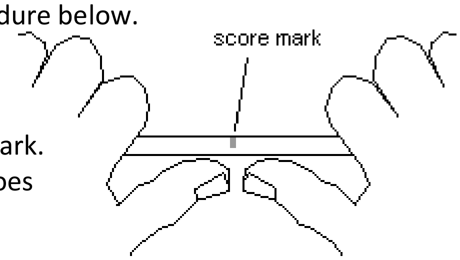
Glass Lab Name: _______________ partner: _____________

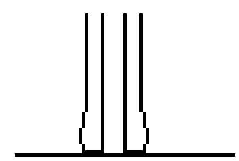
Now that you are familiar with the laboratory burner, let’s use it to learn about glass. Glass is an interesting material. Although it certainly looks and feels like a solid, technically, glass is a liquid – a very slow flowing liquid. In a solid, the particles only vibrate in place and they never move around each other. If they did, a solid couldn’t maintain its shape. In a liquid, the particles can flow around each other and so a liquid has no definite shape. Some liquids (like honey) are very thick -- the correct term to use in chemistry is “viscous.” This happens because the particles are very sticky to each other and therefore have trouble flowing smoothly past one another. Glass is a very, very, very viscous liquid. This means that eventually all glass objects (windows, bottles, etc.) will end up as puddles on the ground! Glass is so slow flowing, however, that this would take trillions of years, so the fact that glass is a liquid really doesn’t mean much in terms of how we experience it... except for one thing. When you heat most substances, they reach a certain temperature where they suddenly melt. Since glass is already a liquid, it can’t melt. But when you heat glass, you do allow the molecules to flow more quickly past one another and this means that the viscosity (thickness) decreases. Thus, instead of melting all at once at a certain temperature, glass just gets softer and more fluid as you heat it up. This makes it pretty easy to work with glass and to modify its shape into...

Caution: Goggles on, hair tied back, sleeves rolled up. Also, glass stays hot for a long time -- much longer than you would think. If you touch hot glass you will get burned... so don’t. You will be making three separate objects out of glass: a restricted right angle bend, an eye dropper tube and a small rimmed test tube. You will be graded on how well your final products match these three templates. To do so, carefully follow the procedure below.



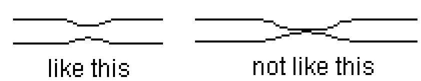
1. Measure 20.0 cm of glass tubing, score it with a triangular file (one hard score mark), wet the score with a drop of water and with the score mark away from you, apply pressure with the thumbs (as shown at right) to break the tubing at the score mark. IMPORTANT: If it doesn’t break easily, don’t force it: you might cut yourself when it does Break. Instead score it again, deeper, and try breaking it again.

2. Place a wing tip on a burner, light it and adjust it to a medium hot flame. “Fire polish” the two ends of the tube. This involves holding the tube horizontally with the tip just barely in the hot part of the flame (the top of the inner blue cone) rotating them slowly. Watch the sharp edges become nice and smooth. Don’t overdo it, or the end will start to sag of even close up. See the example at the front table for a well fire-polished end. Allow this to cool completely on the lab bench. Don’t worry; the lab benches are heat resistant.

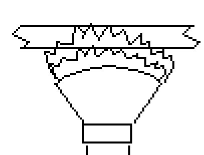


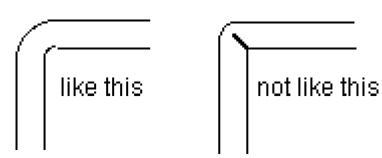
3. While you are waiting for it to cool, measure out a second length of tubing, this time 15.0 cm long, score it and fire polish one end, and while it is still hot, push it gently down on the lab bench to flare it a bit as shown at right. Let this cool.

4. Your 20.0 cm tube should be cool by now. To test for this, hold the back of your hand just over the glass. If you feel no heat, it should be safe to handle. All the same, touch it briefly at first to test it before picking it up. You will now attempt to put a small taper 7.0 cm from one end. Use a permanent marker to make a mark 7.0 cm from one end of the tube. Now heat that spot, rotating the tube while holding it horizontal and perpendicular to the flame. (The mark will burn off pretty quickly, so do your best to remember where the mark was!) It helps to rest your forearm on something to help hold the tube steady. After you feel the glass start to soften, remove it from the flame and pull it very gently to create a small taper as shown at right: Set this down and allow it to cool.



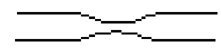
5. Now pick up your 15.0 cm tube and fire polish and flare the other end as you did above. Let this cool.





6. Now pick up your 20 cm tube and make a mark 6.0 cm from the un-tapered end. Heat this mark holding the tube horizontal and parallel to the flame this time (as shown at right). This will heat a wider section of the tube and allow for a smoother bend. When it seems soft enough, remove it from the flame and bend it to a nice smooth 90° angle. Do your best not to crimp the bend as shown at far right. Let this piece cool; you have now completed your tapered right angle bend!

7. Pick up the 15 cm tube, make a mark at the midpoint (7.5 cm from one end). Then at this mark, heat and rotate the tube perpendicular to the flame. Then remove it from the flame and pull it out into a taper as you did in step #4 above. This time make the taper a little longer and narrower than before:
Let this cool. <u>Answer the first follow up questions below while you wait!</u>


like this

8. Once it has cooled, carefully score and break the tube right at the center of the taper. Because it's so small, this is kind of tricky to do. Then briefly fire polish one tapered end. You have now completed your eye dropper tube. If you accidentally leave it in the flame too long it will close up. If this happens, just keep heating as described below and let this piece become your test tube.

9. Fire polish the other tapered end, and then continue heating and rotating it while holding it at a 45° angle upward into the flame. This should seal it off and give a nice rounded bottom to your test tube.

10. Once the three objects have cooled, and you have finished the follow up questions below, tape them to the attached page <u>right on top of their respective templates</u>. Write your name on this sheet and turn this in. If they match the templates fairly closely, you will be getting 15/15 for the lab. (Yeah!) If they do not, oh well, you tried... and at least you didn't burn yourself!

Follow-up questions:
1. To find out why we don't just cool the glass down quickly try taking a small scrap of glass rod (5-7 cm long) and, holding it with steel tongs, heat one end in a burner for about 20 seconds, then stick this hot end in a beaker of water.

Describe what happens: ____
Why do you think this happens? (Hint: recall that things expand when they are heated and contract when they are cooled, and that glass is not a very good conductor of heat.)

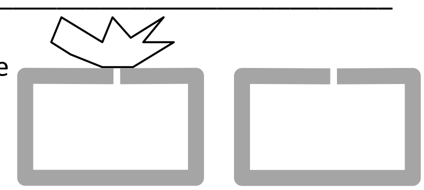
2. Try doing the same thing with an iron nail. What happens? ________________________________
Why do you think the nail didn't do the same thing as the glass tube?

3. Why is glass categorized as a liquid? ____

4. What does viscous mean? ____

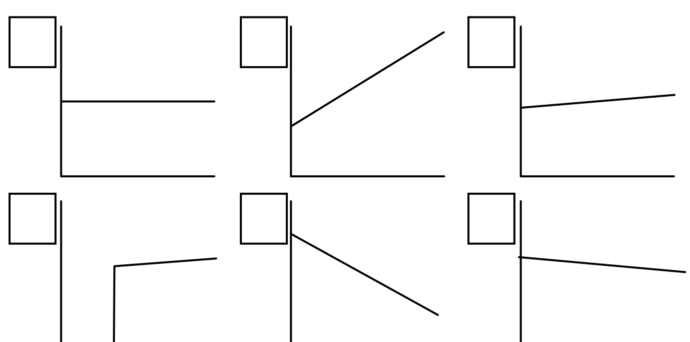
and what causes a substance to be viscous? ____

5. A sharp piece of glass is set over a hole on top of a box as shown at right. If you came back in 100 trillion years, what would the glass look like? Draw in the image at right. Assume the box does not change.


Present day 100 trillion years from now

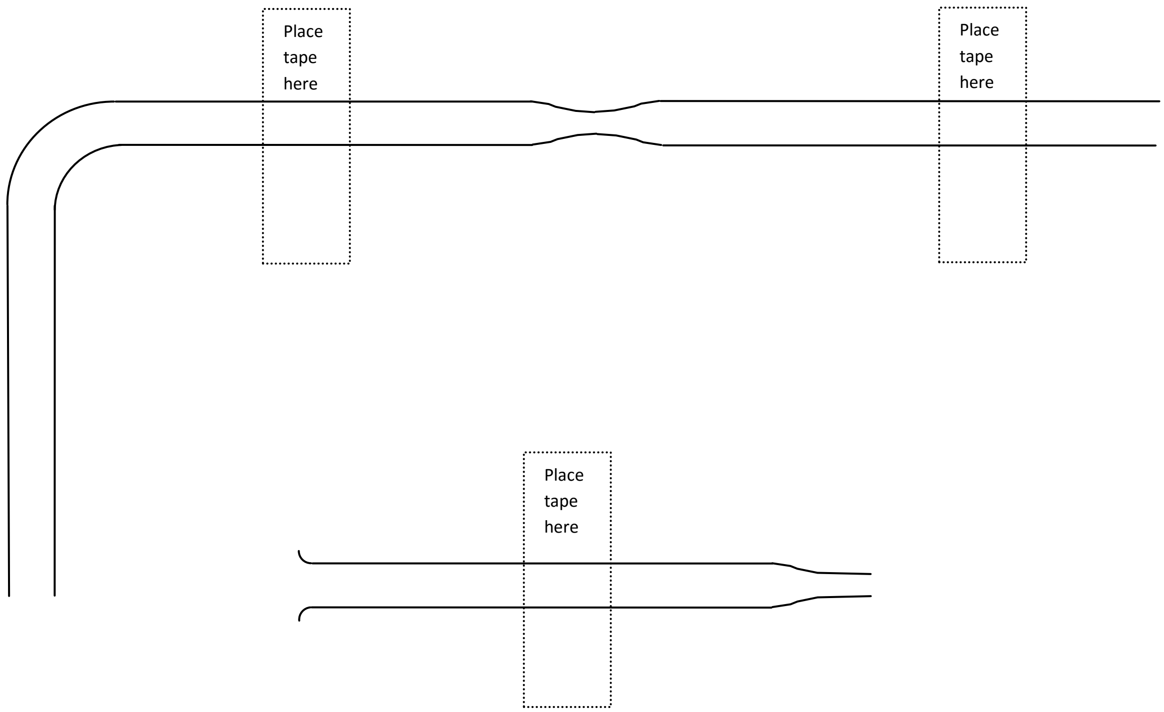
6. A piece of copper and a piece of glass are both being heated up. How do they behave differently – especially concerning their fluidity? (Note: fluidity is essentially the opposite of viscosity.)

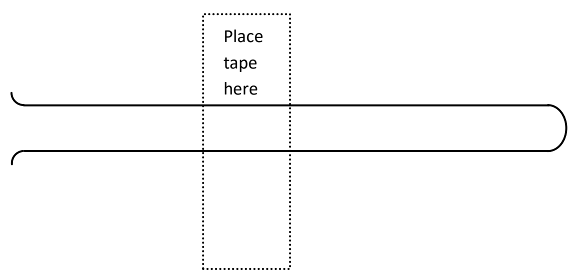
7. Match the six graphs at right to the titles below:
A. Viscosity of a glass sample as a function of temp
B. Mass of a glass sample as a function of temp
C. Volume of a glass sample as a function of temp
D. Density of a glass sample as a function of temp
E. Fluidity of a glass sample as a function of temp
F. Fluidity of a copper sample as a function of temp



Glass Lab Results. Possible score: 15 pts Name: ______________________________ Per: _____

So how well did you follow directions? If you did a good job, then the three glass pieces you made should measure up well against the life-size templates below. Tape them on top of the templates, lining them up as best as you can. And then turn them in to the teacher.

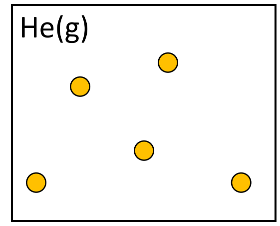


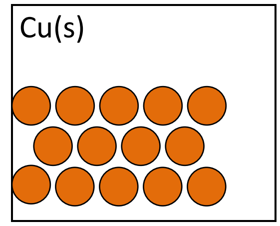


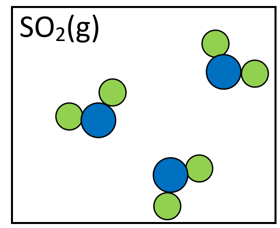
SLG WS (SOLIDS, LIQUIDS & GASES) Name: ________________

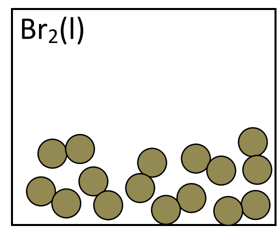
1. First a question: Don't worry about whether you are right or wrong: When water boils, what specific gas or gases are the bubbles made up of? _________________

What makes you think that? ____



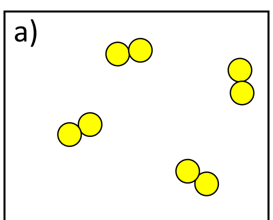




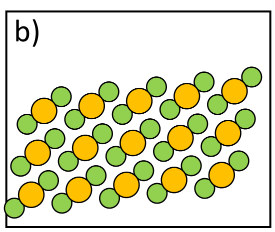


2. First, the basics. Atoms vs Molecules. Atoms are very tiny, individual particles. They have a dense nucleus at their center (containing protons and neutrons), and electrons in various levels, sublevels and orbitals around the nucleus. (We'll learn much more about this later.) The element helium [He(g)] is made up of a individual atoms bouncing around randomly in the gaseous state (as shown at top right). The element copper [Cu(s)] is made up of individual atoms vibrating in place in a precisely organized solid structure (2nd diagram). Molecules are made up of atoms, bonded together into definite clusters. The compound sulfur dioxide [SO_2(g)] is made up of SO_2 molecules -- each molecule comprised of one S atom and two O atoms bonded together in a V-shaped arrangement. These individual molecules are bouncing around randomly in the gaseous state (3rd diagram). The element bromine [Br_2(l)] is made up of Br_2 molecules --each molecule comprised of two Br atoms bonded together side by side. These individual molecules are moving around one another but staying in close proximity in the liquid state (4th diagram). So.... got it? Atoms are atoms, and molecules are made up of atoms. Oh, and one more thing: Elements are all comprised of just one type of atom. Sometimes those atoms are individual (un-bonded) atoms as in He and Cu above, other times they are bonded into molecules as in Br_2 above. Compounds are comprised of two or more types of atoms bonded together in a specific ratio, like SO_2 above.

