Safety

1/30

- no eating / drinking
- don't be crazy
- don't play pratical jokes
- be careful of water + electric
- don't sit on counter
- no flip flops
- clean up broken glass
- check for frayed cords
- only take what you need

Michael Plagmeier

Safety Essay

20/20

1/30

Safety in the lab is important because insures proper protection for ourselves, our peers and the enviroment. Following saftey procedures does not incur a significant detrement which is not over-ridden by a benfit. Thus there is not excuess not to follow the following safety proceedures.

47

Before you start, tie back loose clothing and know the proceedure of the experment you are about to do. In addition, know all of the safety rules, equipment and procedures, as well as the locations of emergency exits.

85

If anything happens while you work, be sure to report it to a teacher. If something got into your eye because you did not wear safety goggles, wash your eyes out in lukewarm water for 15 minutes.

121

Never taste or eat anything in the lab because you don't know if it had been infected or contaminated either on purpose or accedently.

145

In addition do not handle broken or chipped glass. The sharp edges might cut or injure you. If you break

173 any glass, clean it up safely and kee

others away from it. If you are cut, allow the area to bleed for a short while and then run it under cold water. More serious cuts require a trained medical professional.
205

One can not tell the difference, normally, between something which is hot and something which is cold. Be sure to treat all things as if they were hot so one does not get burned.
230
In addition watch out for electrical appliences with broken cords especially around metal or water.
254

If one follows these procedures among others, one can increase the chances of surviving Chemestry Class. (aside from acdemic failure that is)

Michael Plasmeier

Properties of Matter
Questions

1/31

1. Which properties are common to all matter?
   all matter has mass (the amt. of matter) and takes up space (volume).

2. What is the term for these properties?
   These 2 things (mass + volume) are extensive properties. They dependent on the amount of matter in a sample

3. Which type of properties can be used to identify matter
   Intensive properties such as hardness and boiling point indentifies different types of matter.

4. What is a pure substance?
   Matter which is uniform and has a defined composition is a substance.

5. Properties which can be observed without changing a substance's composition are —? Give Examples
   physical properties such as hardness, color, conductivity and malleabilty

6. How is an element different from a compound?
   An element is already at its simplest form and can not be broken down any more
   compound: 2 or more elements

7. What is an example of a chemical change?
   Heating suguar to turn it into carbon and water.

Lab Reports

1/31

1. Title: Seperation of a Compound
2. Purpose: to seperate a compound into its elements
3. Procedurs: see sheet
4. Data: Observations: Color
    Color Changes
    Smell
    Bubles
    Prodduction of any solid

Michael Plasmeier

Seperation of a Compound

$\frac{17}{20}$ & 3 tunning in worksheet

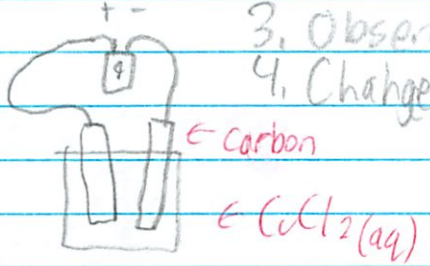
1/31

to seperate a compound into its elements

Equipment: 2 carbon rods
9v Battery
2 alligator cables
small beaker copper (II) choride

$CuCl_{2(aq)} \xrightarrow{dec} Cu_{(s)} + Cl_{2(g)}$

Procedure: 1. Attach alligator clips to rods + battery
2. Place rods into solution
3. Observe products
4. Change terminals



Data: + rod dissolving with bubbles
ground rod turning red-pink (from previous experment)
choride color don't change
smells like Chloride
no solid produced
—change termials—

red flakes of carbon appear
-form circle -but don't touch - black (-) rod
+ rod bubbles again (slightly)
red powder sinking
-rod turning red (~~?oxidizing?~~) - copper
other people's solutions turned lighter blue
-some of the copper in the water changed from liquid to a solid

blue liquid = copper

Seperation of a Compound

1/31

Laboratory Procedure for separating a compound:

1. Obtain the following equipment: a. 2 carbon rods

b. a nine volt battery c. 2 alligator clips d. a small beaker

2. Into the beaker, place a small amount of copper (II) chloride solution.

3. Attach alligator clips to carbon rods and to battery terminals

4. place carbon rods(not touching each other) into the copper (II) chloride solution

5. Observe the products. One of your products is a solid. The other product is a gas.

6. Identify each substance in this activity by a physical property.

Copper (II) choride - bluish color
Carbon - black
bubbles - clear

Chem Talk: The Structure of Matter

electricty provides electrons

1.Why did we use electricity in this experiment?

To seperate the two elements hydrogen (H) and oxygen (O) from each oter -seperate Compounds into elements

2.Describe the reactions of Hydrogen and Oxygen with the glowing splint.

Hydrogen - a flair up or explosion

Oxygen - ignighted a glowing splint causing the fire to grow larger - "re-lit"

3.If the experiment in the book was done(we did a different one), How could the 2 in H_2O be explained?

There are 2 parts of hydrgon for every part Oxygen, in a molecule

4.Hydrogen and Oxygen are examples of elements.

5. An element is: any matter which can not be broken down into simpler materials by chemical means.

6.Everything that we observe in the world is made up of......

matter + elements

7.Chemistry is: the study of how elements combine and the characteristics of these combinations + changes

8.How are elements represented?

by letter symbols

9.When elements combie they form new substances called Compands

10.Water is an example of a Compand.

11.What is an atom?

Smallest particle of an element

12. Which atoms make up water?

hydrogen and oxygen

13. What is their ratio?

2 parts hydrgen for every oxygen atom in a molecule

14. What is the process called which uses electricity to decompose a compound?

electrolysis

15.Explain the differences in properties among, Hydrogen, Oxygen and water.

Oxygen - made fire grow bigger } gasses
hydrogen - explosive

but water puts out fires / liquid

16.Compounds are represented by Chemical Formulas.

17. What is a subscript? (You can give an example) ?

little character below the line which modifies the preceeding characters word subscript

18. Why are subscripts used in formulas?

To tell how many atoms of each element make up a molecule

19. Identify all the elements used in OUR lab.

H - hydrogen
O - oxygen
C - carbon
~~?battery~~ ← don't matter

Copper - Cu
Clorine - Cl

20. Identify all the compounds.(Hint look at the names of substances)

~~?not here?~~
$CuCl_2$

H_2O

21. Do the “Checking up” Questions on p.10 and Chemistry to go on p.11

See other sheet

Activity 1
Questions

2/1

Checking up

White Book p10-11

1. Diff b/w element + compound
   A compound is made up of elements combined.
   An element can not be split up while a compound can

2. Symbols for describing elements?
   One or two letters
   with subscript, saying how many per molecule
   easier then writting words out

3. Symbols:
   Carbon - C
   Copper - Cu
   Gold - Au
   Helium - He

4. What info does formula of compound provide?
   tells what (and how many atoms of each per molecule) elemements make up a compound

p11 Chemestry to Go

1. Sugar - $C_{12}H_{22}O_{11}$
   Carbon - 12
   Hydrogen - 22   45 total atoms/molecule
   Oxygen - 11

   Marble - $CaCO_3$
   Calcium - 1
   Carbon - 1   5 atoms/molecule
   Oxygen - 3

Natural Gas - CH_4 copy error
Carbon - 1
Hydrogen 4 ✓ 5 atoms/molecule

Rubbing Alcohol - C_3H_8O
Carbon - 3
Hydrogen - 8 12 atoms/molecule
Oxygen - 1

Glass - SiO_2
Silicon - 1 3 atoms/molecule
Oxygen - 2

2. Nitrous Oxide (laughing gass) - N_2O

3. Glass - Silicon - hard, brown, pebbles, solid
Oxygen - clear, oderless, tastless gas

Glass - solid, clear, strong

Michael Plasmeier

1 + 2 + 3 = Black

18/20 2/2

Red p56

To make macroscopic observations of chemical reactions and use them to solve problems

Materials
- paper
- metric ruler
- reaction surface
- chemicals in grid
- pipette, medicine droppers, and spatulas

Procedure
1. Draw 2 copies of grid
2. Place reaction surface over one grid
3. Using row + column headers place chemicals
4. Stir each mixture by forcing air from an empty pipette

notice it

	NaClO	H_2O_2	$CuSO_4$ -blue
KI	yellow	lighter yellow	brownish yellow white oil residue
KI starch	dark brown bit of yellow	purplish clear liquid	dark brown yellowish tint white oil residue
KI Paper	light yellow liquid paper turned purpilish black	blackened paper clear pinkish liquid	black paper yellowish liquid white oil residue
KI Cereal	c. mushy and black liquid, clear yellow with black spots	→ some but cereal blacker + more black spots	brownish yellow liquid cereal mushy and black

Analyze / Conclude

1. Color of $NaClO + KI$?
yellow liquid

2. $NaClO + KI$ + starch?
dark brown with a bit of yellow

3. What do $NaClO$, H_2O_2 + $CuSO_4$ have in common?
all liquid, all caused color change when mixed with KI - all compounds

4. What substance is in paper and cereal?
- starch - both turned black (as liquid starch did) with mixed with KI and one of $NaClO$, H_2O_2, $CuSO_4$

You're the Chemist

1. How to tell if something has starch on it.
Drop KI and one of the following on it to see if it turns black: $NaClO$, H_2O_2, $CuSO_4$

2. How to tell if salt has KI in it (if iodized)
Take a sample and add $NaClO$, H_2O_2, or $CuSO_4$ to it to see if it turns yellow - look for black
- iodized salt has about 4% KI - no add too starch!

3. How to tell if tablets have starch?
Add KI and $NaClO$. If it turns brown in addition to yellow it has starch

4. Try $NaClO$ on marker streaks on white paper
It removed some colors like a pen eraser
red tottally gone - black turned brown - blue + green turned light brown

Common Chemical Elements

2/2

aluminum	Al
bromine	Br
calcium	Ca
carbon	C
chlorine	Cl
copper	Cu
gold	Au
helium	He
hydrogen	H
iodine	I
iron	Fe
lead	Pb
magnesium	Mg
mercury	Hg
neon	Ne
nickel	Ni
nitrogen	N
oxygen	O
phosphorus	P
potassium	k
silicon	Si
sodium	Na
sulfur	S
tin	Sn
zinc	Zn

Chem Symbols Test

Michael Plasmeier

2/7

Na - Salt → Sodium

Sn- tin

Hg - Mercury

Si - Silicon

S - Sycrose → Sulfur

Zinc - Zn

Iron - Fe

Copper - Cu

Bromine - Br

Lead - PB

8

10

Name ______________________ Date ________________ Class ________________

2.2 MIXTURES

Section Review

Objectives

- Classify a sample of matter as a substance or a mixture
- Distinguish between homogeneous and heterogeneous samples of matter
- Describe two ways that components of mixtures can be separated

Vocabulary

- mixture
- heterogeneous mixture
- homogeneous mixture
- solution
- phase
- filtration
- distillation

Part A Completion

Use this completion exercise to check your understanding of the concepts and terms that are introduced in this section. Each blank can be completed with a term, short phrase, or number.

A physical blend of two or more substances is a __1__. A mixture has a composition that varies. Mixtures may be identified as __2__ or __3__. Homogeneous mixtures are also known as __4__ and have uniform properties. Any part of a sample with uniform composition and properties is called a __5__.

Many mixtures can be separated into their components by __6__ methods. __7__ is a method of separation that involves boiling a liquid, which is then condensed.

1. mixture
2. homogeneous
3. heterogeneous
4. Solutions
5. phase
6. filtration
7. distillation

Part B True-False

Classify each of these statements as always true, AT; sometimes true, ST; or never true, NT.

ST **8.** Homogeneous mixtures can be separated by distillation.

AT **9.** A solution has a uniform composition.

AT **10.** A heterogeneous mixture contains two or more phases.

ST **11.** Solutions are liquids.

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________________ Date ______________ Class ______________

Part C Matching

Match each description in Column B to the correct term in Column A.

	Column A	Column B
f	12. mixture	a. a mixture that has a uniform composition throughout
c	13. heterogeneous mixture	b. any part of a sample that has uniform composition and properties
a	14. homogeneous mixture	c. a mixture that is not uniform in composition
e	15. solution	d. separation of a liquid by boiling followed by condensation
b	16. phase	e. another name for a homogeneous mixture
d	17. distillation	f. a physical blend of two or more components
g	18. filtration	g. a method for separating a solid from a liquid in a heterogeneous mixture

Part D Questions and Problems

Answer each of the following questions in the space provided.

19. State whether each of the following is a homogeneous or heterogeneous mixture.

a. table salt dissolved in water **a.** homo

b. carbon mixed with sand **b.** hetro

c. filtered apple juice **c.** homo

d. vegetable soup **d.** hetro

e. fresh squeezed lemonade **e.** hetero ? parts in it

20. Classify each of the following as a substance or a mixture.

a. table sugar (sucrose) **a.** substance

b. hot tea **b.** mixture -filtaable

c. table salt (sodium chloride) **c.** substance

d. vinegar **d.** substance

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Section 2.2

Questions

2/6

11. How are mixturs classified?

Mixturs can be classified as heterogeneous and homogenous. Heterogenous mixtures are made up of multiple seperate phases. Homogenous mixtures are made up of substances which can seperate by physcal means but appear to be one with an even composition throughout.

12. What can seperate mixtures?

Physical properties such as filtration or distillation among others can be used to speate mixturs.

13. Explan "phase"

A phase is a section of a mixtures with uniform properties. A homogenous mixture only has one phase while a heterogenous phase has 2 or more phases.

14. Homo. or hetero?

food coloring - homo
ice cubes in water - hetero
mouth wash - homo
mashed unpeeled potatoes - hetero

15. How are a substance and a solution similar + different?

A solution is a homogenous mixture of different substances or elements. Solutions can be seperated by physical means while substances are seperated by chemical changes. Both are the building blocks of other things.

16. When use filtration? distillation?
Filtration can be used to seperate solids of different sizes from each other or solids from liquids. Distillation can be used to seperate liquids of different boiling points.

17. Seperate sand from table salt
Use a filter to seperate the salt from the sand which is smaller and will pass through.

Michael Plasmeier

Structure of Matter Essay

Excellent!

$\frac{26}{26}$ 2/1

Chemestry studies the structure of matter and how it changes. A mixture is a combondation of compounds. This mixture can be seperated with a physical change. In a chemical change the elements shift, changing the makeup of compounds but the amount of matter always stays constant.

47

A combination of compounds is a mixture. Most matter we encanter are mixtures. Mixtures may be seperated into their compounds by physcial changes. Examples of physical changes to seperate the compounds from a mixture include filtration and distillonation among others. All seperation proceedurs take advantage of the different physical propertires of the compounds which make up mixtures.

