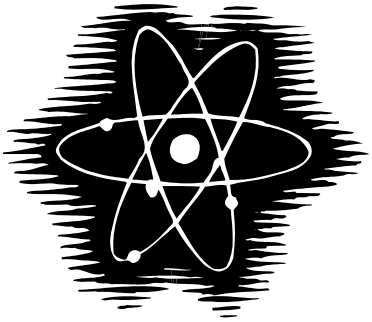


MYP Chemistry Work-in-book

2014 - 2015

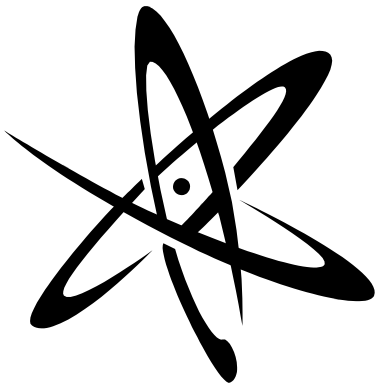


Table of Contents

Unit 1: The Atom

Unit 2: Conservation of Matter

Unit 3: Bonding and Intermolecular Forces

Unit 4: Properties of Matter

STUDENT LAB SAFETY CONTRACT

Rampart High School Science Department

PURPOSE

The Science curriculum at Rampart High School is a hands-on laboratory experience. Students will be asked to participate in some activities which require the use of hazardous chemicals and/or potentially dangerous equipment. Safety in the science classroom is the #1 priority for students, teachers, and parents. To ensure a safe science classroom, a list of guidelines has been developed and provided to you in this student safety contract. These rules must be followed at all times. We ask that you read through the rules carefully and sign a contract agreeing to abide by these guidelines in order to safely participate in our curriculum. **PLEASE RETURN THE SIGNED AGREEMENT TO YOUR SCIENCE TEACHER ON THE NEXT DAY OF CLASSES.**

GENERAL GUIDELINES

1. Conduct yourself in a responsible manner at all times in the laboratory.
2. Follow all written and verbal instructions carefully. If you do not understand a direction or part of a procedure, ask the instructor before proceeding.
3. Never work alone. No student may work in the laboratory without an instructor present.
4. When first entering a science room, do not touch any equipment, chemicals, or other materials in the laboratory area until you are instructed to do so.
5. Do not eat food, drink beverages, or chew gum in the laboratory. Do not use laboratory glassware as containers for food or beverages.
6. Perform only those experiments authorized by the instructor. Never do anything in the laboratory that is not called for in the laboratory procedures or by your instructor. Carefully follow all instructions, both written and oral: Unauthorized experiments are prohibited.
7. Be prepared for your work in the laboratory. Read all procedures thoroughly before entering the laboratory. Never fool around in the laboratory. Horseplay, practical jokes, and pranks are dangerous and prohibited.
8. Observe good housekeeping practices. Work areas should be kept clean and tidy at all times. Bring only your laboratory instructions, worksheets, and/or reports to the work area. Other materials (books, purses, backpacks etc.) should be stored in the classroom area.
9. Keep aisles clear. Push your chair under the desk when not in use.
10. Know the locations and operating procedures of all safety equipment including the first aid kit, eyewash station, safety shower, fire extinguisher, and fire blanket. Know where the fire alarm and the exits are located.
11. Always work in a well-ventilated area. Use the fume hood when working with volatile substances or poisonous vapors. Never place your head into the fume hood.
12. Be alert and proceed with caution at all times in the laboratory. Notify the instructor immediately of any unsafe conditions you observe.
13. Dispose of all chemical waste properly. Never mix chemicals in sink drains. Sinks are to be used only for water and those solutions designated by the instructor. Solid chemicals, metals, matches, filter paper, and all other insoluble materials are to be disposed of in the proper waste containers, not in the sink.
14. Labels and equipment instructions must be read carefully before use. Set up and use the prescribed apparatus as directed in the laboratory instructions or by your instructor.
15. Keep hands away from face, eyes, mouth and body while using chemicals or preserved specimens. Wash your hands with soap and water after performing all experiments. Clean (with detergent), rinse and wipe dry all work surfaces (including the sink) and apparatus at the end of the experiment. Return all equipment clean and in working order to the proper storage area.
16. Experiments must be personally monitored at all times. You will be assigned a laboratory station at which to work. Do not wander around the room, distract other students, or interfere with the laboratory experiments of others.
17. Students are never permitted in the science storage room or preparation room unless given specific permission by their instructor.
18. Know what to do if there is a fire drill during a laboratory period; containers must be closed, gas valves turned off, fume hoods turned off, any electrical equipment turned off.

19. Handle all living organisms used in a laboratory activity in a humane manner. Preserved biological materials are to be treated with respect and disposed of properly.
20. When using knives and other sharp instruments, always carry with tips and points pointing down and away. Always cut away from your body. Never try to catch falling sharp instruments. Grab sharp instruments only by the handles.

CLOTHING

21. Any time chemicals, heat or glassware are used, students will wear laboratory goggles. There will be no exceptions to this rule!
22. Dress properly during a laboratory activity. Long hair, dangling jewelry, and loose or baggy clothing are a hazard in the laboratory. Long hair must be tied back and dangling jewelry and loose or baggy clothing must be secured. Shoes should completely cover the foot. Appropriate foot-ware as required by the teacher.
23. Lab aprons have been provided for your use and should be worn during laboratory activities.

ACCIDENTS AND INJURIES

24. Report any accidents (spill, breakage, etc.) or injury (cut, burn, etc) to the instructor **immediately**, no matter how trivial it may appear.
25. If a chemical should splash in your eye(s) or on your skin, immediately flush with running water from the eyewash station or safety shower for at least 20 minutes. Notify the instructor immediately.

HANDLING CHEMICALS

26. All chemicals in the laboratory are to be considered dangerous. Do not touch, taste, or smell any chemicals unless specifically instructed to do so. The proper technique for smelling chemical fumes will be demonstrated to you.
27. Check the label on chemical bottles twice before removing any of the contents. Take only as much chemical as you need.
28. Never return unused chemicals to their original containers.
29. Never use mouth suction to fill a pipet. Use a rubber bulb or pipet pump.
30. When transferring reagents from one container to another, hold the containers away from your body.
31. Acids must be handled with extreme care. You will be shown the proper method for diluting strong acid. Always add acid to water, swirl or stir the solution and be careful of the heat produced, particularly with sulfuric acid.
32. Handle flammable hazardous liquids over a pan to contain spills. Never dispense flammable liquids anywhere near an open flame or source of heat.
33. Never remove chemicals or other materials from the laboratory area.
34. Take great care when transferring acids and other chemicals from one part of the laboratory to another. Hold them securely and walk carefully.

HANDLING GLASSWARE AND EQUIPMENT

35. Carry glass tubing, especially long pieces, in a vertical position to minimize the likelihood of breakage and injury.
36. Never handle broken glass with your bare hands. Use a brush and dustpan to clean up broken glass. Place broken or waste glassware in the designated glass disposal container.
37. Inserting and removing glass tubing from rubber stoppers can be dangerous. Always lubricate glassware before attempting to insert it in a stopper. Always protect your hands with towels or cotton gloves when inserting glass tubing into, or removing it from, a rubber stopper. If a piece of glassware becomes "frozen" in a stopper, take it to your instructor for removal.
38. Fill wash bottles only with distilled water and use only as intended, e.g., rinsing glassware and equipment, or adding water to a container.
39. When removing electrical plug from its socket, grasp the plug, not the electrical cord. Hands must be completely dry before touching an electrical switch, plug, or outlet.
40. Examine glassware before each use. Never use chipped or cracked glassware. Never use dirty glassware.
41. Report damaged electrical equipment immediately. Look for things such as frayed cords, exposed wires, and loose connections. Do not use damaged electrical equipment.
42. If you do not understand how to use a piece of equipment, ask the instructor for help.
43. Do not immerse hot glassware in cold water; it may shatter.

HEATING SUBSTANCES

44. Exercise extreme caution when using a gas burner. Take care that hair, clothing and hands are a safe distance from the flame at all times. Do not put any substance into the flame unless specifically instructed to do so. Never reach over an exposed flame. Light gas (or alcohol) burners only as instructed by the teacher.

45. Never leave a lit burner unattended. Never leave anything that is being heated or is visibly reacting unattended. Always turn the burner or hot plate off when not in use.

46. You will be instructed in the proper method of heating and boiling liquids in test tubes. Do not point the open end of a test tube being heated at yourself or anyone else.

47. Heated metals and glass remain very hot for a long time. They should be set aside to cool and picked up with caution. Use tongs or heat-protective gloves if necessary.

48. Never look into a container that is being heated.

49. Do not place hot apparatus directly on the laboratory desk. Always use an insulating pad. Allow plenty of time for hot apparatus to cool before touching it.

50. When bending glass, allow time for the glass to cool before further handling. Hot and cold glass has the same visual appearance. Determine if an object is hot by bringing the back of your hand close to it prior to grasping it.

In addition to these general guidelines, ALWAYS abide by any additional safety procedures provided by your instructor at the time of an activity.

Dear Students, Parents, and Guardians,

The Rampart High School Science Department feels that you should be informed regarding the school's effort to create and maintain a safe science classroom/laboratory environment. With the cooperation of the instructors, parents, and students, a safety instruction program can eliminate, prevent, and correct possible hazards.

Please take the time to familiarize yourself with the guidelines established in the *Student Lab Safety Contract* which can be found by visiting the Rampart High School teacher website for your student's teacher. Feel free to print this document for your records.

After reading the Student Lab Safety Contract, please complete the questions below and sign this agreement. Return this form to your science teacher on the second day of school.

QUESTIONS

Do you wear contact lenses? ☐ YES ☐ NO

Are you color blind? ☐ YES ☐ NO

Do you have any allergies? ☐ YES ☐ NO

If YES, list specific allergies:

STUDENT AGREEMENT

I, _____ (student's name) have read and agree to follow all of the safety rules set forth in the Student Lab Safety Contract. I realize that I must obey these rules to insure my own safety, and that of my fellow students and instructors. I will cooperate to the fullest extent with my instructor and fellow students to maintain a safe lab environment. I will also closely follow any oral and written instructions additionally provided by the instructor as part of a specific activity**. I am aware that any violation of this safety contract that results in unsafe conduct in the laboratory or misbehavior on my part, may result in being removed from the laboratory, detention, receiving a failing grade, and/or dismissal from the course.**

_____
Student signature

_____
Date

Parent or Guardian,

Your signature on this contract indicates that you have read this Student Lab Safety Contract, are aware of the measures taken to insure the safety of your son/daughter in the science laboratory, and will instruct your son/daughter to uphold his/her agreement to follow these rules and procedures in the laboratory.

_____
Parent/Guardian Signature

_____
Date

NAME THAT GLASSWARE

(OR OTHER ESSENTIAL LAB TOOL)

At lab tables around the room, there are various pieces of glassware and sometimes the proper name of that glassware. Each of these items is essential for Chemistry. Read the description below to determine the name of the glassware (if needed) and draw a quick sketch of the glassware.

Description	Name	Drawing
• Is good for measuring liquids when accuracy is not very important. Has a small spout to make pouring easy.		
• Its skinny neck means that pouring is difficult, but it makes it easy to swirl to mix chemicals. The small neck also makes it bad for pouring.		
• One of the most accurate tools for measuring liquids as it is dispensed. It must be read in reverse.		
• The flat, concave face of this glassware makes it excellent for evaporating liquids.		
• It can be glass or plastic. It is great for dispensing a liquid drop by drop.		
• This tool is used to accurately obtain the mass (not weight) of a chemical. Sometime it is accurate to $1/1000^{th}$ of a gram.		
• These are the most important safety tool that you should never be without.		
• It is useful for small-scale chemical reactions. You can do more than one reaction at a time, and its clear plastic allows you to notice color changes and solids easily.		
• If you needed to measure out 10.4 milliliters of water and then transfer the water to a flask, this tool (plastic or glass) would be best.		
• Since everyone knows not to stick their fingers in a container of chemicals, we use this tool for mixing instead.		
• This piece of glassware has no markings for measurement, but it can be useful for combining or heating small amounts of chemicals.		

Periodic Table of the Elements

1A	2A	3B	4B	5B	6B	7B		8B			1B	2B	3A	4A	5A	6A	7A	8A
hydrogen 1 **H** 1.00794																		helium 2 **He** 4.002602
lithium 3 **Li** 6.941	beryllium 4 **Be** 9.012182												boron 5 **B** 10.811	carbon 6 **C** 12.0107	nitrogen 7 **N** 14.00674	oxygen 8 **O** 15.9994	fluorine 9 **F** 18.9984	neon 10 **Ne** 20.1797
sodium 11 **Na** 22.98977	magnesium 12 **Mg** 24.3050												aluminium 13 **Al** 26.981538	silicon 14 **Si** 28.0855	phosphorus 15 **P** 30.97376	sulphur 16 **S** 32.065	chlorine 17 **Cl** 35.453	argon 18 **Ar** 39.984
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** 65.409		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
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
caesium 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]
francium 87 **Fr** [223]	radium 88 **Ra** [226]	lawrencium 103 **Lr** [262]	rutherfordium 104 **Rf** [261]	dubnium 105 **Db** [262]	seaborgium 106 **Sg** [266]	bohrium 107 **Bh** [264]	hassium 108 **Hs** [269]	meitnerium 109 **Mt** [268]	darmstadtium 110 **Ds** [271]	roentgenium 111 **Rg** [272]	ununbium 112 **Uub** [285]			ununquadium 114 **Uuq** [289]				

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]

Atomic Theory Organizer

Scientist	Date	Outline the experiment	Experimental Conclu-sions	Name and Sketch of model	Shortcomings of new model
Greeks					
Democritus & Leucip-pus					
John Dalton					

Scientist	Date	Outline the experiment	Experimental Conclusions	Name and Sketch of model	Shortcomings of new model
J.J Thomson					
Ernest Rutherford					
Niels Bohr					

Name _____ Per _____

MODERN ATOMIC THEORY

1. What is a Cathode Ray Tube?

2. What was the significance of the Cathode Ray tube experiments?

3. Briefly describe the set-up of Rutherford's Gold Foil Experiment.

4. What were the two major conclusions that Rutherford made from his Gold Foil Experiment?

Name _____ Per _____

MODERN ATOMIC THEORY PART 2

1. What is wavelength? amplitude? frequency?

2. List the colors of visible light in order of increasing energy.

3a. Label the following wave diagrams with wavelength and amplitude.
b. Which of the waves has a higher frequency?
c. Which of the waves has a lower energy?

4. As wavelength gets shorter, frequency _____. As wavelength gets shorter, energy _____.

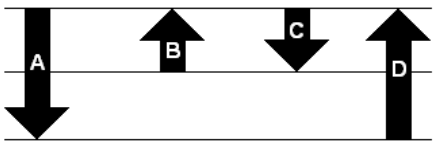
5. Describe the Bohr model of the atom.

6. What is the difference between a bright line spectrum and a continuous spectrum? How do energy levels account for the appearance of bright line spectra?