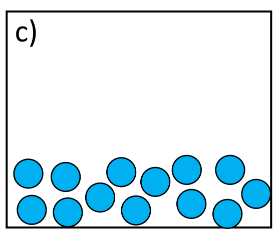
3. So for each diagram below, circle three choices: (1) solid, liquid or gas; (2) an element or a compound and (3) made up of atoms or molecules. The first is done for you.



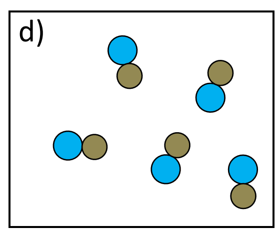
solid liquid gas
element compound
atoms molecules



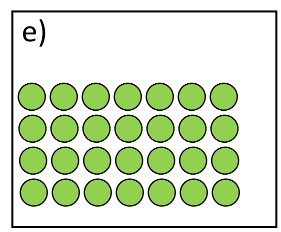
solid liquid gas
element compound
atoms molecules



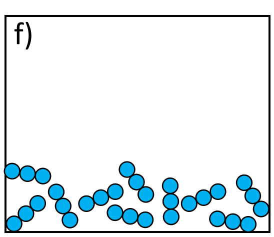
solid liquid gas
element compound
atoms molecules



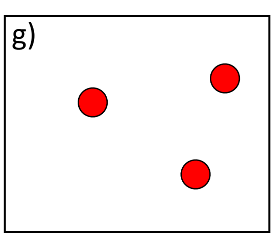
solid liquid gas
element compound
atoms molecules



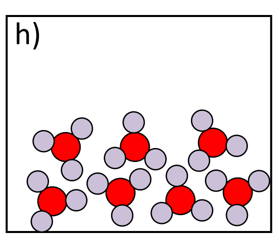
solid liquid gas
element compound
atoms molecules



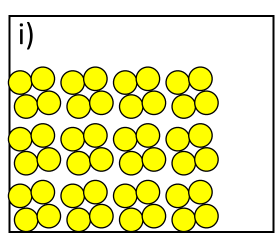
solid liquid gas
element compound
atoms molecules



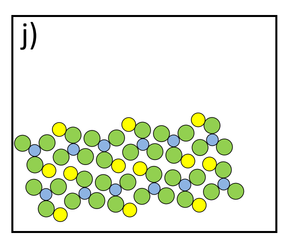
solid liquid gas
element compound
atoms molecules



solid liquid gas
element compound
atoms molecules



solid liquid gas
element compound
atoms molecules



solid liquid gas
element compound
atoms molecules

4. In the spaces below, draw diagrams (like the ones in #3 above) for each substance specified.

a) a molecular, gaseous compound

b) a solid element made of atoms

c) $NF_3(s)$

d) $H_2O(l)$

e) $F_2(g)$

5. For each description or characteristic, fill in the blank with “S” for solid, “L” for liquid, “G” for gas, “SL” for solid & liquid, “LG” for liquid and gas, “A” for all three or “N” for none of the three.

___ particles moving constantly
___ particles moving around each other but staying very close together
___ particles all have the exact same velocity all the time
___ definite shape and volume
___ definite volume
___ definite shape but not definite volume
___ assumes the shape of its container
___ not capable of flowing

___ particles vibrating in place
___ particles moving randomly and far apart
___ particles moving in little circles
___ particles have a range of different velocities
___ particles move faster as temp. increases
___ definite volume but not definite shape
___ definite mass
___ no definite shape or volume
___ assumes the shape and volume of its container
___ considered a “fluid” because it’s able to flow

___ hard ___ usually invisible ___ rigid ___ usually clear (transparent) ___ bendable
___ diffuses quickly across a room ___ wet ___ can be hot ___if colored, then poisonous!

Melting involves ___ changing to ___.
Boiling involves ___ changing to ___.
Sublimation involves ___ changing to ___.
Evaporation involves ___ changing to ___.

Condensing involves ___ changing to ___.
Deposition involves ___ changing to ___.
Freezing involves ___ changing to ___.
Phase changes involve ___ changing to ___.

6. For each of the following descriptions, fill in the blank with “M” for melting, “F” for freezing, “B” for boiling, “C” for condensing, “S” for subliming, or “D” for depositing. IMF = intermolecular forces: the attractive forces that act between molecules and determine phases: strong IMF’s make substances solid, weak ones = gases.

___ Changing from a solid to a liquid
___ Changing from a solid to a gas
___ Changing from a gas to a liquid

___ Changing from a liquid to a solid
___ Changing from a gas to a solid
___ Changing from a liquid to a gas

___ Fast, free-moving particles are slowed down enough that their IMF’s pull them together close but still allow some freedom of movement.

___ Vibration of particles speeds up until enough IMF’s are overcome to allow some free movement, but still in a fairly tight space.

___ Vibration of particles speeds up until enough IMF’s are overcome to allow completely free, random movement.

___ The movement of tightly spaced particles around one another is slowed down until the IMF’s start to lock the particles into fixed positions where only tiny vibrational movements are possible.

___ The movement of tightly spaced particles around one another is sped up until enough IMF’s are overcome to allow them to fly apart from one another and move randomly through space.

___ Fast, free-moving particles are slowed down enough that their IMF’s pull them together close into fixed positions where only tiny vibrational movements are possible

7. Given that phase changes just involve changes in the spacing and speeds of the particles, let’s re-examine your answer to question #1: When water boils,
what specific gas or gases are the bubbles made up of? _________________

ELEMENTS, COMPOUNDS & MIXTURES LAB Name:_______________ Partner: ____________

Purpose: To learn to distinguish between the concepts of element, compound and mixture.

Introduction: Text books often define these three terms in the following way:

element: A fundamental material consisting of only one type of atom [example: Fe (iron)]

compound: A material in which atoms of two or more different elements are bonded to each other in a fixed proportion [example: CO_2 (carbon dioxide)]

mixture: A combination of two or more different substances [example: salt water]

Elements are usually easy to identify, since they are comprised of just one type of atom, but many students have trouble with the distinction between compounds and mixtures, since both involve combinations of different elements. What, for example, is the difference between the compound aluminum sulfide and a simple mixture of aluminum and sulfur? There really are two answers for this question: 1) In the compound aluminum sulfide, the atoms are bonded to one another (either to form molecules or crystal structures), whereas in a mixture of aluminum and sulfur, the atoms are simply mixed together, but not specifically bonded. 2) In the compound aluminum sulfide, the atoms are present in a specific ratio: in this case 2:3 -- that is, there are two aluminum atoms for every three sulfur atoms -- hence aluminum sulfides formula is Al_2S_3. This is true no matter how the aluminum sulfide was made or where it came from... this is what is meant by a “fixed” or “definite” proportion. In a mixture, the aluminum and sulfur could be present in any proportions: 1:1, 1:2, 58:17... in whatever ratio the person who made the mixture decided to combine them.

Procedure:

1. Obtain ten bags: A through J. These bags contain beads which represent atoms: different colors represent different elements. Examine each bag carefully and discuss with your partner whether you think the bag contains an element, a compound or a mixture. Once you agree, sketch the contents of the bag in the appropriate box below; then beneath that, indicate what it represents: “element” “compound” or “mixture,” and then give a brief explanation for your choice.

A Sketch:	**B**	**C**	**D**	**E**
E, C or M?				
Explanation:				

F	**G**	**H**	**I**	**J** Look at this one carefully.

2. Now obtain bottles 1 through 12. These contain the following substance (not necessarily in this order): sulfur, magnesium, lead, polyvinyl chloride, brass, iodine, marble, glycerine, salt water, carbon dioxide, ammonium sulfate and air. Try to determine which bottle contains which substance, and also decide whether that substance is an "element," "compound" or "mixture." [You may not be able to distinguish all twelve substances.] The following information will prove helpful.

sulfur is a yellow powder. It is made up of molecules comprised of eight sulfur atoms bonded into a ring structure.
magnesium is a shiny, silvery low density metal made entirely of magnesium atoms.
lead is a dull, grey, very dense metal made entirely of lead atoms.
polyvinyl chloride is made up of long chain molecules which have two carbon atoms, three hydrogen atoms, one chlorine atom, two carbon atoms, three hydrogen atoms, one chlorine atom, over and over again. It can be formed into thin filaments or flat sheets. .
brass is a gold-colored alloy of the copper and zinc, containing usually between 15 and 40% zinc by mass. [An alloy is a mixture of two or more metallic elements.]
iodine is a grayish brown, brittle solid that looks like little pieces of gravel. It sublimes into a purple gas when heated. (Try putting the vial that you think contains iodine into hot water to see what happens.) Also, iodine is diatomic: that is, its molecules are each comprised of two iodine atoms bonded together.
marble is also known as calcium carbonate: $CaCO_3$. It is comprised of crystals that contain calcium, carbon and oxygen atoms in a 1:1:3 ratio, respectively.
glycerine is a very viscous (slow-flowing) liquid. (Try putting the vial you think contains glycerine into the hot water to confirm that it is a viscous liquid and not a solid.)It is made up of molecules each of which contain three carbon atoms, eight hydrogen atoms and three oxygen atoms bonded together ($C_3H_8O_3$)
salt water, as the name implies, is simply salt (NaCl) dissolved in water (H_2O)
carbon dioxide is a gas with molecules comprised of one carbon atom and two oxygen atoms bonded together
ammonium sulfate $(NH_4)_2SO_4$ is comprised of crystals that contain nitrogen, hydrogen, sulfur and oxygen in a 2:8:1:4 ratio, respectively.
air is comprised of about 78% nitrogen molecules (N_2), 21% oxygen molecules (O_2), and small amounts of argon atoms (Ar), water molecules (H_2O), carbon dioxide (CO_2), and a bunch of other stuff.

Record your answers below: Write the bottle number before the substance, and "element," "compound" or "mixture" after the substance, then A) a brief explanation for how you identified the substance and B) a brief justification for how you categorized the substance.

___ sulfur ____________ A)____________________________ B)__________________________________

___ magnesium ____________ A)____________________________ B)__________________________________

___ lead ____________ A)____________________________ B)__________________________________

___ polyvinyl chloride ____________ A)____________________________ B)__________________________________

___ brass ____________ A)____________________________ B)__________________________________

___ iodine ____________ A)____________________________ B)__________________________________

___ marble ____________ A)____________________________ B)__________________________________

___ glycerine ____________ A)____________________________ B)__________________________________

___ salt water ____________ A)____________________________ B)__________________________________

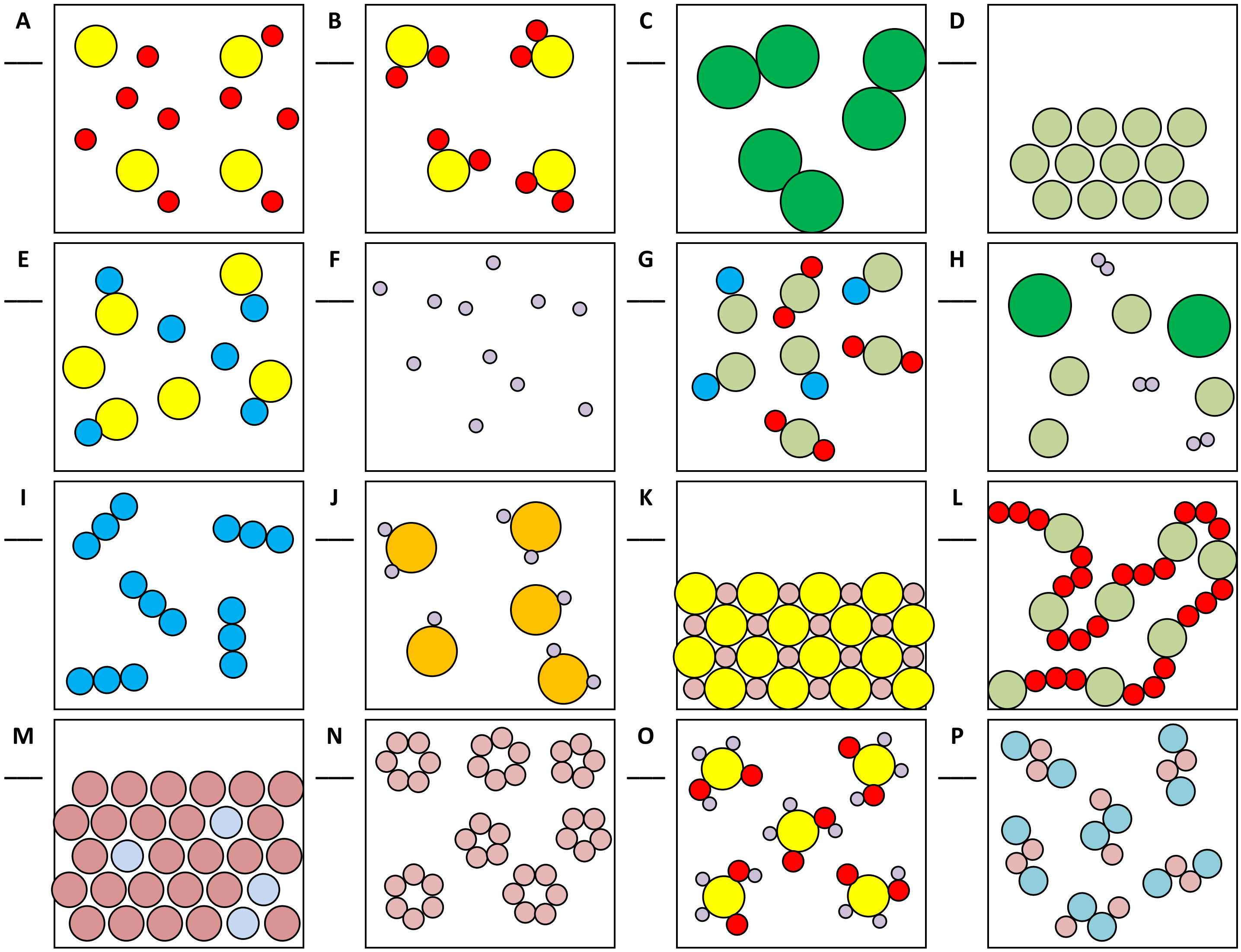
___ carbon dioxide ____________ A)____________________________ B)__________________________________

___ ammonium sulfate ____________ A)____________________________ B)__________________________________

___ air ____________ A)____________________________ B)__________________________________

ELEMENTS, COMPOUNDS & MIXTURES WORKSHEET Name: ____________________

1) Label each of the following as either: **PE** = a pure element **PC** = a pure compound **ME** = a mixture of two or more different elements **MC** = a mixture of two or more different compounds or **MEC** = an element/compound mixture.



