date _____

Partner(s) _____

Be sure to show all steps in calculations when necessary. Include formulas, units, unrounded answers and final rounded answers (with appropriate units) in the calculation table.

Part A: Calibrating the pH Meter

Table 1: Calibrating the pH meter

		Calibration solution 1	Calibration solution 2	Calibration solution 3
A.	pH Value			
B.	Color			

Part B: Measuring pH and Using Universal Indicator

Table 2: Measuring pH with Universal Indicator

	Solution	pH	Color
A.	KCl		
B.	HCl		
C.	NaHCO ₃		
D.	KOH		
E.	H ₃ C ₆ H ₅ O ₇		
F.	Deionized Water		
G.	Tap Water		

Part C: Buffer Solutions

Table 3: Buffer Solutions

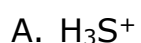
	Solution	pH	Color
A.	Water		
B.	Water with 0.1 M HCl		
C.	Water with 0.2 M NaOH		
D.	Buffer solution		
E.	Buffer solution with 0.1 M HCl		
F.	Buffer solution with 0.2 M NaOH		

Part D: Conjugate Acids and Bases

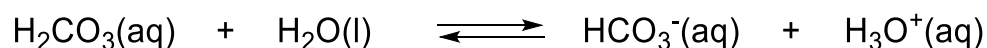
1. Write the conjugate acid for each.



2. Write the conjugate base for each.



3. Use the equilibrium reaction below to answer the following questions.



A. What is the formula of the weak acid? _____

B. What is the formula of its conjugate base? _____

C. Explain what would happen in the reaction if additional acid were added.

Lab 13: Acid Base Lab – Post-Lab Questions

Name _____

Section _____

Date _____

Post-lab Questions

For calculations, all work must be shown. The work shown must include the formula, the unrounded answer, and the correctly rounded final answer. Answers without work may not receive full credit.

1. What is the conjugate base of acetic acid? _____
2. What is the conjugate acid of nitrite? _____
3. Explain how strong acids differs from weak acids.
4. Is citric acid, $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$, a strong or weak acid? What experimental evidence supports this answer?
5. Write the neutralization reaction that occurs between perchloric acid and strong barium hydroxide. Physical states may be ignored.
6. When universal indicator was added to an aqueous solution, the color of the solution was dark purple. Would the solution be considered a strong acid, a weak acid, a strong base or a weak base. Briefly explain.

7. List the compounds in Part B that would be considered a neutral salt or a strong/weak acid/base.

A. Neutral salt _____

B. Strong acids _____

C. Weak acids _____

D. Strong bases _____

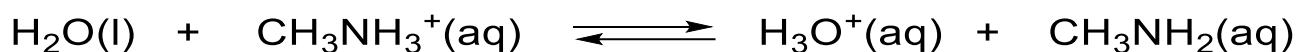
E. Weak bases _____

F. Briefly explain the experimental results that support your answer.

8. A NaOH solution has a $[\text{OH}^-] = 1.58 \times 10^{-4} \text{ M}$. Calculate the pH of this solution.

Would this solution be considered acidic or basic? _____

9. In the reaction below, identify the Bronsted-Lowry acid (A), the Bronsted-Lowry base (B), the conjugate acid (CA) and the conjugate base (BA) by writing the appropriate label under the correct formula.



10. In Part C, explain the effect observed on the pH when the

A. Strong acid and strong base were added to water.

B. Strong acid and strong base were added to the buffer solution.

Common Equivalence Statements

Length

1 kilometer (km)	=	0.6214 mile (mi)
1 meter (m)	=	39.37 inches (in)
1 meter (m)	=	1.094 yards (yd)
1 foot (ft)	=	30.48 centimeters (cm)
1 inch (in)	=	2.54 centimeters (cm)

Mass

1 kilogram (kg)	=	2.205 pounds (lb)
1 pound (lb)	=	453.59 grams (g)
1 ounce (oz)	=	28.35 grams (g)
1 ton	=	2,000 pounds (lb)

Volume

1 liter (L)	=	1.057 quarts (qt)
1 gallon (gal)	=	3.785 liters (L)
1 milliliter (mL)	=	1 cm ³

Pressure

1 atmosphere (atm)	=	760 mmHg or torr
1 atmosphere (atm)	=	101,325 Pa
1 atmosphere (atm)	=	14.70 psi
1 atmosphere (atm)	=	1.013 bar
1 atmosphere (atm)	=	29.92 in Hg

Temperature

K	=	°C + 273.15
°C	=	K - 273.15
°F	=	1.8(°C) + 32
°C	=	$\frac{(^{\circ}\text{F} - 32)}{1.8}$