7. What is the electromagnetic (EM) spectrum? What is the highest energy wave? Lowest?

8. Explain how electron movement between energy levels produces photons of light.

9. Consider this diagram of an atom with arrows representing electron movement.



Which two arrows correspond to energy absorption by the atom? _____
Which two arrows correspond to energy emission by the atom? _____
If violet and green light are produced by the movement illustrated here, which arrow represents emission of violet light? _____ green light? _____

Electron Configuration Practice

In the space below, write the electron configurations of the following elements:

1. sodium _____
2. iron _____
3. bromine _____
4. barium _____
5. molybdenum _____
6. cobalt _____
7. silver _____
8. tellurium _____

Determine what elements are denoted by the following electron configurations:

9. $1s^2 2s^2 2p^6 3s^2 3p^4$ _____
10. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$ _____
11. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^3$ _____
12. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$ _____

Determine which of the following electron configurations are not valid and explain why:

13. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4d^{10} 4p^5$ _____
14. $1s^2 2s^2 2p^6 3s^3 3d^5$ _____
15. $1s^2 2s^2 2p^6 3s^2 3p^6\ 3d^{10} 4s^2$_____

Write the electron configurations of the following symbols:

16. Mg^{2+} _____
17. F^{-1} _____
18. Ne _____

Atomic Structure

Go to the following website: http://phet.colorado.edu/en/simulation/build-an-atom **Click run now.**

1. **Drag a proton onto the x in the middle of the atom. (Make sure to click on the green + symbols to the side)**
   a. What element did you make? _____
   b. What charge does this atom have? _____
   c. What does this tell you about the charge of a proton? _____
   d. What is the mass of this atom? _____
   e. What does this tell you about the mass of a proton? _____

2. **Add a neutron to the center of the atom next to the proton.**
   a. Did the identity of the element change? _____
   b. Did the charge of the atom change? _____
   c. What does this tell you about the charge of a neutron? _____
   d. Did the mass of the atom change? If so, by how much? _____
   e. What does this tell you about the mass of a neutron? _____

3. **Add an electron to the outside of the atom.**
   a. Did the identity of the element change? _____
   b. Did the charge of the atom change? If so, by how much? _____
   c. What does this tell you about the charge of an electron? _____
   d. Did the mass of the atom change? _____
   e. What does this tell you about the mass of an electron? _____

4. **Fill in the following table (if you are unsure about the answers add more protons, electrons, and neutrons)**

Subatomic Particle	Charge	Mass	Location in atom
Proton			
Neutron			
Electron			

5. **Make an atom with 6 protons, 6 neutrons, and 6 electrons.**
    a. What element did you make?_____
6. **Add another proton this atom.**
    a. What element did you make?_____
7. **Take two protons away from this atom.**
    a. What element did you make?_____
8. **Remake an atom with 6 protons, 6 neutrons, and 6 electrons.**
9. **Add one more neutron.**
    a. Did the element change?_____
10. **Add one more electron.**
    a. Did the element change?_____
11. **Which subatomic particle determines an element's identity? _____**
12. **Remake an atom with 6 protons, 6 neutrons, and 6 electrons. Add one more neutron.**
    a. What changed?_____
    b. What stayed the same? _____
13. **Add another neutron to this atom.**
    a. What changed?_____
    b. What stayed the same?_____
14. **Make an atom with 1 proton and 1 electron.**
    a. What element did you make? _____
    b. What is its mass?_____
15. **Add a neutron to this atom.**
    a. Did the identity of the element change? _____
    b. Did the mass change?_____
16. **Atoms from numbers 12 and 13 are called isotopes. Atoms from number 14 and 15 are also called isotopes.**

    Based on this, how would you define the term isotope?_____

_____

17. Make an atom with 6 protons, 6 neutrons, and 6 electrons.

a. What element did you make? _____

b. What charge does this atom have? _____

18. Add another electron to this atom.

a. Did the element's identity change? Why or why not? _____

b. What is the new charge of this atom? _____

19. Add another electron to this atom.

a. Did the element's identity change? Why or why not? _____

b. What is the new charge of this atom? _____

20. Take three electrons away from this atom.

a. Did the element's identity change? Why or why not? _____

b. What is the new charge of this atom? _____

21. Atoms in numbers 19-21 are called ions. Based on this, how would you define the term ion?

_____

22. Make an atom with 7 protons, 7 electrons, and 7 neutrons.

a. Draw the symbol that is shown in the box to the right (also called a nuclear symbol).

23. Add another neutron to this atom.

a. Draw the new symbol that is shown in the box to the right.

24. The upper left number is called the mass number. How is the mass number calculated?

_____

25. What does the lower left number represent? (hint add and substrate particles to figure this out)

_____

26. Add 2 more electrons to this atom.

a. Draw the new symbol for this atom.

27. What does the upper right number represent? _____

How are charges determined? _____

Play the game when you are done— What's your score? _____

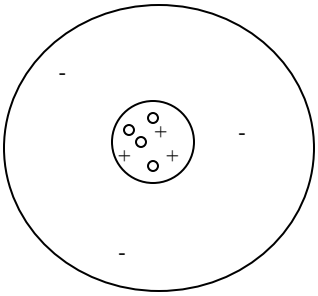
Name _____ Per _____

ATOMIC STRUCTURE

1. Fill in the following table:

Particle	Charge	Mass (amu)	Location in atom
Proton			
Electron			
Neutron			

2. Label the diagram of the atom with the following terms: nucleus, electron cloud, electron, proton, neutron.



Which element does this diagram represent? _____

3. Define the terms atom, ion, and isotope.

4. Complete the blanks in the table.

Element	Mass Number	Atomic Number	# of Protons	# of Electrons	# of Neutrons
Oxygen	17				
			10		12
		17			15
				9	9
	71		35		

Atomic Structure and Mass

Fill in the table with the appropriate information.

Isotope	Mass Number	Atomic Number	# of Protons	# of Electrons	# of Neutrons	Nuclear Symbol
Sodium-23	23			11		
			8	10	9	
		24		24	27	
			12	10	12	
	32		15	18		

Determine the number of protons and electrons in the following ions:

Element	Ion Symbol	Number of Protons	Number of Electrons
	K^{+}		
	P^{3-}		
	Ca^{2+}		
	I^{-}		

Calculate the atomic mass of a sample that contains 5% ^{79}Br and 95% ^{80}Br.

Calculate the average atomic mass of Sulfur.

Isotope	Abundance
Sulfur — 32	95
Sulfur — 33	0.76
Sulfur — 34	4.22
Sulfur — 35	0.014

Atom Review

Answer the following questions:

1. Draw **Dalton's model** of the atom:

2. What are some shortcomings of Dalton's model?

3. **Describe Thomson's experiment**. Provide a diagram to help with your explanation:

4. What subatomic particle did **Thomson** discover with his experiment? _____

5. What was **name of Thomson's model** of the atom? _____

6. **Draw AND label Thompson's model** of the atom:

7. **Describe Rutherford's experiment** (provide a diagram to aid your explanation):

8. What two things did **Rutherford** discover about the atom with his experiment?

   _____ and _____

9. **Draw AND label Rutherford's model** of the atom:

10. How did Bohr improve Rutherford's model of the atom? (hint: what is different about the electrons?)

11. **Draw AND label** the Bohr model of an oxygen atom:

12. Draw AND label the Bohr model of a chlorine atom:

13. Draw a Bohr model of a neon atom. Show an electron moving from n = 1 to n = 3 and an electron moving from n = 2 to n = 3. Which of these is more likely to emit red light? Green light?

14. Define wavelength, amplitude, and frequency?

15. List the colors of visible light in order of increasing energy.

16. As wavelength gets shorter, frequency _____.

17. As wavelength gets shorter, energy _____.

18. What is the difference between a bright line spectrum and a continuous spectrum? How do energy levels account for the appearance of bright line spectra?

19. Explain how electron movement between energy levels produces photons of light.

20. How do waves of red light and blue light differ with respect to frequency? wavelength? energy?

21. Compare the Bohr model to our current model of the atom (quantum model). How are they different? How are they similar?

22. Use the quantum mechanical model to explain how electrons move about the nucleus.

23. What are atomic orbitals?

24. What do the letters s, p, d, and f represent in the current quantum model of the atom?

25. "s" sublevels can hold a total of ___ electrons. p sublevels can hold ___ electrons, while d sublevels can hold ___, and f sublevels can hold ___ electrons.

26. Write the electron configuration for the following:

    Cu

    Cu^{2+}

    I

    I^{-1}

    Ba

    Ba^{+2}

27. In an electron configuration, there is a whole number, followed by a lower case letter, and then a number as a superscript to the letter. (Example: $4p^3$) What do each of the symbols represent?

28. Define the following terms (or write the equation that defines them):

    Atomic Number:

    Mass Number:

    Average Atomic Mass:

    Atom:

    Ion:

    Isotope:

29. Provide an example of an ion:

30. Provide an example of two or more isotopes:

31. Fill in the following table:

Subatomic Particle	Mass (amu)	Charge	Location
Proton			
Neutron			
Electron			

32. What two subatomic particles contribute to an atom's mass?

33. What two subatomic particles contribute to an atom's charge?

34. Fill in the following table:

Atomic Symbol	Atomic #	# of protons	# of neutrons	# of electrons	Mass #
$^{57}_{26}Fe$					
$^{24}_{12}Mg^{+2}$					
		35	45	36	
		14	14	17	
$^{235}_{92}U$					
	14			14	29

Put a star next to all the ions in the table above

35. Fill in the atomic symbol for the following table:

Atomic Symbol	# of protons	# of neutrons	# of electrons
	5	6	2
	1	1	1
	1	0	0
	5	5	5

36. Using the table above, which elements are isotopes of each other (2 pairs)?

37. Write the atomic symbol for an ion with 17 protons, 19 neutrons, and 18 electrons.

38. Write the atomic symbol for an atom with 6 protons, 8 neutrons, and 6 electrons.

39. Calculate the average atomic mass of a sample that contains 80% Iodine-127, 17% Iodine-126, and 3% Iodine-128 (the number after the element is the mass number).

40. Calculate the average atomic mass of a sample that contains 50% Gold-197 and 50% Gold-198 (the number after the element is the mass number)

41. Fill in the following statements with the appropriate elemental symbol:

_____ is the 7th period alkaline earth metal

_____ is the 3rd period noble gas

_____ is the 6th period alkali metal

_____ is the 2nd period halogen

_____ is the transition metal with the lowest atomic number

_____ is the 9th Lanthanide element

_____ is the only non-metal in the boron family

_____ is the 1st inner transition metal

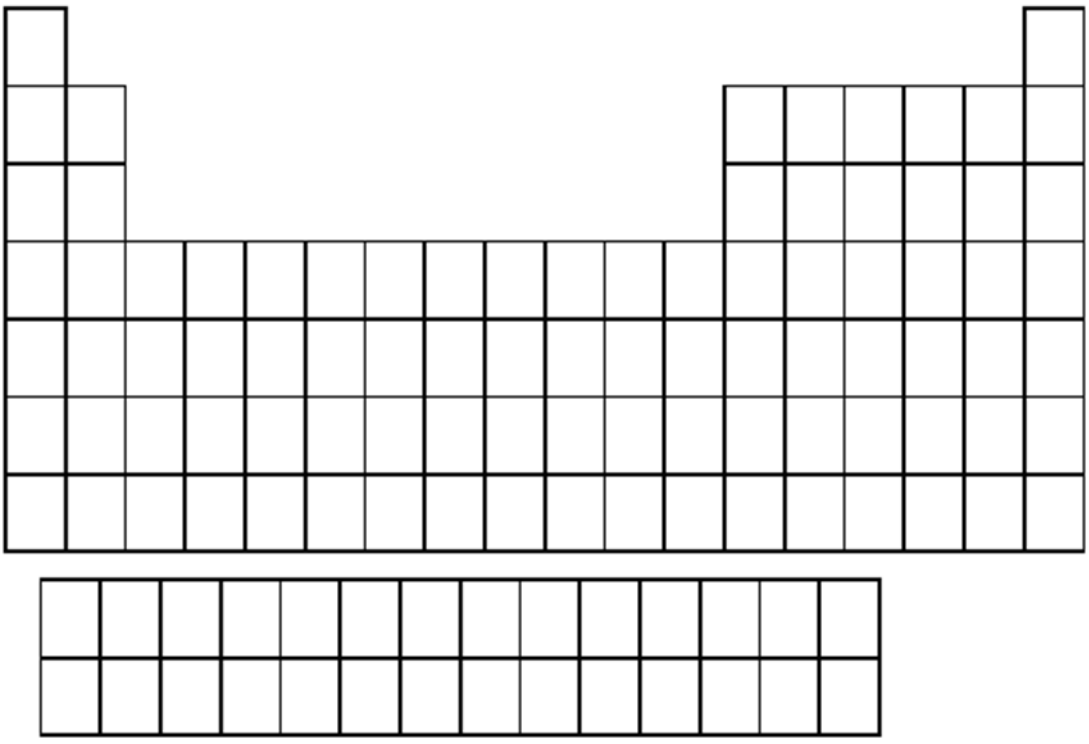
_____ is the element in the nitrogen family that is in the 5th period

_____is the element in the 3rd period and 2nd family

Use the blank periodic table to complete the following questions.

42. Label the chart with the following groups: Hydrogen Family, Transition Metals, Halogens, Alkaline Earth Metals, Alkali Metals, Noble Gases, Inner Transition Metals, Boron Family, Carbon Family, Nitrogen Family, and Oxygen Family.

43. Label metals and nonmetals on the chart below. Label periods and groups on the periodic table (write their numbers in).



Name _____ Per____

Naming and Writing Formulas for Compounds
An Inquiry Activity

Part 1: Naming Rules for Binary Covalent Compounds

You and your partner were given a stack of eight cards. On one side of each card there is a chemical formula (ex. H_2S) of a binary (two element) covalent compound and on the other side a name (ex. Dihydrogen monosulfide).

Using the information on these cards you are to work together to develop a list of rules for writing the name of a binary covalent compound when given its chemical formula. Be sure that your rules would allow you to appropriately name all of the compounds you have been given.

Use the space below to write a rough draft of your rules based on the formulas and names you see on the cards.

Now use your rules to name the compounds below:

Formula	Name		Formula	Name
B_2F_6			IBr	
H_2O			C_3H_8	
CO_2			F_3N	

After having used your rules, write a final, more concise draft (if necessary) in the space below.

Part 2: Writing Formulas for Covalent Compounds

Use the space below to change your rules to work backwards – i.e. so that you can use them to write a formula for a covalent compound if you are given its name.

Apply your rules to the names in the following table:

Formula	Name		Formula	Name
	Carbon Dioxide			Nitrogen Monoxide
	Diboron Trisulfide			Disulfur Decaiodide
	Phosphorus Pentachloride			Tetrasulfur Dioxide

If necessary, rewrite your rules for writing formulas of binary covalent compounds below in a more concise manner.

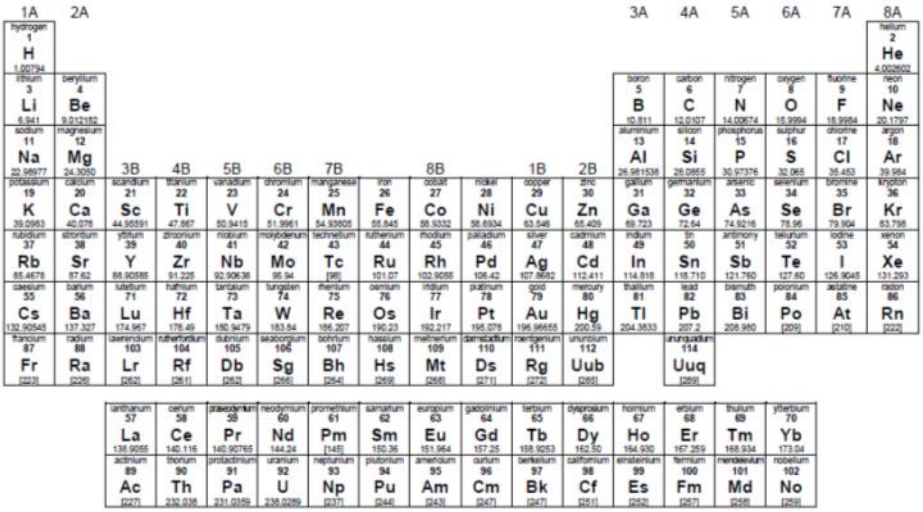
Part 3: Comparing Covalent and Ionic Compounds

Examine both the covalent compound cards and the ionic compound models. In the chart below, list the similarities and differences that you notice. It may help to complete the periodic tables (next page) first.