2) Which diagrams above represent **crystals**? (In crystals, the atoms are in specific repeating patterns) ___ ___ ___

3) Many elements are made up of individual atoms (like neon: Ne). Others bond together into small distinct clusters -- perhaps pairs like oxygen (O_2) or nitrogen (N_2) or maybe foursomes like phosphorus (P_4). Which of the diagrams above show elements made of small distinct clusters? ___ ___ ___

4) **Polymers** are defined as long chain molecules made of atoms strung together in a repeating pattern. They make up most of the plastics we see and use every day. Polyethylene, for example is used for making soda bottles and its molecules are made up of carbon and hydrogen atoms in a ...$CH_2CH_2CH_2$... arrangement. Which diagram above represents a polymer? ___

5) **Allotropes** are defined as two different forms of the same element. Oxygen for example comes in two forms: O_2 -- two oxygen atoms bonded together into one molecule, which we and most other animals depend on for life, and O_3 -- three oxygen atoms bonded into one molecule known as ozone which protects us form uv radiation but which would be fatal if you tried to breathe it. Diamond and graphite are two different allotropes of carbon: the exact same atoms, just bonded in a different arrangement. Which of the diagrams above represents a mixture of two different allotropes of the same element? ___

6) **Isomers** are defined as two different molecules both made from the same combination of atoms. Ethanol (AKA: grain alcohol) is the "active ingredient" in beer and wine, and its formula is C_2H_6O. Dimethyl ether has the same formula (C_2H_6O), but it is a gas at room temperature and forms explosive mixtures with air!
Which of the diagrams above represents a mixture of two different isomers? ___

ANS: PE PE PE PE PE PC PC PC PC ME ME ME ME MC MC MC MEC C D I K L M N N P

Physical vs. Chemical Changes Lab Name: ___________________ Partner: ____________________

In the world around us, changes are happening all the time. Chemists like to make the distinction between those changes that are chemical -- in which bonds between atoms are being broken and new bonds formed -- and those changes that are purely physical -- in which the molecules stay intact, they merely spread out or mix with other molecules. Sometimes it is difficult to tell which kind of change is taking place -- chemical or physical -- without knowing exactly what is happening on a molecular level. Appearances can be deceiving. For example: when shiny dark gray crystals of iodine are heated, a purple gas appears. It seems that this would be a chemical reaction -- that something new has been produced. It turns out, however, that the solid iodine has simply turned into a gas. The solid was made of tiny I_2 molecules all packed together in a crystalline pattern, and the heat has simply spread them out into a purple gas -- still I_2 molecules. Thus, this is merely a physical change. On the other hand, when solid sodium bicarbonate (baking soda) is heated, no visible change can be seen or heard, but a chemical reaction does occur as the sodium bicarbonate ($NaHCO_3$) decomposes into two new compounds: sodium hydroxide (NaOH) and carbon dioxide (CO_2).

In this lab, you will witness twenty different changes, and you will be asked to consider which ones represent chemical changes and which ones are only physical. Here are some things to look for.

1) Something chemically new is produced. This is always evidence for a chemical reaction.
When wood burns, the ashes left behind are very different chemically than the wood from which they came. This is a good example of a chemical change. When water freezes, the ice cubes may look and feel different than the water from which they came, but we know the ice molecules have the exact same make-up (H_2O) as the water molecules. Phase changes like freezing and melting are simply considered physical changes.

2) Heat is given off (or taken in). This can be evidence of a chemical reaction or a physical reaction.
When a piece of magnesium burns ("burns" means to react with oxygen), a lot of heat is given off and this is a chemical reaction. When a sample of sodium hydroxide dissolves in water, a lot of heat is given off but dissolving processes are merely physical changes.

3) A color change occurs. This is often evidence of a chemical reaction, but not always...
When red paint and yellow paint are mixed together, they turn orange. This is not really a color change as much as it is just a mixing of two pigments. The individual molecules haven't changed at all, and nothing new is being formed. They are still each reflecting their own color of light, but the mixture appears orange. This is obviously a physical reaction. (This is sort of like mixing a bag of red beads and a bag of yellow beads: the mixture would appear orange form a distance, but certainly nothing new has been made.) When a drop of ammonia solution (colorless) is added to some red cabbage juice it turns it bluish green. Here, something chemically new must have formed to result in this kind of a change, and so this is a chemical reaction.

4) The change is easily reversed. This is usually (but not always) evidence that a change is only physical.
When calcium chloride is dissolved in water, just letting the water evaporate brings you back your calcium chloride crystals. Both the dissolving and the recrystallizing are considered physical changes.
When a piece of paper is torn in half, nothing new is really formed, and this is definitely an example of a physical change, but the process is not very easy to undo. On the other hand, electricity may be used to separate water into hydrogen gas and oxygen gas. If these two gases are mixed and sparked, they turn instantly back into water. Both of these changes are chemical processes, and yet they are easy to reverse.

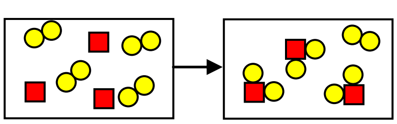
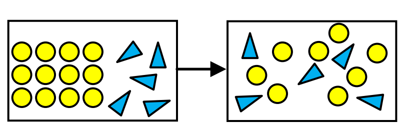
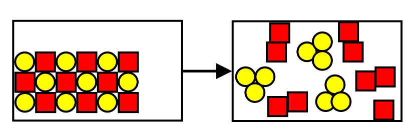
Pre-Lab questions:
Consider the following changes and decide whether they represent Physical or Chemical changes.

__ 1. A wooden match is burned. __ 2. Steam condenses on a mirror __ 3. An iron nail rusts

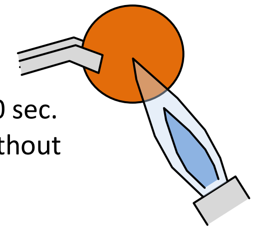
__ 4. A light stick glows for hours __ 5. A gold wire shrinks as it cools __ 6. An ice cube is melted

__ 7. A lump of clay hardens __ 8. Mold grows on a slice of bread __ 9. Mustard stains a shirt

__10. Gasoline ignites in the engine cylinder __11. An oil drop creates a rainbow puddle on a wet street

__12.  __13.  __14. 

Go to each of the following thirteen stations and follow the procedures there. Record your observations and then decide whether you think the change is a physical one or a chemical one. (Circle P or C). Then explain your decision with a sentence or two. Several of the stations have two experiments to conduct. To keep you on the right track, only about eight or nine of the changes you will observe are chemical.



1 & 2) Old penny heated in a flame, then buffed by steel wool:

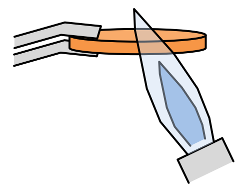
Using tongs, hold a penny by the edge vertically in the hottest part of a Bunsen burner flame for about 30 sec. When done, let it cool in the air for about 1 minute and observe. DO NOT TOUCH. Note any changes. Without Letting go, quench it in water, then take some steel wool and buff both sides until it looks like new again.

1) Heated:

Observations:	**P** or **C**	Explanation:

2) Buffed:

Observations:	**P** or **C**	Explanation:



3 & 4) New penny in a flame, then quenched in water:

Using tongs, hold a penny by the <u>very</u> edge horizontally in the hottest part of a Bunsen burner flame for 20-40 sec. Ignore any color changes this time. As soon as a very noticeable change occurs, quench the penny in water for 5 seconds, observe, and then discard it in the trash.

3) Heated:

Observations:	**P** or **C**	Explanation:

4) Quenched:

Observations:	**P** or **C**	Explanation:

5) a) Light bulb and electricity and b) Molybdenum glass

a) Turn the light bulb on and observe it. (It is a low wattage light bulb so it will not be too bright to look at.) Touch it briefly to determine if there is an intake or output of heat. b) Observe the glass bowl under fluorescent light, then turn the fluorescent light off and turn the incandescent light on. Observe. Leave it with the fluorescent light back on.

a) Observations:	**P** or **C**	Explanation:
b) Observations:	**P** or **C**	Explanation:

6) Potassium chlorate and sugar:

Measure out a very small scoop (about 0.1- 0.3 mL) of potassium chlorate/sugar mix onto the metal plate. Use the butane lighter to ignite the mixture. Observe the remains.

Observations:	**P** or **C**	Explanation:

7) Blaster balls:

Knock the two blaster balls together hard. Just once, please! Observe with all your senses!

Observations:	P or C	Explanation:

8) & 9) Iodine crystals heated, then cooled:

Place the test tube of iodine crystals in the hot water bath for 20 sec. Observe. Then place them in the cold water bath (half submerged) for 20 sec. Carefully observe the bottom inside surface of the test tube.

8) Heated:

Observations:	P or C	Explanation:

9) Cooled:

Observations:	P or C	Explanation:

10) Lead(II) nitrate and potassium iodide:

Place 1-2 drops of each together in the same spot on the petri dish. After observing, use a dropper of water to rinse off the sheet into the large beaker labeled "WASTE", then wipe the sheet dry with a paper towel for the next group to use.

Observations:	P or C	Explanation:

11) & 12) Supersaturated sodium acetate solution "activated," then heated in hot water bath:

Gently squeeze the disc in the sealed bag to activate the sodium acetate solution inside. Observe with eyes and hands! Then heat the bag by placing it in the large beaker of hot water. Observe.

11) Activated:

Observations:	P or C	Explanation:

12) Heated:

Observations:	P or C	Explanation:

13) & 14) Red oil and blue water shaken, then allowed to settle:

Vigorously shake the bottle, then allow the bottle to stand undisturbed for 1 minute.

13) Shaken:

Observations:	P or C	Explanation:

14) Settled:

Observations:	P or C	Explanation:

15) Potassium permanganate and water:
Fill a small test tube half way with water. Add 2-3 tiny crystals of potassium permanganate. Stopper and shake for 10 seconds. Rinse the test tube out in the sink.

Observations:	P or C	Explanation:

16) & 17) NaOH added to phenolphthalein, then swirled:
Add 2-3 drops of NaOH solution to the flask filled with phenolphthalein solution. Observe. Then stopper the flask and swirl to mix the contents.

16) Added:

Observations:	P or C	Explanation:

17) Swirled:

Observations:	P or C	Explanation:

18) Hot wax and copper wire:
Quickly dip the end of a copper wire into the beaker of molten wax. Remove the wire and observe. Peel off the coating and place it back in the beaker.

Observations:	P or C	Explanation:

19) & 20) Magnesium and HCl, then exposed to a flame:
Place 4-5 mL of HCl solution in a 25 mL graduated cylinder. Take a 1-cm length of magnesium ribbon and drop it in the cylinder. After 10 seconds, insert a flaming wood splint down into the open end of the cylinder. Lower it gradually in, and do not drop it. Wait for the "POP!" Pour the contents out into the large beaker labeled "waste."

19) Mg & HCl

Observations:	P or C	Explanation:

20) Exposed to a flame

Observations:	P or C	Explanation:

Law of Conservation of Mass Lab Name: __________________ Partner: __________________

During chemical and physical changes, substances may be broken down, new substances may be produced, colors may change, solids may turn into gases… but at no point in time are atoms ever created; at no point in time are atoms ever destroyed. Atoms are simply rearranged as old bonds between them are broken and new bonds are formed. Since the mass of a system depends only on the atoms present, it should make sense then that this mass must always remain constant during any kind of change. A block of ice will have the exact same mass as the water from which it was made. When silver reacts with sulfur to form a black silver sulfide coating known as tarnish, this tarnish will have the exact same mass as the silver and sulfur had together before they reacted. This phenomenon is known as the **Law of Conservation of Mass**, and all substances and systems must follow this law. The products of a reaction must always have the same total mass as the reactants had to start off with. And yet, consider the burning of a log. After the fire has burned out, the ashes remaining do not weigh anywhere near as much as the log from which those ashes came. Is this a violation of the law? Or is there some other explanation? Write down your thoughts below:

In this lab we will attempt to confirm this law of conservation of mass, and perhaps learn that it isn't always as easy as it may seem to observe this law actually being followed!

Procedure

Part I:

1. Place 1 level spoonful of baking soda in one clean dry cup. Place 25-30 mL of vinegar in the other. Put <u>both</u> cups side by side on the scale. Record the initial total mass of this "system" in the table at right.

initial total mass	
final total mass	

2. Being careful not to spill anything, slowly pour the vinegar into the cup with the baking soda. Observe the reaction – with eyes, ears, nose and fingers (on the outside of the cup)! Record all those observations:

3. Before you put the cups back on the scale, make a prediction. What do you think the final total mass of the system will be? Give an exact number:______________

Explain your prediction: ____

4. When the reaction appears to be over, place <u>both</u> cups on the scale, again side by side. Record this final total mass of system in the table above.

5. So, what happened to the mass of the system? ______________ How close was your prediction? __________

6. Why did the mass of the system do this? ____

Does this mean the Law of Conservation of Mass is wrong? ______

What could be done differently – using this same reaction—to show that mass really is conserved? ____

Part II:

1. Place 1 level spoonful of baking soda in the corner of a ziplock bag. Twist the bag and then wrap a twist tie tightly around it. Be careful not to tear or puncture the bag in any way.