Metric Prefix Table

Prefix	Symbol	Meaning	Decimal Equivalent
Tera-	T	1×10^{12}	1,000,000,000,000
Giga-	G	1×10^9	1,000,000,000
Mega-	M	1×10^6	1,000,000
Kilo-	k	1×10^3	1,000
Hecto-	h	1×10^2	100
Deca-	da	1×10^1	10
Deci-	d	1×10^{-1}	0.1
Centi-	c	1×10^{-2}	0.01
Milli-	m	1×10^{-3}	0.001
Micro-	μ	1×10^{-6}	0.000001
Nano-	n	1×10^{-9}	0.000000001
Pico-	p	1×10^{-12}	0.000000000001
Femto-	f	1×10^{-15}	0.000000000000001

Appendix III

Simple and Polyatomic Ions

1 ⁺ Cations	
Hydrogen ion	H ⁺
Lithium ion	Li ⁺
Sodium ion	Na ⁺
Potassium ion	K ⁺
Rubidium ion	Rb ⁺
Cesium ion	Cs ⁺
Silver ion	Ag ⁺
Ammonium	NH ₄ ⁺
Hydronium	H ₃ O ⁺

1 ⁻ Anions			
Hydride	H ⁻	Nitrate	NO ₃ ⁻
Fluoride	F ⁻	Nitrite	NO ₂ ⁻
Chloride	Cl ⁻	Acetate	C ₂ H ₃ O ₂ ⁻
Bromide	Br ⁻	Hydroxide	OH ⁻
Iodide	I ⁻	Cyanide	CN ⁻
Perchlorate	ClO ₄ ⁻	Bicarbonate	HCO ₃ ⁻
Chlorate	ClO ₃ ⁻	Permanganate	MnO ₄ ⁻
Chlorite	ClO ₂ ⁻	Bisulfate	HSO ₄ ⁻
Hypochlorite	ClO ⁻	Bisulfite	HSO ₃ ⁻

2 ⁺ Cations	
Beryllium ion	Be ²⁺
Magnesium ion	Mg ²⁺
Calcium ion	Ca ²⁺
Strontium ion	Sr ²⁺
Barium ion	Ba ²⁺
Cadmium ion	Cd ²⁺
Zinc ion	Zn ²⁺
Mercury (I) ion	Hg ₂ ²⁺

2 ⁻ Anions			
Oxide	O ²⁻	Chromate	CrO ₄ ²⁻
Sulfide	S ²⁻	Dichromate	Cr ₂ O ₇ ²⁻
Selenide	Se ²⁻	Carbonate	CO ₃ ²⁻
Sulfite	SO ₃ ²⁻	Silicate	SiO ₃ ²⁻
Sulfate	SO ₄ ²⁻	Oxalate	C ₂ O ₄ ²⁻
Thiosulfate	S ₂ O ₃ ²⁻	Peroxide	O ₂ ²⁻
Biphosphate	PO ₄ ²⁻		
Biphosphite	PO ₃ ²⁻		

3 ⁺ Cations	
Aluminum ion	Al ³⁺
Gallium ion	Ga ³⁺
Indium ion	In ³⁺

3 ⁻ Anions			
Nitride	N ³⁻	Phosphate	PO ₄ ³⁻
Phosphide	P ³⁻	Phosphite	PO ₃ ³⁻
Arsenate	AsO ₄ ³⁻	Borate	BO ₃ ³⁻

Solubility Rules

Soluble Ionic Compounds	Important Exceptions
Alkali metal ions	None
Ammonium, NH_4^+	None
Nitrate, NO_3^-	None
Acetate, $\text{C}_2\text{H}_3\text{O}_2^-$	None
Halides, Cl^- , Br^- , I^-	Compounds of Ag^+ , Hg_2^{2+} and Pb^{2+}
Sulfate, SO_4^{2-}	Compounds of Sr^{2+} , Ba^{2+} , Hg_2^{2+} and Pb^{2+}

The exceptions in this part of the table mean the compound will be insoluble and have a solid physical state. For example, magnesium sulfate is soluble or $\text{MgSO}_4(\text{aq})$, but strontium sulfate is insoluble or $\text{SrSO}_4(\text{s})$.

Insoluble Ionic Compounds	Important Exceptions
Carbonate, CO_3^{2-}	Compounds of ammonium and alkali metals
Phosphate, PO_4^{3-}	Compounds of ammonium and alkali metals
Hydroxide, OH^-	Compounds of alkali metals and Ca^{2+} , Sr^{2+} and Ba^{2+}
Sulfide, S^{2-}	Compounds of ammonium, the alkali metals and Ca^{2+} , Sr^{2+} and Ba^{2+}

The exceptions in this part of the table mean the compound will be soluble and have an aqueous physical state. For example, magnesium sulfide is insoluble or $\text{MgS}(\text{s})$, but calcium sulfide is soluble or $\text{CaS}(\text{aq})$.