	Similarities	**Differences**
Covalent		
Ionic		

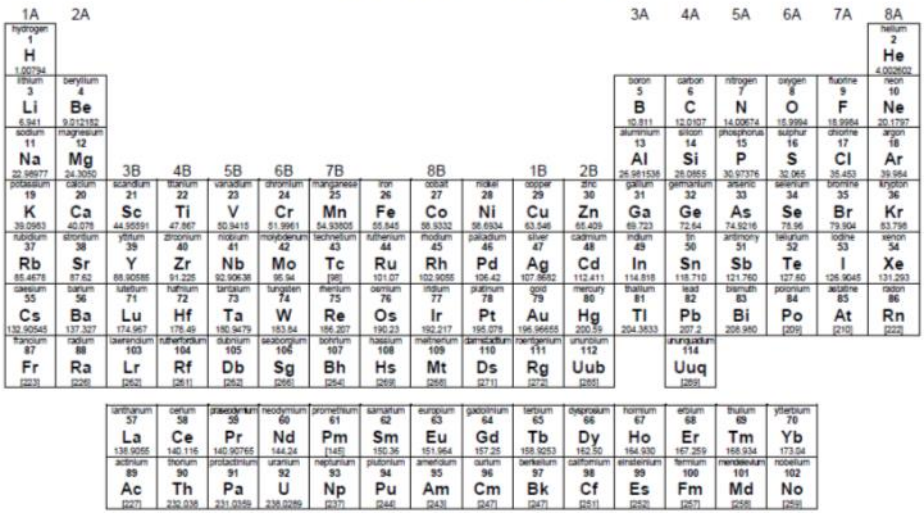
On the first periodic table below, highlight any elements that are found in your eight covalent compound examples on the cards.

Periodic Table of the Elements



On the second periodic table below, highlight any elements that are found on your six ionic compound models.

Periodic Table of the Elements



Concluding Questions:

What similarities do you notice between the elements highlighted on the two periodic tables?

What differences do you notice between the elements highlighted on the two periodic tables?

Based on your answers to the previous two questions, provide separate definitions for covalent and ionic compounds.

Part 4: Naming Simple Ionic Compounds

You and your partner were also given six models of binary ionic compounds that look like puzzles pieces put together. On one side of each card there are ion symbols and names (ex. Li^+; Lithium Ion and S^{2-}; Sulfide Ion) and on the other side both a chemical formula (ex. Li_2S) of an ionic compound and a name (ex. Lithium Sulfide).

Using the information on these models you are to work together to develop a list of rules for writing the name of a binary ionic compound when given its chemical formula. Be sure that your rules would allow you to appropriately name all of the compounds you have been given.

Use the space below to write a rough draft of your rules based on the formulas and names you see on the models.

Now use your rules to name the compounds below:

Formula	Name		Formula	Name
KF			$BaBr_2$	
Rb_2O			Ga_2S_3	
CaS			Cs_3N	

Looking once more at the models you were given – there are three more compounds that these six ions (sodium, magnesium, aluminum, chloride, oxide, and nitride) could form of which you were NOT given models. Please list the names and formulas of these three compounds.

After having used your rules, write a more concise draft (if necessary) in the space below.

Part 5: Writing Formulas for Ionic Compounds

A few guiding questions:

In order to write a formula for an ionic compound there is **one very important piece of information** that you need to know or be given in some way about each ion in the compound. What is that?

For most ions (i.e. those that do not require a roman numeral as seen later in part 6), how do you obtain this information?

Once you have this information about each ion, how do you proceed to write a single formula from the two ions?

Use the space below to change your rules from part 4 as well the information from the previous questions to work backwards – i.e. so that you can use them to write a formula for an ionic compound if you are given its name.

Apply your rules to the names in the following table:

Formula	Name		Formula	Name
	Calcium Iodide			Gallium Oxide
	Beryllium Sulfide			Potassium Bromide
	Sodium Phosphide			Barium Nitride

If necessary, rewrite your rules for writing formulas of ionic compounds below in a more concise manner.

Part 6: More Naming and Writing Ionic Compounds

These models look a lot like the ones you and your partner used in Parts 3-5, but they have a few small changes. The models still have ion symbols and names (ex. Li^+; Lithium Ion and S^{2-}; Sulfide Ion) on one side and on the other side both a chemical formula (ex. Li_2S) of an ionic compound and a name (ex. Lithium Sulfide). However you will notice that now there are two additions: roman numerals (ex. Copper (II)) and polyatomic ions (ions composed of more than one element; ex. Sulfate – SO_4^{2-}). To help you out the table below lists some common polyatomic ions that you will come across throughout the year.

Formula	Name of Ion
NH_4^+	Ammonium
CO_3^{2-}	Carbonate
OH^-	Hydroxide

Formula	Name of Ion
NO_3^-	Nitrate
PO_4^{3-}	Phosphate
SO_4^{2-}	Sulfate

Using the information on these models you are to work together to **modify** your previous list of rules for writing the name of an ionic compound when given its chemical formula to apply to **ALL** ionic compounds, not just binary. Be sure that your rules would allow you to appropriately name all of the compounds you have been given.

A few guiding questions:
All of the elements that have roman numerals appear where on the periodic table (what group)?

How is the roman numeral determined?

When using multiple polyatomics in a compound (ex. two sulfate ions), how is that represented in the final chemical formula?

Do polyatomics change how a chemical formula is written? If so, how?

Use the space below to write a rough draft of your rules based on the formulas and names you see on the various models from parts 4 and 6 as well as the guiding questions above.

Now use your rules to name the compounds below:

Formula	Name
KOH	
$(NH_4)_2O$	
CoS	
$CaSO_4$	

Formula	Name
PdO_2	
$GaPO_4$	
$Ba(NO_3)_2$	
Cu_2CO_3	

After having used your rules, if possible, write a more concise draft in the space below.

Adjust your rules from Part 5 for writing ionic formulas to add in steps for compounds containing roman numerals and polyatomic ions.

Apply your rules to the names in the following table:

Formula	Name
	Beryllium Sulfate
	Copper(II) Chloride
	Nickel(II) Phosphate

Formula	Name
	Ammonium Carbonate
	Mercury(I) Nitrate
	Calcium Hydroxide

If necessary, rewrite your rules for writing formulas of ionic compounds below in a more concise manner.

Common Polyatomic Ions

Cations

+1	+2
NH_4^+ Ammonium	Hg_2^{2+} Mercury (I) or Mercurous

Anions

-1	-2	-3	-4
$C_2H_3O_2^-$ Acetate	CO_3^{2-} Carbonate	AsO_4^{3-} Arsenate	$Fe(CN)_6^{4-}$ Hexacyanoferrate (II) or Ferrocyanide
AlO_2^- Aluminate	CrO_4^{2-} Chromate	AsO_3^{3-} Arsenite	
HCO_3^{--} Bicarbonate, Hydrogen Carbonate	$Cr_2O_7^{2-}$ Dichromate	$Fe(CN)_6^{3-}$ Hexacyanoferrate (III) or Ferricyanide	
BrO_3^- Bromate	$C_2O_4^{2-}$ Oxalate	PO_4^{3-} Phosphate	
ClO_4^- Perchlorate	O_2^{2-} Peroxide	PO_3^{3-}Phosphite	
ClO_3^-Chlorate	SiO_3^{2-}Silicate		
ClO_2^- Chlorite	SO_4^{2}Sulfate		
ClO^- Hypochlorite	SO_3^{2-} Sulfite		
CrO_2^-Chromite	$C_4H_4O_6^{2-}$ Tartrate		
CN^- Cyanide	$S_2O_3^{2-}$ Thiosulfate		
OH^- Hydroxide	MoO_4^{2-}Molybdate		
IO_3^- Iodate.			
NO_3^-Nitrate			
NO_2^-Nitrite			
MnO_4^- Permanganate			

4. Molecular Compounds

First Element	Second Element	Formula	Formula Name
		CO_2	
			Sulfur Hexaflour-ide
		N_2O	
		CO	
			Carbon Tetrahydride
		PO_3	
		H_2O	
			Dinitrogen Trioxide
		CS_2	

Name _____ Per _____

Ionic and Covalent Compounds Worksheet

1. Complete the table of ionic compounds with all missing information:

Name of Cation	Name of Anion	Formula of Cation	Formula of Anion	Formula of Compound	Name of Compound
Calcium ion	Chloride ion				
Iron (III) ion	Phosphide ion				
		Na^{1+}	S^{2-}		
		Al^{3+}	Br^{1-}		
					Lithium Sulfide
					Platinum (IV) Oxide
Magnesium ion	Nitride ion				
		Ca^{2+}	NO_3^{1-}		
				HgS	

2. Write the names of these ionic compounds

CsBr	SnF_2
BaS	CaI_2
Hg_2S	Li_3N
CeO_2	BeO
$NiCl_3$	TiO_2

3. Write the formulas for these ionic compounds

Chromium (IV) Oxide	Magnesium Chloride
Chromium (III) Oxide	Nickel (II) Nitride
Aluminum Oxide	Iron (III) Phosphide
Nickel (II) Sulfide	Potassium Hydride
Silver (I) Sulfide	Lead (IV) Iodide

Name _____ Per _____

Introduction to Chemical Names and Formulas

Understanding Simple, Binary Ionic Compounds

Chemical terminology refers to the process of naming chemicals. This assignment focuses on the most basic rules and ideas involved in the naming of compounds.

The simplest compounds contain just two elements. Sodium chloride, NaCl, is an example of a binary compound. Several other examples are listed below.

Potassium bromide	KBr
Calcium bromide	$CaBr_2$
Lithium fluoride	LiF
Lithium oxide	Li_2O

In naming binary compounds, follow these rules:

1. The element with the positive ionic charge is written first.
2. The second word is formed by changing the ending of the name of the element to "ide". For example, bromine changes to bromide, fluorine changes to fluoride, and oxygen changes to oxide.

Metals usually have positive ionic charges, or states, while nonmetals (when combined with metals) have negative ionic charges.

Using the rules given above, name the compounds listed below.

1. MgO 1. _____
2. BaS 2. _____
3. K_3P 3. _____
4. Na_3N 4. _____

5. Below are the symbols for selected elements. Circle the symbols of the elements that usually have positive ionic charges.

Fe	C	N	Na	Sr
Se	Mn	Mg	Al	As
H	O	Ca	Ag	At

Practice Problems

Some transition metals have more than one positive ionic charge. Look at the formulas below.

Cu_2O	copper(I) oxide
CuO	copper(II) oxide
$FeCl_2$	iron(II) chloride
$FeCl_3$	iron(III) chloride

Note that roman numerals follow the names of the positive elements. The numerals indicate the ionic charge of the element in the compound.

Transition Metals in Ionic Compounds

Name the following compounds.

6. $SnCl_4$ 6. _____
7. MnO_2 7. _____
8. PbS 8. _____
9. Fe_2O_3 9. _____

Practice Problems

Polyatomic Ions in Ionic Compounds

In nature there are certain elements whose atoms combine to form charged groups called polyatomic ions. These polyatomic ions combine with other ions or other polyatomic ions to form compounds. To name such compounds, you must know the names of the polyatomic ions. Some common polyatomic ions are listed on the following page.

NH_4	1+	ammonium	SO_4	2-	sulfate
H_3O	1+	hydronium	NO_3	1-	nitrate
OH	1-	hydroxide	CO_3	2-	carbonate
$C_2H_3O_2$	1-	acetate	HCO_3	1-	hydrogen carbonate (bicarbonate)
ClO_3	1-	chlorate			
			PO_4	3-	phosphate

Name the following ionic compounds.

10. $BaSO_4$ 10. _____
11. $BaSO_3$ 11. _____
12. Na_2CO_3 12. _____
13. $NaHCO_3$ 13. _____
14. $(NH_4)_3PO_4$ 14. _____
15. NH_4OH 15. _____

Some nonmetallic elements may form more than one compound with another nonmetal. The names of these compounds must show the differences between them. Look at the following examples.

CO	carbon monoxide
CO_2	carbon dioxide
SO_2	sulfur dioxide
SO_3	sulfur trioxide
N_2H_4	dinitrogen tetrahydride

In the examples, the prefixes mono-, di-, tri-, and tetra- indicate the number of atoms of the nonmetal in the molecule. These and other prefixes with their meanings are listed below. Note that the prefixes are not used when naming compounds formed between a metal and nonmetal.

mono-	one	hexa-	six
di-	two	hepta-	seven
tri-	three	octa-	eight
tetra-	four	nona-	nine
penta-	five	deca-	ten

Formula	Charge	Name	Formula	Charge	Name

Name the following compounds.

16. N_2O_3 16. _____
17. PCl_3 17. _____
18. SiO_2 18. _____
19. P_2O_5 19. _____
20. CS_2 20. _____
21. B_2O_3 21. _____

Practice Problems

Name or give the formula for the following.

22. Iron (III) chloride 22. _____
23. Silver (I) acetate 23. _____
24. $Ca(OH)_2$ 24. _____
25. CCl_4 25. _____
26. Mercury (II) oxide 26. _____
27. $NaClO_3$ 27. _____
28. $(NH_4)_2SO_4$ 28. _____
29. Fe_2S_3 29. _____

30. Polyatomics are ions composed of two or more nonmetals. Most (with exception of ammonia) have a negative charge since they are composed of nonmetals. List the ten formulas for the ten polyatomics you are responsible for MEMORIZING:

a. ammonium_____ e. sulfate _____

b. hydroxide_____ f. phosphate _____

c. carbonate_____ g. bicarbonate _____

d. nitrate_____

31. When naming an ionic compound, what element is always named first, the metal or the nonmetal?

32. What ending is added to the name of the anion (the nonmetal)?

33. When a transition metal is present in an ionic compound, it must be written with a roman numeral. Why is the roman numeral necessary and what does it represent?

34. Write the formula of the following ionic compounds. First write out the ions for each element in the compound, and then combine them to make a NEUTRAL compound.

a. calcium chloride b. magnesium oxide

c. sodium iodide d. magnesium hydroxide

e. potassium nitrate f. aluminum chloride

g. iron (II) oxide h. iron (III) oxide

35. Name the following compounds:

a. MgS g. KBr

b. Ba_3N_2 h. Al_2O_3

c. NaI i. SrF_2

d. Li_2S j. Ca_3P_2

e. $CaSO_4$ k. $Pb(NO_3)_2$

f. $Fe(OH)_3$ l. Na_3PO_4

Name _____ Per _____

MIXED UP NAMING

Determine whether the compound is ionic or molecular. Write the chemical formula for the named compound.

Name of Compound	Molecular or Ionic?	Chemical Formula
iron (III) iodide		
strontium nitride		
boron trichloride		
Fluorine dioxide		
aluminum sulfide		
calcium carbonate		
diphosphorous pentasulfide		
mercury (I) oxide		
silver sulfate		
dinitrogen tetraoxide		
zinc nitrate		
dichlorine heptaoxide		
carbon monoxide		
magnesium phosphate		
ammonium hydroxide		

Determine whether the compound is ionic or molecular. Write the correct name for the formula provided.