2. Pour 25-30 mL of vinegar in the bag. Then zip it closed, trying to get as much of the air out of the bag as you can. Then place the bag on the scale and record the total initial mass of this new system in the table:

initial total mass	
final total mass	

3. Now (and you probably saw this coming!) undo the twist tie, and allow the baking soda and vinegar to react with one another – this time in a system that is closed off so that no gases can enter or escape. Again, observe the reaction – with eyes, ears, nose and fingers! Record all those observations:

4. Before you put the system back on the scale, make a prediction. What do you think the final total mass of the system will be? Give an exact number: _______________

Explain your prediction: ____

5. When the reaction appears to be over, place the bag <u>and the twist-tie</u> on the scale. Record this final total mass of system in the table above.

6. So, what happened to the mass of the system? _______________ How close was your prediction? ___________

7. Why did the mass of the system do this? (There may be a slight leak out of the bag, but there is no way that a little leak could account for the large change in weight you observed… so come up with a better answer. (Hint: think about what you have learned about apparent weight loss due to fluid displacement.)

Does this mean the Law of Conservation of Mass is wrong? ______ You should be able to think of <u>two</u> different modifications you could make – again using this same reaction—to show that mass really is conserved. One would be very easy, the other much more difficult. What would be the easy modification? ___________________________________

What would be the more difficult modification? ____

Part III:

Look at the remaining equipment you have available to you. Think of a procedure that you think will overcome both shortcomings from part I and II and finally allow you to verify that during this chemical reaction, the mass of the system does in fact remain constant. Write down this procedure in number steps below. Then show it to your instructor. Once he/she has approved your procedure, go ahead and try it.

Record your data in the table at right:

initial total mass	
final total mass	

If there was a decrease, even a slight one, can you think of two possible reasons why the mass may have done that?

____ ____

After a few minutes have past, go back and re-weigh the bottle: ____________
Has the mass decreased any further? ______ What does this tell you concerning your two possible explanations

above? ____

Now observe the two "identical" soda bottles on the middle lab station. One is pressurized, the other isn't. Do they look like they are the same size? ____ Look carefully at their labels; and see if they are tightly wrapped or is there any slack. What difference do you notice?

What does this tell you concerning
your two possible explanations above? ____

CLEAN-UP: Rinse out the cups, the bags and the bottle. Dry out the cups and the bag as best as you can and leave them open to the air to dry some more.

Follow-up questions:

1. Precisely 10.00 g of salt are dissolved in a few teaspoons of water. This is then poured into a 5.00 g petri dish and set on a window sill to allow the water to evaporate. Three days later, with all the water apparently evaporated away, the petri dish – now containing a layer of recrystallized salt – is weighed: 15.73 g. What should it weigh? _______
What do you think happened?

2. Joey and Jimmy both weigh precisely 95.00 kg. Joey eats 1.00 kg of chocolate, while Jimmy drinks 1.00 kg of water. Now how much does Joey weigh? __________ and how much does Jimmy weigh? ____________. If they continue doing this day after day for the next four months along with the rest of their daily routine, will their weights continue to be the same? _______
Explain:

3. If mass is conserved, how is it possible for a person on a low Calorie diet to lose weight?

4. If a person on a diet were in an air tight room, would it be possible for that person to lose weight? ______ (Assume the room contained a sufficient supply of oxygen.) If the entire room (with the person in it) is weighed, will it increase, decrease or remain the same? _________________.
So when a person does lose weight, what
does that mean about the mass of the rest of the universe? ____

5. How might you modify the third procedure (the one with the bottle) to get rid of even the tiniest loss in weight?

6. Sally's results for part II (with the Ziploc bag) were: initial total mass: 34.67g final total mass: 34.19 g. The density of air is 0.00112 g/mL. Assuming no gas leaked out of the bag, what volume of gas (carbon dioxide) was produced from the baking soda & vinegar reaction? Hint: Use Archimedes Principle and focus on the apparent weight lost by the bag. Show your work.

Ans: ___________

7. Sally's results for part III (with the bottle) were: initial total mass: 72.875g final total mass: 72.871g. Again using Archimedes Principle and assuming no gas leaked out, by how much did the plastic bottle expand due to the increased pressure pushing outward?
Show your work.

Ans: ___________

8. A large open flask weighs 468.34 g. 11.76 g of water are added to the flask. Then a 9.28 g cork is placed inside. It floats on the water. a) So how much does the entire system (flask and contents) weigh now? b) Eventually, all the water evaporates away and the cork now sits on the bottom of the flask; what is the total mass of the system now? c) A 0.27 g fly flies into the flask. How much will the entire system weigh with the fly resting on the bottom of the flask? d) How much will the system weigh while the fly is hovering in mid-air inside the flask? e) The fly gradually flies upward and out of the flask. At what point does the mass drop back down to its previous (fly-less) weight?

a: ____________

b: ____________

c: ____________

d: ____________

e: ____

9. Along those same lines… A man weighing 0.100 tons is driving a truck weighing 10.00 tons and is hauling 2.00 tons of live chickens inside the truck. How much does the entire system weigh? ___________He comes to a bridge which he has to get across but there is a sign that says: "Weight limit: 11.00 tons." He thinks he will have to turn back, but then he gets this great idea: he goes out and bangs on the sides of the truck, and that gets all the chickens inside to start flapping their wings and flying around inside the truck. He quickly hops in the truck and starts to drive it across the bridge. Will this work? Explain.

10. 17.52 g of zinc and 38.47 g of iodine are placed together in a 117.29 g container. Zinc and iodine react to form zinc iodide. After one hour, 8.60 g of the zinc has reacted, and 29.88 g of zinc iodide were produced. a) How much iodine was left in the container? b) How much did the entire system (container and contents) weigh to begin with? c) How much did the entire system weigh when the reaction was finished? d) 67.90 g of water are then added and all of the zinc iodide dissolves in the water. How much does the system weigh now? e) How much does just the zinc iodide solution in the container weigh?

a: ____________

b: ____________

c: ____________

d: ____________

e: ____________

Law of Conservation of Mass WS Name: ________________________

1. State the Law of Conservation of Mass.

2. From an atomic point of view, why does this law have to be true?

3. Some copper and sulfur are mixed together in a beaker and the entire system (copper, sulfur and beaker) is weighed: 145.78 g. They are then heated and the copper and sulfur react to form copper sulfide. a) How much will the system weigh after the reaction has finished? b) The beaker weighed 88.32 g. What was the mass of the contents of the beaker? c) There was no sulfur left in the beaker, but it turns out that there was 17.57 g of unreacted copper left in the beaker. What mass of copper sulfide was produced?

a: ________

b: ________

c: ________

Ans (iro+1): 0 3.58 4.70 14.59 36.00 38.36 39.89 57.46 75.66 88.00 145.78
Units (iro+1) g g g g g g g g g mg kg

4. Calcium nitride can be decomposed into its two elements: _____________ and ____________. A 18.89 kg sample of calcium nitride is decomposed. If 15.31 kg of calcium are produced, how much nitrogen would be produced?

Ans: _______

5. 32.00 g of methane react with 128.00 g of oxygen. They produce 72.00 g of water vapor and __________ g of carbon dioxide.

6. 16.78 mg of element X is reacted with 34.64 mg of element Y. They produce 46.72 mg of XY compound. All the X got used up during the reaction, but there is some Y left over. How much?

Ans: _______

7. Calcium carbonate is a compound made up of the elements: calcium, carbon and oxygen. 75.00 g of calcium carbonate are decomposed. If 30.00 g of calcium and 9.00 g of carbon are produced, how much oxygen is produced?

Ans: _______

8. 24.32 g of aluminum and 65.93 g of fluorine were placed together in a flask and allowed to react to produce the compound aluminum fluoride. All the aluminum reacted, but only 51.34 g of fluorine reacted. a) How much unreacted fluorine was left in the flask? b) How much aluminum was left in the flask? c) How much aluminum fluoride was produced?

a: _______

b: _______

c: _______

SEPARATION TECHNIQUES Name:____________ Partner: ___________

Purpose: To investigate a range of different separation techniques. At each station, sketch and label the set up and answer the questions.

Station A: Distillation:

A1) Describe how distillation would work to separate salt from water.

A2) Which way does the water flow through the outer sleeve of the condensing tube: uphill or downhill? ____________
Why?

A3) A boils at 60*C; B boils at 90*C. If a 50-50 mixture of A & B is placed in the distillation set-up, will the distillate contain mostly A or B? ____ Explain:

A4) If the distillation described in #2 above continues to run for a very long time, what will eventually be left in the heating flask?

And what will end up in the collection flask?

Label: boiling flask, mixture, distillate, condensing tube (& direction of water flow), heat source, collection flask

Station B: Filtration, Sedimentation and Centrifugation:

B1) Whereas distillation relies on the components having different boiling points, what physical property does filtration rely on?

B2) Why would filtration not be effective at separating a mixture of salt and water?

B3) Would sedimentation work to separate a mixture of salt and water?

B4) What physical properties do sedimentation and centrifugation rely on?

B5) Which would still work in zero gravity:
Filtration? Y / N Sedimentation? Y / N
Centrifugation? Y / N
Explain each answer:

Filtration	Centrifugation

Label: funnel, filter paper, residue, filtrate, collection flask, centrifuge, centrifuge tube, precipitate, supernate

Station C: Paper Chromatography

C1) If the black ink had been a pure substance rather than a mixture of different dyes, how would the chromatography turn out?

C2) What physical property does paper chromatography rely on to separate one dye from another?

C3) If the chromatography is not stopped, what will you eventually end up with?

C4) Describe a situation in which this technique might be useful in a forensics (crime) lab?

Label: dye mixture, stationary phase, mobile phase (& direction of mobile phase movement)

Station D: Electrolysis

D1) What is the chemical formula for water? ______

D2) How can you tell by looking at the set-up which gas is hydrogen and which is oxygen?

D3) At which electrode (+) or (-) did the hydrogen gas form. ___

Knowing what you know about charges, why does this make sense?

D4) In what way is this electrolysis very different than each of the other separation techniques?

D5) When the two gases are mixed back together and reacted (by the instructor) what compound is produced?

Label: positive electrode (anode), negative electrode (cathode), power supply, water, hydrogen gas, oxygen gas

HOME LAB #2: MIXTURE SEPARATION LAB Name: ______________

Do the following experiment with one or both of your parents or guardians. Have him/her/them sign it below to confirm that they worked with you from beginning to end. NO SIGNATURE, NO CREDIT!

Signature of Parent(s)/Guardian(s): ________________________

Introduction: Separating out the different components in a mixture can often prove quite challenging, yet separation and recovery are extremely important operations both for research and for industry. In this lab you will have an opportunity to develop and then implement your own self-designed procedure in an attempt to separate out and recover the components of a given mixture. The mixture contains the following four substances: salt, sand, iron filings, and poppy seeds, all essentially in dry granular form.

Materials (anything you have around the house may be used). Below is a list of some materials that you may or may not want to use (just to get you started): cellophane, plastic cup, spoon, tape, paper clip, cotton, aluminum foil, coffee filter, plastic straw, clothes pin, refrigerator magnet, scissors, water, paper towel, eye dropper, plastic bag, wax paper...

Procedure

1. Obtain from the instructor a sample of the mixture and examine it carefully. Place a small, representative portion of the mixture (about 20% of what you are given) in the corner of a plastic sandwich bag, twist it closed, cut it off just above the twist and tape this small plastic sack to the lab sheet in the space below.

2. Using any materials you want, develop and implement a procedure that will enable you to separate out the remaining 80% of the mixture and recover all four components, each in as pure and dry a state as possible.

[Helpful Hint: If water is part of your procedure, be careful not to use too much!]

3. As each component is separated off, place it in a small plastic sack (as described above) and tape it to this lab sheet in the appropriate space below.

SEPARATED AND RECOVERED COMPONENTS

representative portion of mixture	salt	sand	iron filings	poppy seeds

QUESTIONS:

1. In short, concise, numbered steps, write out the procedure you followed. It should be written clearly and thoroughly enough so that anyone not familiar with the lab could reproduce your results simply by reading and following your procedure. Include at least one diagram at right.

2. In your estimation, how successful were you (on a scale of 1-10) in separating and recovering each of the four components?

Sand: ___ Salt: ___ Iron filings: ___ Poppy seeds: ___

3. What made you decide to do your procedure steps in the order that you did them? Would any order have worked?

4. If you were to do the lab over again, what specifically might you do differently?

5. For each of the four components, describe a specific physical property that enabled you to separate it from the rest of the mixture.

salt -

sand -

iron filings -

poppy seeds -

6. What might you do to determine the purity of each of your recovered components – give at least two answers – Hint: think intensive properties?

7. How might you separate out each of the following 2-part mixtures:

a) lead filings and iron filings?

b) sand and gravel

c) sand and styrofoam (both very finely ground)

d) salt and sugar?

e) alcohol and water?

f) nitrogen and oxygen?

Law of Definite Proportions Lab Name: _______________ Partner: _____________

What evidence is there that matter is actually made up of atoms? In the early 1800's, chemists like John Dalton were proposing a theory that matter was made up of very, very tiny indestructible particles called atoms, but what evidence did they have back then that atoms might exist? There were no scanning tunneling microscopes or computer generated images to support this theory. But one thing they did have was an observation made by Joseph Proust in 1796 that compounds could be broken down into their component elements and that when they were, the same mass ratio of these elements would always be observed. When the compound zinc sulfide, for example, is decomposed into its two elements: zinc and sulfur, it doesn't matter how much of the compound you start with, the proportions will always be the same: 2.04 parts of zinc for every 1 part of sulfur. How can this be? How can the elements "know" the proportions in which they are supposed to combine? And why is this evidence supporting the theory that matter is made up of atoms? Let's see if we can find out.

In this lab, you will do two different experiments to try to confirm the Law of Definite Proportions: In the first, you will decompose copper(II) oxide, and in the second, you will produce lead(II) iodide. In both cases, you will carefully weigh the pieces and the whole. Then you will compare your mass ratios to ones obtained by other groups – who used completely different amounts – to see how consistent this proportion really is.