Distillation separates the vapor of compounds which boil at different tempratures. The compound which boils first when a mixture is heated sepreates from the compound which has a higher boiling point and then condences and collects in another area separating the two compounds.

147

A mixture can be classifed as heterogeneous or homogenous. A homogeneous mixture, or solution as an equal distribution of the compounds in the mixture, for example vinegar. Hetrogeneous mixtures have varying proportions of the compounds. Layers of

100 a compound with a uniform composition

Excellent!

30/30

and property are called phases. For example when one mixes olive oil and vinegar, the olive oil phase floats on the vinegar phase
217 making a hetrogenous mixture.

Chemical changes cause elements in compounds to rearrange. Chemical changes include burning, rotting, rusting, decomposing, fermenting, exploding and corroding. During these changes chemical properties of substances are exposed as the composition of matter changes. The substances which are present at the start of the reaction are the reactants while the result is the product.

271

There are some clues that a chemical change has taken place. However they do not prove that a chemical, not physical change has occured. Every chemical change involves a transfer of energy. In addition reactants may change color, or produce gas or a precipitate. A precipitate is a solid which forms on a mixture

325

After a chemical reaction the mass (amt. of matter) in the products always equals the mass of the reactants. This is the same with physical changes and makes up the law of the conservation of mass. Matter can not be created nor destroyed. However the products might not always be obvious, for example burning
386 wood creates carbon dioxide and water vapor.

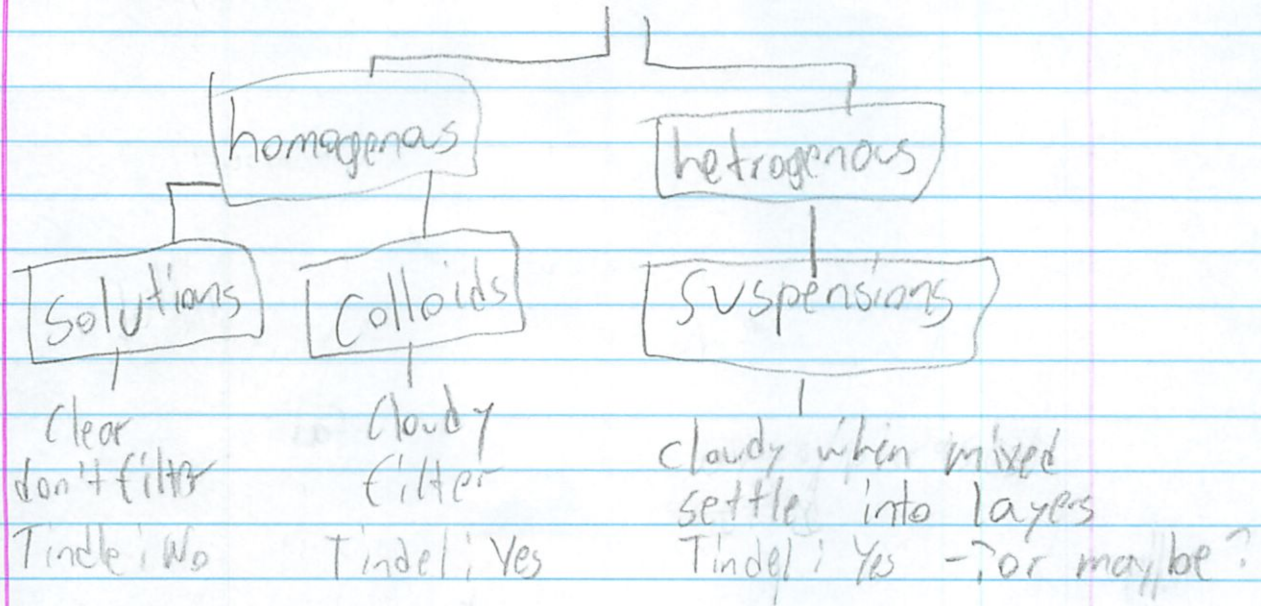
Thus the structure of matter can be changed by physical and chemical changes. However the amount of mass never changes. Combonations of compounds are mixtures. Mixtures can be heterogeneous or homogeneous. They can be seperated by physical, not chemical changes

427

Mixtures

2/7

all (have variable composition
components retain their properties



Clear
don't filter
Tindle: No

Cloudy
filter
Tindel: Yes

cloudy when mixed
settles into layers
Tindel: Yes - or maybe?

Separation of a Mixture Laboratory

Purpose: To separate a mixture:

Procedure: You will be given the names of the components of a mixture. You are to devise a procedure to separate it. You **must record the steps you followed to separate this mixture, so anyone can follow this procedure and get your results.**

Data: Record all data upon which you based calculations and results.

Calculations: Should show sample calculations and must **include masses and percentages** of all components.

Error analysis. If you do not all to 100%, explain sources of probable error.

Michael Plagner

Solutions, Suspensions + Colloids Lab

2/7

White book p22

20/20

Purpose: To explore ways mixtures can be made and test materials to test mixture types

Procedure

1. Fill 6 test tubes with water and
   1. nothing (just water)
   2. .5 g sugar
   3. few drops milk
   4. .5 g $CuSO_4$ (copper sulfate)
   5. 2mL Olive oil
   6. .5 g soil
2. Fill Chart ↓ mix filter ↴

ll still
quid

#	Observations b/f mixing	Ob. after mix	Type	Filter
1.	clear pure substances (homo)	looks same - but bubbles	Solution Tindel - No	all thragh
2.	clear - homo looks just like water	looks same - some bubbles	Solution T - No	all thragh
3.	white - milk like homo opaque	looks same - bubbles at top	Colloniod T - Yes	all thragh
4.	clear with bluish tint homo	looks same - some bubbles	Solution T - No	all thragh
5.	Clear with yellow layer at top (oil) hetero	liquid turns white more homo like	Suspension T - Yes	kinda white oil stays liquid kind

	b/f mix	after mix	Kind	Filter
6.	Clear with brown tint - appears homo(?) but some solid dirt at bottom so hetero	see lots of solid dirt particals det. hetero looks clouder	Suspension	solid particals stay in filter
			Tindle - Yes	liquid clearer

Questions

Chem Talk p24 1. How to distinguish solution, colloid, and suspension?

First I would make observations or try to filter the particals to see if it was a heterogenous suspension. If it all filtered or appeared homogenous I would shine a laser in to see if I observed the tyndall Effect. If I did I would know that the liquid mixture was a colloid and not a solution.

2. What is Tyndall Effect?

The Tyndall Effect uses a focused laser light to see if the particles are large enough to break up and distribute the light as in a colloid or suspension.

Chem to Go p26 1. Classify:

a. entire mixture passes though filter - its a homogenous mixture either a solution or colloid

b. Particals settle in mixtures - its a heterogenous mixture meaning it can only be a suspension

c. small particals visable in microscope - it could ba a colloid or a solution - or depending on type of microscope - even a solution

d. Tyndal Effect - its a colloid or suspenion before it settles

e. Mixtue blue + transparent - Need more information but prob. a solution

2. How to seperate;
   a. solution - use distillation if parts have different boiling points or one's a solid
   b. colloids; same a sollution or possibly a very good filter
   c - suspension - Use a fiter

3. Tyndal Effect on cloudy days from sun tere is a colloid gas mixture dispersing light

4. These mixtures; homo or hetero
   See chart Homo; #1, 2, 3, 4
   Hetero; #5, 6

5. Pick 5 kitchen items - What type? If diff type - what diff;
   Salid dressing - suspension - if it was a colloid tere wouldn't be different phases or little pieces in it
   ↑ can be filtered →
   if oil, it floats on top

clear, unfilterable

Water - its a solution - if it was a colloid it would look cloudy

(Flat) Sprite® - solution - clear - looks 1 piece - can filter - light passes through - if it was a suspension, you would need to shake it to seperate pieces

Milk - its a colloid - can't be filter, looks cloudy, - Tyndal effect - if it was a sollution it would be clear

Pure orange juice - suspension - has filterable orange pieces - if I filtered it, it would be a colloid (? or solution?)

States of Matter
Questions

2/8

Red
p408

50. What happens to avg kinetic energy of your water molecules when you have a fever?
You are warmer, which means that the average kinetic energy increases making them move faster and evaporate by turning into gas on your shirt where they condense into sweat.
- molecules break free easier w/↑ avg kinetic energy

55. In a liquid as intermolecular force of attraction strenghtens - would vapor preasure ↑ or ↓?
It would decrease because the molecules would be held together more and thus not as much would evaporate lowering the vapor preasure.

56. What state? (see chart p408 red)
a. phenol 99°C - liquid
b. ammonia) -25°C - gas
c. methanol in ice water bath - liquid
d. methanol in boiling water - turn to gas
e. ammonia at -100C - ~~solid~~ liquid
f. phenol 25°C - solid

pressure = $\frac{\text{force}}{\text{area}}$

note middle is a liquid

freeze boil
solid ⇔ liquid ⇔ gas

58. What causes atmospheric presure + why lower on mountains?
Athmospheric pressure results from collisions of atoms and molecule in the air with objects. It decreases on mountains beacuse the density of earth's athmosphere also decreases.

59. Why does paring +196 °C nitrogen on a baloon make it shirnk and why does it expand then?

It must be because the nitrogen absorbs energy from the air molecules slowing the average kinetic preascre pushing at againts a baloon.

- molecules store potential energy so some of that must be lost to slow down particals

Charle's law - volume + temp propotinal

Michael Plassmeier

Seperating a Mixture

Salt, Sand, Iron

20/20 2/7

Purpose: to seperate a mixture of salt, sand and iron

Procedure: Weigh the dish with and without mixture to find mixture weight (use g)

2, Use a magnet to seperate the iron from the mixture. (Make sure to wrap magnet to make it easy to weigh iron) Weigh it,

3, Pour ≈ 40 mL (weigh water + beaker 1st) of water into the salt and sand mixture, Stir for 4 minutes

4, Pour the water into another dish and weigh it. The salt represents the amount over just the water + beaker, mass,

5, Weigh the dish with sand (Dry it 1st) (leared before 3rd trip)

Weight empty dish - 45,75 g
Weight dish + mixture - 49,451 g

Mixture - 3,5 g

Weight dish + iron - 46,49 g

Iron - ,74 g

water (40 mL) + beaker - 66,76 g
water + beaker + salt - 65,94 g

Salt = -,80 g

↑ something wrong

dish + sand (+ some water) = 49,33 g

Sand - 3,58 g

↑ too much

Revision

The salt + sand mesurment didn't go well becase not all sand disolved and water remaind around the sand mixture when I pored the water and dissolved sand out

So I am boiling the disolved salt + water + beaker to boil out the water and weight it again.

Also I am mixing anoter 40 mL of water in a beaker to try and disolves rest of sand, I then set it to boil out the water

Beaker$_2$ + 40mL water - 64.23g
Beaker$_2$ + 40mL water + disolved salt - 63.37g
-Some water splached out

Beaker$_1$ + dried salt - 29.47g   .8g salt
Beaker$_1$ Empty - 29.39 g   Beaker$_1$

Beaker$_2$ + dried salt - 29.10g   -.7g salt
Beaker$_2$ empty - 29.17g   Beaker$_2$

no visable salt

.8g Salt
.74g Iron
1.96 Sand ← subtraction

experment didn't really work
Confidene = 0% so error = 100%

Could also evaporate the left over water from the sand

Going to try again

Dish	38,01 g	Mixture
Dish + mixture	47,36 g	9,35
Dish + Iron	42,12g	Iron - 4,11

some salt b/c it stuck togeter b/c it was wet

Beaker	30,03g	Diso
Beaker + 40 ml water	70,90g	
Beaker + water + disolved salt	70,09g	
Dish + Sand + some water	47,79g	
Beaker + Dried Salt	30,47g	salt - ,46 g
Dish + Dried Sand	46,57g	Sand · 8,56g

Componant 8,56 + ,46 + 4,11 = 13,13 gram

Supposed total 9,35 grams

30% error

Iron	31%
Salt	4%
Sand	65%

Better this time

Closer but still a bit off - might have been chemicals in beaker

~~Cleaned out beaker = 29.96 g~~

~~So were really .51 grams of salt~~

~~But that~~

Michael Plasmeier

States of Matter

20/20

White Activity 2

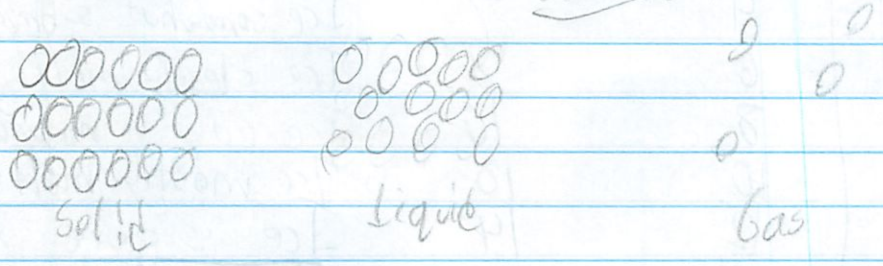
Part 1: Boiling Freezing Water (Heating Curve of water)

2/4

White p12

Purpose: To investigate states of matter and accompanying energy changes

What Do Ya Think: see red p41



1. Fill a beaker with a 450mL ice and water mixture

2. Position hot plate with beaker on it and a therometer not touching the bottom

3. Wait for 5 minutes for the therometer to cool off, Record at the 0 time

4. Turn on hot plate to medium and record temp every to minutes while stiring. Continue till water boils for 5 minutes. Also record your observations, including boiling point + point where ice gone