Chemical Formula	Molecular or Ionic?	Name of Compound
$SnCl_2$		
SiS_2		
NH_4MnO_4		
HgS		
XeO_4		
BaS		
BrF_5		
$Ca(NO_3)_2$		
$TeCl_4$		
Li_3PO_4		
N_2F_2		
P_6S_{10}		
$CrCl_3$		
Zn_3N_2		
PbS_2		

Name_____ Per _____

Compound Naming Worksheet

1. Identify as ionic or covalent, then write the correct name for the following ionic and covalent compounds.

a. $AlPO_4$

b. $NaC_2H_3O_2$

c. CS_2

d. KNO_3

e. Cu_2SO_4

f. OCl_2

g. Li_2CO_3

h. NH_4NO_3

i. $Mn(OH)_2$

j. N_2O

k. H_2S

l. $SnCO_3$

m. $Pb(NO_3)_2$

n. $Ni_3(PO_4)_2$

o. Mn_2O_3

p. S_2O_6

q. $(NH_4)_3N$

r. CF_4

2. Identify as ionic or covalent, and then write the formulas for these ionic and covalent compounds.

a. sodium hydroxide

b. cobalt(II) chloride

c. ammonium oxide

d. tribromine tetrafluorine

e. silver carbonate

f. copper(I) phosphate

g. manganese(II) acetate

h. silicon dioxide

i. zinc hydroxide

j. iron(III) sulfate

k. tin(IV) phosphide

l. ammonium hydroxide

m. calcium phosphate

n. tetraphosphorus heptanitride

o. dicarbon tetrahydride

p. aluminum nitride

q. titanium(II) nitride

r. titanium(II) nitrate

Demystifying an Ingredient Label

Ingredient labels are filled with chemical names, most of which are highly technical and quite intimidating to the everyday consumer. Some of these ingredients are inorganic salts (ionic compounds), while others are organic compounds, meaning they contain carbon.

This ingredient label was taken from a bar of all-natural soap

This one is from a can of spray paint:

Both contain a significant amount of organic and inorganic compounds, although their identity and properties are a complete mystery to most of the general public. Manufacturers are legally required to report the identities of all components within their products. Some of these are listed under the guise of "natural flavors" or other simple names like "salt." Others are chemical trade names that are patented by the company in order to keep the actual chemical identity a trade secret. Most, however, are named according to the agreed upon chemical nomenclature standards, which are maintained by a group called the International Union for Pure and Applied Chemistry, or IUPAC.

Your goal will be to choose an ingredient from a label on some product within your house. You will produce an informative brochure about this compound in order to be shared with other consumers.

1. Choose an ingredient whose identity and structure would not be known to a common consumer
2. Produce a tri-fold brochure listing:
    a. The name as listed on the label (front of brochure), and what the IUPAC name would be
    b. A drawing of the compound showing its chemical structure (front of brochure)
    c. Desired uses of the compound, and why it is added to certain products (inside the brochure)
    d. Potential hazards or dangers of exposure to the compound (inside the brochure)
    e. Any special instructions regarding the compounds disposal should the product need to be disposed of (inside the brochure)

Be creative! Make these artistic and visually pleasing so that a consumer would be interested in reading it. I'm not interested in black and white brochures with cut-and-pasted google-isms. Such submissions will score low. An exemplary submission will include drawings or photos, multiple colors, and information that has been supplied **in your own words – not something downloaded and copy/pasted from the web**.

Introduction to Matter and Reactions

Procedure:

1. Use a pipette to fill one test tube ¾ full of sodium carbonate.
2. Use a pipette to fill a different test tube ¾ full of calcium chloride.
3. Gently slide both test tubes into an Erlenmeyer flask. Place a stopper into the flask so it is sealed tightly.
4. Determine and record the mass (grams) of flask and contents.
5. Slowly invert the flask to allow the two solutions in test tubes to mix.
6. Determine and record the final mass (grams) of the flask and reacted contents. Make sure to record your observations of the reaction.

Prior to Completing the Lab:

1. Predict how you believe the mass (grams) will change during the experiment. Will it increase, decrease, or remain constant? Explain why you predicted this result.

Raw Data Table:

Design a data table that allows you to collect all the data you will need to collect during this experiment (use the procedure to do this).

Processing Data

1. What was the total mass (grams) change? Show your calculations.

Analysis

2. What does this tell you about the starting mass and final mass of a reaction?

3. The following reaction occurred in this lab: $1\ Na_2CO_3 + 1\ CaCl_2 \rightarrow 1\ CaCO_3 + 2\ NaCl$ Based on this equation and your data, which of the following was conserved during the chemical reaction? <u>Explain</u> your reasoning.

   Number of atoms?

   Amount of grams?

   Number of $CaCl_2$ molecules?

4. Based on your data, try to explain what the law of conservation of matter is.

Balancing Reactions Practice

Write in numbers in front of the formulas, to balance the following equations.

___ CaO + ___ HNO_3 → ___ $Ca(NO_3)_2$ + ___ H_2O

___ NH_3 + ___ H_2SO_4 → ___ $(NH_4)_2SO_4$

___ HCl + ___ $ZnCO_3$ → ___ $ZnCl_2$ + ___ H_2O + ___ CO_2

___ SO_2 + ___ Mg → ___ S + ___ MgO

___ Fe_3O_4 + ___ H_2 → ___ Fe + ___ H_2O

___ K + ___ C_2H_5OH → ___ KC_2H_5O + ___ H_2

___ $Fe(OH)_3$ → ___ Fe_2O_3 + ___ H_2O

___ CH_3CO_2H + ___ O_2 → ___ CO_2 + ___ H_2O

___ $Pb(NO_3)_2$ → ___ PbO + ___ NO_2 + ___ O_2

___ $NaMnO_4$ + ___ HCl → ___ $NaCl$ + ___ $MnCl_2$ + ___ Cl_2 + ___ H_2O

More Balancing Practice

1. _____ Ba + _____K_2O → _____BaO + _____ K

2. _____S_8 + _____O_2 → _____SO_2

3. _____C_2H_4 + _____O_2 → _____CO_2 + _____H_2O

4. ____Fe_2O_3 + _____$BaCl_2$ → _____ BaO + _____$FeCl_3$

5. _____$CaCO_3$ → _____CaO + _____CO_2

6. _____$KClO_3$ → _____KCl + _____ O_2

7. _____ Al + _____FeO → _____ Al_2O_3 + _____Fe

8. _____$CaCl_2$ + _____$AgNO_3$ → _____AgCl + _____$Ca(NO_3)_2$

9. _____ Na + _____Al_2O_3 → _____Na_2O + _____Al

10. _____Na + _____ O_2 → _____ Na_2O

11. _____K + _____Cl_2 → _____KCl

12 _____H_2O_2 → _____ H_2O + ____O_2

13. _____CH_4 + _____O_2 →_____CO_2 + _____H_2O

14. _____$AgNO_3$ + _____$AlCl_3$ → _____AgCl + _____$Al(NO_3)_3$

15. _____$(NH_4)_3PO_4$ + ____$Mg(NO_3)_2$ →____$Mg_3(PO_4)_2$ + ____NH_4NO_3

Write balanced equations for the following reactions.

Copper (II) carbonate forming copper (II) oxide and carbon dioxide.

Nickel (I) oxide reacting with sulfuric acid (H_2SO_4) to form nickel (I) sulfate and water.

Iron and bromine reacting to give iron(III) bromide.

Lead (IV) oxide and carbon monoxide forming lead metal and carbon dioxide.

Iron (II) chloride reacting with chlorine to form iron (III) chloride.

Ethanol (C_2H_5OH) burning in air (O_2) to form carbon dioxide and water.

Silver reacting with nitric acid (HNO_3) to form silver nitrate and nitrogen dioxide and water.

Manganese (IV) oxide reacting with hydrochloric acid (HCl) to form manganese (II) chloride and chlorine and water.

Sulphur dioxide reacting with hydrogen sulfide (H_2S) to form sulphur and water.

Ammonia (NH_3) reacting with oxygen to form nitrogen monoxide and water.

Reactions in your Life!

Complete each reaction and record your observations. Read about how the reaction works, and then write the balanced equation.

1.	**Rub a small amount of Icy Hot on your arm or leg.**
a.	Observations:

How it works: When applied to the body, Icy Hot quickly changes the temperature felt on the skin from cold to hot. This is done by the chemical reaction between the skin and two active ingredients (Methyl Salicylate and Menthol). Pain is normally sent from the muscles through surrounding nerve receptors and nerve fibers. The pain signal is then sent up these fibers, through the nervous system and into the brain. Once the pain signal reaches the brain, the brain responds by creating the sensation of pain. Sensations of hot and cold also travel through nerve fibers to the brain for response. Icy Hot works by blocking the pain signals by introducing new sensation signals in response to the change in temperature on the skin. This prevents many of the pain signals from reaching the brain, relieving some of the pain being felt after application of the Icy Hot.

Active Ingredient Reaction: Salicylic acid ($C_7H_6O_3$) reacts with methanol (CH_3OH) to produce methyl salicylate ($C_8H_8O_3$) and water.

b.	Balanced equation:
2.	**Add about 75-100ml of warm water into a flask. Next, add about a teaspoon of yeast and teaspoon of sugar. Place a balloon over the mouth of the flask.**
a.	Observations:

How it works: Baker's yeast is of the species *Saccharomyces cerevisiae*. This is a single-celled microorganism that is also found on and around the human body. Active dry yeast is a common commercial product that consists of live yeast cells encapsulated in a thick jacket of dry, dead cells with some growth medium. Under most conditions, active dry yeast must first be rehydrated. It can be stored at room temperature for a year, or frozen for more than a decade. You activated the yeast by adding it to warm water. Once activated, the yeast began to metabolize the sucrose you fed it.

Reaction produced by yeast: sucrose ($C_{12}H_{22}O_{11}$) reacts with oxygen to produce carbon dioxide and water.

b.	Balanced equation:

3.	**Light a candle.**
a.	Observations:

How it works: Candles really are an amazing lighting system -- the fuel itself is the package. There are two parts that work together in a candle:

- The fuel, made of some sort of wax
- The wick, made of some sort of absorbent twine

The **wick** needs to be naturally **absorbent**, like a towel, or it needs to have a strong capillary action. This absorbency is important in a candle because the wick needs to absorb liquid wax and move it upward while the candle is burning. **Paraffin wax** is a heavy hydrocarbon that comes from crude oil. When you light a candle, you melt the wax in and near the wick. The wick absorbs the liquid wax and pulls it upward. The heat of the flame vaporizes the wax, and it is the **wax vapor** that burns.

Combustion Reaction: Paraffin wax ($C_{20}H_{42}$) reacts with oxygen to produce water and carbon dioxide.

b.	Balanced equation:
4.	**Strike a match.**
a.	Observations:

How it works: Phosphorus is the star of the show when it comes to igniting a match. Phosphorus is unstable, sensitive and volatile. Modern matches use an element called red phosphorus. The friction that results from striking against powdered glass generates enough heat to convert some of the red phosphorus to white phosphorus which then reacts with oxygen. To sustain itself, the flame needs something to burn and more oxygen than it can get from the air. The fuel, sulfur and the oxidizing agent, potassium chlorate, work together in their capacities. The heat generated from the phosphorus is enough to break down potassium chlorate, and in the process, it releases oxygen. This oxygen combines with sulfur, allowing the flame to thrive long enough for us to light a candle or barbecue. And in case you're wondering, the matchstick isn't a natural redhead: Red dye is added to the tip to lend a dash of drama to the ensemble.

General Reaction: tetraphosphorus reacts with oxygen to produce tetraphosphorus hexaoxide.

	Balanced equation:

5.	**Add one Alka-Seltzer tablet into a beaker of water.**
a.	Observations:

How it works: If you look at the ingredients for Alka-Seltzer, you will find that it contains **citric acid** and **sodium bicarbonate** (baking soda). When you drop the tablet in water, the acid and the baking soda react -- this produces the fizz. You can think of an Alka-Seltzer tablet as compressed baking powder with a little aspirin mixed in. Because Alka-Seltzer is already in solution when you take it, it is absorbed quickly. This means you get rapid relief from your headache and heartburn, acid indigestion or sour stomach.

Antacid Reaction: Citric acid ($C_6H_8O_7$) reacts with sodium bicarbonate ($NaHCO_3$) to produce water, carbon dioxide, and sodium citrate ($Na_3C_6H_5O_7$).

b.	Balanced equation:
6.	**Use a pipette to add a small amount of methanol onto a watch glass. Using a wood splint, light the methanol on fire.**
a.	Observations:

How it works: Methanol is used in transportation in 3 main ways - directly as fuel or blended with gasoline, converted in dimethyl ether (DME) to be used as a diesel replacement, or as a part of the biodiesel production process. Methanol is a fatal poison. Small internal doses, continued inhalation of the vapor, or prolonged exposure of the skin to the liquid may cause blindness. As a result, commercial use of methanol has sometimes been prohibited. Methanol is used as a solvent for varnishes and lacquers, as an antifreeze, and as a gasoline extender in the production of gasohol.

Combustion Reaction: Methanol (CH_3OH) reacts with oxygen to produce carbon dioxide and water.

b.	Balanced equation:

7.	**Place a Tums tablet on the bottom of a beaker. Using a pipette, add several drops of HCl directly onto the Tums tablet.**
a.	Observations:

How it works: Tums assorted fruit antacid tablets contain the active ingredient calcium carbonate, which is a type of medicine called an antacid. Calcium carbonate works by binding to excess acid produced by the stomach. This neutralizes the acid and decreases the acidity of the stomach contents. Antacids relieve the pain and discomfort of indigestion, heartburn and other symptoms related to excess stomach acid, by reducing the acidity of the stomach contents. They work as soon as they reach the stomach

Antacid Reaction: Calcium carbonate reacts with hydrochloric acid (HCl) to produce calcium chloride, carbon dioxide, and water.

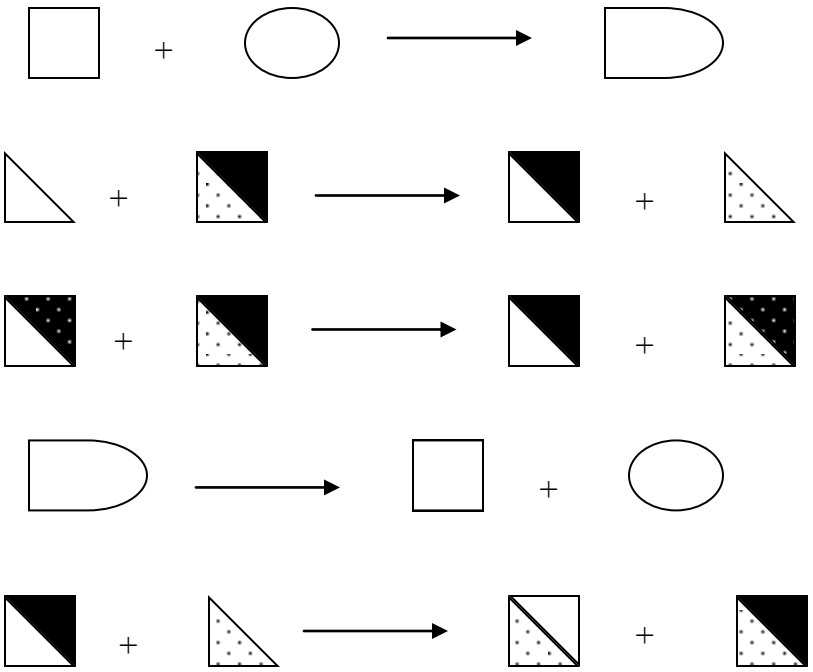
b.	Balanced equation:

Name _____ Per ____

Types of Reactions

Match the following shapes in set 1 with a type of reaction in set 2.

Set 1:



Set 2:

Decomposition: $MgO \rightarrow Mg + O_2$

Combustion: $CH_4 + O_2 \rightarrow CO_2 + H_2O$

Single Replacement: $Al + Na_2CO_3 \rightarrow Al_2(CO_3)_3 + Na$

Double Replacement: $KCl + Ca(OH)_2 \rightarrow KOH + CaCl_2$

Synthesis: $Li + S \rightarrow Li_2S$

Questions:

1. Balance the reactions above using coefficients (large numbers that go in front of a chemical formula) to ensure that there are the number of each element on both sides of the reaction.

   For example - **2** $MgO \rightarrow$ **2** $Mg + O_2$

   The two coefficients in the reaction above make it so that there are two magnesiums on both sides of the arrow and two oxygens on both sides of the arrow.