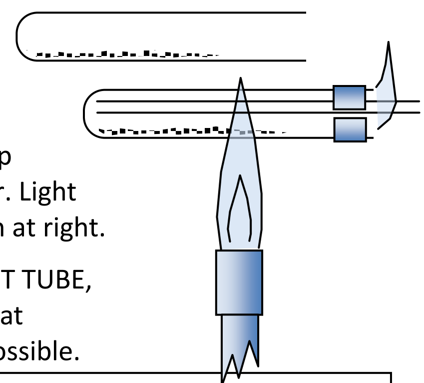
Procedure: Goggles must be worn at all times:

Part I:

1. Obtain an empty test tube, and weigh it on an electronic balance. Record this mass in the data table below.

2. Obtain a sample of copper(II) oxide. Pour it in the test tube and weigh the test tube again. Record this mass in the table. [Note: these samples were portioned out for you to make sure a wide range of different masses are used.] Describe the appearance of this copper(II) oxide:

3. Holding the test tube horizontally, carefully spread the sample out to evenly cover the bottom two thirds of the test tube in a thin layer as shown at right. If any spills out, you will need to reweigh the test tube and copper(II) oxide.



4. Then clamp the test tube to the ring stand. Make sure it is clamped right up near the lip of the test tube and that it is oriented horizontally about 3 cm above the top of the burner. Light the burner and adjust it to a flame with the hottest part just hitting the test tube as shown at right.

5. Then, MAKING SURE YOU ARE NOT STANDING DIRECTLY IN THE PATH OF THE OPEN TEST TUBE, hold a flame by the mouth of the test tube and VERY slowly turn the other gas jet on so that methane is pumping slowly across the hot copper(II) oxide. Make this flame as small as possible.

6. Observe the reaction for a while, then move the flame to a new section of the sample and heat that area for a while. Continue heating for 10-12 minutes. Go on to part II of the procedure.

Record observations about the reaction in the space below:

7. After the sample has completely changed color, turn off the burner but keep the flame burning at the mouth of the tube for three minutes, then turn that flame off, let the tube cool for one more minute, then weigh it, and record the final mass in the table at right.

mass of empty test tube	
init. mass of tt & sample	
final mass of tt & sample	

Part II:

1. Obtain an empty 100 mL beaker and weigh it. Record this mass in the table below. Obtain a sample of lead(II) nitrate. Add it to the beaker and weigh it again. Note: lead(II) nitrate is a compound that is 62.62% lead, 8.44% nitrogen and 28.94% oxygen by mass.
Describe the appearance of this compound:
Also, go over and pick up in one hand the bottle filled with this compound and in the other hand, a bottle filled with potassium nitrate. What difference do you notice?

2. Add 15-20 mL of distilled water to the beaker. Distilled water is pure water, unlike tap water which has some impurities in it. Swirl the beaker for a minute or so and observe what happens.

3. Add 3-4 mL of sodium iodide solution from the bottle labeled NaI and observe the reaction: Swirl the beaker.
The **lead** from the lead(II) nitrate is combining with the **iodine** from the sodium iodide solution. This makes the bright yellow compound: lead(II) iodide. The other two components, the sodium and the nitrate, stay dissolved and are therefore invisible in the surrounding water.

4. Write your name and period number on the edge of a piece of filter paper. Then weigh it. Record that mass in the table below. Then fold it in quarters and place it in a glass funnel, supported by a ring stand over a 250 mL beaker. Swirl the small beaker containing the lead(II) iodide and then pour it gradually into the funnel, careful not to let the liquid level rise above the filter paper. Describe the color and clarity of what comes through the filter paper:

5. Allow the liquid to drain for a while into the big beaker, then pour a few mLs of distilled water into the small beaker, swirl it around to pick up any lead(II) iodide that may be left in there, then pour it into the filter. Repeat this process a few more times to get as much lead(II) iodide as possible transferred from the small beaker into the filter funnel. [Go back to procedure part I: Is it time to turn off the burner??]

6. After the filtration appears to have stopped, pour 5 more milliliters of distilled water into the funnel and allow that to drain. [Is it time to turn off the flame at the mouth of the test tube?]

7. Just so you are familiar with the terms: the yellow pasty substance that was caught in the filter paper is called the "residue." The clear liquid that dripped through the funnel and collected in the large beaker beneath is called the "filtrate." As a quick test add a few drops of the sodium iodide solution to the filtrate in the large beaker. Record what you observe:

8. When the filtration has finished dripping, carefully remove your filter paper from the funnel and place it on a petri dish and place this on a sunny window. In a few days, when it is dry, we will weigh it again, and record that final mass in the table below. [Go back and complete part I.]

mass of empty beaker	
mass of beaker and lead(II) nitrate	
mass of empty filter paper	
mass of filter paper and dried residue	

Calculations: part I:

Show your work here:

Determine the mass of copper left in the test tube: _________

Determine the mass of the oxygen that had been combined with it: _________

Determine the mass ratio copper:oxygen: ___________

Record the mass copper:oxygen ratio obtained by three other lab groups: ___________ ___________ ____________

Calculations: part II: **Show your work here:**

Determine the mass of lead(II) nitrate placed in the beaker: ___________

Determine the mass of just the lead in that lead(II) nitrate:___________
Hint: Go back and look at the percent lead in lead(II) nitrate.

Determine the mass of lead(II) iodide collected in the filter paper: __________

Determine the mass of just the iodide in that lead(II) iodide: ___________

Determine the mass ratio lead:iodide __________

Record the lead:iodide mass ratio obtained by three other lab groups: ___________ ____________ ___________

Follow-up questions:

1. State the Law of Definite Proportions: ____

____

____

2. Do you think your results for part I supported this Law? ______ How about part II? ________

Justify your answers:____

3. Was the change taking place inside the test tube chemical or physical? __________

Justify your answer:____

4. What evidence was there that the element copper was left in the test tube? _____________________________

5. How do you think the copper oxide was produced in the first place? _________________________________

6. With that in mind, why do you think you were instructed to leave the methane blowing across the hot copper until it had completely cooled down? ____

7.* Consider your copper:oxygen mass ratio. Would it have been made H (too high) L (too low) or N (not affected) by each of the following potential error sources:

___ The sample was not heated enough and so not all of the copper oxide was converted to copper.

___ As you were first pouring the copper oxide into the test tube, a little spilled out.

___ Just after you weighed the test tube and copper oxide, a little spilled out.

___ The test tube had a little moisture in it that got boiled off during the heating.

8.* Look at a periodic table. Although we haven't yet covered it in class, the decimal numbers shown there are the relative masses of the atoms. So for example, the average mass for an atom of nitrogen (N) is 14.007 amu. What are the average atomic masses for Cu ? ____________ O? _____________.
Consider the following three possible formulas for copper oxide: CuO (meaning one atom of Cu for every one atom of O), Cu_2O (meaning two atoms of Cu for every one atom of O), and CuO_2 (meaning one atom of Cu for every two atoms of O). Calculate the mass ratio for each of these possible formulas. (For example, in NH_3, 14.007 amu of N would combine with 3 x 1.008 = 3.024 amu of H, so the mass ratio would be 14.007 ÷ 3.024 = **4.632)**
CuO = _________ Cu_2O = ___________ CuO_2 = ____________

Based on your results, which formula do you think is correct for the copper oxide you used in the lab? __________

9. In part II, step #1, what evidence was there that lead was present in the original white crystalline compound you started with? ____

10. In part II, step #2, was the change taking place inside the beaker chemical or physical? ______________

Justify your answer:____

11. In part II, step #3, was the change taking place inside the beaker chemical or physical? ______________

Justify your answer:____

12. In part II, step #4, was the change taking place inside the funnel chemical or physical? ______________

Justify your answer:____

13. In part II, step #8, as the filter paper set overnight, was the change chemical or physical? ______________

Justify your answer:____

14. The lead and iodide formed a compound lead iodide which was not soluble in water, and that's why it instantly came out of solution as a cloudy yellow powder. The other two components, potassium and nitrate, remained dissolved in solution and really didn't do much during the reaction. These are known as "spectators." Explain why this name is appropriate: ____

15. Why do you think you were instructed to add a few more drops of potassium iodide solution to the filtrate in step #7? ____

16. If the filtrate had turned yellow, what would that have meant? _______________________________________

17.* Consider your lead:iodide mass ratio. Would it have been made H (too high) L (too low) or N (not affected) by each of the following potential error sources:

___ The sample was not given enough time to fully react before being poured through the filter.

___ While you were swirling the lead(II) nitrate and water, a little spilled out.

___ As you were first pouring the filtrate and residue into the funnel, a little spilled out.

___ The small beaker had a little distilled water in it from the very beginning.

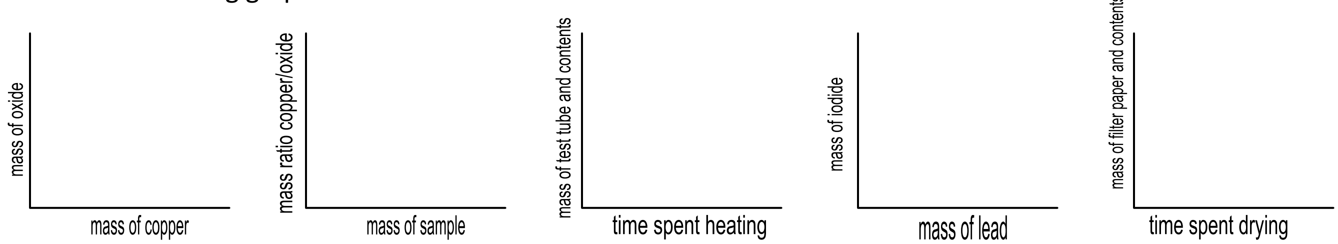
___ The sample was not given enough time to fully dry out on the filter paper before the final weighing.

18.* Look at a periodic table. What are the average atomic masses for Pb ? ____________ I? _____________. Consider the following three possible formulas for lead iodide: PbI (meaning one atom of Pb for every one atom of I), Pb_2I (meaning two atoms of Pb for every one atom of I), and PbI_2 (meaning one atom of Pb for every two atoms of I). Calculate the mass ratio for each of these possible formulas.

PbI = _________ Pb_2I = ___________ PbI_2 = ____________

Based on your results, which formula do you think is correct for the lead iodide you produced in the lab? ________

19.* Sketch the following graphs:



20. In part I, what was the independent variable? _______________ dependent variable? ______________

21. Is elemental composition an intensive property or extensive property? ___________________

Justify your answer: ____

BEVERAGE DENSITY LAB: Sugar content analysis Name: _________________ Partner: ________________

So, how well do you know the beverages you drink? Try to rank the following four from lowest sugar content to highest: Coke, Welch's 100% Grape Juice, Gatorade, Orange Soda

______________________ ______________________ ______________________ ______________________

(lowest sugar content) (highest sugar content)

Along with adding Calories to a beverage, dissolved sugar also increases the density of the solution. In this lab, you will first determine the density of five known (standardized) sugar solutions: 0% (which is just plain water, not colored), 5% (colored yellow), 10% (colored green), 15% (colored blue), and 20% (colored purple). You will then plot these densities on graph of density as a function of sugar content. Finally, you will determine the densities of the four beverages, then use the graph to approximate their sugar contents.

Procedure:

The instructor will demonstrate the proper technique for using a volumetric pipet.

Sample	Mass(g)	Density (g/mL)
0% sugar		
5% sugar		
10% sugar		
15% sugar		
20% sugar		

Coke		
Grp Juice		
Gatorade		
Or. soda		

1. Place the beaker on the balance and hit the "tare" (re-zero) button. The scale should read "0.00 g." Draw up a precisely measured 10.00 mL of 0% sugar solution into the pipet. Then empty it into the beaker, touching the tip of the pipet to the inside wall of the beaker to help get out most of the liquid in the tip. Do not try to shake out any liquid that remains there. The pipets are designed "TD" ("to deliver") 10.00 mL and that remaining drop should not be squeezed out. Since the beaker has already been zeroed out, the mass is that of the liquid alone. Record this mass reading in the data table at right, then push the "tare" button to re-zero the scale for the next reading.

2. Touch the pipet to a paper towel to get rid of the excess liquid in the tip. Then repeat step 1 with each of the remaining sugar solutions, and then with each of the four beverages. Do not put any of the solutions back into the cups from which they came, just leave them in the beaker. When the beaker gets full, simply empty it into the sink, set it back on the scale and push the "tare" button.

Calculations:

Calculate the density of each solution. Enter these values in the boxes alongside the data table. (Hint: Since the volume is always 10.00 mL, this should be easy; you don't even need a calculator!)

Graph:

On the last page, carefully plot a graph of the calculated densities of the five known sugar solutions as a function of their sugar content (%). Then use a ruler to draw a best fit straight line through the point (DO NOT PLAY CONNECT-THE-DOTS!!!). Then use the densities of the four beverages to approximate their sugar contents (%). To do this, start on the y-axis at a density of one of the beverages, then follow the line over to where it hits the best fit line you drew, then go straight down to the x-axis to determine the corresponding sugar content value. As with measurements, read your graph one place past the scale.

Results:

Below, list the sugar content values (%) you approximated from your graph. Also, list the values obtained by some other lab group (their names: ________________ ________________)

your values: Coke _________ Grape Juice _________ Gatorade _________ Orange Soda _________

their values: : Coke _________ Grape Juice _________ Gatorade _________ Orange Soda _________

Follow-up Questions:

1. How well did you do in guessing the sugar-content rank of the four beverages?

2. Water has a density of 1.00 g/mL at 4*C. Was your room-temperature value for the density of water (0% sugar) higher or lower than 1.00? Why do think it was off in that direction?

3. Why were you told not to put solutions back into the cups from which they came?

4. Why was it OK to leave liquid in the beaker from one trial to the next?

5. In this lab, you discarded all of the solutions into the sink. Can you think of situations where this would not be appropriate?