Water Vapor Pressure Table

Temperature °C	Pressure mmHg
17.0	14.5
17.1	14.6
17.2	14.7
17.3	14.8
17.4	14.9
17.5	15
17.6	15.1
17.7	15.2
17.8	15.3
17.9	15.4
18.0	15.5
18.1	15.6
18.2	15.7
18.3	15.8
18.4	15.9
18.5	16.0
18.6	16.1
18.7	16.2
18.8	16.3
18.9	16.4
19.0	16.5
19.1	16.6
19.2	16.7
19.3	16.8
19.4	16.9
19.5	17.0
19.6	17.1
19.7	17.2
19.8	17.3
19.9	17.4
20.0	17.5
20.1	17.6
20.2	17.7
20.3	17.8
20.4	17.9
20.5	18.0
20.6	18.1
20.7	18.2
20.8	18.3
20.9	18.4

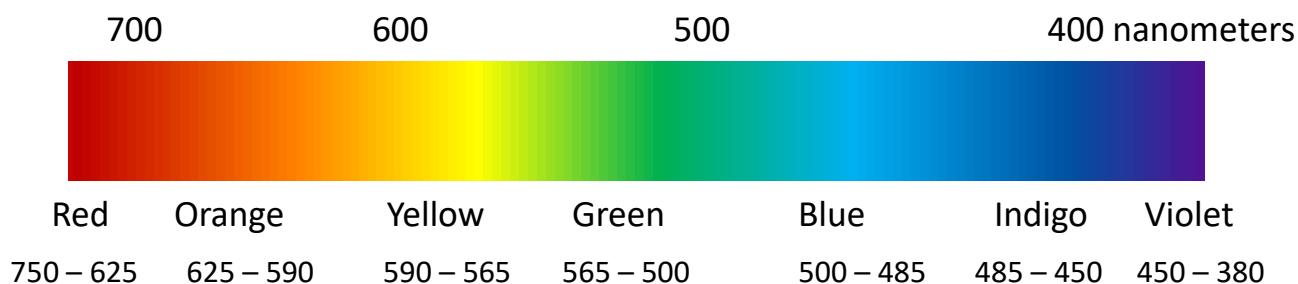
Temperature °C	Pressure mmHg
21.0	18.5
21.1	18.6
21.2	18.7
21.3	18.8
21.4	18.9
21.5	19.0
21.6	19.1
21.7	19.2
21.8	19.3
21.9	19.4
22.0	19.5
22.1	19.6
22.2	19.7
22.3	19.8
22.4	19.9
22.5	20.0
22.6	20.1
22.7	20.2
22.8	20.3
22.9	20.4
23.0	20.5
23.1	20.6
23.2	20.7
23.3	20.8
23.4	20.9
23.5	21.0
23.6	21.1
23.7	21.2
23.8	21.3
23.9	21.4
24.0	21.5
24.1	21.6
24.2	21.7
24.3	21.8
24.4	21.9
24.5	22.0
24.6	22.1
24.7	22.2
24.8	22.3
24.9	22.4

Temperature °C	Pressure mmHg
25.0	22.5
25.1	22.6
25.2	22.7
25.3	22.8
25.4	22.9
25.5	23.0
25.6	23.1
25.7	23.2
25.8	23.3
25.9	23.4
26.0	23.5
26.1	23.6
26.2	23.7
26.3	23.8
26.4	23.9
26.5	24.0
26.6	24.1
26.7	24.2
26.8	24.3
26.9	24.4
27.0	24.5
27.1	24.6
27.2	24.7
27.3	24.8
27.4	24.9
27.5	25.0
27.6	25.1
27.7	25.2
27.8	25.3
27.9	25.4
28.0	25.5
28.1	25.6
28.2	25.7
28.3	25.8
28.4	25.9
28.5	26.0
28.6	26.1
28.7	26.2
28.8	26.3
28.9	26.4

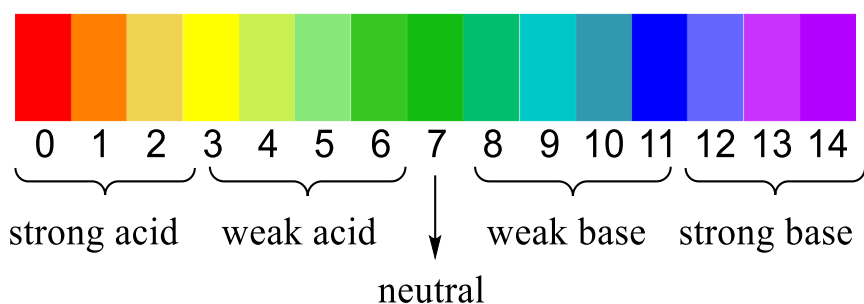
Appendix VI

Visible Spectrum, pH Scale, Conductivity

Visible Spectrum



pH Scale



Conductivity and Electrolyte Table

Scale	Red LED	Green LED	Conductivity	Electrolyte
0	Off	Off	Low or None	Nonelectrolyte
1	Dim	Off	Low	Nonelectrolyte
2	Medium or Bright	Off or Dim	Medium	Weak
3	Bright	Dim	High	Strong
4	Very bright	Medium	Very High	Strong