2. For each of the five types of reactions, list some unique attributes of each.

Decomposition:

Combustion:

Single Replacement:

Double Replacement:

Synthesis:

3. Label each of the reactions below with one of the five types of reactions. Balance

__ C_3H_8 + __ O_2 → __ CO_2 + __ H_2O _____

__ Al + __ Fe_3N_2 → __ AlN + __ Fe _____

__ Na + __ Cl_2 → __ NaCl _____

__ H_2O_2 → __ H_2O + __ O_2 _____

__ $C_6H_{12}O_6$ + __ O_2 → __ H_2O + __ CO_2 _____

__ $NaClO_3$ → __ NaCl + __ O_2 _____

__ $(NH_4)_3PO_4$ + __ $Pb(NO_3)_4$ → __ $Pb_3(PO_4)_4$ + __ NH_4NO_3 _____

__ BF_3 + __ Li_2SO_3 → __ $B_2(SO_3)_3$ + __ LiF _____

Name _____ Per _____

CLASSIFYING EQUATIONS

Balance the following reactions and state what type of reaction is occurring.

1. _____ ___ C_3H_8 + ___ O_2 → ___ CO_2 + ___ H_2O

2. _____ ___ Al + ___ Fe_3N_2 → ___ AlN + ___ Fe

3. _____ ___ Na + ___ Cl_2 → ___ NaCl

4. _____ ___ H_2O_2 → ___ H_2O + ___ O_2

5. _____ ___ $C_6H_{12}O_6$ + ___ O_2 → ___ H_2O + ___ CO_2

6. _____ ___ H_2O + ___ CO_2 → ___ C_7H_8 + ___ O_2

7. _____ ___ $NaClO_3$ → ___ NaCl + ___ O_2

8. _____ ___ $(NH_4)_3PO_4$ + ___ $Pb(NO_3)_4$ → ___ $Pb_3(PO_4)_4$ + ___ NH_4NO_3

9. _____ ___ BF_3 + ___ Li_2SO_3 → ___ $B_2(SO_3)_3$ + ___ LiF

10. _____ ___ C_7H_{17} + ___ O_2 → ___ CO_2 + ___ H_2O

11. _____ ___ $CaCO_3$ + ___ H_3PO_4 → ___ $Ca_3(PO_4)_2$ + ___ H_2CO_3

12. _____ ___ Ag_2S → ___ Ag + ___ S_8

13. _____ ___ KBr + ___ $Fe(OH)_3$ → ___ KOH + ___ $FeBr_3$

14. _____ ___ KNO_3 + ___ H_2CO_3 → ___ K_2CO_3 + ___ HNO_3

15. _____ ___ $Pb(OH)_4$ + ___ Cu_2O → ___ PbO_2 + ___ CuOH

16. _____ ___ $Cr(NO_2)_2$ + ___ $(NH_4)_2SO_4$ → ___ $CrSO_4$ + ___ NH_4NO_2

17. _____ ___ KOH + __ $Co_3(PO_4)_2$ → ___ K_3PO_4 + ___ $Co(OH)_2$

18. _____ ___ $Sn(NO_2)_4$ + ___ Pt_3N_4 → ___ Sn_3N_4 + ___ $Pt(NO_2)_4$

19. _____ ___ B_2Br_6 + ___ HNO_3 → ___ $B(NO_3)_3$ + ___ HBr

20. _____ ___ ZnS + ___ AlP → ___ Zn_3P_2 + ___ Al_2S_3

Questions 21-30: For each of the following
a) classify the reaction
b) write chemical equation

21. ethyl alcohol (CH_3CH_2OH) burns in air _____

22. sodium reacts with aluminum carbonate _____

23. zinc reacts with oxygen _____

24. sodium sulfide reacts with barium nitrate _____

25. copper (II) carbonate reacts to form copper (II) oxide and a gas _____

26. potassium reacts with iodine _____

27. ethane (C_2H_6) burns in air _____

28. barium reacts with zinc (II) chloride _____

29. sodium carbonate reacts with zinc (II) fluoride _____

30. lead (II) oxide decomposes _____

Predicting Products Practice

For each of the following:

a) classify the reaction
b) write the equation with correct formulas
c) balance the equation

a. A strip of calcium metal is combined with oxygen in a Bunsen burner flame. Classification:_____

b. A piece of sodium metal is placed in an aqueous solution of lead (II) oxide. Classification:_____

c. Aqueous lithium nitrate solution is combined with aqueous aluminum sulfate. Classification: _____

d. Glucose ($C_6H_{12}O_6$) is heated (combined with oxygen) in a flame. Classification:_____

e. An aqueous solution of barium hydroxide is mixed with aqueous iron (III) sulfate. Classification:_____

f. Magnesium metal is burned in nitrogen gas. Classification:_____

g. Ethanol (C_2H_5OH) is burned completely in air. Classification:_____

h. Solid sodium nitride is heated to decomposition. Classification:_____

Classifying Reactions

Procedure: Wear Goggles

1. Magnesium metal and hydrochloric acid: (CAUTION: acid is corrosive) Fill a medium sized test tube with about 1 inch of hydrochloric acid. Place the test tube in a test tube rack and add a small piece of Mg. Once the Mg is added, cover the test tube with another test tube to trap the gas. Collect gas for 5-10 seconds, then hold a burning splint at the mouth of the test tube to help identify the gas formed. Record any observations.
2. Put small squirt of silver nitrate in a test tube and add small squirt of sodium chloride. Record observations.
3. Al and O: Hold a piece of aluminum foil in flame briefly. Record observations. Ethanol reaction: place 5 drops of ethanol (C_2H_5OH) on a watch glass. Using a wood splint, light the ethanol on fire. Record observations.
4. Small scoop of $CuCO_3$ into test tube, and heat it. Stick a flaming splint into the test tube.

Data Table:

Reaction	Observation
Mg and HCl	
$AgNO_3$ and NaCl	
Al and O	
C_2H_5OH and O_2	
$CuCO_3$ and heat	

Conclusion:

Write a complete reaction, including:

Formulas, states-o-matter, classify, balance

Conservation of Matter Review

Balancing

__ C_7H_{16} + ___ O_2 → ___ CO_2 + ___ H_2O

___ Na_2O_2 + ___ H_2O → ___ NaOH + ___ O_2

___ SiO_2 + ___ HF → ___ SiF_4 + ___ H_2O

___ C + ___ H_2O → ___ CO + ___ H_2

___ $KClO_3$ → ___ KCl + ___ O_2

___ H_3AsO_4 → ___ As_2O_5 + ___ H_2O

___ $KClO_3$ → ___ $KClO_4$ + ___ KCl

___ $Al_2(SO_4)_3$ + ___ $Ca(OH)_2$ →___ $Al(OH)_3$ + ___ $CaSO_4$

___ P_4O_{10} + ___ H_2O → ___ H_3PO_4

___ $FeCl_3$ + ___ NH_4OH → ___ $Fe(OH)_3$ + ___ NH_4Cl

___ Sb + ___ O_2 → ___ Sb_4O_6

___ $Ca_3(PO4)_2$ + ___ SiO_2 → ___ P_4O_{10} + ___ $CaSiO_3$

___ C_3H_8 + ___ O_2 → ___ CO_2 + ___ H_2O

___ N_2O_5 + ___ H_2O → ___ HNO_3

___ Fe_2O_3 + ___ CO → ___ Fe + ___ CO_2

___ Al + ___ HCl → ___ $AlCl_3$ + ____ H_2

For each of the following:
a) Write correct formulas b) Balance the equation using coefficients c) Classify the reaction

sodium reacts with barium chloride to form barium and sodium chloride.

isopropyl alcohol (C_3H_8O) reacts with oxygen to form water and carbon dioxide

silver (I) nitrate reacts with magnesium chloride to form silver chloride and magnesium nitrate.

calcium carbonate reacts when heated to form calcium oxide and carbon dioxide.

hydrogen reacts with oxygen to form water

For each of the following:
a) classify b) Write correct formulas including states of matter c) balance the equation

sodium reacts with aqueous aluminum carbonate

potassium reacts with oxygen

aqueous sodium sulfide reacts with aqueous lead (II) nitrate

liquid ethane (C_2H_6) burns in air

Solid lead (II) oxide reacts

Ionic compounds

Cation	Anion	Formula	Name
Na^+	**Cl^-**		
		K_3P	
Al	**S**		
Ca	**F**		
			Barium iodide
Mg	**N^{3-}**		

			Iron(II) oxide
			Iron(III) oxide
		CuO	
		CuF	
			sodium fluoride
K^+	**NO_3^-**		
Mg^{2+}	**NO_3^-**		
Na^+	**SO_4^{2-}**		
NH_4^+	**PO_4^{3-}**		
		Cu_2SO_4	
		$(NH_4)_2S$	
			Aluminum sulfate
			Sodium carbonate
			Iron (III) hydroxide
			Silver nitrate
			Zinc hydroxide

Covalent Compounds

Name	Formula
	SiO_3
	SO_3
	Cl_2O_5
Carbon tetrachloride	
Trinitrogen pentasulfide	
	NO_2
	AsO_3
	P_2O_5
Tricarbon octachloride	
Dinitrogen hexasulfide	
	NO_4
Tetraxenon heptaflouride	

Name _____ Per _____

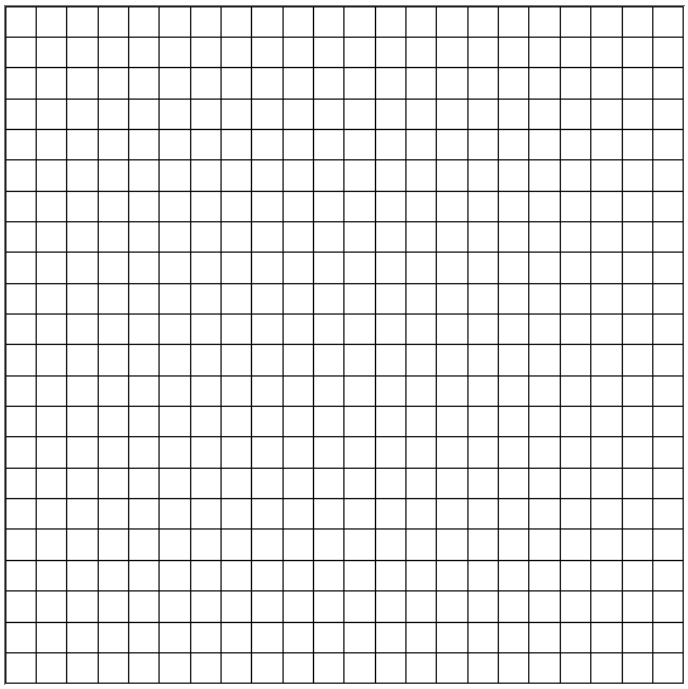
GRAPHING PERIODIC TRENDS

Purpose: Observe the patterns in periodic trends.

Graphing:

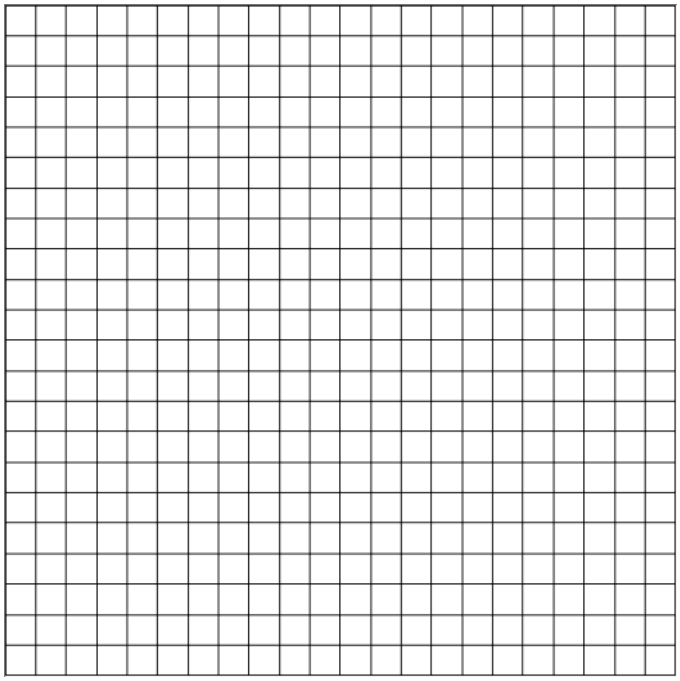
Use the table below to graph atomic size (radius). Label your x-axis with atomic number and the y-axis with atomic size.

Element	Atomic Size (A)	Element	Atomic Size (A)
Hydrogen	0.79	Neon	0.51
Helium	0.49	Sodium	2.23
Lithium	2.05	Magnesium	1.72
Beryllium	1.40	Aluminum	1.62
Boron	1.17	Silicon	1.44
Carbon	0.91	Phosphorus	1.23
Nitrogen	0.75	Sulfur	1.09
Oxygen	0.65	Chlorine	0.97
Fluorine	0.57	Argon	0.88



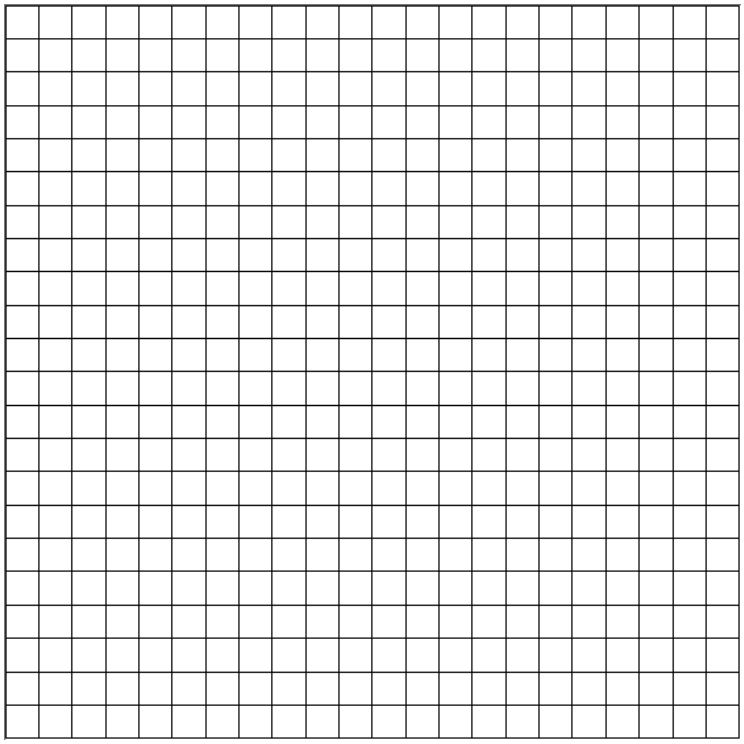
Use the table below to graph ionization energy (energy required to remove an electron). Label your x-axis with atomic number and the y-axis with ionization energy.

Element	Ionization Energy (kJ/mol)	Element	Ionization Energy (kJ/mol)
Hydrogen	1360	Neon	2156
Helium	2459	Sodium	514
Lithium	539	Magnesium	765
Beryllium	932	Aluminum	599
Boron	530	Silicon	815
Carbon	1126	Phosphorus	1049
Nitrogen	1453	Sulfur	1036
Oxygen	1362	Chlorine	1297
Fluorine	1742	Argon	1576



Use the table below to graph electronegativity (pull of an atom on an electron in a neighboring atom). Label your x -axis with atomic number and the y-axis with electronegativity.