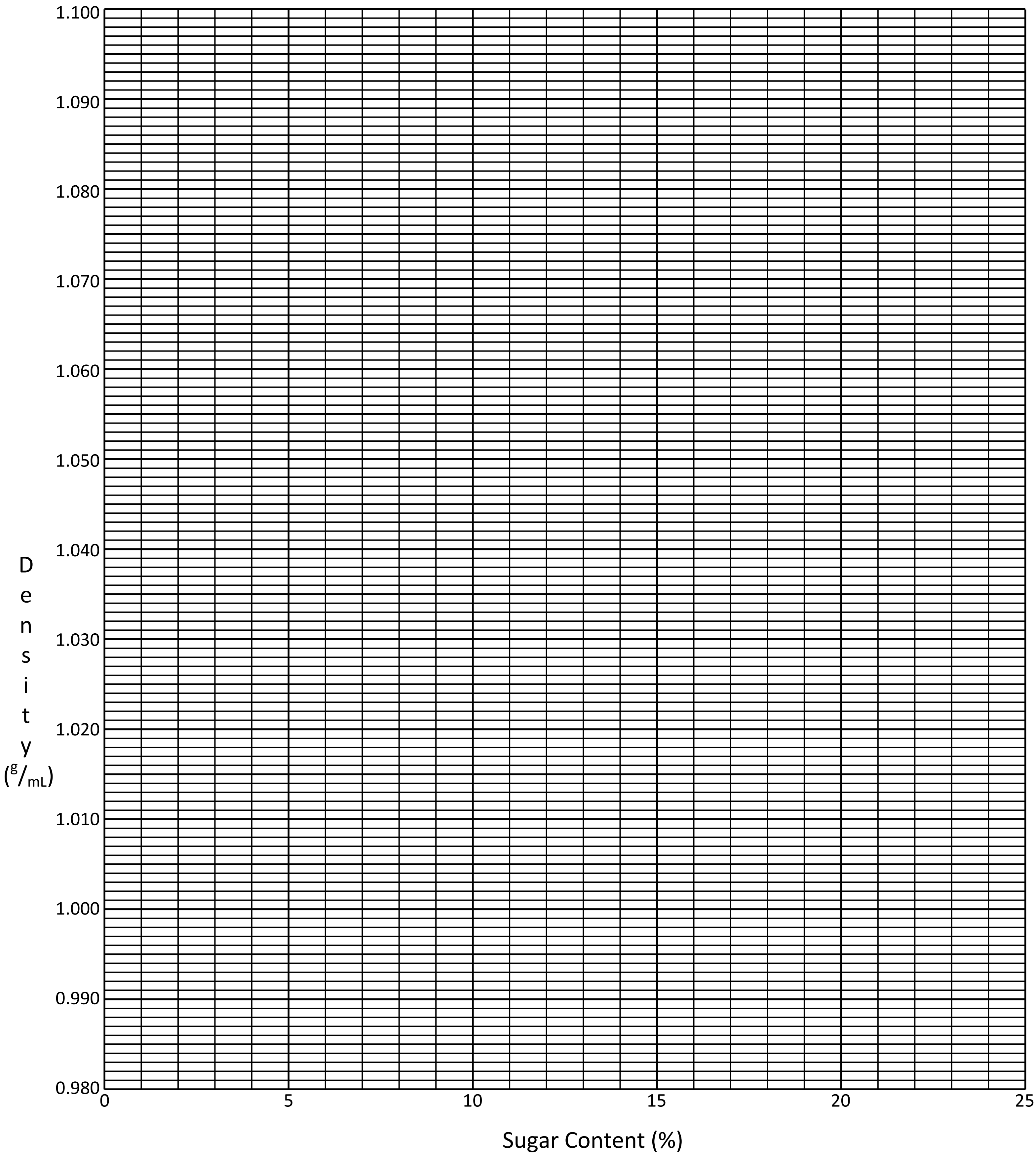
6. List three liquids found around your house that should not be discarded into the sink.

7. This lab focuses on sugar content. What assumption does the lab make concerning all the other ingredients in each of the beverages?

8. When plotting data such as these, why is it not appropriate to connect the dots in the graph you draw? If you were to redo the lab, do you think you would get the same results?

9. How precisely would you go about making up some 12% (by mass) sugar solution? Precisely how much sugar would you use, and how much water? [This is an easy question: just think what you would do if you were told to make up a group of people that was 12% male.]

10. Draw a sketch of the volumetric pipet. What makes it so precise? Why do you think it bulges in the middle?


1.100
1.090
1.080
1.070
1.060
1.050
1.040
1.030
1.020
1.010
1.000
0.990
0.980
Density (g/mL)
0
5
10
15
20
25
Sugar Content (%)

Law of Definite Proportions and Law of Multiple Proportions WS Name: ___________________

Once again, the Law of Definite Proportions states that whenever a compound is analyzed, it will always contain the same proportions (same ratio) by mass of elements. This was significant because it provided strong evidence in support of the Atomic Theory. Elements A and B combined in fixed proportions because (at a level that was so small it couldn't be seen) atoms of A were bonding to atoms of B in fixed whole number ratios. Then along came a closely related observation: Under different experimental conditions, the same two elements could be made to combine in a different proportion and thus form a different compound. But this different proportion was always a whole number multiple of the first proportion. This came to be known as the Law of Multiple Proportions.

For example, let's say 8.00 g of A combine with 1.00 g of B to make a certain AB compound. The mass ratio A/B would be 8.00 g/ 1.00 g = 8. Under different experimental conditions, 8.00 g of A might combine with only 0.50 g of B to form a completely different compound with completely different properties. This gives a mass ratio A/B of 8.00g/0.50g = 16. Note how this new proportion (16) is precisely twice as big as the first proportion (8). This was explained within the context of the atomic theory by hypothesizing that in this new compound, the atoms were simply combining in a different whole number ratio than they were in the first compound. For example, water can be made from the elements H and O and its formula is H_2O. But under different circumstances, these same two elements can be made to form a completely different compound: hydrogen peroxide whose formula is H_2O_2.

For the following problems, decide whether the experimental evidence supports…
D: the Law of Definite Proportions, **M**: the Law of Multiple Proportions or **N**: neither (maybe the experiments were so poorly done that the data are too inaccurate to support either law). Important: Always look at element ratios (one mass divided by the other) and not at things like percentages which can be misleading.

D D D
M M
M M
N N N

1. Exp #1: 45.62 g of P combines with 17.83 g of Q. Exp #2: 21.87 g of P combines with 8.58 g of Q. ___

2. Exp #1: 45.62 g of P combines with 17.83 g of Q. Exp #2: 37.82 g of Q combines with 48.38 g of P. ___

3. Exp #1: 45.62 g of P combines with 17.83 g of Q. Exp #2: A PQ compound is found to be 16.83% P . ___

4. Exp #1: 17.76 g of X combines with 27.88 g of Y. Exp #2: 12.50 g of XY is found to contain 9.94 g Y . ___

5. Exp #1: 17.76 g of X combines with 27.88 g of Y. Exp #2: An XY compound is found to be 38.9% X . ___

6. Exp #1: 17.76 g of X combines with 27.88 g of Y. Exp #2: An XY compound is found to be 24.2% X . ___

7. Exp #1: A JK compound is found to be 32.8% J. Exp #2: A JK compound is found to be 32.8% J. ___

8. Exp #1: A JK compound is found to be 32.8% J. Exp #2: A JK compound is found to be 10.9% J. ___

9. Exp #1: A JK compound is found to be 32.8% J. Exp #2: A JK compound is found to be 59.4% J. ___

10. Exp #1: A JK compound is found to be 32.8% J. Exp #2: A JK compound is found to be 46.4% J. ___

11. State the Law of Definite Proportions:

12. A copper sulfide compound is analyzed and found to contain 17.5 g of copper combined with 8.83 g of sulfur. a) What percent copper and what percent sulfur are present in the compound? b) A second sample of the same compound is analyzed. If it contains 12.6 g of copper, how much sulfur would be in it? c) A third sample is analyzed and found to contain 148.4g of sulfur. How much copper would there be? d) If a 3.67 mg sample of the same compound were analyzed, how much copper and sulfur would it contain? Which Law does this illustrate?

a: ________ ________

b: __________

c: __________

d: _________ ________

e: ___________________

13. A 67.5 g sample of the compound glycerol is analyzed and found to contain 35.2 g of oxygen (O) and 5.9 g of hydrogen (H). a) If the only other element it contains is carbon (C), what mass of carbon would it contain? b) A second sample of glycerol is analyzed and found to contain 27.6 g of C. How much O and how much H would it contain? c) What are the percents of each element present in the compound? d) How much C, H and O would be present in a 32.8 kg sample of glycerol? e) If a sample of glycerol is analyzed and found to contain 42.7 g of H, how much did the original sample weigh?

a: ________

b: _________ ________

c: ________ ________ _______

d: ________ ________ _______

e: __________

14. How does the Law of Definite Proportions relate to concept that matter is made up of atoms?

Ans (IRO + 1): 1.23
2.44 2.9 6.2 6.36
8.7 12.8 17.1
26.4 33.5 36.8
39.1 44.3 52.1
66.5 294 490

Un its (IRO + 1): %
% % % % mg mg

Build-a-Bohr Lab Name: ______________________

In this activity, you will visit a computer simulation at the website: http://scratch.mit.edu/projects/beckerr/867623. In this simulation, you will have a chance to build your own atoms – deciding for yourself how many protons and neutrons to put in the nucleus and then seeing what's stable and what's not. You will find out what happens to an unstable nucleus, and you will also see how electrons organize themselves in distinct energy levels around the nucleus.

Read through all the questions on this worksheet first, then just start exploring and see what you uncover!

1. Which particles (protons, electrons or neutrons) determine what type of element an atom is? _____________

2. What two subatomic particles are found in the nucleus? _______________ ______________

3. What particle exists in specific energy levels around the nucleus? ______________

4. What charges do the three particles have? Protons: _____ Neutrons: _____ Electrons: _____

5. The mass number is always equal to the number of _______________ plus the number of _________________

6. Generally speaking, there are three types of nuclei: a) those that are too unstable to exist for any time at all, b) those that are somewhat unstable, but they last for a while and then change into something more stable, and c) those that are stable just as they are. Using isotope symbols, list four examples of each type. Make sure you use a different element for each of the twelve examples you give. Three have been done for you:

a) $^{7}_{2}\mathrm{He}$ ______ ______ ______ b) $^{16}_{7}\mathrm{N}$ ______ ______ ______ c) $^{23}_{11}\mathrm{Na}$ ______ ______ ______

7. Multiple choice:

__ Among the first twelve to fourteen elements, nuclei with the same number of protons and neutrons are
a) never stable b) rarely stable c) usually stable d) always stable

__ Nuclei which are unstable because they have more protons than neutrons are most likely to become more stable by undergoing: a) alpha decay b) beta decay c) electron capture d) proton emission e) positron absorption

__ Nuclei which are unstable because they have more neutrons than protons are most likely to become more stable by undergoing: a) alpha decay b) beta decay c) electron capture d) proton emission e) positron absorption

__ Very large nuclei which simply have too many protons all repelling each other are likely to become smaller and more stable by undergoing: a) alpha decay b) beta decay c) electron capture d) proton emission e) positron absorption

__ An alpha particle is actually the same thing as…
a) a hydrogen-1 nucleus b) a helium-4 nucleus c) a lithium-7 nucleus d) an electron e) a neutron

__ A beta particle is the same thing as... (same choices as above)

__ A proton is the same thing as… (same choices as above)

8. T or F

__ When a nucleus is unstable, it just has to undergo one type of decay to become stable.

__ The charge on the nucleus is always positive.

__ A neutral atom must have the same number of electrons as it has protons.

__ Uranium-235 nuclei undergo several changes and eventually end up right back where they started.

__ Radioactive radon gas is actually present in the air we breathe.

__ An alpha particle is comprised of two protons and two electrons

__ When a nucleus undergoes electron capture

9. Define radioactivity? ____

What causes an atom to be radioactive?

Protons Neutrons and Electrons

10. Fill in the following table. Note: the top five are all ones you can actually build in the Build-a-Bohr program. The bottom five require that you really understand the concepts and can apply them to problems beyond the program. A periodic table will prove helpful.

	element	isotope symbol	atomic number	mass number	charge	# of protons	# of neutrons	# of electrons
a)		$^{14}_{7}N$						
b)						5	6	5
c)	lithium			7	0			
d)		$^{23}_{11}Na^{1+}$						
e)	magnesium				2+		25	
f)						20	22	20
g)			80	202	0			
h)		$^{33}_{16}S^{2-}$						
i)	barium			137				56
j)				59	3+		32	

11. How many protons are there in a $^{20}_{9}F^{1-}$ particle? _____ How many neutrons are there in a $^{39}_{19}K^{1+}$ particle? _____

12. The only difference between a $^{52}_{24}Cr^{2+}$ particle and a $^{53}_{24}Cr^{2+}$ is the number of _______________

13. The only difference between a $^{52}_{24}Cr^{2+}$ particle and a $^{52}_{24}Cr^{3+}$ is the number of _______________

14. The only difference between a $^{52}_{24}Cr^{2+}$ particle and a $^{53}_{25}Mn^{3+}$ is the number of _______________

Ions:

Simply put ions are particles which have positive or negative charges—like Mg^{2+} or F^{1-}. Neutral particles (ones that have a charge = 0) are called "atoms." Neutral atoms can bond together small clusters called "molecules."

15. Which four particles in the table above would be considered ions: ___________________________

16. Among protons, neutrons and electrons; which two affect the charge of a particle? _____________ ______________

17. Considering your answer above, what must be true of all ions? ____

Isotopes:

Although they use the term "isotope" in the simulation, they never really define it. Consider the following statements to see if you can figure out what the term means.

"Some isotopes of carbon are stable, some are radioactive, and others can't exist at all. Fluorine has only one stable isotope: fluorine-19. Nitrogen-12 and nitrogen-13 are considered isotopes. Nitrogen-12 and carbon-12 would not be considered isotopes.

Define isotope: ____

Fill in the blanks below:

18. Isotopes are particles of the same ________________ that have different ___________. Another way of thinking about it is that isotopes are particles with the same number of ________________ but different number of _________________.

Atomic radii:
19. When you have built a Bohr diagram for a stable lithium atom, measure its radius in cm: __________
(The radius is the distance from the center nucleus out to the outermost occupied level.)