5. Graph + answer: Whats happening at different points along the graph: What does the heat do?

20/20

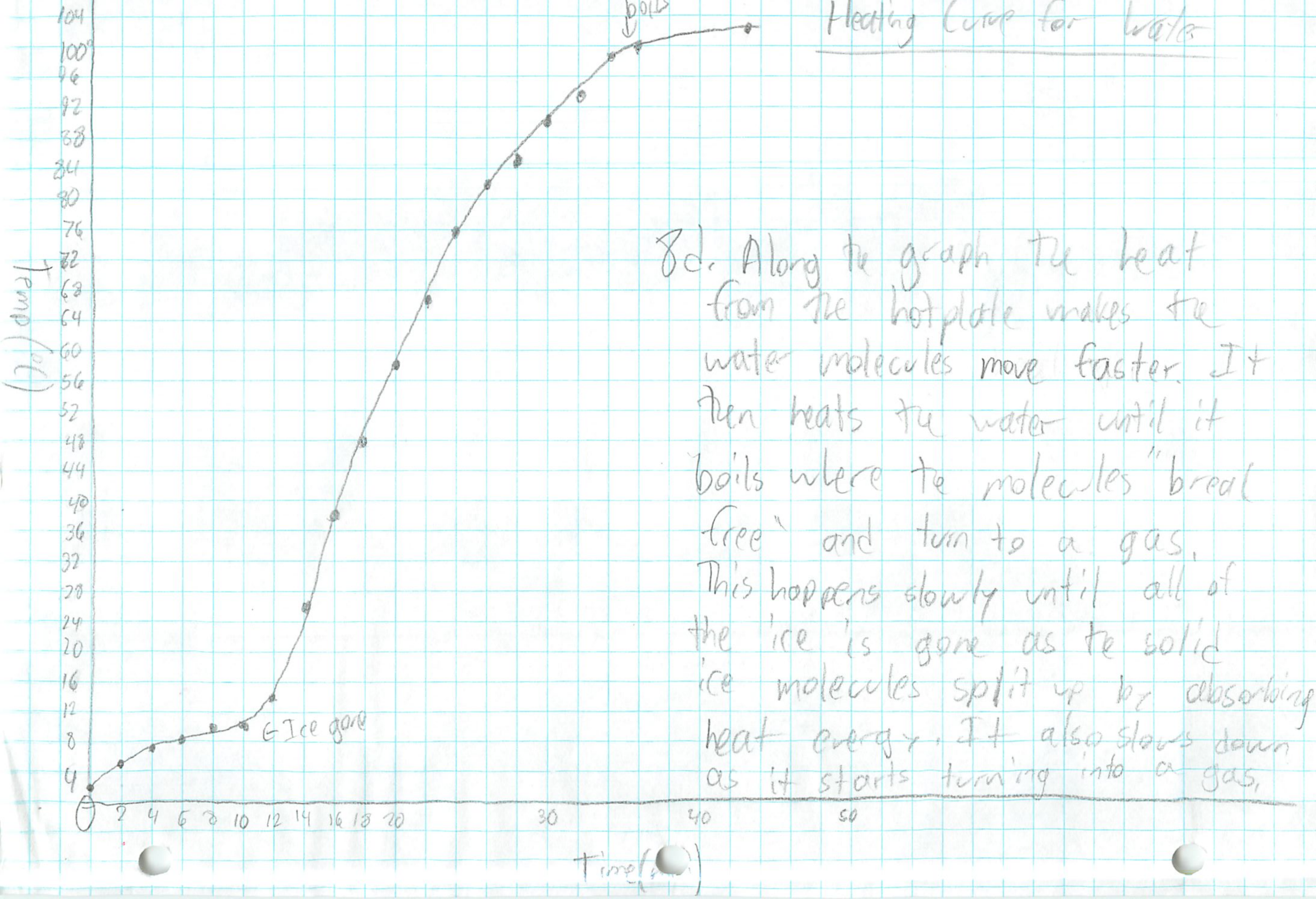
~175 mL water

Time (min)	Temp (°C)	Observations
0	2	Ice floats top of water, thermom looks frozen
2	5	Ice still above ~~water~~ - very visable under water
4	7	Ice somewhat submerged - begining to melt
6	8	Ice clearly melting - somewhat gone
8	10	Ice easily moveable - some cubes melted
10	10	Ice mostly melted
12	14	Ice is gone *
14	27	Temp 2xed - just water
16	38	Temp continues to ↑ rapidly
18	48	
20	58	condensation on glass
22	67	condensation
24	76	condensation
26	83	heavy condesation + slight steam
28	89	larger condensation + steam visable
30	91	small bubbles on bottom + thermometer
32	95	small bubbles + steaming
34	99	more, bigger bubbles
36	100	**boiling** - lots of condensation + bubble fast
38		
43	102°	Bubbling vigorsly, condensation + steam

Part 1 Boiling Frozen Water

Michael Plasmeier

Heating Curve for Water



8d. Along the graph the heat from the hotplate makes the water molecules move faster. It then heats the water until it boils where the molecules "break free" and turn to a gas. This happens slowly until all of the ice is gone as the solid ice molecules split up by absorbing heat energy. It also slows down as it starts turning into a gas.

Part C Volume Changes

2/12

1-3. Submerge a pipet with suspendend water plugs into warm and cool water. Observe what happens. Why?

The warm water makes the air in the pippet heats up and expands because of absorbed energy from the warm water. This forces the water plug upwards because of the increased vapor presure. In cool water the opposet happend. The cool water took energy from the particals making them stick closer together. This reduced the vapor pressure and caused the plug to return downward quickly.

4a. What happens to the preasure on the wall of the pippet as the air heats up.

The presure increased as the extra gas pushed out more againts the tube.

b-c. Draw box + show what would happen if pressure under piston would be increased.

The plug/piston would be forced upwards if the preasure would be increased.

Part D: Iodine

1. Place solid iodine in a closed tube. Submerge in hot water. What happens?
   The iodine sublimes and turns to a gas. The volume of solid particals is reduced and the walls of the tube would turn purple. Heat is being transferred to the iodine as its elements speed up and turn into a gas.

Chem to Go: Changing State

See seperate sheet 1st.

8 cont. and turn to vapor. It would then be condensed on the other side where it collects.

9. Materal - white solid - clear liquid - melts 146°C
   Is it sugar? Its a white solid and clear liquid. It also melts around 146°C From my experience with it I think its sugar

Chem Talk

MSE Activity 2 States of Matter

1.Matter is made up of tiny particles. The particles of different kinds of matter are different. (the same, [✓ differrent])

2.What are the relationships between energy and particle speed and particle spacing and attractive forces?

↑ energy = ↑ speed

Closer particles togeter the ↑ attractive forces

3.What is temperature?

measure of avg kinetic energy of all particles in a material

4. How can you make a thermometer?

Suspend a plug of water in a pippet

As the water moves up it gets warmer

5.What is kinetic energy? What is the relationship between kinetic energy and temperature?

energy related to motion of particles = $\frac{1}{2}mv^2$

proportional relationship b/w KE + temp

6.What happens to particles in a solid as it is heated?

The avg. kinetic energy increase + particles shake more

7.Define the normal melting point and normal freezing point.

point where material goes from solid to liquid/ liquid to solid at 1 atmosphere of presure

8.How can you tell if there is a change in kinetic energy?

Substance heats up or cools down

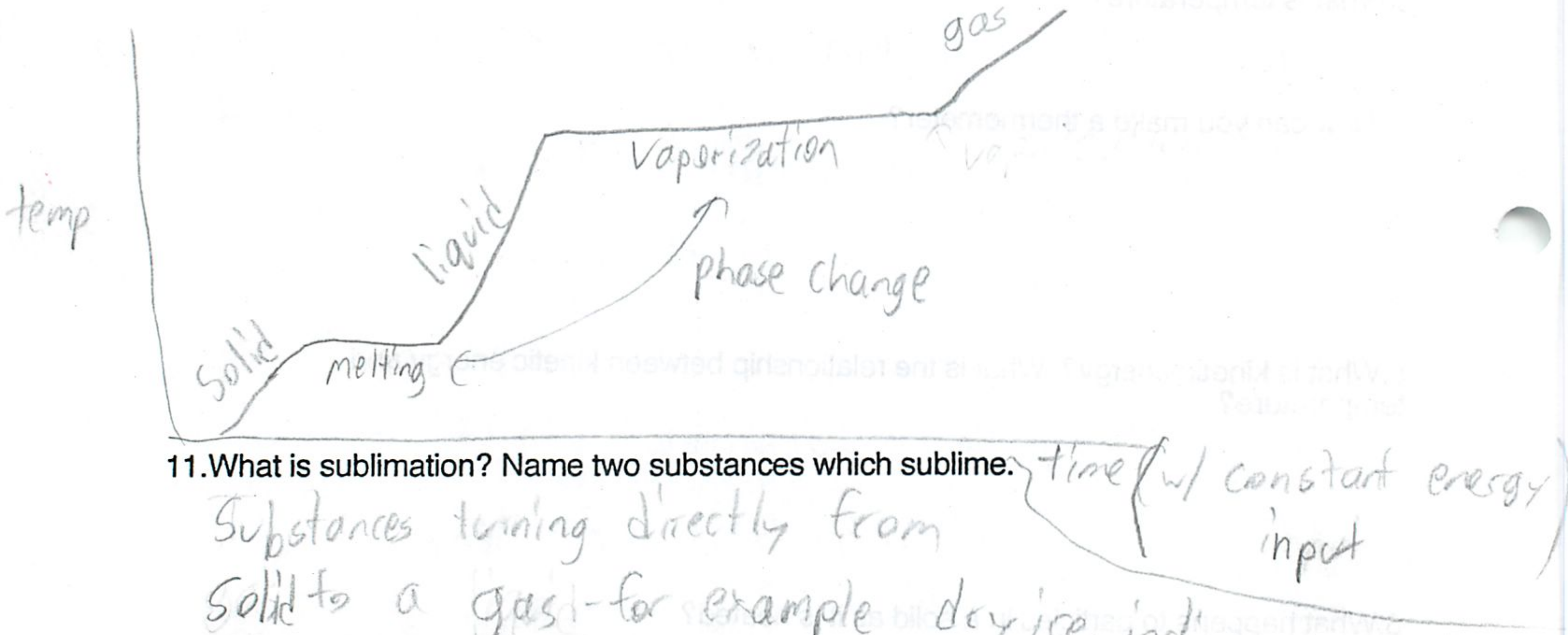
How can you tell if there is a change in potential energy?

Substance changes state

9. What happens to boiling points when atmospheric pressures are lowered?

Boiling tempratures for substances decrease

10. Sketch a heating curve of water and label all parts.



11.What is sublimation? Name two substances which sublime.

Substances turning directly from solid to a gas-for example dry ice, iodine, or moth balls

^ no water

DO CHEM TO GO 1-9

Chem to Go

Changing State

2/12

white p20

1. Copy chart:

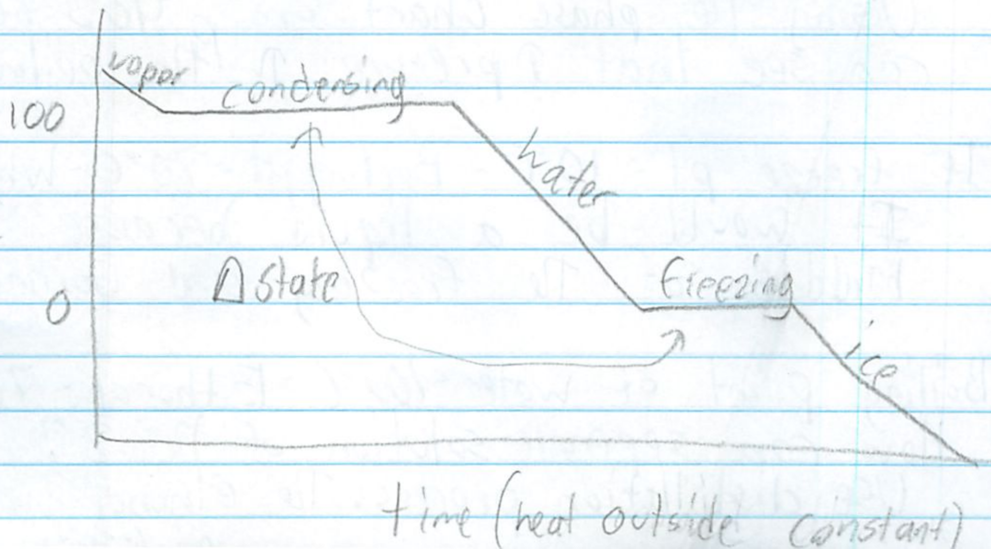
same!

Changing State	From	To	Heat Δ
boiling	liquid	gas	In - endothermic
condensing	gas	liquid	out - exothermic
evaporating	liquid	gas	in - end
freezing	liquid	solid	out - ex
melting	solid	liquid	ex
deposition	gas	solid	ex
sublimation	Solid	gas	end
vaporization	liquid	gas	end

2. Copy Chart!

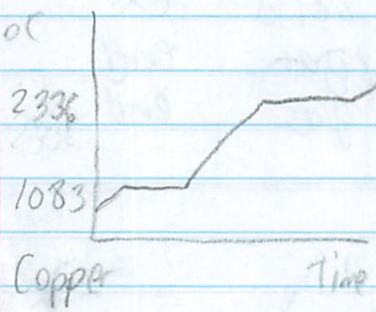
Definate?	Solid	Liquid	Gas
shape	Yes	No	No
volume	Yes	Yes	No

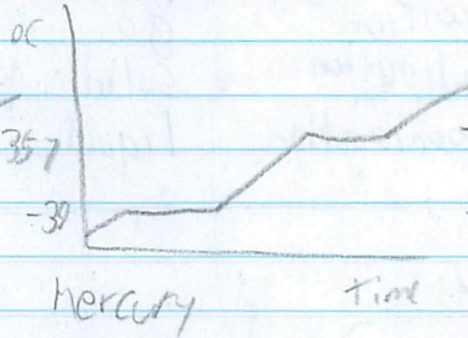
3. Draw Cooling Curve for Water

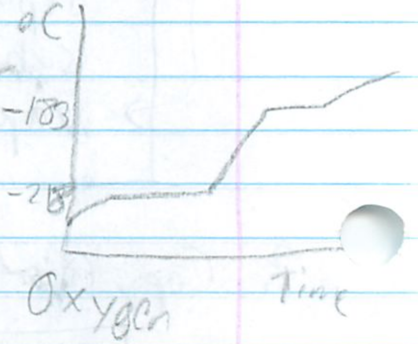


4. How is the heating curve for water different if you have 2x amt of water?
The time for everything lenghtens. The phase changes grow longer and the heating slopes decrease because the heat can disperse more among to water

5. Draw 3 heat curves from info on p21 White
Explain why all have diff states at room temp







Each substance is different and has a different state at room temp, Because boiling and melting points differ due to the elements makeup they are in different states,

6. If presure > 1atm how does boiling pt, of water Δ,
Using the phase chart on p403 red, I can see that ↑ presure ↑ the boiling point.

7. If freeze pt = 10°C + Boiling pt = 70°C - What state @ 22
It would be a liquid because it's in the middle of the freezing and boiling points.

8. Boiling point of water = 100°C - Ethonal = 78.5°C
How can seperate solution of the 2?
Use distillation process. The ethonal would boil 1s

Notes from Chem Tulla

what causes Changing States

2/13

solids - orderly arrangement of particals

gas - disorderly particals have little contact with other particals

liquid - particals move somewhat, not locked in place

gasses are compressible - you can squeeze them together

non polor - both ends are the same ++ or --

polar-ends are different + -

The larger the particles, the more attraction there is

polar particals have more attraction than non polar particles

boiling point is indicator of attraction

water has a high boiling points b/c its polarity

Name: Michael Plasmeier

20/20

Date: ________

Block: _______

Active Chemistry Unit 1.5: Solid, Liquid, or Gas?

What Do You Think?

What makes gases different from liquids and solids?

no/not as strong attraction b/w particles

The particles are much more spread out - tey don't have a defined shape or volem - ??