Element	Electro-negativity	Element	Electro-negativity
Hydrogen	2.1	Neon	-
Helium	-	Sodium	0.9
Lithium	1.0	Magnesium	1.2
Beryllium	1.5	Aluminum	1.5
Boron	2.0	Silicon	1.8
Carbon	2.5	Phosphorus	2.1
Nitrogen	3.3	Sulfur	2.5
Oxygen	3.5	Chlorine	3.0
Fluorine	4.0	Argon	-



Questions:

1. Predict the values for the atomic radius, first ionization energy, and electronegativity of potassium (element 19).

2. Which group of elements appears to occupy the major peaks for atomic radius? _____

3. Which group of elements appears to occupy the major peaks for ionization energy? Why might that be?

4. Why wouldn't noble gases have electronegativity?

5. Do metals or non-metals have the highest electronegativities? Why might that be?

6. What might cause the decrease in ionization energy for elements in group 3A compared to those in 2A? (Hint: think about electron configuration)

7. Compare and contrast the trends for atomic radius and ionization energy.

8. Compare and contrast the trends for atomic radius and electronegativity.

9. Compare and contrast the trends for ionization energy and electronegativity.

Name _____ Per _____

PERIODIC TRENDS

1. A group is a _____ of elements. The group number tells you how many _____ _____ are in the atom. Elements in the same group (family) have very similar chemical properties.
2. A period is a _____ of elements. Elements in the same period all have the same number of _____ _____ in their atoms.
3. _____ is the 5th period alkaline earth metal
4. _____ is the 6th period noble gas
5. _____ is the 4th period alkali metal
6. _____ is the 3rd period halogen
7. _____ is the most reactive metal in the 4th period
8. _____ is the most reactive nonmetal in the 4th period
9. _____ has the largest radius of the 6th period
10. _____ is the largest atom with two valence electrons
11. _____ is the most reactive element with 6 valence electrons
12. _____ is the most reactive halogen
13. _____ is the least reactive alkaline earth metal
14. _____ is the most reactive element with 1 valence electron
15. _____ has 10 electrons and 13 protons
16. _____ has 18 electrons and 17 protons
17. _____ has 2 valence electrons in the 3rd energy level
18. _____ has 3 valence electrons in the 6th energy level
19. _____ has the greatest ionization energy of the alkaline earth metals
20. _____ is the element in the boron family with 4 energy levels
21. _____ is the element in oxygen family with 3 energy levels
22. _____ is the 3rd period element with lowest ionization energy
23. _____ is the noble gas with a 2nd energy level valence shell
24. _____ is the 5th period atom most likely to lose an electron
25. _____ is the 5th period atom least likely to react
26. _____For elements 39 through 48, additional electrons are added to the _____ sublevel
27. _____Which element has the greatest radius: S Cl Se Br
28. _____ is the third period element which is most likely to form an ion with a +3 charge
29. _____ is the second period element most likely to form an ion with a –1 charge
30. _____The most reactive metal on the periodic table
31. The most reactive nonmetal is _____

32. Electrons in the outer energy level of an atom are called _____ electrons

33. _____ is the third period element which is most likely to form an ion with –2 charge

34. _____is the third period metal that forms an ion with +2 charge.

35. The Noble gases, except He, have ___ valence electrons.

36. _____ react by gaining electrons while _____ react by losing electrons.

37. Which has the largest radius? Rb Sr K Ca

38. Which has the greatest reactivity? Rb Sr K Ca

39. Which is the most reactive? Se Br S Cl

40. Which has the highest electronegativity? Se Br S Cl

41. Which has the lowest electronegativity? Rb Sr K Ca

43. Which has the larger atomic radius, K or Zn? Explain.

44. Which has the larger atomic radius, Rb or Pt? Explain.

45. Which has a higher ionization energy, Li or Cs? Explain.

46. Which has a higher electronegativity, Na or N? Explain.

47. Based on periodic trends, which would you expect to be more reactive, Rb or K? O or F? Explain each.

Name _____ Per _____

Lewis Electron Dot Models

Purpose:

Use electron dot structures to demonstrate the sharing of electron pairs between non-metals to form covalent bonds.

Procedure:

For each molecule in the Data Table:
Count the total number of valence electrons available.
For cations: subtract one electron for each unit of positive charge
For anions: add one electron for each unit of negative charge
Construct a skeleton structure. Join atoms so that each atom shares **one pair** of valence electrons (drawn as a line).
Central atom tends to be carbon, or else the atom with the fewest valence electrons
Hydrogen and halogens will **NEVER** be a central atom
3. From the total number of valence electrons, subtract the number of electrons used to make the skeleton structure to determine how many electrons remain to be distributed.
Follow the octet rule to distribute **all** remaining electrons.
Hydrogen only need two electrons (one pair)
If there are too few valence electrons to give each atom an octet, a multiple bond is probably required.
Double bond (=) is 2 pairs of electrons, 4 total electrons
Triple bond (º) is 3 pair of electrons, 6 total electrons
Always double check that
each atom has a full octet (8 electrons) except hydrogen (2 electrons)
total electrons as dots and bonds equal total valence electrons

Example:

Draw the Lewis dot structure for F_2 –

Step 1: each fluorine contains 7 valence electrons, 7 + 7 = 14 total valence electrons

Step 2: skeleton structure $F \overset{x}{\underset{x}{}} F$ which is better written as F – F

Step 3: 14 total electrons – 2 used in bond = 12 electrons left

Step 4: Lewis dot structure $\overset{x}{\underset{x}{}} \overset{x\,x}{\underset{x\,x}{F}} - \overset{x\,x}{\underset{x\,x}{F}} \overset{x}{\underset{x}{}}$

Step 5: Not necessary

Step 6: Used 14 electrons, 14 electrons total

Formula	Name	Valence Electrons for each atom	Total number of v. e.	Lewis Dot Structure
H_2	Hydrogen	1 + 1	2	H – H
Cl_2				
HCl				
CH_4				
CH_3F	-----			
NCl_3				
H_2O				
H_2O_2				*Hint – two center atoms
O_2				
CO				
SF_2				
PH_3				
OH^-				
NH_4^+				

Name _____ Per _____

LEWIS STRUCTURES

1. Label each element with its number of **valence electrons**, and draw a **Lewis structure** of the element.

   a. Na

   b. P

   c. B

   d. F^-

   e. Cu

   f. Ar

   g. Zn

   h. Eu

   i. Pb

   j. Ca^{2+}

   k. Br

   l. Cl^-

2. Draw the Lewis dot structures for the following compounds. Denote the total number of valence electrons somewhere near the diagram.

   a. H_2S

   b. CF_4

   c. N_2

   d. CO_3^{2-}

VSEPR GEOMETRY CHART

Number of "Groups"	Number of		Lewis Structure	3-D Drawing	Angle(s)	Geometry (shape)	Examples
	At-oms	Lone Pairs					
n/a							
2							
3							
4							

Name _____ Per _____

THREE DIMENSIONAL MODELS OF COVALENT MOLECULES

Purpose:

Build three-dimensional models of some simple covalent molecules. Predict their shapes, bond angles, and polarities from knowledge of bonds and molecule polarity rules.

Background:

A single covalent bond is formed when two atoms share a pair of electrons. Each atom provides one of the electrons to the pair. If the two atoms are alike, the bond is said to be *nonpolar covalent*. If the two atoms are unlike, one exerts a greater attractive force on the electrons, and the bond is *polar covalent*. More than one pair of electrons can be shared. This results in a double or triple bond.

A group of atoms held together by covalent bonds is called a molecule. Molecules can be either polar or nonpolar. If bonds are nonpolar, the molecule is nonpolar. If bonds are polar, molecules can still be nonpolar if the distribution through the molecule is symmetrical. A molecule's symmetry depends on its shape, that is, the positions in space of the atoms making up the molecule. Some possible shapes are liner, bent, pyramidal and tetrahedral.

Procedure:

1. Obtain a molecular model building set. The set should contain the following components:

Element	Color	Number
Hydrogen		
Carbon		
Nitrogen		
Oxygen/Sulfur		
Fluorine/Chlorine		

2. Observe that the following atoms have one hole (also known as a bonding site): hydrogen, fluorine, and chlorine. The atoms with two holes are oxygen and sulfur. A nitrogen atom has three holes and a carbon atom has four.

3. Draw the Lewis dot structure for each molecule listed in the data table.

4. From the Lewis structure, construct a 3-D model. (Steps 3 and 4 can be done in reverse order if you are having problems drawing the Lewis structure.

5. Verify that your Lewis structure and model correlate.

6. Draw a 3-D depiction of the structure with proper bond angles in the data table.

7. Using the 3-D structure, predict the shape, bond angles, and polarity of the molecule.

8. Answer the questions on the last page of the lab.

Data Table:

Name/ Formula	Lewis Structure (include VE#)	3-D Drawing	VSEPR Geometry	Bond Angle	Polar or Non-Polar	Reso-nance (Yes/No)
Hydrogen H_2						
Water H_2O						
Methane CH_4						
Chlorine Cl_2						
Ammonia NH_3						
Hydrogen Fluoride HF						
Ethyne C_2H_2						
Dichloro-methane CH_2Cl_2						

Name/ Formula	Lewis Structure (include VE#)	3-D Drawing	VSEPR Geometry/ Geometries	Bond Angle	Polar or Non-Polar	Reso-nance (Yes/No)
Nitrogen N_2						
Carbon Dioxide CO_2						
Methanol CH_3OH						
Hydrogen Peroxide H_2O_2						
Oxygen O_2						
Hydrogen Sulfide H_2S						
Sulfur Di-oxide SO_2						
Sulfate Ion SO_4^{2-}						
Formalde-hyde CH_2O						

Questions:

1. What factors (more than one) determine the geometry of a molecule?

2. What factors (more than one) determine the polarity of a molecule?

3. What must be present in the structure of a molecule in order for a resonance structure to be possible?

4. What will be the geometry of carbon tetrachloride? Is this molecule polar or nonpolar?

5. To have best answered question #4 or any question like it, what should you always do first?

6. Why does a carbon "atom" contain four holes while a nitrogen "atom" contain only three?

7. What does VSEPR stand for? Define the theory in your own words.

8. Using your answer to #7, explain why water (H_2O) has an H-O-H bond angle of 105 while a hydronium ion (H_3O^+) has an H-O-H bond angle of 107.

Name _____ Per _____

LEWIS STRUCTURES AND VSEPR PRACTICE

Compound	Lewis Dot Structure	Geometry	Drawing	Polarity Y or N
F_2				
BI_3				
CCl_4				
H_2O				
CO_2				
CO				
NH_3				
SO_4^{2-}				

Compound	Lewis Dot Structure	Geometry	Drawing	Polarity
				Y or N
AsF_3				
SO_3				
SO_2				
C_2H_2 (2)				
C_2H_5OH (3)				
CH_3COOH (3)				
CH_3COCH_3 (3)				

States of Matter Simulation

Log on to: http://phet.colorado.edu/simulations/index.php?cat=Chemistry

Click on States of Matter. Open simulation

Part One: Solid, Liquid, Gas

1. Click on a molecule of choice. _____

2. Look at the chamber: for each phase draw and describe what is going on.

Solid	Liquid	Gas
_____	_____	_____
_____	_____	_____
_____	_____	_____

3. When the molecule is in a solid state, add heat. Describe what occurs.

   _____

   _____

4. What is kinetic energy?

   _____

5. At what temperatures does the solid start to behave like a liquid, and a liquid starts to behave like a gas.

   _____

6. Remove the heat. Describe what occurs.

   _____

   _____

7. Is there a temperature in which the molecules stop vibrating? Why or why not?

   _____

   _____

8. Does any of these observations change with the molecule you are working with? Why?

   _____

   _____

Part Two: Phase Changes

1. Choose a molecule to work with. _____

2. Note the temperature and pressure prior to starting the simulation.

   _____

3. There are three variables is which we can work with, what are they?

   _____

4. Working with one variable at a time, what happens with the molecules?

   Variable 1:

   _____

   _____

   Variable 2:

   _____

   _____

   Variable 3:

   _____

   _____

5. Reset the program, each trial. Work with two variables at a time, what happens?

   Variables: _____

   _____

   Variables:_____

   _____

   Variables:_____

   _____

6. Reset the program. Work with all three variables at a time, what happens with the molecules?

   _____

   _____

   _____

7. The colorful graph on the side, what information does that graph tell us?

   _____

   _____

Name _____ Per _____

TYPES OF SOLIDS AND THEIR INTER- AND INTRA-MOLECULAR FORCES

Compound	Description of Bond	Functional Unit (drawing)	Elements Combined	Electronegativity Differences	Properties
Ionic					
Non-polar and polar Covalent					
Metal					
Network					

Name _____ Per _____

PROPERTIES OF VARIOUS TYPES OF COMPOUNDS

1. List the three types of intermolecular forces found between covalent molecules in order of weakest to strongest. Describe each briefly.

2. Explain why covalently bonded substances are poor conductors of heat and electricity.

3. Why do covalent solids have low melting points?

4. What are London forces?

5. What are the structural features of substances that only have London forces between the molecules? (i.e. what would you see in their Lewis structures?)

6. In which of the following compounds would hydrogen bonding occur? Explain your answer and remember that Lewis structures are a great place to start.
   a) $COCl_2$ b) PH_3 c) H_2CO d) CH_3OH

7. In which of the following substances would there be the strongest forces between the molecules? Explain your answer.
   a) SiH_4 b) $H_2C{=}O$ c) $CH_3\text{-}CH_3$ d) O_2

8. Explain why at room temperature chlorine is a gas, bromine is a liquid, and iodine is a solid. Be sure to first assign each an intermolecular force.

9. Explain why water (H_2O) is liquid at room temperature by hydrogen sulfide (H_2S) is a gas. Be sure to first assign each an intermolecular force.

10. What are three key properties of solids containing ionic bonds?

11. Describe the bonding that occurs in a metallic solid.

12. Use the commonly accepted model of metallic bonding to explain why:
    -metals are malleable
    -metals conduct electricity in the solid state

13. How do network solids differ from covalent crystals in structure? in hardness? in melting point?

14. Classify these substances as ionic, covalent, network, or metallic solids:
    a. A structure composed of atoms covalently bonded to neighboring atoms
    b. A solid only at extremely low temperatures
    c. A good conductor of heat and electricity
    d. A good electric conductor only in solution

15. Germanium (Ge) is a solid whose atoms are all covalently bonded to each other, much like those of carbon atoms in a diamond crystal. Glycerol [$C_3H_6(OH)_3$] is an alcohol. Potassium chloride is a white crystalline solid. Methane is a gas which can only be liquefied under high pressures and low temperatures. Rubidium is very malleable and is an excellent conductor of electricity. Which has:

    a. hydrogen bonding?

    b. the greatest hardness in the solid phase?

    c. the highest melting point?

    d. the lowest boiling point?

    e. conductivity of electricity when in solution?

    f. particles held together primarily by van der Waals forces?

16. When ethanol (C_2H_5OH) boils, the gas consists of
    a. a mixture of carbon dioxide and water.
    b. carbon, hydrogen, and oxygen gases.
    c. water and ethanol.
    d. ethanol only.

17. Elements A,B, C, and D have consecutive atomic numbers. Element D is a monatomic gas with low melting and boiling points. All efforts to form compounds of D in the laboratory have failed.

a) Which of the remaining elements, A, B, or C has the strongest affinity for an additional electron?

b) A compound of an alkali metal M with element C has a formula of MC. Does this compound have ionic or covalent bonds? Predict other properties of MC such as melting point and solubility in water.

c) Write the formulas for hydrides of elements A, B, and C.

d) Predict the shape of each hydride molecule from above.

e) Predict the conductivity of solid B.