20. Do the same thing for beryllium: __________ and boron: ___________ and carbon: __________

21. Find these four elements in the periodic table What do you notice about the size of the atomic radius as you go left to right across a period in the periodic table? ____

22. Why do you think the atomic radius does this? (Hint: think about what causes the electrons to be attracted to the nucleus in the first place.) ____

23. Now measure the atomic radius for the Bohr diagram for sodium (which is just below lithium on the PT): __________

24. Do the same thing for magnesium (which is just below beryllium): ___________

25. What do you notice about the size of the atomic radius as you go top to bottom down a group in the periodic table? ____

26. Why do you think the atomic radius does this? ____

27. Based on the patterns you observed, which would be bigger? (Circle your choice.) K or Ca Ca or Mg Sr or Kr

28. The actual radius of a lithium atom is 152 pm (pm = picometer, which is a trillionth of a meter!). Use the measurements you made above to set up proportions to determine the radius (in pm) of

Be: __________ C: ___________ Na: __________

Historical figures:
29. You met several important historical figures at the Build-a-Bohr workshop. For each of the following, specify what contribution they made.
a) Ernest Rutherford

b) Niels Bohr

c) Marie Curie

d) John Dalton

e) JJ Thomson

f) James Chadwick

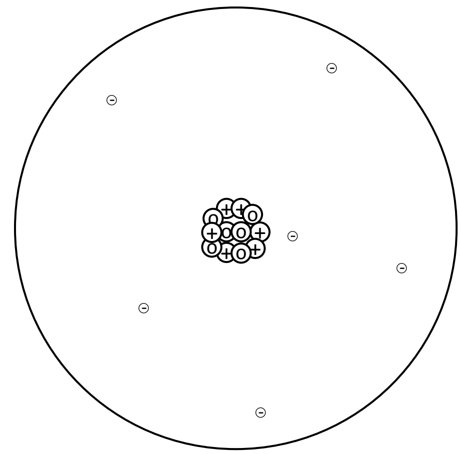
30. What element did Marie Curie discover ______________. What was it named after? _______________

31. What objection did John Dalton have to the Build-a-Bohr workshop?

Atomic Structure WS Name: ____________________ Per:_____

For a while, chemists believed that atoms were absolutely the smallest particle in nature. They eventually discovered, however, that atoms themselves are made up of even smaller subatomic particles: protons, neutrons and electrons. The protons and neutrons each have about the same mass: 1.67×10^{-24} g, which is equal to 1 amu ("amu" stands for "atomic mass unit"). The big difference between a proton and a neutron is that protons have a charge of 1+ and neutrons do not have any charge at all -- they are neutral. The protons and neutrons bind together into a very tight, very, **very** dense core of the atom known as the "nucleus." The nucleus is so small that if an atom were the size of a large football stadium, the nucleus would only be the size of a golf ball sitting on the 50 yard line, and yet the nucleus is so dense that that golf ball would weigh 40,000 times more than the rest of the stadium!!! Kind of hard to imagine...

And what about the electrons? They are much smaller than protons and neutrons: they weigh only 0.0005 amus -- which we will round to 0 amus and consider them essentially weightless! They move at high speeds in the empty space around the nucleus. In the above football stadium comparison, electrons would be like tiny salt grains buzzing around the empty stadium. The other important thing to remember about electrons is that they have a charge of 1-. It is this negative charge that keeps them close to the positively charged nucleus. Our model of a typical atom looks like the drawing at right. Note the drawing is not even close to being to scale.



So, what we know about these three subatomic particles is summarized in this table:

	mass	charge
proton	1 amu	1+
neutron	1 amu	0
electron	0 amu	1-

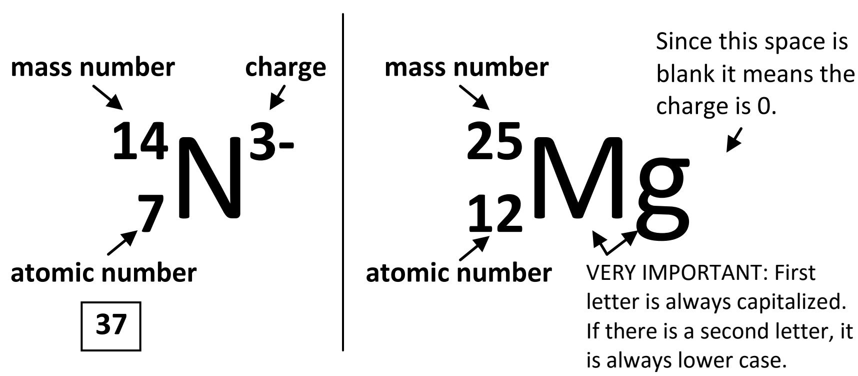
Four important rules to remember:

1) # protons = atomic number Look at a periodic table. Each element has two numbers associated with it -- one is a whole number and one is a decimal. The whole number is called the atomic number, and it tells you how many protons there are in each atom of that element. For example: carbon (C) has an atomic number 6. That means that every carbon atom in the universe has exactly 6 protons -- no more and no less. If an atom had 7 protons, it wouldn't be carbon; it would have to be nitrogen (N). This is true 100% of the time for all atoms of elements. No exceptions.

2) # electrons = # protons - charge Atoms are neutrally charged; that is, they have a charge of 0. This means they must have the same number of electrons as they do protons. Thus, since a carbon atom has 6 protons, it also has 6 electrons to balance out the charges. Ions are what atoms are called when they have a charge; sometimes the charge is positive, sometimes it is negative. Ions get their charge by gaining or losing electrons... not protons. So for ions, the number of electrons does not equal the number of protons. For example the C4- ion has four extra electrons -- that's what gives it a 4- charge. Since we already know it has 6 protons, it must have 10 electrons. Or... using the equation above: # electrons = # protons (6) - charge (-4) = 6 - -4 = 10

3) # neutrons = mass number - # protons An atom's mass is determined by the number of protons and neutrons only (not the number of electrons). Recall that protons and neutrons each weigh 1 amu. So if, for example, an atom weighs 25 amus (we say its mass number is 25), it must have some combination of protons and neutrons that adds up to 25. And if we also told, for example, that the atom is a magnesium atom (Mg), we can see from the periodic table that Mg has atomic number 12 and therefore has 12 protons. If it has 12 protons, it must have 13 neutrons to give it a total mass of 25. Or...using the equation above # neutrons = mass number (25) - # protons (12) = 25 - 12 = 13

4) Particle notation: When atoms are being described, the most common way is to write their particle notation: this shows their element symbol with the atomic number to the left and below, the mass number to the left and above, and the charge (if there is a charge) to the right and above. Look at the examples shown at right:



Now apply the rules above to answer the following questions. (You will need a periodic table)

Ans: 1- 1- 1- 0 0 0 0 1 1 1 1+ 2+ 2+ 3+ 5 5 6 6 6 8 9 10 11 11 12 13 13 13 14 15 16 16 16 18 18 18 19 19 19 19 19 20 20 20 22 22 27 35 36 46 46 47 50 62 75 76 80 81 82 82 118 126 208 ion Al As Au F H Pb Ti V

1. What is the atomic number of potassium (K)? _____ How many protons does a K atom have? _____

2. If a K atom has no charge, how many electrons must it have? _____
If a K ion has a charge of 1+, how many electrons must it have? _____

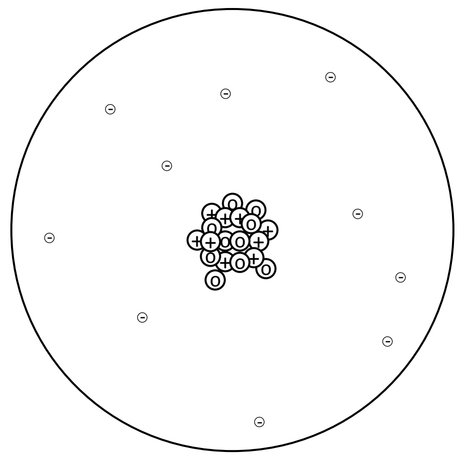
3. If a K atom has a mass number of 41, how many neutrons must it have? _____
If a K atom has a mass number of 39, how many neutrons must it have? _____

4. Write the particle notation for an aluminum (Al) atom with a mass number of 27. [ ]
Write the particle notation for a lead (Pb) ion with a mass number of 208 and a charge of 2+. [ ]

5. How many protons, electrons and neutrons are in the Al atom described above? p:___ e:___ n:___
How many protons, electrons and neutrons are in the Pb ion described above? p:___ e:___ n:___

6. For each of the particle notations below, determine how many protons, electrons & neutrons there are.

$^{14}_{6}C$	$^{32}_{16}S^{2-}$	$^{42}_{20}Ca$	$^{109}_{47}Ag^{1+}$	$^{31}_{15}P^{3-}$
p: _____	p: _____	p: _____	p: _____	p: _____
e: _____	e: _____	e: _____	e: _____	e: _____
n: _____	n: _____	n: _____	n: _____	n: _____



7. Consider the particle pictured at right. Is it an atom or an ion? _____

How can you tell? ____

What is its charge: _____ What is its mass number: ______

What would its particle notation be: [ ]

What element is it: ________________ (make sure you spelled it correctly)

8. In the box at right, NEATLY draw a sketch (like the one above) for a sodium (Na) ion (with a mass number of 23 and a charge of 1+).

How many protons does it have: ______Describe how you figured this out:

How many electrons: ______ Describe how you figured this out:

How many neutrons: _______ Describe how you figured this out:

9. Complete the tables below by filling in the missing numbers or symbols:

particle notation	charge	mass number	# p's	# e's	# n's
$^{11}_{5}B$					
$^{81}_{35}Br^{1-}$					
			33	33	42

particle notation	charge	mass number	# p's	# e's	# n's
			22	20	28
	3+	197	79		
	1+	1			

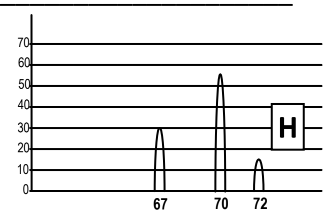
Mass Spectrometry Worksheet Name: ____________________

Use the atomic masses from the periodic table to answer the following questions: The first three require no calculations... just logical thinking!

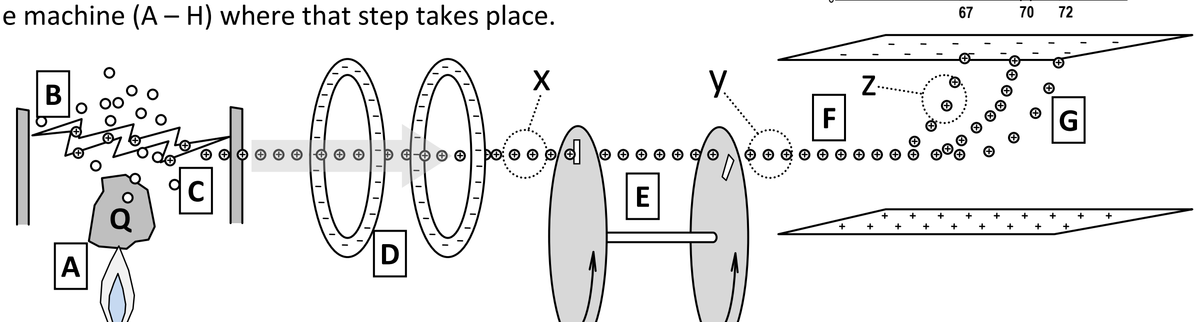
1. Indium has only two stable isotopes: In-113 (mass = 112.904) and In-115 (mass 114.904). Which of these two isotopes do you think is more abundant ("abundant" means common or plentiful)? ____________
How can you tell? ____

2. Silver has only two stable isotopes: Ag-107 (mass = 106.905) & Ag-109 (mass 108.905). Which of the following best describes the relative proportions of these two isotopes? (circle your choices)
a) exactly equal: 50%-50% b)slightly more of one than the other c) a lot more of one d) almost entirely one -- very little of the other. Which is more abundant? Ag-107 Ag-109 neither
How can you tell? ____

3. Oxygen has three stable isotopes: O-16 (mass = 15.995), O-17 (mass = 16.999) and O-18 (mass = 17.999). Which of the following best describes the relative proportions of these three:
a) exactly equal: 33%-33%-33% b) slightly more of one than the other two c) a lot more of one
d) almost entirely one -- very little of the others. Which is most abundant? O-16 O-17 O-18
How can you tell? ____



4. Below is a schematic diagram of a mass spectrometer, the machine used to separate an element into its isotopes and to determine its isotopic composition. Below the diagram, in random order, are descriptions of the steps involved in the process. Match each step to the part of the machine (A – H) where that step takes place.



__ a series of negatively-charged rings accelerates the positively-charged ions (opposites attract)
__ the beam is then passed between two oppositely-charged parallel plates
__ a sample of the element (Q) is heated to vaporize it -- converting it into individual gaseous atoms
__ particle detectors keep count and translate the information into a bar graph known as a mass spectrum; it shows the isotope masses on the x-axis and their relative percentages on the y-axis.
__ the beam of ions is passed through two rotating disks with staggered slits -- this part of the machine allows only ions with a certain velocity to make it through
__ high voltage electric current is passed through the gaseous atoms
__ the ions are deflected toward the negative plate -- the lighter ones get deflected more and hit the plate sooner; the heavier ones have more momentum and hit the plate farther down.
__ this knocks electrons off the atoms and changes them into positively-charged ions

Where do the particles possess a variety of masses but all the same velocity? x y z nowhere
Where do the particles possess a variety of velocities but all the same mass? x y z nowhere
Where do the particles possess a variety of masses and a variety of velocities? x y z nowhere
Where do the particles possess all the same mass and all the same velocity? x y z nowhere

According to the mass spectrum shown above, what are the three isotopes of Q, and what are their corresponding percent abundances? _____ _____ _____ _____ _____ _____

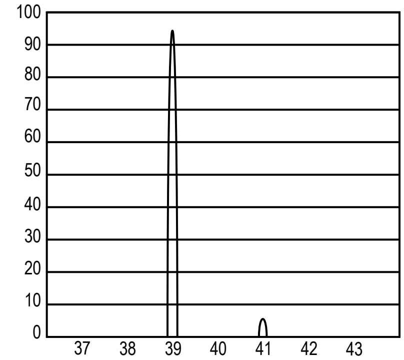
(continued on back side)

For the following problems, show all work neatly in the space provided.

5. Based on your reading of the relative abundances for element Q in problem #4 above, make an approximation of the average atomic weight for Q. ______

Now calculate the average using the equation:

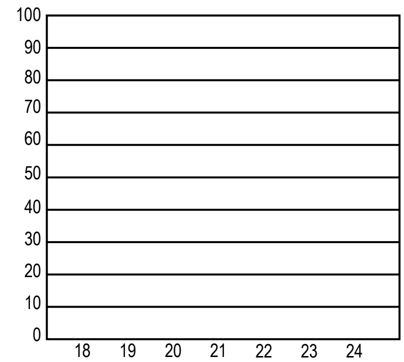
average = % x mass + % x mass + ...