Describe what you think gases look like at the molecular and atomic level.

Particles spread out bancing around

interspersed w/ air

Describe how the molecules and atoms move and interact with each other.

Molecules move is a straight line until they collide in an elastic collisions

kinetic energy

Goals

- Describe how size and shape of molecules affect physical phase
- Classify molecules as polar or nonpolar

Investigate

1. Your group will make ball and stick models of a few of the molecules listed below. Your teacher will assign the molecules.

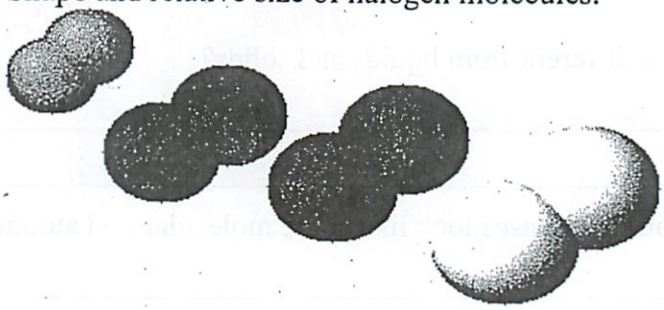
Station 1: F_2, Cl_2, Br_2, I_2
Station 2: CH_4, CF_4, CCl_4, CBr_4,
Station 3: CH_4, C_2H_6, C_3H_8, C_4H_{10}
Station 4: CH_3OH, C_2H_5OH, C_3H_7OH, C_4H_9OH

2. After you make your molecules, put each model in the correct station that your teacher has set up around the room.

When each group has completed their models, you will visit each station to answer the following questions.

For station 1, halogen molecules:

Shape and relative size of halogen molecules:



3. What do you notice about the shape of the molecules in station 1?

size

They have varrying sizes with an almost 3x larga radius from F_2 to I_2

4. What about the sizes?

shape

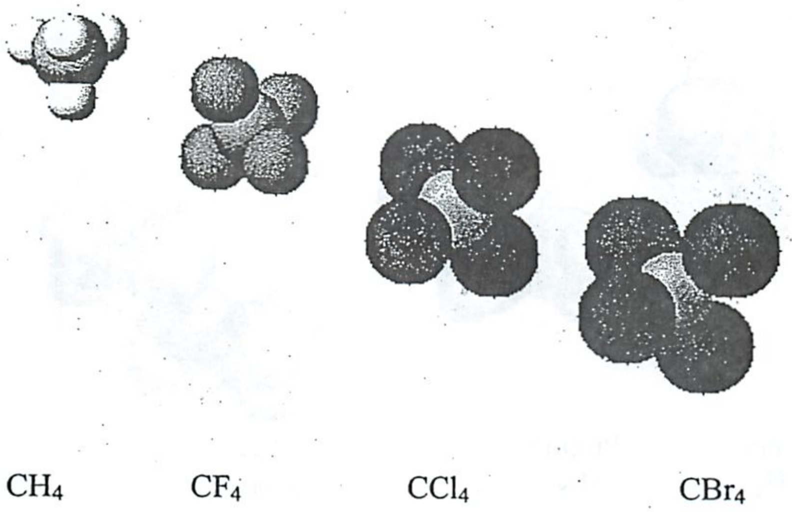
All round, all attached and attracted togeter

5. Make a guess as to which ones represent a solid, liquid, or a gas.

Are they all solids because they are taching — or perhaps there is no way to tell

For station2, CX_4 molecules:

Shape and relative size of CX_4 molecules (where X = H, F, Cl, or Br):



6. What do you notice about the shape of the molecules in station 2?

All round, all touching; always one in center with 4 on the outside

7. What about the sizes?

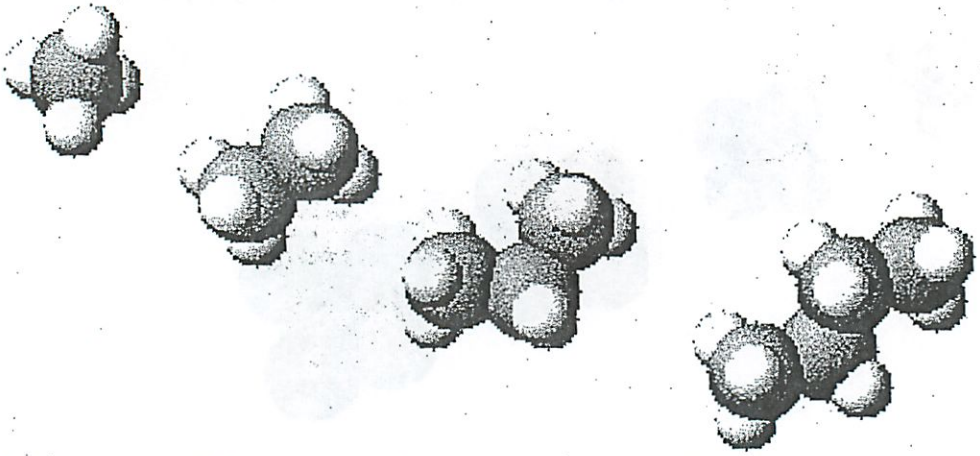
Central molecule all the same size (the carbon) - outside molecules varry again

8. Make a guess as to which ones represent a solid, liquid, or a gas.

no way to tell

At stations 3 and 4, the molecules are similar. CH_4 is methane and CH_3OH is methanol. C_2H_6 is ethane and C_2H_5OH is ethanol. The other pairs are propane and propanol, and butane and butanol.

Shape and relative size of hydrocarbons:



Methane CH_4

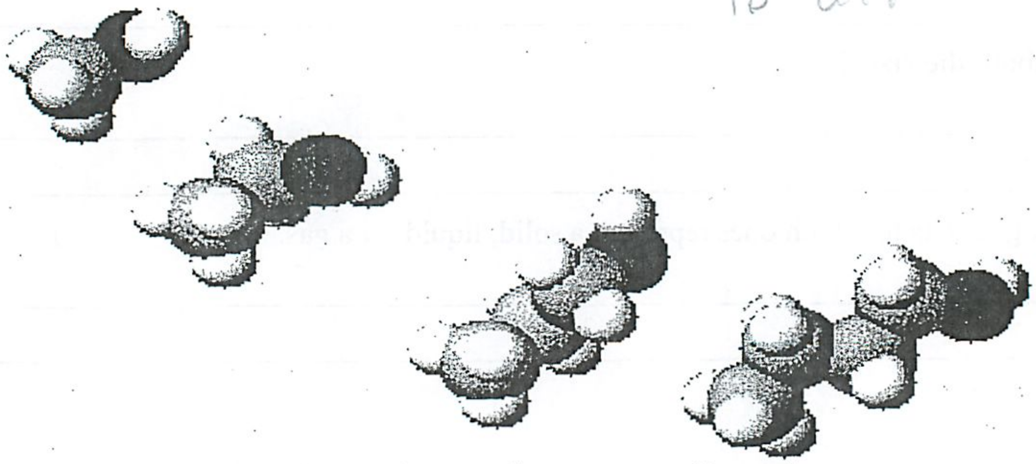
Ethane C_2H_6

Propane C_3H_8

Butane C_4H_{10}

add an oxygen and move an h to after that

Shape and relative size of alcohols:



Methanol CH_3OH

Ethanol $C_2H_5\ OH$

Propanol $C_3H_7\ OH$

Butanol $C_4H_9\ OH$

9. Looking carefully at the shape, describe any differences that you notice between the molecules at the two stations.

They add an oxygen molecule to the end and
bump an "H" molecule to the end of the oxygen

10. Make a guess as to which ones represent a solid, liquid, or a gas.

Again no way to tell by looking at a molecule

The tables and graphs below list the boiling points of the groups of molecular substances. Room temperature is about 25° C. If a substance has a boiling point below this temperature, it must be a gas at room temperature. If its boiling point is above 25° C, then it is a liquid (or possibly a solid) at room temperature. If its melting point is above 25° C, then it is a solid at room temperature.

Melting and Boiling Points of Halogens (group 7A on periodic table)

Substance	Melting Point, °C	Boiling Point, °C
	Solid / Liquid	Gas
F_2	-220	-188 R
Cl_2	-101	-35 R
Br_2	-7 R	59
I_2	R 113	184

larger ↓

11. Which of the Halogens are gases at room temperature?

$F_2 + Cl_2$

12. Since all of the halogens have the same shape, what is the difference between those that are gases and the ones that are liquids or solids? (Hint: see the diagrams of the halogens on page 2 of this packet)

The larger the molecule the more likely its a liquid then
a Solid, ^ and then

Melting and Boiling Points of CX_4 Compounds		
Substance	Melting Point, °C Boils 1st	Boiling Point, °C
CH_4 methane	-113	-164 G
CF_4 carbon tetrafluoride	-184	-128 G
CCl_4 carbon tetrachloride	-23 L	77
CBr_4 carbon tetrabromide	S 90	190

13. Of the CX_4 compounds, which are gases, liquids, and solids at room temperature?

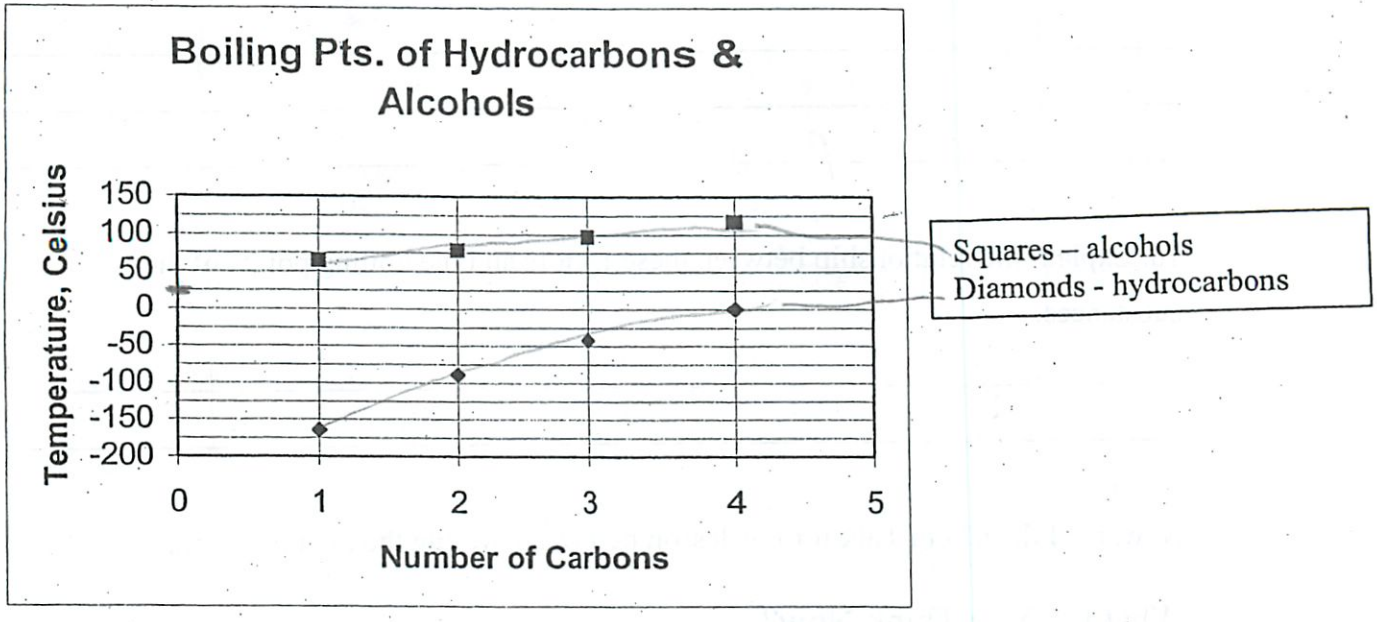
Solids: CBr_4

Liquids: CCl_4

Gases: CH_4, CF_4

14. Since all of the CX4 compounds have the same shape, what is the difference between those that are gases and the ones that are liquids or solids? (Hint: see the diagrams of the halogens on page 2 of this packet)

Outside (X) molecules are of different sizes



15. Which of the hydrocarbons and alcohols are gases at room temperature?

All of the hydrocarbons but none of the alcohols

16. Comparing the size of methane and methanol, ethane and athanol, and so on, we can see that there is not much difference between each pair. What other factor might have an influence on the boiling points? Explain your answer.

Addition of an oxygen molecule

17. Based on the information that you have from this lesson so far, what two factors seem to be a general rule that might be used to determine if a substance is a solid, liquid, or a gas?

Size and number of particles of a molecule
- force of attraction
→ polarity

18. Explain the relationship between these factors and the boiling points of the substances.

The less the molecules or the smaller they are the lower the boiling point,
- b/c less force of attraction

Now, read the Chem Talk for this lesson before answering the next question.

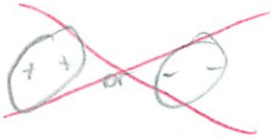
What Do You Think Now?

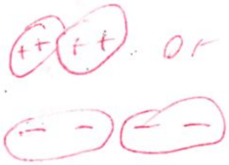
Revisit your response to the What Do You Think question. Now add to your thoughts about the microscopic nature of gases. Be sure to include the effects of size and shape of gas molecules.

The more polar particles attract other particles more as well as other molecules determining how "easily" they unjoin or join from one another as they turn from one state to another.

The compounds which are "easily" (at room temp, 1 atm) gasses are the ones which are symetrical and have small particles.
- little attraction + lots of space

Chem Words

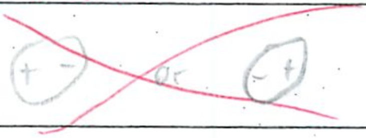


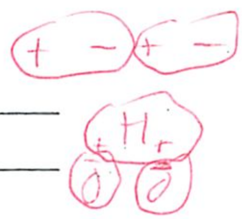


- nonpolar

both ends ^ are the same
on touching moleles

- polarity





ends are different

- polar

? having a charge

Be sure to attach the Chem to Go to this packet. (See worksheet)

1. ? The smallness of the particals in relation to the law of gravatational attractions as well as their less polarity. Lastly the small size leaves from for air

2. See page

3. F_2 -188°C -NP
CO_2 Mid -NP
C_3H_7OH ~95°C -P

liquid room temp
boiling point

? (4. The diethylether because it has more particles making it bigger. This also must mean that the polarity varries more on each side. Plus you know it has a higher boiling point meaning the force of attraction must be stronger.

What Do You Think Now?

Revisit your response to the What Do You Think Now question. Now add to your thoughts about the microscopic nature of gases in your *Active Chemistry* log. Be sure to include the effects of size and shape of gas molecules.