18. What intermolecular forces are present between the molecules or atoms of the following substances?

a. Ne b. O_2 c. Br_2

d. P_4 e. HI

19. For each of the following liquids, list the type of intermolecular forces you would expect to find.

a. water, H_2O b. bromine, Br_2 c. carbon tetrachloride, CCl_4

20. Explain the reasons for the difference in boiling points between

a. HF (20 C) and HCl (-85 C) b. HCl (-85 C) and (LiCl (1360 C)

21. Match each of the solids in the first column with two properties in the second column. Try to use each property at least once.

a. metallic solid
b. covalent network solid
c. ionic solid
d. covalent solid

I. low melting point
II. high melting point
III. conducts electricity in solution
IV. brittle
V. hard
VI. malleable

Intermolecular Forces Design Lab

Overall Expectation:

- Determine what characteristics of covalent molecules effect how well the molecules stay together
- Could be measured by:
  - ◊ Surface tension
  - ◊ Evaporation
  - ◊ Vaporization
  - ◊ Viscosity
  - ◊ Different temperatures
  - ◊ Or a variety of other methods

Available Chemicals:

- Water (H_2O)
- Methanol (CH_3OH)
- Ethanol (C_2H_5OH)
- Iso-propyl alcohol (C_3H_7OH)
- n-propyl alcohol (C_3H_7OH)
- Butyl alcohol (C_4H_9OH)
- Pentanol ($C_5H_{11}OH$)
- Hexane (C_6H_{14})
- Acetone (CH_3COCH_3)
- Ethylene Glycol ($HOCH_2CH_2OH$)
- Glycerine ($HOCH_2CHOHCH_2OH$)
- Hydrogen Peroxide (H_2O_2)

- Your investigation should include three to five data points (whether three to five different chemicals being measured or one chemical being measured three different ways).

Design—NO PERSONAL STATEMENTS IN THE LAB

Objective/research question: Be sure to use your dependent and independent variable,

Variables:

Independent Variable:

Identify the independent variable that will be manipulated. Describe how it will be manipulated. You need to identify a trend in your choice that can be analyzed. There should be NO random choices.

Dependent Variable:

Identify the dependent variable that will be measured in this experiment. Describe how it will be measured. How does it relate back to IMFs?

Controls:

Provide a list of controls you will employ to ensure that the only factor that is altered between successive trials is the independent variable.
Explain why each control is selected by providing a brief description of what effect changing this factor would have on the dependent variable.

Hypothesis:

Formulate a hypothesis addressing your research question. For this investigation, your hypothesis should provide a prediction of how you believe your independent variable will affect the dependent variable. Your hypothesis should be supported by a well developed rationale based on sound chemical principles. (If . . . then . . . because . . .)

Method: The will be different based on what you are testing.

- ♦ Step-by-step method
- ♦ Use command terms (measure, record, pour)
- ♦ Be sure to tell what chemical you are working with and how many trials (IV and DV need to be represented in method)

Data Collection and Processing

Raw Data: design a data table to record your measurements from lab.

- Quantitative data: any measurements from lab
  - Title must describe your data (IV and DV)
- Qualitative data: any observations you made that could help you in the analysis.

Processed Data: these are <u>any</u> calculations you perform with your raw data. Remember what your goal is.

- Processed data must be in a table
  - Title must describe your processed data (IV and DV)
- Must show 1 sample calculation

Presenting Data: Place your processed data into a graph (IV on x-axis, DV on y-axis).

- Line graphs mean both axis's are quantifiable
- Bar graphs means the IV is a category (descriptor)
- Label the axis, include units
- Title must describe your presented data

Conclusion and Evaluation

Conclusion Section:

- A discussion of the data ***explaining*** trends it contains based on your measurement of intermolecular forces
- Draws specific conclusion about what the data (processed data) shows even if it is inconclusive
- Specifically reference the data, even if you got unexpected results. It is still data and part of your conclusion
- What does your data tell you about the intermolecular forces of your compounds

Analysis:

- Identify and describe at least up to three true sources of error. One needs to be method based, one needs to be errors in the way the data was measured (source of errors in the equipment).
  - The sources of error described actually have had an impact on your data.
- For each source of error identified state how you expect it to have affected your results. (Data being skewed is not a proper response. Explain what caused the outliers or bad data)
- For each source of error identified describe a specific, detailed, and realistic improvement to your method.
- Make a judgment based upon your error analysis describing how reliable (repeatable) the results of the experiment were
- Make a judgment based upon your error analysis describing valid (trustworthy) the results of the experiment were
- Describe any other possible experiments you could run as a result of what you have learned in your first experiment. Think about how the environment relates to this lab.

- Error, affect on data, improvement x3

Bonding and States of Matter Review Sheet

1. Why do atoms form chemical bonds?

2. Define the following terms:
   - octet rule
   - covalent bond
   - ionic bond
   - polar covalent bond
   - network covalent
   - metallic bond
   - van der Waals forces
   - dipole forces
   - hydrogen bond

3. What is a Lewis dot structure?

4. Write dot structures for: potassium, arsenic, bromine, silicon, tellurium, aluminum, oxygen, sodium, and radon.

5. Two atoms each provide 3 electrons that are shared by the two atoms. This is an example of a:
   a. single covalent bond b. double covalent bond
   c. triple covalent bond d. quadruple covalent bond

6. Two atoms each provide 1 electron that are shared by the two atoms. This is an example of a:
   a. single covalent bond b. double covalent bond
   c. triple covalent bond d. quadruple covalent bond

7. A certain nonmetal usually forms two covalent bonds in its compounds. How many valence electrons does this element have?

8. Noble gases do not form chemical compounds because:

9. Which of the following compounds contains ionic and covalent bonds?
   a. SiO_2 b. BaF_2 c. Na_2CO_3 d. Cl_2O

10. Complete the table for the following molecular compounds

Formula	Lewis Structure	Molecular geometry	Polar?	Intermolecular forces
SO_2				
SCl_2				
CS_2				
CCl_4				
PCl_3				
O_2				
SO_3				
$COCl_2$				

11. In order to melt solids of each of the following substances, tell what type of bonds or forces of attraction must be disrupted.

    - Zinc
    - methane (CH_4)
    - ammonia (NH_3)
    - potassium chloride

12. Give three characteristics of each of the three states of matter.

13. How do intermolecular forces affect the state of matter of a substance at room temperature?

Write the symbol of the element that best completes the statement or answers the question.

_____ is the 3rd period alkaline earth metal

_____ is the 5th period noble gas

_____ is the most reactive metal in the 5th period

_____ is the most reactive nonmetal in the 2nd period

_____ has the smallest radius of the 3rd period

_____ is the largest atom with 3 valence electrons

_____ is the most reactive element with 7 valence electrons

_____ is the least reactive alkaline earth metal

_____ is the most reactive element in group 6a

_____ has 12 electrons and 12 protons

_____ has an electron configuration that ends in $6s^2$

_____ has an electron configuration that ends in $4d^4$

_____ is the 3rd period element with lowest ionization energy

_____ is the third period element which is most likely to form an ion with a +3 charge

_____ the most reactive nonmetal on the periodic table

_____ has 1 valence electrons in the 1st energy level

_____ is the alkaline earth metal with 3 energy levels

_____ has 5 valence electrons in the 4th energy level

_____ is the 2nd period atom most likely to lose an electron

_____ is the 2nd period atom least likely to react

_____ is the 3rd period element with the highest electronegativity

_____ is the halogen with the highest electronegativity

_____ is the element with the largest atomic radius

_____ is the alkaline earth metal with the highest ionization energy

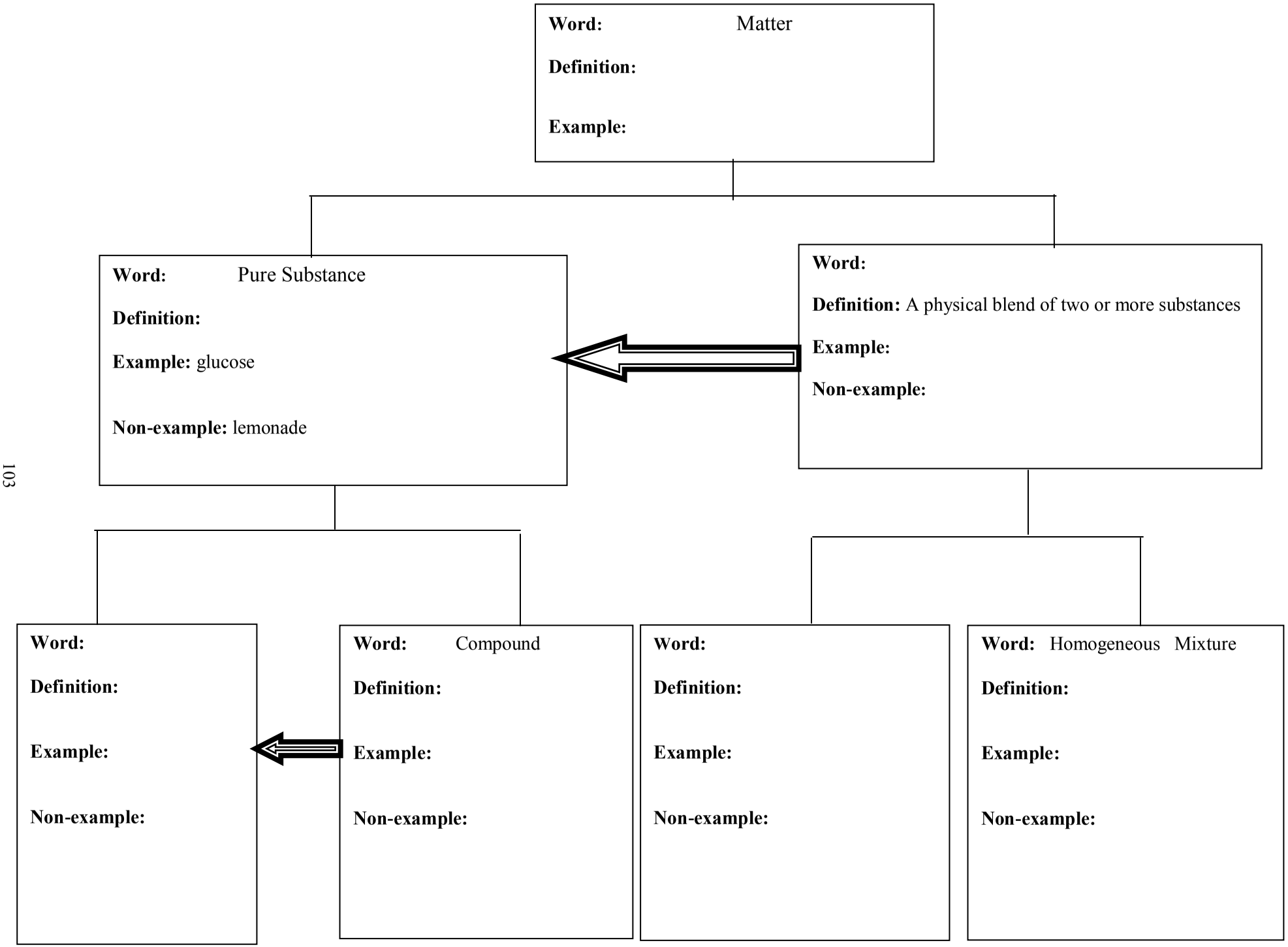
Define and explain the trend for:

- Atomic Radius

- Ionization Energy

- Electronegativity

- Reactivity of metals and nonmetals


Word: Matter
Definition:
Example:
Word: Pure Substance
Definition:
Example: glucose
Non-example: lemonade
Word:
Definition: A physical blend of two or more substances
Example:
Non-example:
Word:
Definition:
Example:
Non-example:
Word: Compound
Definition:
Example:
Non-example:
Word:
Definition:
Example:
Non-example:
Word: Homogeneous Mixture
Definition:
Example:
Non-example:

Matter Terms

Definitions with examples:

Physical Separation (1 examples)

Chemical Separation (1 examples)

Physical Change (1 examples)

Chemical Change (1 examples)

Physical Properties (1 examples)

Chemical Properties (1 examples)

Matter

Substance

Element

Compound

Mixture

Heterogeneous Mixture

Homogeneous Mixture

Classification of Matter Lab

Part 1: After completing each of the following activities, use your observations to determine whether a chemical or physical change occurred. Explain your answer.

1. Fill a test tube halfway full of water and add several drops of food coloring. Clean test tube when done.

2. Take a small piece of aluminum and crush it as small as you can.

3. Add about 20-25 drops of potassium carbonate onto a watch glass and then add 20-25 drops of copper chloride onto the same watch glass. Rinse the contents down the drain and clean the test tube.

4. Add about 20-25 drops of vinegar into a test tube and then add a small scoop of baking soda into the same test tube. Rinse the contents down the drain and clean the test tube.

5. Place few crystals onto watch glass. Place on hot plate (setting 4), watch for change. Remove watch glass from hot plate, let cool. Add couple drops of water to crystals. Put in trash can.

6. Add about 20-25 drops of water to a test tube and then add a small scoop of sodium chloride salt into the same test tube. Rinse the contents down the drain and clean the test tube.

Reaction	Observations	Physical or chemical change?	Explain why it's a physical or chemical change
Water and food coloring			
Crushed aluminum			
Potassium Carbonate and Copper Chloride			
Baking soda and vinegar			
Copper Sulfate hydrate and heat			
Water and NaCl			

Part 2: List all the substances on your lab bench and describe their different chemical and physical properties. Explain how you distinguish between chemical and physical properties

Properties of Matter Practice

Determine if the descriptor is a physical or chemical property and <u>why</u>

	Physical Property	**Chemical Property**
Blue Color		
Density		
Flammability		
Solubility		
Reacts with acid to form H_2		
Supports combustion		
Sour taste		
Melting point		
Reacts with water to form a gas		
Hardness		
Boiling point		
Luster		
Odor		

Classify each of the following as an **element, compound, heterogeneous mixture, or homogeneous mixture.**

Gold	
Sand and iron filings	
Kool-aid completely dissolved in water	
Baking Soda (sodium bicarbonate or $NaHCO_3$)	
Sulfur	
Sodium chloride (NaCl)	
Starbucks Frappuccino	
Water	
Oxygen	

Name _____ Per _____

THE MOLE

1. Determine the molar mass of each of the following compounds. SHOW WORK!

a. CH_4

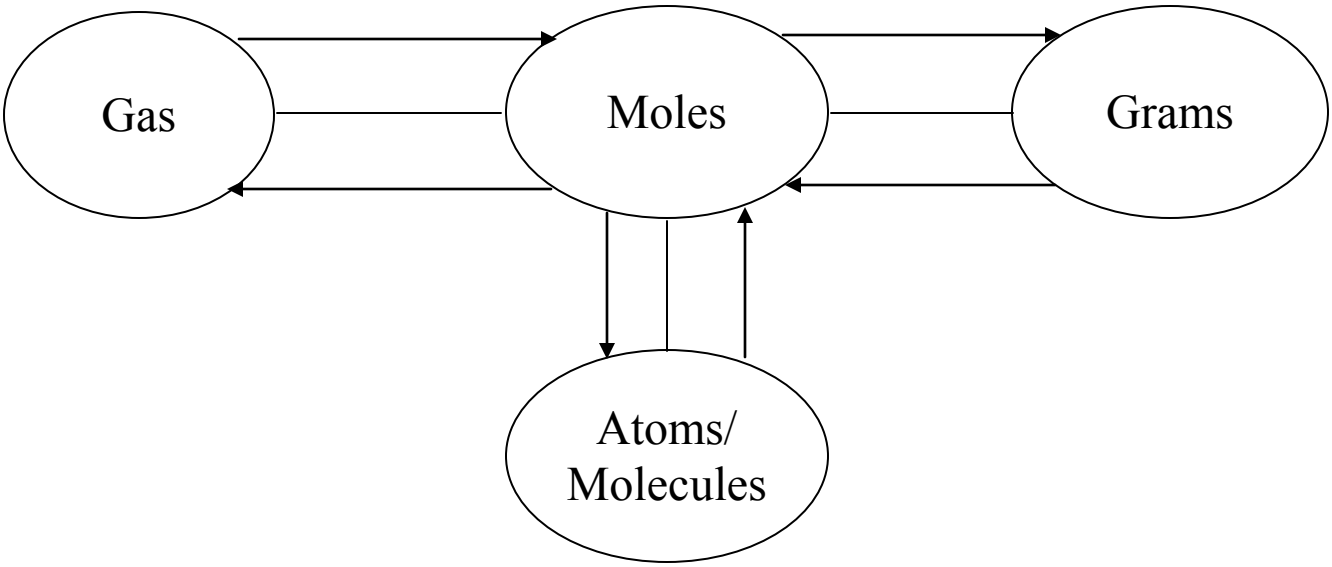
b. C_2H_2O

c. $NaNO_3$

d. $Mg_3(PO_4)_2$

e. C_6H_5OH

f. $LiAlH_4$



How many moles are in 5 g of carbon dioxide?