6. Element Y has a mass spectrum shown at right. What are the two isotopes of Y, and what are their corresponding % abundances? _____ _____ _____ _____,

Make an approximation of the average atomic weight for Y. _______

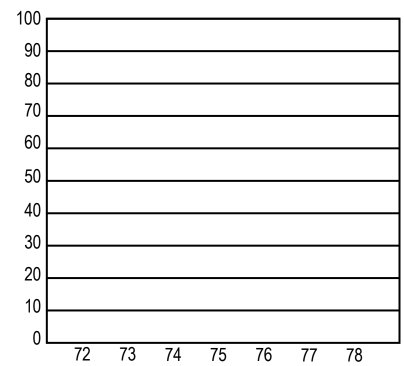
Now calculate the average using the equation above: ________



7. Neon has three isotopes: 90.48% is Ne-20 (mass = 19.992), 0.27% is Ne-21 (mass = 20.994), and the remaining 9.25% is Ne-22 (mass = 21.991). Sketch a mass spectrum for Ne at right: Without looking at the periodic table, approximate Ne's average atomic weight: _________

Now calculate the average. (Then compare it to the weight in the per. table)

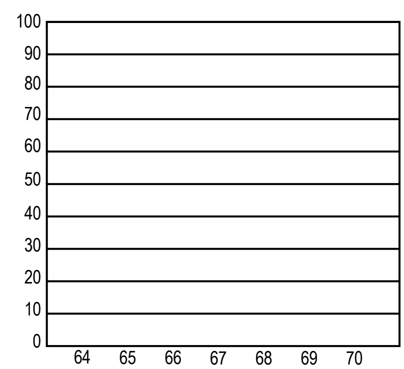
Ans: ______



8. Arsenic has only one stable isotope: 100.00% is As-75 (mass = 74.922), Sketch a mass spectrum for As at right: Without looking at the periodic table, approximate As's average atomic weight: _________

Now calculate the average. (Then compare it to the weight in the per. table)

Ans: ______



9. Zinc has five isotopes: 48.63% is Zn-64 (mass = 63.929), 27.90% is Zn-66 (mass = 65.926), 4.10% is Zn-67 (mass = 66.927), 18.75% is Zn Zn-68 (mass = 67.925) and the remaining ______% is Zn-70 (mass = 69.925). Fill in the final % above. Then sketch a mass spectrum for Zn at right: Without looking at the periodic table, approximate Zn's average atomic weight: _________

Now calculate the average. (Then compare it to the weight in the per. table)

Ans: ______

10. In a previous worksheet, you categorized substances as being either elements, compounds or mixtures. In view of what this worksheet is all about, is it possible for something to be a pure element and a mixture at the same time? ____

Explain:

Are any of the elements on this worksheet not mixtures?______ Which one(s)? ___________

Is it possible for something to be a pure compound and a mixture at the same time? ____ Explain:

Isotopic Composition Worksheet Name: ________________________

1. One element has five stable isotopes. The precise masses and relative abundances are shown in the table at right. Estimate in your head what you think the average might be to four sig figs: ______________. Now calculate it. ________________
What element is this? _________

63.9291	48.63 %
65.9260	27.90 %
66.9271	4.10 %
67.9248	18.75 %
69.9253	0.62 %
______	______

The kind of question above is pretty easy: you just plug the numbers into the equation… so let's kick it up a notch:

2. Chlorine has only two stable isotopes. One is Cl-35 (actual mass = 34.969) and it accounts for 75.78% of all chlorine. What is a) the percent, b) the mass number and c) the actual mass of the other isotope?

a: _______ b: _______ c: __________

3. Gallium has only two stable isotopes. One is Ga-71 (actual mass = 70.9247) and it accounts for 39.892% of all gallium. What is a) the percent, b) the mass number and c) the actual mass of the other isotope?

a: _______ b: _______ c: __________

4. Magnesium has only three stable isotopes. One is Mg-24 (actual mass = 23.985) and it accounts for 78.99% of all magnesium. A second is Mg-25 (actual mass = 24.986) and it accounts for 10.00% of all magnesium. What is a) the percent, b) the mass number and c) the actual mass of the third isotope?

a: _______ b: _______ c: __________

5. Europium (#63) has two stable isotopes: Eu-151 (actual mass = 150.9197) and Eu-153 (152.9212). What are the relative abundances (%'s) to 4 sig figs of these two isotopes?

Eu-151: _______ Eu-153: _______

6. Copper (#29) has two stable isotopes: Cu-63 (actual mass = 62.9296) and Cu-65 (64.9278). What are the relative abundances (%'s) to 4 sig figs of these two isotopes?

Cu-63: _______ Cu-65: _______

7. Silicon (#14) has three stable isotopes: Si-28 (actual mass = 27.9769), Si-29 (28.9765) and Si-30 (29.9738). Si-30 has a relative abundance of 3.087%. What are the relative abundances (%'s) to 4 sig figs of the other two isotopes?

Si-28: _______ Si-29: _______

(continues on back side, but you'll need access to the internet!)

Go to the website: http://www.webelements.com/isotopes.html and click on any of the elements shown there. It will show you what the stable isotopes are, what their precise masses are and what their relative abundances are. For example, when you click of Zr, it gives you the following:

Isotope	Atomic mass (m_a/u)	Natural abundance (atom %)	Nuclear spin (I)	Magnetic moment (μ/μ_N)
^{90}Zr	89.9047026 (26)	51.45 (40)	0	
^{91}Zr	90.9056439 (26)	11.22 (5)	$^{5}/_{2}$	-1.30362
^{92}Zr	91.9050386 (26)	17.15 (8)	0	
^{94}Zr	93.9063148 (28)	17.38 (28)	0	
^{96}Zr	95.908275 (4)	2.80 (9)	0	

This tells you that Zr-90 (with an actual mass of 89.9047026) accounts for 51.45% of all Zr atoms, and Zr-91 (with an actual mass of 90.9056439) accounts for 11.22%, etc, etc. If we wanted to calculate the average atomic mass from this, the calculation (with no rounding from the table) would be:

$$\frac{(89.9047026 \times 51.45) + (90.9056439 \times 11.22) + (91.9050386 \times 17.15) + (93.9063148 \times 17.38) + (95.908275 \times 2.80)}{100}$$

= 91.22364607 (which matches the 91.224 from the text book periodic table)

8. Repeat the above calculation using the information from the web site for <u>one</u> of the following: **Rb, Sb, B, or Li**.

9. Repeat the above calculation using the information from the web site for <u>one</u> of the following: **Pb, Ce, Cr or Fe**

10. Repeat the above calculation using the information for the web site for <u>one</u> of the following: **Cs, Al, Bi or Au**

Relative Atomic Mass WS Name: _______________

Chemists came up with atomic weights for the elements long before they could isolate individual atoms. But how did they do that? How did they weigh individual atoms? Simple: they didn't! Instead, they weighed samples of those elements and reacted them with samples of other elements also carefully weighed out. That enabled them to come up with what they believed were relative masses of the various elements.

Three important concepts to keep in mind:

* The lightest element was arbitrarily assigned a mass of 1.00.
* All the other elements were then assigned masses relative to this 1.00 based on their mass combining ratios.
* The assumption was always made that the compounds formed were always in a 1:1 atom ratio, thus their sample combining mass ratio would be the same as their atomic mass ratio.

29.41 g of A react with 5.18 g of B
17.91 g of A react with 67.98 g of D
17.80 g of C react with 36.71 g of B

1. The fictitious data at right were collected about the hypothetical elements A, B, C and D. Assume each compound is made up of a simple 1:1 ratio of atoms. For example, the first compound is AB – not A_2B or AB_3 Assign the lightest element a mass of 1.00, then determine the relative atomic masses for the other three elements.

A = _______ B = _______ C = _______ D = _______

It turns out that the first compound above is actually AB_2 (not AB).
What would be the relative masses in light of this new information: A = _______ B = _______ C = _______ D = _______

37.1 g of F react with 67.8 g of G
4.73 g of E react with 3.46 g of F
124.8 g of E react with 13.6 g of H

2. The fictitious data at right were collected about the hypothetical elements E, F, G and H. Assume each compound is made up of a simple 1:1 ratio of atoms Assign the lightest element a mass of 1.00, then determine the relative atomic masses for the other three elements.

E = _______ F = _______ G = _______ H = _______

It turns out that the first compound above is actually FG_3 (not FG).
What would be the relative masses in light of this new information: E = _______ F = _______ G = _______ H = _______

37.10 g of M react with 59.80 g of K
37.91 g of J react with 279.8 g of L
24.78 g of K react with 73.67 g of J

3. The fictitious data at right were collected about the hypothetical elements J, K, L and M. Assume each compound is made up of a simple 1:1 ratio of atoms Assign the lightest element a mass of 1.00, then determine the relative atomic masses for the other three elements.

J = _______ K = _______ L = _______ M = _______

It turns out that the first compound above is actually K_4M (not KM).
What would be the relative masses in light of this new information: J = _______ K = _______ L = _______ M = _______

Ans (IRO+1): 6x(1.00) 1.61 2.06 2.06 2.48 2.97 4.09 4.79 6.71 6.71 9.18 9.18 11.7 12.3 15.4 21.9 23.4 35.4 44.4 88.8

Relative Atomic Mass Lab Name: ____________ Partner: _____________

Chemists came up with atomic weights for the elements long before they could isolate individual atoms. But how did they do that? How did they weigh individual atoms if they weren't even sure they existed!? Simple answer: they didn't! Instead, they carefully weighed out samples of those elements and then reacted them with one another. This enabled them to come up with *mass combining ratios*, and from these ratios, they were able to determine what they believed were atomic masses of the various elements – not actual *absolute* masses, just *relative* masses. That is: hydrogen seemed to be the lightest, so they assigned that a mass of 1.00 amu and fluorine seemed to be 19 times heavier so they assigned fluorine a relative mass of 19.0 amu. Because these relative masses were based on the assumption that all the compounds were 1:1 compounds, some of the relative masses were way off, For a long time, for example, chemists believed that water's formula was just HO, and that gave them a relative mass for oxygen of 8.0 amu instead of the 16.0 amu we believe it to be today.

The elements you will use today: A = hexnutium B = nailium C = q-tipium D = screwium

Part I: Setting up the lab for another group.

1) Pick any one of the four elements and place some number of those "atoms" (between 3 and 10) in one of the envelopes. Then use a pencil to write that symbol on the envelope. [Say for example you picked B (nailium), and you put six atoms (nails) in an envelope and then wrote a "B" on the outside.]

2) Pick a different element (say C, q-tipium) and place the <u>same</u> number of its atoms in a separate envelope. Use a pencil to write its symbol on the envelope. Now paper clip these two envelopes together. This is essentially saying: "Since these two elements react in a 1:1 ratio, this much of element B (6 atoms) reacts with this much of element C (6 atoms)."

3) Repeat steps 1 and 2 for a different pair of elements and a different number of atoms. This time perhaps you put eight A in one envelope and 8 C in another. Again label these envelopes.

4) Repeat steps 1 and 2 one more time for a third pair (make sure that in your three pairs, all four elements are used at least once), BUT this time use something other than a 1:1 ratio – perhaps 1:2 or 1:3. (For example, six A atoms in one envelope and twelve D atoms in the other). Make note which combination this was by writing its formula here: ______ (for example AD_2 to show that A and D were combined in a 1:2 ratio.) Do NOT indicate this on the envelopes: the envelopes should just have "A" and "D". You should now have 3 sets of envelopes to hand off to another group.

Part II: Determining Relative Masses.

1) Place your one empty envelope on the scale and zero it. (We will assume for this lab that all the envelopes weigh the same.) Obtain three pairs of envelopes from another group. Now, without looking inside, weigh each envelope. Record these in the table at right. Make sure to keep paired envelopes paired.

Element	Mass with envelope zeroed out	
_____		=
_____		
_____		=
_____		
_____		=
_____		

2) You now have the combining masses for these four elements. Assume all the compounds are in a 1:1 ratio. (You know that one of them is not 1:1, but we will ignore that for now!) Based solely on the combining masses – and not on your having handled the individual "atoms" in part I – determine which element has the lowest atomic mass. Assign it a value of 1.00. Then use proportions to figure out the other element's relative masses: Show your work at right.

A: ________ B: __________ C: _________ D: __________

4) Return the envelopes to the group that gave them to you and find out from them which compound was not 1:1. Record its formula (like AD_2) here: ________. Use this new information (the formula of the non-1:1 compound) to revise your relative masses. A: ________ B: __________ C: _________ D: __________

5) One more thing: Re-zero the scale with nothing on it, then weigh one "atom" of each type:

A = _____ B= _____ C = ______ D = _____ (We'll come back to these masses in follow-up questions 4-6.)

Clean-Up: When you get the envelope back from the group you gave it to, put the "atoms" back in their appropriate cups and erase the labels from the envelopes.

Follow-Up Questions:

1) How did your first calculated relative masses compare to other groups? Were any the same? ____________________

Why were some of them different?

2) How do your second calculated relative masses compare to other groups? __________________________________ (They should match up pretty well!).

3) What error sources can you think of that might cause these to be off a bit?

4) From part II, step five, what were the masses of the "atoms?" A = _____ B= _____ C = ______ D = _____

5) Divide each of these by the smallest one. What values does that give: A = _____ B= _____ C = ______ D = _____

6) How do these numbers compare with the second set of calculated relative masses?

7) This lab was just a simulation. It was meant to model how early chemists were able to come up with relative masses of the atoms without ever isolating individual atoms. In part II of the lab, you were not allowed to open the envelopes. Why was this restriction in place?

8) In part I of the lab, you played the role of the chemical reaction that had the atoms combining into compounds. What law were you enforcing?

9) Look at the final relative masses you determined in step #4 above. What real-life elements have relative masses similar to those in the lab? A = ____ B = ____ C = ____ D = ____

10) a) What mass of nailium would combine with 24.12 g of hexnutium to form the compound AB? __________

b) What mass of nailium would combine with 7.93 g of q-tipium to form the compound BC_3? ___________

c) What mass of screwium would combine with 3.12 mg of hexnutium to form the compound A_2D_3? __________