Reflecting on the Activity and the Chapter Challenge

You now understand that, at the microscopic level, the size and shape of molecules determine whether or not the substance is a gas at room temperature (which is what you see at the macroscopic level. The molecules of most gases have very little attraction for one another and there is space between the molecules. Because of this, gases are easily compressed, something that is not easily done with most liquids and solids. This can easily be represented symbolically by the use of drawings.

Many toys take advantage of the compressibility of gases.

Chem to Go

See packet

1. Explain why most gases have very little attraction between their molecules.

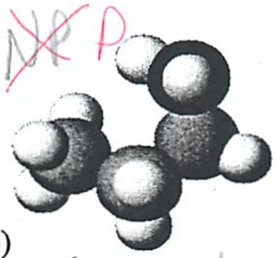
2. The following molecules are drawn to scale. Label each as polar or nonpolar.

a)

NP

CO_2

b)



~~NP~~ P

C_3H_7OH

c)

NP

F_2

Polarity
dependent
on end or
whole thing
if diff

3. Rank the molecules listed in #2 from lowest to highest boiling points. Only one is a liquid at room temperature. Which one is it?

4. Dimethyl ether has a boiling point of -23 °C, while diethyl ether boils at 34.5 °C. Which molecule has a greater polarity? Explain

CH_3-O-CH_3	CH_3CH_2-O-CH_2CH_3
dimethyl ether	**diethyl ether**

→

Preparing for the Chapter Challenge

Prepare a list of all of the toys that you know use a gas in some manner. List the gases used and the purpose of each gas.

Michael Plosmaier

Scientific Measurements

Essay $\frac{20}{20}$ 2/15

red p62

500 words

Measuring is very important in science. It lets us get quantitative data, which can be easily compared to different results, in addition to qualitative observations. But measuring has its rules and is expressed some what differently from everyday measurements.

39

For starters, scientific notation lets us express numbers which are very large or very small. This makes reporting and recording numbers eaiser and more accuratly. In scientific notation, a given number is written as the product of two numbers a coefficient and 10 raised to a power. The coefficient is always between 1 and 10. For example

$$7,305 \cdot 10^7 = 73050000$$

This also works in reverse for negitive numbers

$$7,305 \cdot 10^{-7} = 0,0000007305.$$

108

In addition we must know the difference between accuracy and precision. Accuracy is the measurement of how close a single measurement compares to the actual or expected value. Precision in science is the measurement of how close a series of measurements are to one another. For example if darts on a dart board land far from the bull's eye but close to one another, they are said in science to have high precision but low accuracy. Outside of science the two are often considered to be the same.

196

No measuring tool is perfect. All have some error. You can test the error by measuring a known quantity (determaind by a more accurate measurement device) and subtracting the resolved value from the known quantity. In this example the resolved value means the experimental value and the known quantity you are measuring by is the accepted value. The percent error can be calculated by multipling 100% by the quantity of the error figured above by the known or accepted value. An error can be positive or negitive. 283

When measuring something, you normaly know most of the siginificant figures in a measurement by seeing that the mark is between two calibration or measuring marks. You then estimate the location between the two measuring marks to get the final significant figure. 325

To determain if a digit is significant, you must follow the following rules

1. Every nonzero digit is assumed significant.
2. Zeros b/w non zero digits are significant
3. Leftmost zeros in front of non zero digits are not significant
4. Zeros at the end of a number and to the right of a decimal point are always significant ?(b/c show precision)? 383

5. Zeros at the rightmost end of a measurement to the left of a decimal point are not significant unless they show the magnitude of a number.

6. Defined quantities have an unlimited number of significant figures. For example
423 100cm = 1m

When making calculations, you must be careful that you do not add precision where there is non. For example when multiplying $7.7 \cdot 5.4$ you get 41.58. However because the factors have only 2 siginificant figures, the answer can only be reported at that accuracy, in this case to two siginificant figures or 42.

To adjust numbers to be the proper siginifant figures, first round the number counting numbers from the left to the right. Then check the next number. If it is 4 or below keep the last diget the same. If it is 5 or above round the last didget up to the next integer.

When adding 2 numbers round to the minimum number of decimal places as indicated by the number with the least number of decimal places.

In multiplication and divison round the product to the factor with the least number of significant figures reguardless of decimal points.

Measurement Notes

2/16

Measuring

- need accuracy
  - closeness to "true" value
- need reproduceabilty
  - or precision

read at eye level at miniscus (water + solution)
↑ lowest point

each line = 1 mL
on this 100 mL - estimate to x,8, x,5 or x,2 plac

- determain vale at each mark on measuring device
- estimate to one more decimal place

Rounding

0-4 - drop
5-9 - add 1 to next decimal place

Zeros + siginificant figures

- between nonzero digits $7\underline{0}$ g
- when they result of randing or actual measurements $\underline{10,0g}$
  or 3,2012 → 3.20

significant

not significant

- holding decimal places - 57000000
  - ,00075
- to clean, put in exponental notation

$$\underbrace{5,7000000}_{7\;6\;5\;4\;3\;2\;1} = 5,7 \times 10^7$$

Density Notes

2/16

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{g}{mL}$$

$$\frac{14.37\text{ g}}{6.7\text{ mL}} = 2.14477 = 2.1\text{ g/mL}$$

2 sig. figuers

2 sig figures

round to least accurate measurement (least. sig. figures w/ × ÷)

$$\begin{array}{r} 3.217 \\ 2.96 \\ 1.05 \\ +\ .092 \\ \hline 7.269 \end{array} \Rightarrow 7.27$$

least acc figuers (by decimal places for + −)

Michael Plaender Measuring

Mass and Volume
→ Density Lab

20/20

2/20

rounding
& graduated cylinder - significant figures

white
act 5
p34

learn proper scientific measurement techniques by measuring the mass and densities of objects

What do you think: A kg of lead and feathers have the same mass, How do they appear different?
They have different densities and physical properties. A kg of feathers spreads its mass out because less particals are packed into a certain volume, In addition when dropped, feathers encounter more air resistance

Part A

Procedure: 1. Copy the chart on the top of p 35 white

2. Measure + record the mass of a graduated cylinder

3. Put 10 mL of water in the cylinder + record this

4. Measure + record the mass

5. Repeat every 10 mL until 100 mL

6. Graph data + answer qu + record slope + predict line

7. Divide mass by volume - notice anything?

3 sig figures
↓

Mass cylinder (g)	Volume water (mL)	Mass cyl+water (g)	Mass water (g)	Density Mass/volume (g/mL)	Density → sig figures (3)
42.8	00.0	42.8	0	—	—
42.8	10.0	52.9	10.1	1.01	1.01
42.8	20.0	62.9	20.1	1.005	1.01
42.8	30.0	73.2	30.4	1.013	1.01
42.8	40.0	82.9	40.1	1.0025	1.00
42.8	50.0	92.7	49.9	.998	1.00
42.8	60.0	102.7	59.9	.9983	1.00
42.8	70.0	112.7	69.9	.9985	1.00
42.8	80.0	122.8	80	1	1.00
42.8	90.0	132.7	89.9	.9988	1.00
42.8	100.0	142.7	99.9	.998	1.00

max error 30.4 - 30 = .4 g

$\frac{.4}{40}$ ·100 = 1%

See attached graph paper for graph + questions

Mystery Liquid

Mass (g) cylinder	Volume Liquid	Mass cyl+liq	Mass liquid	Density (raw)	Desity 2-3 sig fig
14.6	0.0	—	—	—	—
14.6	5.0	18.8	4.2	.84	.84
14.6	10.0	22.6	8.0	.8	.80
14.6	15.0	26.7	12.1	.8066	.806
14.6	20.0	30.8	16.2	.81	.810
14.6	25.0	34.9	20.3	.812	.812

Part B Mass + Volume of Solids

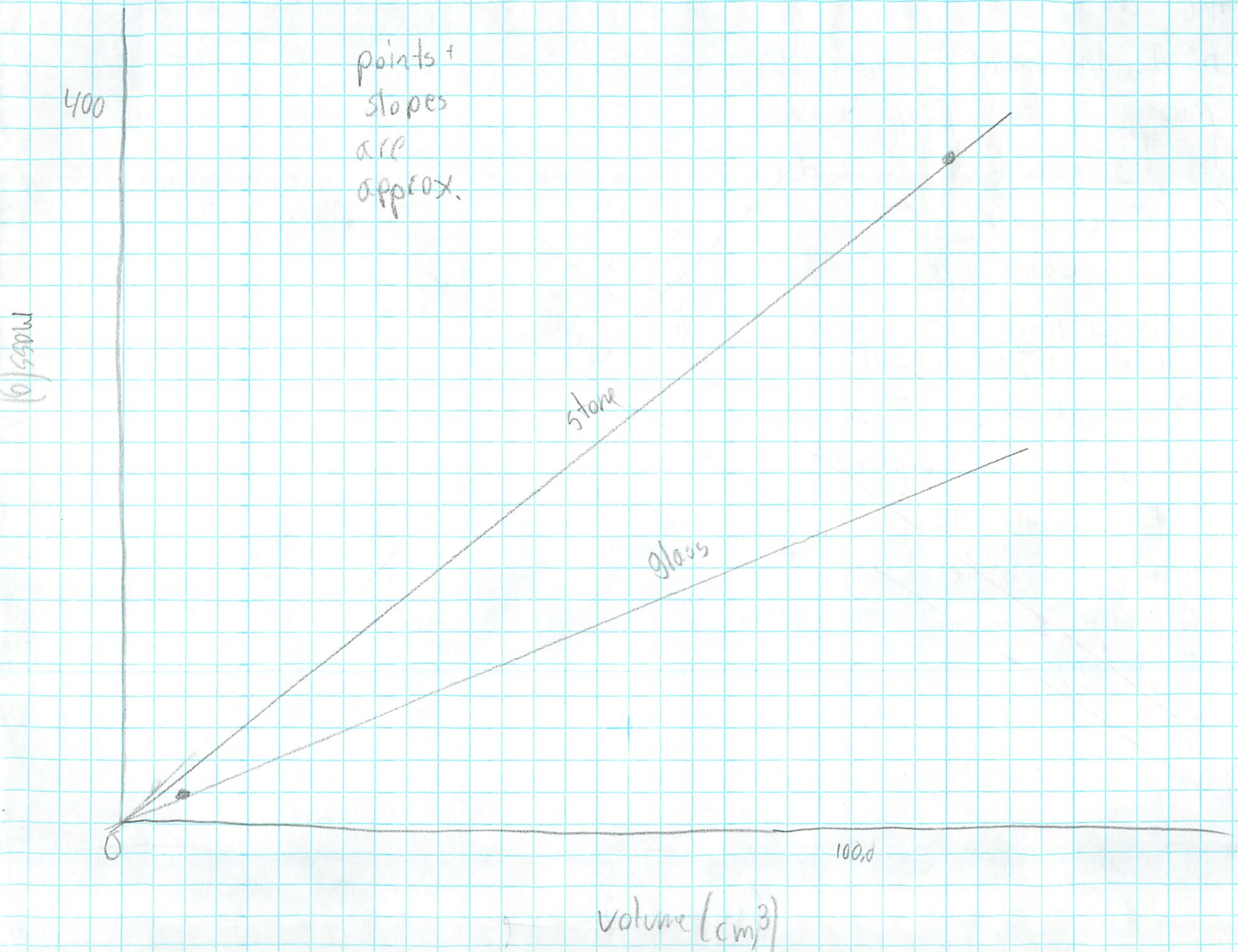
1, Calculate mass + volume + density of three solids. Use cubic centimeters instead of mL - ($1 cm^3 = 1 mL$). To calculate volume measure the overflow from an overflow can by catching it in a graduated cylinder

Solid	Mass (g)	Volume (cm^3)	Density (g/cm^3)	Sig Fig (2)
little glass thing	15,1	6,5	2,32	2,3
stone (c)	370,6	112,0	3,31	3,3

4, Calculate the density and plot point on graph

4b, The slope for the stone is steeper because it is more dense then the glass thing,

Part



points + slopes are approx.

Part 1

#6 a:

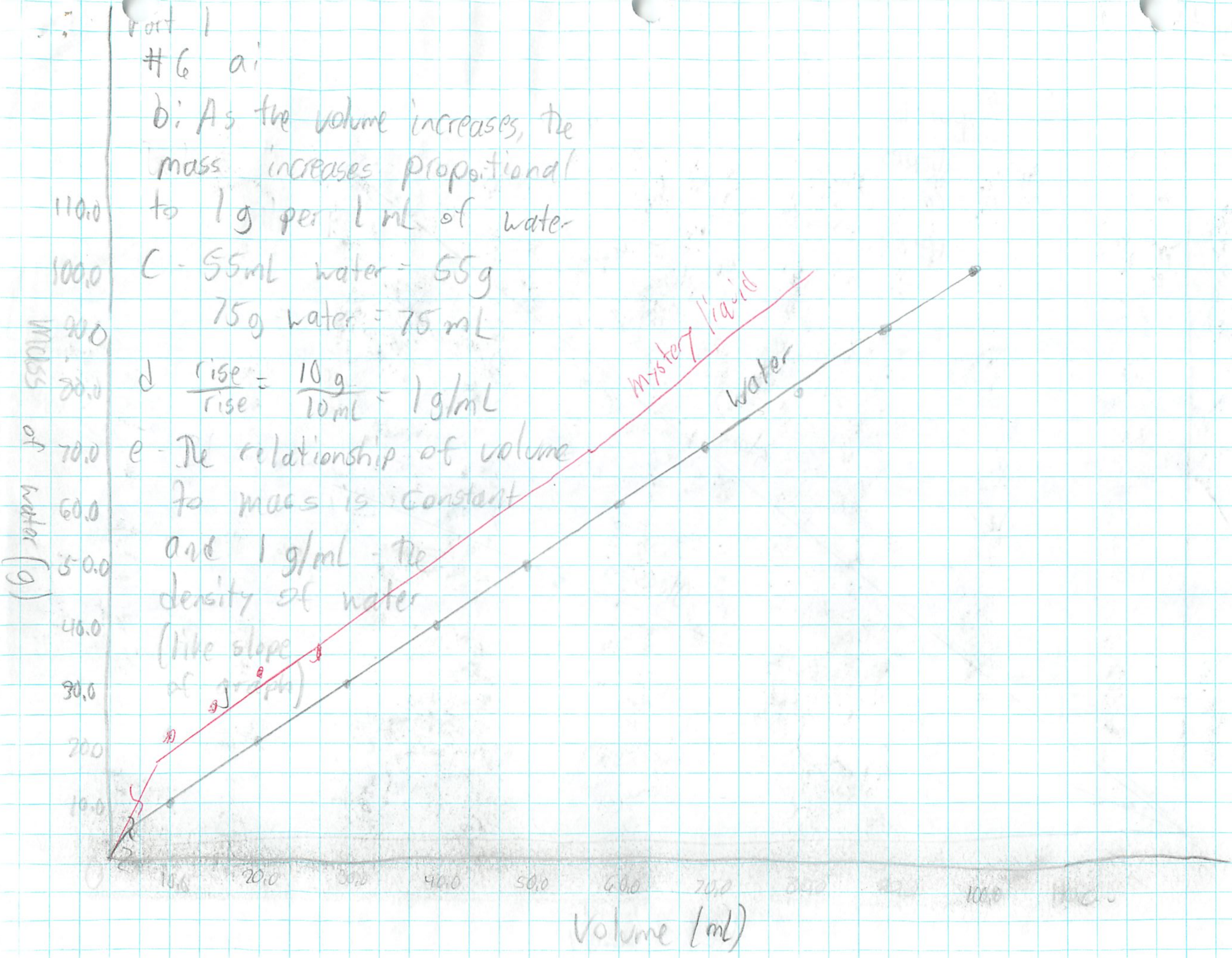
b: As the volume increases, the mass increases propotional to 1g per 1mL of water.

c - 55mL water = 55g

75g water = 75mL

d $\frac{rise}{rise} = \frac{10\,g}{10\,mL} = 1\,g/mL$

e - The relationship of volume to mass is constant and 1 g/ml the density of water (like slope of graph)



Act 5 - Mass + Volume
Chem Talk + To Go

2/3/

white
p38-41

Vocab

density - the mass per unit of volume of a material

Checking Up

1. Explain the meaning of Density
Density is the amount of mass or matter in a certain volume of an object. This is a fair comparison of "heavyness" of an object because it removes volume as a factor.