How much mass are in $3.5x10^{25}$ atoms of fluorine?

How many molecules are in 2.6 moles of sodium chloride?

How much mass is 34.6 L of C_3H_8?

How many grams are in 24.3 moles of sodium nitrate?

How much gas (in liters) are in 10.5 grams of argon?

How much gas is in 7.8 moles of C_2H_5OH?

Name _____ Per _____

MORE MOLE CONVERSIONS

1. How many moles are there in 35.5 g H_2CO_3?	2. How many moles are there in 1559 g barium chloride?
3. How many molecules are there in 0.0151 moles of water?	4. How many moles are in 44.8 g aluminum sulfate?
5. How many grams are in 2.5 moles of sodium hydrogen carbonate?	6. How many moles are in 1.00 g NaCl?

LEMONADE ANYONE?

Objective: To determine the concentration of six different bottles of lemonade using qualitative and quantitative measurements.

Procedure:

Qualitative Analysis

- You may begin at any of the six stations.
- Get a drinking cup and pour a small amount of lemonade in your cup.
- Taste and record your observations on the data table.
- Go to all the stations and repeat the procedure.
  MAKE SURE THAT YOU BRING YOUR OWN CUP TO EACH TASTING.
- Which is more concentrated? Record.
- Rank the six solutions from 1-6, with number 1 being the most concentrated.
- Share your results on the blackboard with the class.

Quantitative Analysis

- The instructor will now give you the grams of lemonade mix that was dissolved in each bottle
- Work together with a partner
  - ◊ Lemonade's primary ingredients include sucrose ($C_{12}H_{22}O_{11}$) and citric acid: ($C_6H_8O_7$). Assume that the lemonade is a 50% / 50% mass mixture of sucrose and citric acid. Calculate the molecular mass of each of the 2 components and the grams of each component in the bottles.
  - ◊ Based on the grams of each component dissolved, determine the number of moles of each component in each bottle.
  - ◊ Determine the molar ratio of each component. How does this compare to the mass ratio?
  - ◊ Record the volume (liters) of lemonade solution in each bottle.
  - ◊ Use the formula below to calculate the molarity of lemonade (concentration) of the six different bottles of lemonade you tasted.

Data/ Results:

Bottle	Taste*	Liters	g of Mix	g of Sucrose	Moles of Sucrose	g of Citric Acid	Moles of Citric	Moles of Mix	Molarity of Mix
A									
B									
C									
D									
E									
F									

Calculations

- Show one sample calculation for finding the moles of sucrose.

- Show one sample calculation for finding the moles of citric acid.

- Show one sample calculation for finding molarity.

- Determine the mass percent of sample A.

- What is the mole fraction of sample C?

Conclusions:

- Which bottle was the most concentrated? Explain how you know this.

- Which bottles were very close in taste?

- How accurate were your taste results when compared with the molarity? Give specific examples/ numbers.

- If you were off with your taste results compared to the molarity, what might have helped to make your taste test more accurate?

SOLUTIONS

1. Define the terms solute, solvent, and solution.

2. Determine the solute and solvent in each of the following solutions.

a. Orange Juice
Solute(s)-

Solvent-

b. Coffee
Solute(s)-

Solvent-

c. Air
Solute(s)-

Solvent-

3. Why does alcohol dissolve in water but oil does not?

<u>Molarity</u>

1.) If 3.4 moles of sugar are dissolved in 250 mL of solution, what is the concentration?

2.) What is the concentration when 14.5 grams of carbon dioxide is dissolved in 15.5 L of solution?

3.) A solution has a concentration of 12.2 M and a volume of 150 milliliters. How many moles of solution were added?

4.) 25.0 grams of HCl were added to an unknown volume to make a solution with a concentration of 6.5 M. What is the volume of the solution?

5.) How many grams are in a 17 L solution of calcium oxide with a concentration of 17 M?

6.) 0.5 moles of sodium chloride was dissolved to make a 0.05 L solution. What is the molarity?

7.) 0.5 grams of sodium chloride is dissolved to make 0.05 L of solution. What is the molarity?

8.) 0.9 grams of lithium chloride is dissolved to make 2 L of solution. What is the molarity?

9.) How many moles of potassium iodide are in 5 L of a 6 Molar solution?

10.) A 21 M solution has 7 moles of sodium bromide. How many liters are present?

Dilutions

11.) How many liters of 12 M H_2SO_4 are needed to make 550 mL of a 3.5 M solution of H_2SO_4?

12.) 1.5 L of a 2.0 M solution of HCl is made from 250 mL of the original solution. What was the concentration of the original solution?

13.) Describe how you would prepare 1.0 L of 0.20 M sulfuric acid, H_2SO_4 solution from a 18.0 M stock solution of the acid.

MYP Chemistry: Final Review

Unit: Conservation of matter
- ⇒ Polyatomic ions (formula and charge)
- ⇒ Writing formulas
- ⇒ Naming compounds
- ⇒ Reading and writing chemical equations
- ⇒ Balancing equations
- ⇒ Br i n cl h o f
- ⇒ Classifying reactions (synthesis, decomposition, combustion, single replacement, double replacement)
- ⇒ Predicting products of a reaction

Unit: The Atom
- ⇒ Basic atomic structure; locations/charges/masses of subatomic particles
- ⇒ Atomic symbols, calculating atomic mass
- ⇒ Wave relationships: wavelength, frequency, and energy
- ⇒ EM spectrum, continuous spectrum vs. Bright line spectrum
- ⇒ Bohr model, relationship to bright line spectra
- ⇒ Electron dancing, identifying elements with spectra
- ⇒ Electron configuration
- ⇒ Organization of the Periodic Table – Periods and Groups
- ⇒ Names of common families: alkali metals, alkaline earth metals, boron family, carbon family, nitrogen family, oxygen family, halogens, noble gases

Unit: Bonding
- ⇒ Periodic trends: atomic radius, ionization energy, electronegativity, reactivity
- ⇒ Ionic bonding: electron transfer
- ⇒ Covalent bonding: electron sharing,
- ⇒ Lewis dot structures, exceptions to octet rule, double/triple bonds, ions
- ⇒ Shapes of molecules (VSEPR), bond angles, resonance, effect of shape on molecular polarity
- ⇒ Intermolecular Forces

Unit: Properties of Matter
- ⇒ Matter definition
- ⇒ The mole (moles/molecules/liters)
- ⇒ Atomic mass; formula mass (g/mole)
- ⇒ Solutions
- ⇒ Molarity Calculations

PRACTICE PROBLEMS

Nomenclature:

Identify each compound as either ionic or molecular. Then **name** it and determine formula mass.

Compound	Ionic/Covalent	Name	Molar Mass
1. MgO			
2. SO_4			
3. Fe_2S_3			
4. N_2O_5			
5. $BaSO_3$			
6. $Zn(SO_4)$			
7. CCl_4			
8. SF_6			
$Ni_3(PO_4)_2$			

Identify each compound as either ionic or molecular. Then **write** a correct formula and determine formula mass.

Compound	Ionic/Covalent	Name	Molar Mass
10. magnesium sulfate			
11. nitrogen triiodide			
12. lead (II) phosphate			
13. ammonium nitrate			
14. dichlorine monoxide			
15. carbon dioxide			
16. aluminum chloride			
17. chromium (III) oxide			
18. potassium iodide			
19. diphosphorus tetraoxide			

Reactions and equations:

For each of the following: write a balanced equation and classify the reaction.

- sodium reacts with aqueous aluminum carbonate
- aqueous sodium sulfide reacts with aqueous barium nitrate
- copper (II) carbonate reacts to form copper (II) oxide and a gas
- potassium reacts with iodine
- ethane (C_2H_6) burns in air
- barium reacts with aqueous zinc (II) chloride
- aqueous sodium carbonate reacts with aqueous zinc (II) fluoride
- lead (II) oxide decomposes

Atomic Structure:

- Sketch a diagram of an atom including locations, charges, and masses of all three subatomic particles.
- Define (or write the equation that defines them): atomic number, mass number, average atomic mass, atom, ion, and isotope.
- Provide an example of an ion.
- Provide an example of two or more isotopes:
- What two subatomic particles contribute to an atom's mass?
- What two subatomic particles contribute to an atom's charge?
- Fill in the chart:

Isotope	# of pro-tons	# of neutrons	# of electrons
$^{57}_{26}Fe$			
$^{14}_{6}C$			
$^{12}_{6}C$			
$^{35}_{17}Cl^{-1}$			

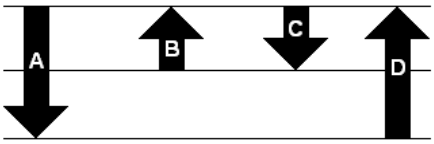
- Calculate the atomic mass of a sample that contains 20% ^{6} Li and 80% ^{7}Li.
- What experiment did J.J. Thomson run? What did he discover?
- What experiment did Ernes Rutherford run? What did he discover?
- How did Bohr contribute to Rutherford's atomic model?
- Determine the average atomic mass of Iron if the following information is known:

Isotope	Percent abundance
Iron – 54	5.85
Iron—56	91.75
Iron—57	2.12
Iron—58	0.28

- What is the electron configuration of sodium? Of chlorine?
- What element has the electron configuration $1s^2 2s^2 2p^3$? … $[Ne]3s^2 3p^3$?
- How are elements in the same group (column) related? How are the alkali metals all related? The noble gases?

Nature of Light:

- As wavelength gets shorter, frequency _____.
- As wavelength gets shorter, energy _____.
- What is the electromagnetic (EM) spectrum? What is the highest energy wave? Lowest?
- What is the difference between a bright line spectrum and a continuous spectrum?
- Explain how electron movement between energy levels produces photons of light.
- Consider this diagram of an atom with arrows representing electron movement.



  - Which two arrows correspond to energy absorption by the atom? _____
  - Which two arrows correspond to energy emission by the atom? _____
  - If violet and green light are produced by the movement illustrated here, which arrow represents emission of violet light? _____ green light? _____

- How do waves of red light and blue light differ with respect to frequency? wavelength? energy?

Periodic Trends:

Define the terms electronegativity, ionization energy, atomic radius. Describe their trends on the periodic table.

_____ is the 5th period alkaline earth metal
_____ is the 3rd period halogen
_____ is the most reactive nonmetal in the 4th period
_____ has the largest radius of the 4th period
_____ is the largest atom with two valence electrons
_____ is the least reactive alkaline earth metal
_____ has 18 electrons and 17 protons
_____ has 2 valence electrons in the 3rd energy level
_____ has the greatest ionization energy of the alkaline earth metals
_____ is the element in oxygen family with 3 energy levels
_____ is the 3rd period element with lowest ionization energy
_____ is the noble gas with a 2nd energy level valence shell
_____ is the 5th period atom most likely to lose an electron
_____ is the 5th period atom least likely to react
_____For elements 39 through 48, additional electrons are added to the _____ sublevel
_____Which element has the greatest radius: S Cl Se Br
_____ is the third period element which is most likely to form an ion with a +3 charge
_____The halogen with the highest melting point
_____ react by gaining electrons.
_____ react by losing electrons.
_____Write the symbol that has the largest radius Rb Sr K Ca
_____Write the symbol that has the greatest reactivity Rb Sr K Ca
_____Write the symbol that is the most reactive Se Br S Cl
_____ Which has the highest electronegativity? Se Br S Cl
_____ Which has the lowest electronegativity? Rb Sr K Ca

- Why do elements form chemical bonds?

Covalent Bonding:

For each of the following: a) draw a dot structure b) give the shape of the molecule c) does the mode have resonance?

- H_2O
- CH_4
- CO_2
- NH_3
- BCl_3
- NH_4^+
- C_2H_4
- C_2H_2
- CH_2O
- HCl
- F_2

- What are some distinctive properties of ionic compounds?
- What makes an ionic bond different than a covalent bond?
- What is VSEPR theory?
- Draw a Lewis dot structure for each of the following molecules: O_2, H_2O, CH_2Cl_2, NI_3
- In the question above, which molecules are polar? Which are nonpolar? Assign geometries to each.
- What are the three types of intermolecular forces? What type(s) of molecules is each one present in?
- What are some properties of the types of compounds?
- Describe the differences (on the molecular level) between solid, liquid, and gas phases.

Matter:

- Define the following terms: matter, pure substance, homogeneous mixture, heterogeneous mixture, element, atom, and compound.
- What are the two types of mixtures? Describe each and say how we can tell them apart.
- Classify each of the following as an element, compound, heterogeneous mixture, or homogeneous mixture.
  a. Gold (Au) b. Kool-Aid completely dissolved in water c. Sodium bicarbonate (H_2CO_3) d. Sulfur (S) e. NaCl f. Salsa
- Classify each of the following changes as either chemical or physical**.**
  a. Cutting a sheet of aluminum foil into 4 pieces b. Burning of coal c. Cooling a liquid until it freezes d. A white solid and sulfuric acid are mixed and an orange gas is produced e. Dissolving sugar in tea f. A piece of copper is hammered into a thin sheet
- Give three examples of physical properties and three examples of chemical properties.

Mole calculations

- Find the mass in grams of 5.6L of F_2.
- Find the number of moles of argon in 481 g of argon.
- Calculate the number of moles in 6.0 grams of carbon.
- Penicillin F has the formula $C_{14}H_{20}N_2SO_4$. How many molecules of this medicine are in a 2.0 g dose?
- Calculate the volume of 100.0 g of oxygen gas.
- How many moles of sodium are in 1 mol of Na_2O?

Solutions

- Define the terms solute, solvent and solution.
- Detail the differences between a heterogeneous mixture and a homogeneous mixture. Which category does a solution fall into?
- Define the term concentration. List a few of the units and formulas used to determine concentration (particularly molarity).
- What is the molarity of a solution containing 2.5 grams of HCl in 3.5 liters of solution?
- How many moles are contained in 3.4 liters of a 2.5 M solution?

Periodic Table of the Elements

1A	2A											3A	4A	5A	6A	7A	8A
hydrogen 1 H 1.00794																	helium 2 He 4.002602
lithium 3 Li 6.941	beryllium 4 Be 9.012182											boron 5 B 10.811	carbon 6 C 12.0107	nitrogen 7 N 14.00674	oxygen 8 O 15.9994	fluorine 9 F 18.9984	neon 10 Ne 20.1797
sodium 11 Na 22.98977	magnesium 12 Mg 24.3050	3B	4B	5B	6B	7B		8B		1B	2B	aluminium 13 Al 26.981538	silicon 14 Si 28.0855	phosphorus 15 P 30.97376	sulphur 16 S 32.065	chlorine 17 Cl 35.453	argon 18 Ar 39.984
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 65.409	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
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
caesium 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]
francium 87 Fr [223]	radium 88 Ra [226]	lawrencium 103 Lr [262]	rutherfordium 104 Rf [261]	dubnium 105 Db [262]	seaborgium 106 Sg [266]	bohrium 107 Bh [264]	hassium 108 Hs [269]	meitnerium 109 Mt [268]	darmstadtium 110 Ds [271]	roentgenium 111 Rg [272]	ununbium 112 Uub [285]		ununquadium 114 Uuq [289]				

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]