2. Difference b/w feathers + lead
Lead is so much more denser than feathers because so much more matter or molecules are packed in a certain area. This is why lead feels heavier than a feather

3. Why use balsa not birch wood for airplanes
Balsa wood is less dense, so the same amount of it would weigh less, making easer and less costly (for fuel) to get the plane off the ground.

4. $\frac{253g}{80.0cm^3}$ = 3.1625 → 3 sig fig → 3.16 g/cm

white
42-43

Chem to Go

1. Using chart on white p39, find what substances we investigated

Stone - not listed - 3.3g/cm³

Water - 1 g/cm³ (✓) Glass - not listed - 2.3 g/cm³

2. Calculate Density

$V_{w+s} = 62,7$ $M_s = 123,4g$

$- V_w - 48,4$

$V_s = 14.4cm^3$

$\frac{123,4g}{14.4cm^3} = 8,63g/cm^3$

3. Calculate Density

$M_{C+L} = 40,14$ g $V = 13,3mL$

$M_C - 33,79$ g

$M_L = 6,35$ g

$\frac{6,35g}{13,3mL} = ,477$ g/mL

4. Calculate Mass

Density = ,79 g/mL $589 \cdot ,79 = 465$ g

V = 589 mL

5. Calculate Volume

Density = 8,90 g/cm³ $746 / 8,90 = 83.8$ cm³

Mass = 746 g

6. a Can you replace gold w/ sand of same volume?
No it will not be the same weight - the trap will go off

b What's mass w/ 1000 ml of 19.3g/mL gold?
$1000 \cdot 19.3 = 19300$ g or 19.3 kg - no that's too heavy to easily hold

c - Mass of 454 g is 4.45N (≈1lb) - What's the statue
around 42 lbs - again too heavy to hold

d - one gallon of milk is 3,7 kg (≈8lbs) - how many gallons of milk is the statue - around 5,2 gallons of milk weighs the statue

7. Which has a greater mass?
a. 1 kg lead or 1 kg feathers - both the same (tricky)
b. 1 L gold or 1 L water - water - 1000 mL • 1 g/mL = 1000 g
gold - 1000 mL • 19.30 = 19,300 g - the gold has ↑ mass
more dense thing
c. 1 L copper or 1 L silver - copper - 1000 mL • 8.90 g/mL = $8.90 \cdot 10^3$
silver - 1000 mL • 10.50 g/mL = $1,050 \cdot 10^4$ - silver has ↑ mass

8. Which has greater volume?
a. 1 kg lead or 1 kg feathers → common sense says feathers
data not given
less dense thing
b. 1 kg gold or 1 kg water - water - 1000 g / 1.00 g/mL = $1.00 \cdot 10^3$
gold - 1000 g / 19.30 = 51.81 mL - the water has > volume
c. 1 kg copper or 1 kg silver - silver - 1000 g / 10.50 = 95.23 mL
Copper - 1000 g / 8.90 g/mL = 112 mL - copper has a > volume

How certain where you?

9. I was pretty certain to a mL - better measuring tools would make answers more certain

10. Fix measurments for significant figures
I had that in mind the whole time

Inquiring Further

1. Is it real gold?
You could measure the mass + volume to calculate the density to determain if something is real gold

Compressing
Δ volume
- can measure by pouring water in container

2. Density of air
Could ya compress the gas to calculate its volume when max compressed and weigh the container w/ + w/o the gass and then continue as usual

Test 1 Review — 25 multiple choice q/v 2/23

Active Chem Review of Movie Special Effects

1.In activity 1 you separated a compound into its elements.

2. Describe the properties of the gases Hydrogen, Oxygen and Chlorine.

H - clear, colorless odorless -explosive
O - clear, colorless odorless relit glowing splint
C - clear, yellow green - smells like bleach
Poison

gas at room temp

3.How are these gases different from the compounds which produced them?

water - liquid - puts out fire

copper$_2$ cloride - ~~solid, black green, odorless~~
liquid, clear, colorless — Odor ← ? same

4.You learned the symbols for the elements and how to interpret chemical formulas.

Give the name for each symbol: Ca Calcium

Cl cloride

Hg ~~Hydrogen~~ Mercury

Ni nickel

Iron Fe

Sulfur S

Sodium Na

Potassium K

P phosphrous

Cu Copper

lead Pb

tin Sn.

u? sp

Florine = F ≠ Fl

5.How many of each kind of atom are in this formula: $(NH_4)_3PO_4$

N H H H H

P O O O O

N H H H H

N H H H H

6.Describe the difference between an element and a compound.

element-Can't be seperated furter by chemical means

Compound - made of elements - sep, by chem means

has chemical formula

7.Describe the arrangement of particles in solids, liquids and gases.

Solid - stuck ridgetly togeter, vibrate

Liquid - bance around a bit - still stuck together free to flow around each other

Gas - porticals have free reign

8. How are pressure, temperature and volume affected by changes in energy.Discuss Kinetic and potential energy.

temprature changes the kinetic energy in matter

pressure changes the boiling + melting point

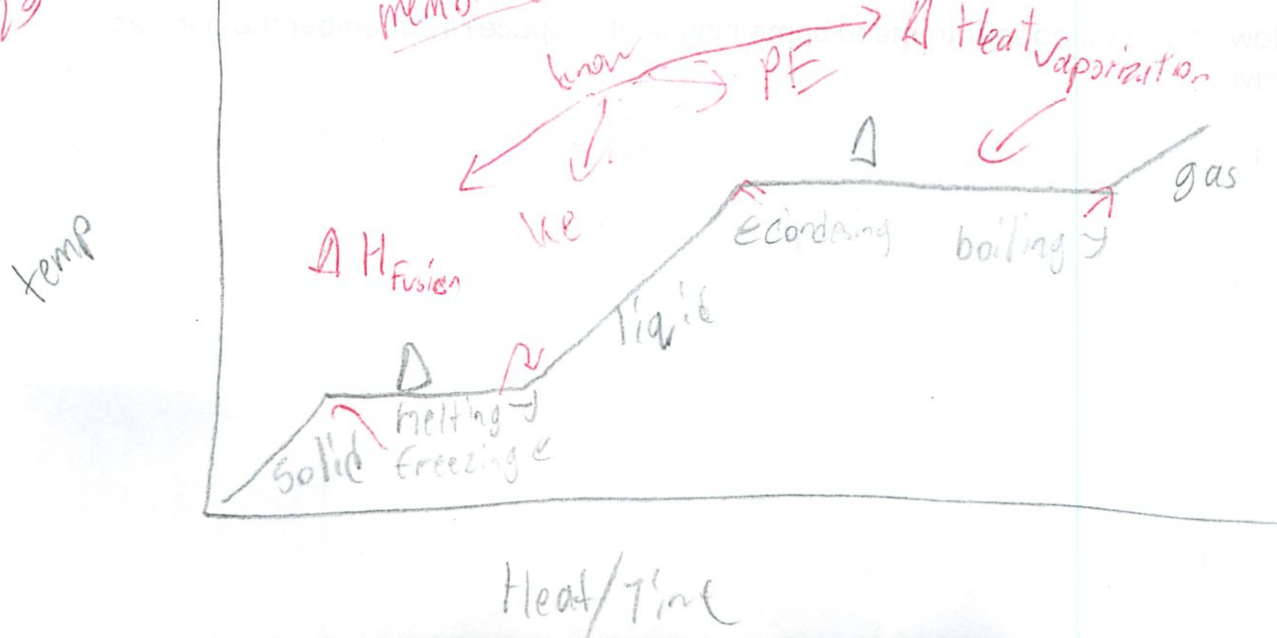
? larger volumes take longer to absorb energy

? potential-energy of position - bonds held in akward positions makes more likly to catch on fire

kinetic-energy of ~~position~~ motion

9.Sketch a heating curve and label each phase and each change of phase.

Red p523



10.What is sublimation? How can this be used in a special effect?

Solid -> gas no liquid

dry ice creates fog to create spookiness

11.Describe the liquid mixtures and explain how to distinguish among them.

Combo of 2 elements/compounds - seperate physically, not chemically

hetero -> suspensions - Tyndal - yes - filter yes
2+ P.

Colloids - Tyndal - yes } filter - no

homo
1 phase -> Solutions - Tyndal - No! ← clear

12.How could you set the mood for a horror movie with a suspension or colloid?

~~blood in water~~ ?

Smoke } in air
fog

13.How could you appear to make something float in space? Remember the pen cap in activity 5.

Have it in a liquid with the same density as the surrounding liquid

14. Explain how to round numbers that result from calculations.

5-9 - round up

0-4 - same

Sig fig

round to least accuracy when × ÷ -example 4digit × 2digit = 2digit

round to least decimal places when + −

7,32

353,7

15. Using your knowledge of density and materials, how could you make an ant appear to lift a large piece of metal?

~~liquid density trick~~

or light metal-looking substance.

Yes

16. How can a mixture of salt , sand and ammonium chloride(a solid which sublimes) be separated?

heat until ammonium chloride sublimes

add water then mix then filter to remove salt

(condense if needed)

Sand is left over

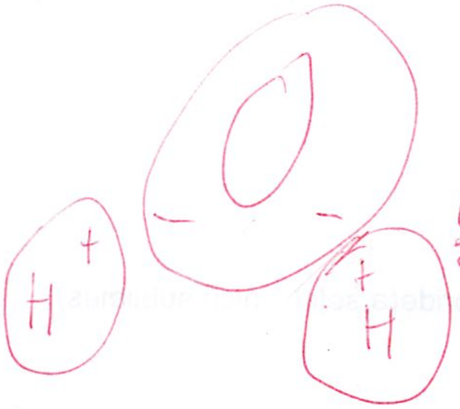
17.In your own words,explain why some materials exist as solids, liquids or gases at room temperature.

needed reminder

diff levels of attraction b/c polarity + size

18. Polarity results when ......

diff charge particals attract



~~both ends different~~

~~so not that~~

○ RESCORE ○ MARK ✓ ○ TOTAL ONLY

USE NO. 2 PENCIL ONLY

Use a No. 2 Pencil

Ⓐ ● Ⓒ Ⓓ Ⓔ Fill circle completely

Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ Erase cleanly

T F

1 ● Ⓑ Ⓒ Ⓓ Ⓔ
2 ● Ⓑ Ⓒ Ⓓ Ⓔ
3 Ⓐ Ⓑ Ⓒ ● Ⓔ
4 Ⓐ ● Ⓒ Ⓓ Ⓔ
5 Ⓐ Ⓑ ● Ⓓ Ⓔ
6 ● Ⓑ Ⓒ Ⓓ Ⓔ
7 Ⓐ ● Ⓒ Ⓓ Ⓔ
8 Ⓐ Ⓑ ● Ⓓ Ⓔ
9 Ⓐ Ⓑ Ⓒ ● Ⓔ
10 Ⓐ Ⓑ ● Ⓓ Ⓔ
11 Ⓐ Ⓑ ● Ⓓ Ⓔ
12 Ⓐ Ⓑ ● Ⓓ Ⓔ
13 Ⓐ ● Ⓒ Ⓓ Ⓔ
14 Ⓐ Ⓑ ● Ⓓ Ⓔ
15 Ⓐ Ⓑ Ⓒ ● Ⓔ
16 Ⓐ ● Ⓒ Ⓓ Ⓔ
17 Ⓐ Ⓑ ● Ⓓ Ⓔ
18 Ⓐ ● Ⓒ Ⓓ Ⓔ
19 Ⓐ Ⓑ ● Ⓓ Ⓔ
20 Ⓐ Ⓑ ● Ⓓ Ⓔ
21 Ⓐ ● Ⓒ Ⓓ Ⓔ
22 Ⓐ Ⓑ ● Ⓓ Ⓔ
23 Ⓐ ● Ⓒ Ⓓ Ⓔ
24 Ⓐ Ⓑ ● Ⓓ Ⓔ
25 Ⓐ Ⓑ ● Ⓓ Ⓔ
26 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
27 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
28 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
29 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
30 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
31 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
32 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
33 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
34 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
35 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
36 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
37 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
38 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
39 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
40 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
41 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
42 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
43 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
44 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
45 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
46 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
47 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
48 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
49 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
50 Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ

#8 - wouldn't solution be the best answer pure substance is one but not all of possibilities

a Solution isn't a pure substance

#17 - I've heard it refered to mostly as law of conservation of mass

Correct

#25 - Did we learn this?

opps
- I must have meant element
- But dosn't fit problem
(I see know

SCORE	25 100	# CORRECT % CORRECT
RESCORE		# CORRECT % CORRECT
ROSTER NUMBER	68	SCORE RESCORE

NAME Michael Plasmeier

COURSE Science - Chemistry Test 1

HOUR 3-52 06-07 DATE 2/28/07

TEST RECORD	
PART 1	
PART 2	
TOTAL	

TEST ANSWER SHEET J
Form No. 72512

ACCU-SCAN™ #26180-TR 10/00 70976 APPERSON PRINT MANAGEMENT SERVICES (800) 473-6761 05/05

DO NOT WRITE IN THIS AREA

Michael Plasmeier

Charles' Law

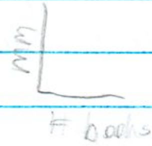
Pressure + Volume Lab

20/20

2/26

to determine the relationships between presure and volume at constant tempratures

1. Get a filled pipette. Shake all liquid into the bulb.
2. Measure the air space in the stem in mm
3. Carefully place a book on the bulb. Measure the air space in the stem.
4. Add up to 5-7 of the same size book + do #3
5. Graph data



Books	Airspace (mm)
0	110
1	75
2	67
3	53
4	48
5	39
6	35
7	31
8	29

QV Board

1. What does the mm of air represent?
2. What do the books provide?
3 Describe the shape of the graph
4. What is the relationship b/w pressure + volume?

Read p413-425 + do problems #1-21 (end of section only)
red

Ans

1. The mm of air is the volume of air not in the pipette and not occupied by liquid. It is ⓘcompressed as books are added.

2. The books provide a force, reducing the overall volume inside the pipette. They force the water away from the bulb and compress the air in the ti

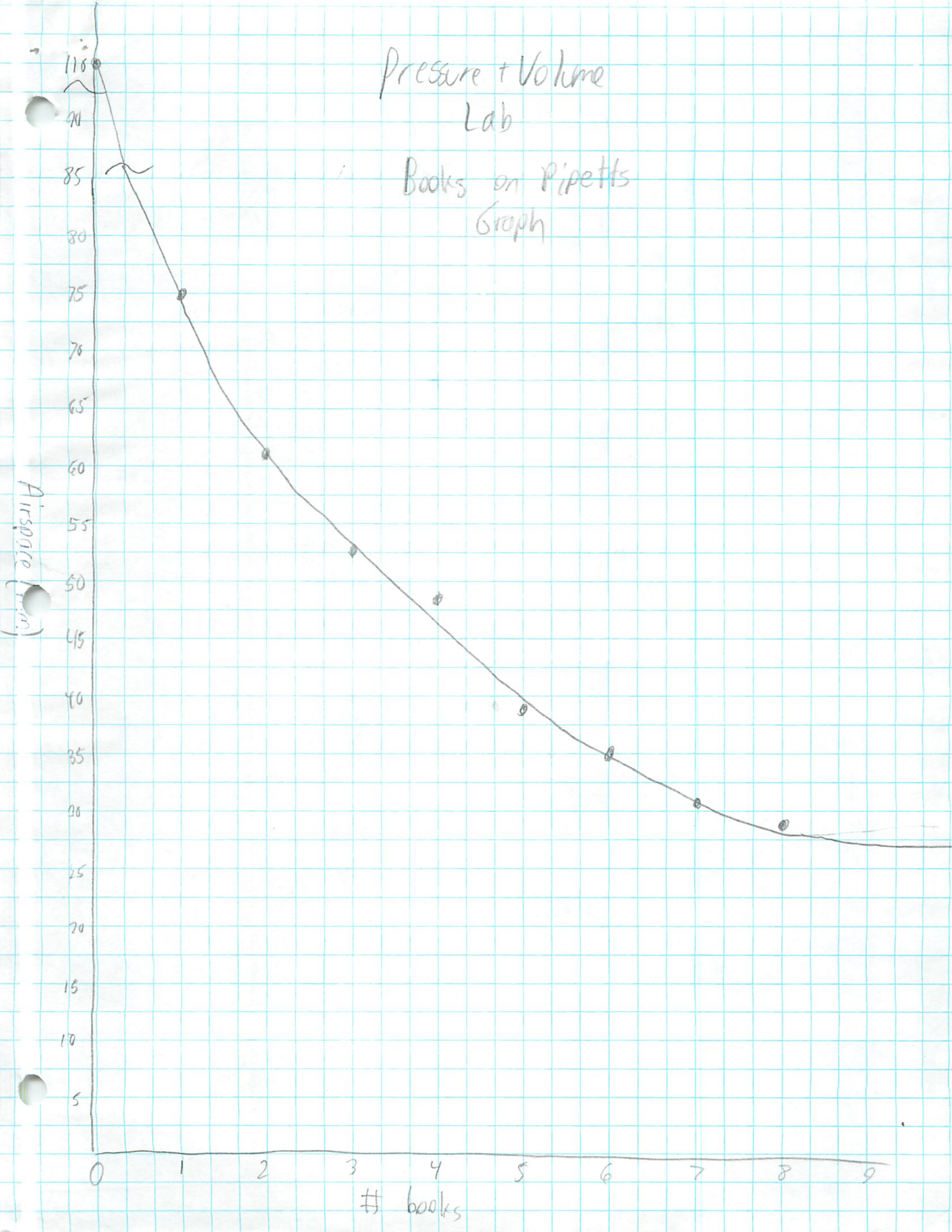
3. Its a down sloping convex graph. As more books are added, their marginal force is reduced as the marginal decreasement of air is reduced.

4. ? The smaller the volume, the higher pressure will become. Preasure increases as volume a substance (usually gas) decreases.

temp ↑ = pressure ↑ = ↑ avg KE

½ volume = 2x pressure
2x temp = 2x pressure

Boyle's Law $P_1 \times V_1 = P_2 \times V_2$



Pressure + Volume
Lab

Books on Pipetts
Graph

Properties of Gas Qu

2/27

red p 413-425 red book

14,1
p417 red
ection
ssement

1, Why is gas easy to compress?
Gas is easily compressed because of the space between the particles in a gas. Solids + Liquids have their particles Closer together.

2, 3 Factors that can affect gas pressure.
Gas, volume, and temprature all affect gas presure.

3, Why use airbags?
Airbags are better to crash into then a solid steering wheel because the gas can be eaisly be compressed which absorbs most of the energy of the crash. In a solid, the particles are closer together and don't move as eaisly.

4, What does ↓ temp. do to presure.
A decrease in tempeature decreases the average kinetic energy of the particles causing them to slow down and express less vapor presure againts the walls

5, What volume Δ would cause pressure to go to ¼ of value.
If the volume would be expanded 4 times, the pressure would be ¼ of its orignal value

6, but ↑ 100 x
If the volume would be compressed to 1/100 of its orignal value, the pressure would be increased 100 times.

Gas Laws

14.2
p425 red
Section
Assessment

15. How are pressure + volume related at constant temp?
As pressure ↑, volume ↓ - Boyle's

16. How does a change in temp affect volume of gas with constant pressure? - Charles
As temp ↑, volume ↑

17. What relationship b/w temp + pressure at constant volume
As temp ↑, pressure ↑ -Gay-Lussac

18. When use combined gas law?
When only amount of gas is constant.

19. Boyle's Law equation
$P_1 \times V_1 = P_2 \times V_2$
↑pressure ↑volume

20. Certain mass of air has 6.00 L at 101 kPa. At 25.0 kPa what volume? (constant temp)
6.00L × 101 kPA = 606 = 25 kPa × 24.24 L → 3 sig fig
24.2 L of volume

21. How Charles law derived from combined gas law?
Charles law assumes pressure is constant, so if you ignore the Ps in the combined gas law you get Charles law $\frac{P_1 \cdot V_1}{T_1} = \frac{P_2 \cdot V_2}{T_2}$

Temprature + Volume

2/28

$\frac{20}{20}$

to explore the relationship between volume and temprature while deriving the kelvin scale

1. Obtain 5 thin transfer pipetts, weights and a beaker
2. Completely fill one pipette with water
3. Count the drops in a pipette as you empty it
4. Place other 4 in beaker with water + low-mid heat
5. When the bath reaches 30°C, remove a pipette, place finger over end, sumerge without squeezing into bucket of water complotly for 10sec
6. Count # of drops which entered pipette
7. Repeat for 45°C, 60°C, 90°C
8. Graph data from -300°C to 100°C

°C	Drops		Avg
	1st	2nd	
23 - start room	84	using 118	118
30	+6 = 90	+7 = 125	
45	+14 = 98	+12 = 130	
60	+8 = 92	+31 = 149	
90	—	+50 = 168	

9. To what temp need to heat .750L of gas @ 25°C to get volume = 2.5L?

$$\frac{.750}{25} = \frac{2.5}{x}$$

$$25 \cdot 2.5 = .750x$$

$$\frac{63}{.750} = \frac{.750}{.750}$$

$$\boxed{x = 87\tfrac{2}{3}}$$

1. What vale does te x-intercept value represent
Te place were te temprature would ?
-200°C

0 k = absolute zero - no matter moves :)

2. What temprature scale does this define
The kelvin scale where 0 k = -273°C

3. Equation °C → k
k = °C + 273.15

4. Why values only in quad 1+2?
You can't have negitive drops or negitive matter.

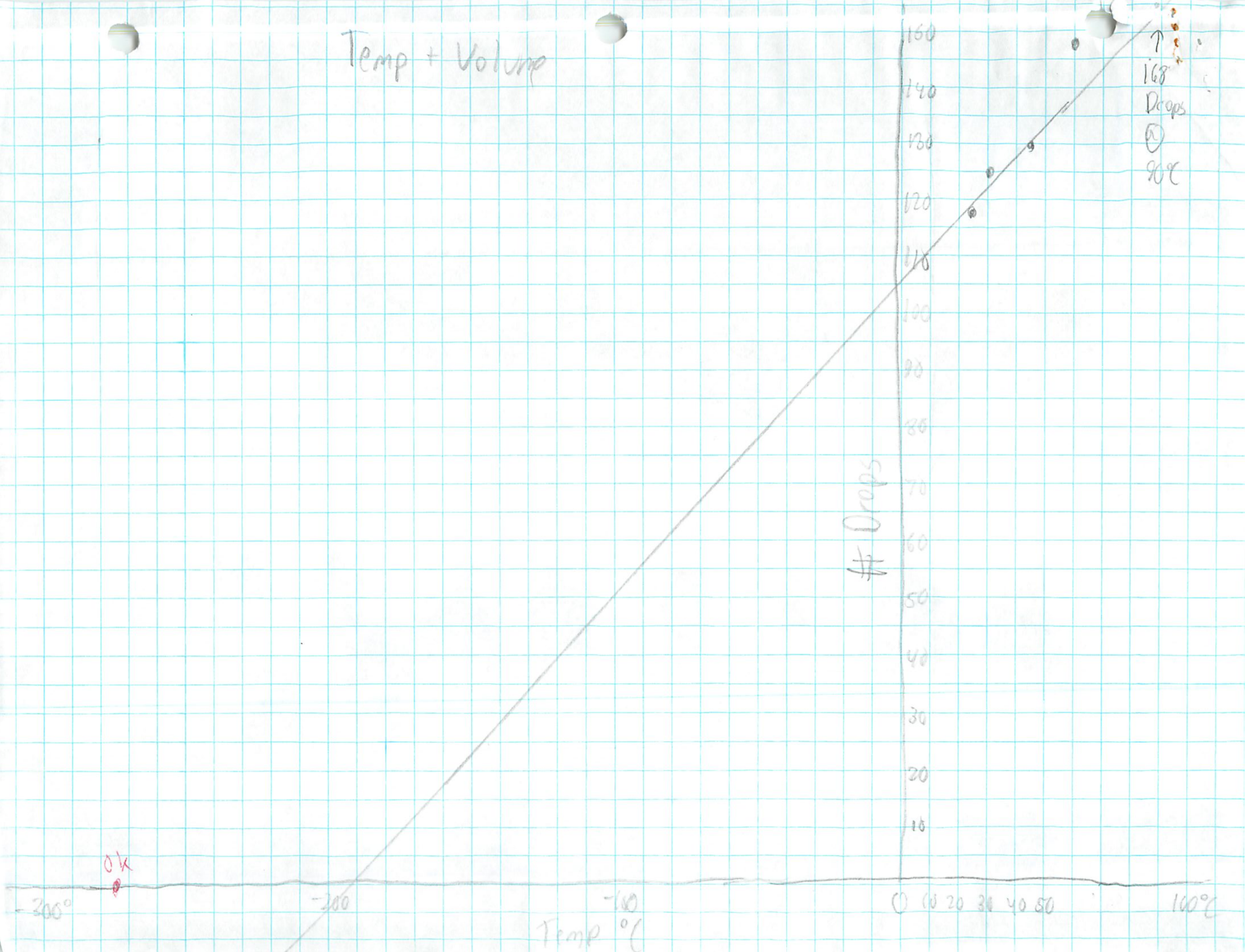
5. Relationship b/w temp + volume at constant pressure?
Charles's Law $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

6. Explain why a hot air ballon flies?
Both te temprature + volume decrease as te altitude rises. Te decrease in pressure makes te gas expand, te decreasing temprature causes it to decrease a little bit. But always te helium gas is less dense ten air.

7. Solve: Volume of 2.5 L gas at 20°C when temp ↓ to -20°C
$\frac{2.5}{20} x = \frac{\ }{-20}$ $-20 \cdot 2.5 = 20x$ $/20 = -2.5L$
$-50 = 20x$
^ dosn't work

#9
8. Why do auto tires appear flat in cold weather?
Te gas inside shrinks in te cold weater making tem appear flat

Temp + Volume



Drops
160
140
130
120
110
100
90
80
70
60
50
40
30
20
10
168
Drops
@
90°C
0K
-300°
-200
-100
0 10 20 30 40 50
100°C
Temp °C

Gas Laws
Notes

3/2

Boyle's Laws (P, V at constant T)

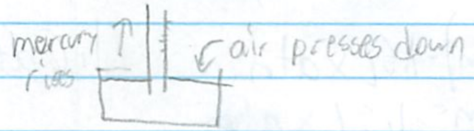
$\boxed{P_1V_1 = P_2V_2}$, if presure ↑, volume ↓
if volume ↑, presure ↓

inverse relationship

we did books on bulb

Barometers

- measure pressure



760 mm Hg = 1 atm = 101,3 kPa

low barometer = bad weather coming
high pressure = good weather coming

Charles's Law (V, T at constant P)

$\boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$ if temprature (in kelvin) ↑, volume ↑
if temp. ↓, volume ↓

proportional relationship

Absolute Zero

- 0K
- -273,15°C
- temp where all matter is a solid
  - kinetic energy stops

Gay-Lussac's Law (P, T at constant V)

$\boxed{\frac{P_1}{T_1} = \frac{P_2}{T_2}}$ ↑pressure = ↑temp
↓pressure = ↓temp

propotional relationship

- pressure cookers

Gas Law
Practice

1. Convert 520 atm to kPa

   1. Start with given

$$\frac{520\,atm}{1} \cdot \frac{101.3\,kPa}{1\,atm} = 52676\ kPa \quad \} \text{ 3sig fig}$$

$$\boxed{52700\ kPa}$$

2. What pressure is nessassry to condense 2L of a gas to 1.5L (constant T) with gas at 1.2atm

Boyle: $2L(1.2atm) = 1.5L\,x\,atm$ $\boxed{1.6atm}$

$$\frac{2.4\,L/atm}{1.5L} = \frac{1.5L\,x\,atm}{1.5L}$$

3. What is the volume of 2.0L of gas at 20.°C when temp is ↓ to −20.°C

Convert to k 293 k 253 k

↗ or else -volumes

$$\frac{2L}{293k} = \frac{x}{253k}$$

$253k(2L) = 293k\,x$

$$\frac{506\,kL}{293} = \frac{293k}{293}$$

$L = 1.7269$

2sig fig ↳ $\boxed{1.7L}$

4. What is pressure of 2.0L of gas when temp rased from 283k to 500k. The inital P is 1.5atm.

$$\frac{1.5\,atm}{283\,k} = \frac{x\,atm}{500k}$$

$1.5\,atm \cdot 500\,k = 283k\,x\,atm$

$$\frac{750\,k\,atm}{283k} = \frac{283k\,x\,atm}{283k}$$

2.65 atm → 2sig fig ↲

$\boxed{2.7\,atm}$

T	P	V	Law
C	inc dec	dec inc	$P_1V_1 = P_2V_2$ Bayle's
inc dec	C	inc dec	$\frac{V_1}{T_1} = \frac{V_1}{T_2}$ Charle's
inc dec	inc dec	C	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$ Gay-Luccas

Combined

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Ideal Gas Laws

3/6

When a gas obays Boyle's, Charle's and Gay-Lussac's laws all the time (Δ temp + pressure) it is considered _ideal_

Perfect

Ideal gasses _do not exist_ because it assumes particles have no volume and there is no attraction between them. When gasses are very cold or very compressed (↓Temp ↑pressure) the volume of particles comes into play and breaks the law.

$$\boxed{PV = nRT}$$

P = Pressure (atm, mm Hg, kPa)
V = Volume (L)
n = num moles
R = ideal gas constant (varries w/ units)
T = Temp (kelvin _not_ °C)

↑ use when
- no inital or final conditions
- need # moles

○ RESCORE ○ MARK ✓ ○ TOTAL ONLY

USE NO. 2 PENCIL ONLY

Use a No. 2 Pencil

Ⓐ ● Ⓒ Ⓓ Ⓔ
Fill circle completely

Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ
Erase cleanly

T F
1 ○A ●B ○C ○D ○E
2 ○A ●B ○C ○D ○E
3 ●A ○B ○C ○D ○E
4 ●A ○B ○C ○D ○E
5 ○A ●B ○C ○D ○E
6 ○A ○B ●C ○D ○E
7 ○A ●B ○C ○D ○E
8 ○A ○B ●C ○D ○E
9 ○A ●B ○C ○D ○E
10 ○A ●B ○C ○D ○E
11 ○A ○B ●C ○D ○E
12 ○A ○B ●C ○D ○E
13 ○A ○B ●C ○D ○E
14 ○A ●B ○C ○D ○E
15 ●A ○B ○C ○D ○E
16 ●A ○B ○C ○D ○E
17 ●A ○B ○C ○D ○E
18 ○A ●B ○C ○D ○E
19 ●A ○B ○C ○D ○E
20 ○A ○B ○C ●D ○E
21 ○A ○B ○C ●D ○E
22 ●A ●B ○C ○D ○E duh
23 ○A ●B ○C ○D ○E
24 ○A ●B ○C ○D ○E
25 ○A ○B ○C ●D ○E
26 ○A ○B ○C ○D ○E
27 ○A ○B ○C ○D ○E
28 ○A ○B ○C ○D ○E
29 ○A ○B ○C ○D ○E
30 ○A ○B ○C ○D ○E
31 ○A ○B ○C ○D ○E
32 ○A ○B ○C ○D ○E
33 ○A ○B ○C ○D ○E
34 ○A ○B ○C ○D ○E
35 ○A ○B ○C ○D ○E
36 ○A ○B ○C ○D ○E
37 ○A ○B ○C ○D ○E
38 ○A ○B ○C ○D ○E
39 ○A ○B ○C ○D ○E
40 ○A ○B ○C ○D ○E
41 ○A ○B ○C ○D ○E
42 ○A ○B ○C ○D ○E
43 ○A ○B ○C ○D ○E
44 ○A ○B ○C ○D ○E
45 ○A ○B ○C ○D ○E
46 ○A ○B ○C ○D ○E
47 ○A ○B ○C ○D ○E
48 ○A ○B ○C ○D ○E
49 ○A ○B ○C ○D ○E
50 ○A ○B ○C ○D ○E

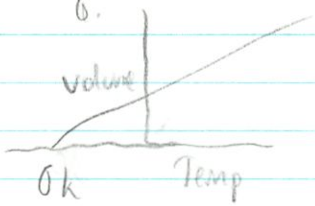
$P \times V$ Boyles
$\frac{V}{T}$ - Charles
$\frac{P}{T}$ - Gay-Lussiac
$\frac{PV}{T}$ - Combined
$\frac{PV}{Tn}$ = Ideal

4. $3.2\,atm = \frac{1\,atm}{101.3\,kPa} = 324.16 \rightarrow 320$

3.

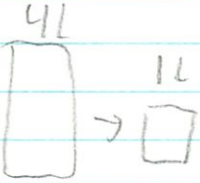
← Why isn't change constant?

6.



9. $101.3 - 101.2 = 0.1\,kPa$
But sig fig is done wrong for B

13.



14. °C

16. .3L 363k 102kPa
→ 173k
~0.476

$\frac{.3L}{363k} \rightarrow \frac{?L}{173k}$ ← .143L
5 143mL
×.476

18. Isn't it attraction and size? Size is part of looks and polarity is its personality figurativly.

19. .3L 423k
→ .6L
2x

$\frac{.3L}{423k} = \frac{.6L}{?k}$ ← 846k
2x

20. Say "are" not "is" when talking about plural things

In Gay-Lussac's law I see how temp changes pressure but not other way around

22. Doesn't seem to make sense... but law says it will

wasn't thinking

23. I thought evaporating was only if it wasn't boiling, so aren't no answers correct - and only if exernal pressure →

SCORE	24 96	# CORRECT % CORRECT
RESCORE		# CORRECT % CORRECT
ROSTER NUMBER	1	SCORE RESCORE

NAME Michael Plasmeier
COURSE Gasses Chemistry
HOUR 06-075283 DATE 3/8/07

TEST RECORD	
PART 1	
PART 2	
TOTAL	

TEST ANSWER SHEET J
Form No. 72512

ACCU-SCAN #26180-TR 10/00 70976 APPERSON PRINT MANAGEMENT SERVICES (800) 473-6761 05/05

are less then the maximum kietic energy pressure to allow particles to escape.

DO NOT WRITE IN THIS AREA

Michael Plasmeier

Organic Substances
Lab

$\frac{20}{20}$ 3/8

to combust materials present in fruit rinds
learn about hydrocarbons

What do you think - Why pay more for organic produce and fruit?

People buy them because they don't like chemicals in their food. Less as important is the small-farm business support.

Organic - molecular compound based on carbons
in 19th century
ment came from living (not mineral) sources

Organic means something different in chemistry. I think it means that it was not ~~man made~~ It has nothing to do with using pesticides/steriods.

Procedue

1. Obtain a orange peel. Note observations
2. Fold orange peel in half. With tongs hold and pinch peel over an open flame Obserbe.
3. Try with other fruit peels

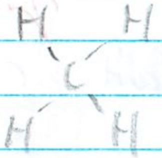
Orange peel - juice ~~drops~~ vapor catch on fire for ~½ sec like a spark

other fruits behave the same way - like a mini flame thrower

First 10
Straight-Chain Alkanes

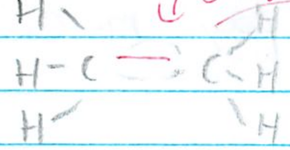
4a.

hydro carbon = made of hydrogen + carbon

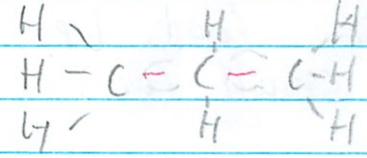


methane CH_4

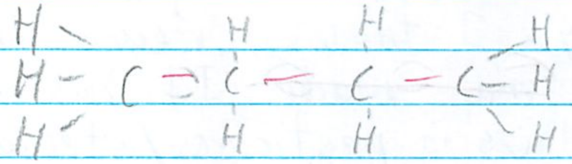
single chain only in alkanes



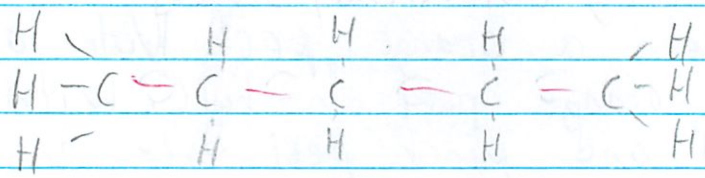
Ethane C_2H_6



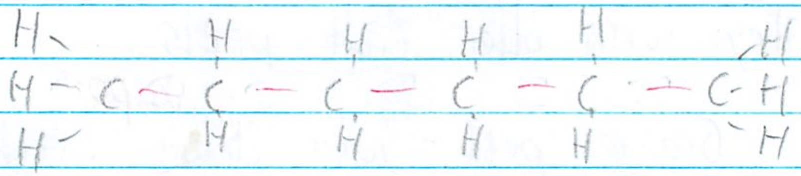
Propane C_3H_8



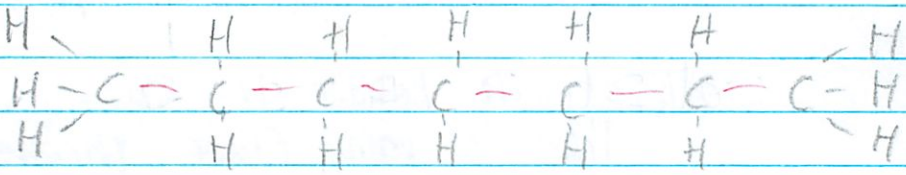
Butane C_4H_{10}



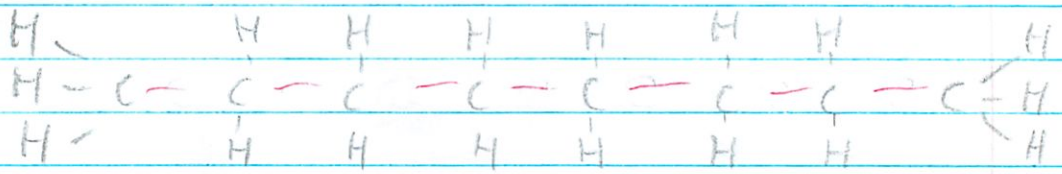
pentane C_5H_{12}



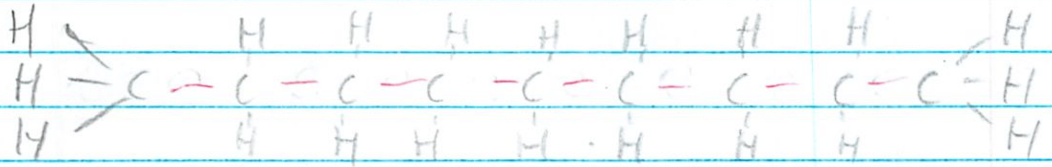
Hexane C_6H_{14}



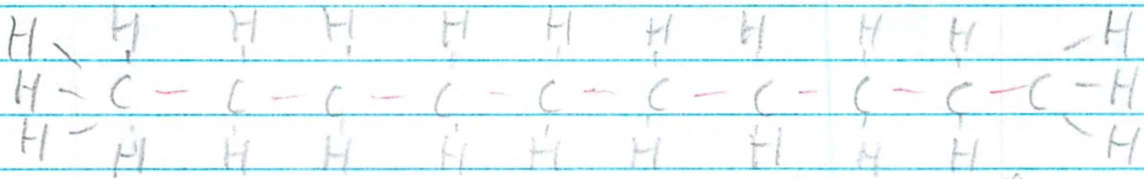
Heptane C_7H_{16}



Octane C_8H_{18}



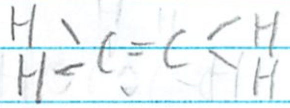
Nonane C_9H_{20}



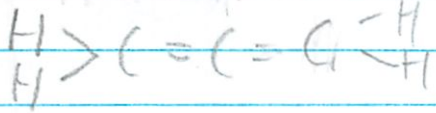
Decane $C_{10}H_{22}$

4b.

Alkenes (-ene)



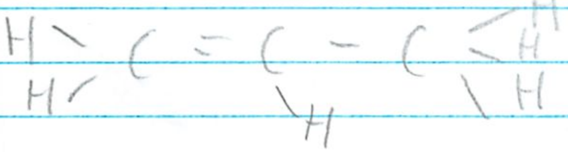
Ethyleen (ethene)



?

?

4 bonds per C



?

MSE Activity 9 Reading Guide

1.What was the evidence of a gas?

? The fact that the spark looking thing appeared to come from the fruit via the fire

2. Which gas was produced?

Ethylene

3.What is the formula for ethylene gas?

C_2H_4

4.What is an “organic” compound?

made partially w/ carbon molecules

5.What are inorganic compounds? Give an example of an inorganic compound.

everything not organic

H_7H_5 ← just made that up

6. Organic Chemistry is the study of........

Carbon ~~molecules +~~ compounds ~~it forms~~

7.Give some examples of organic compounds.

gasoline, artifical sweetners

propane

8.What is a hydrocarbon?

organic Compounds made only of carbon + hydrogen

9. Most hydrocarbons are useful as fuels. They undergo combustive reactions.

Do Chem to Go p.65 1-7

1. Examples of how organic means natural

Organic apples made with out pesticides

Organic animals

trees could be considered organic

2. Organic or Inorganic?

Salt - NaCl - Inorganic

Gasoline - C_8H_{18} - Organic

Sugar - $C_{12}H_{22}O_{11}$ - Organic

Sand - SiO_2 - Inorganic

Vinegar - CH_3COOH - Organic

Oxygen - O_2 - Inorganic

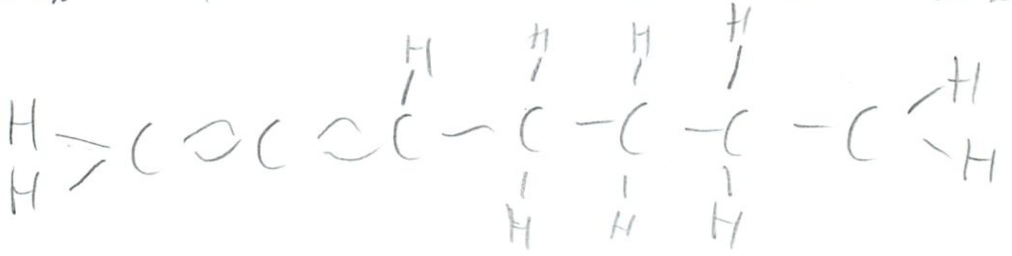
3. Hydrocarbon w/ 5 carbons + all single bonds

Pentane

C_5H_{12}

→

4. Hydrocarbon w/ 7 carbons + ≥ 2 carbon-carbon double bonds

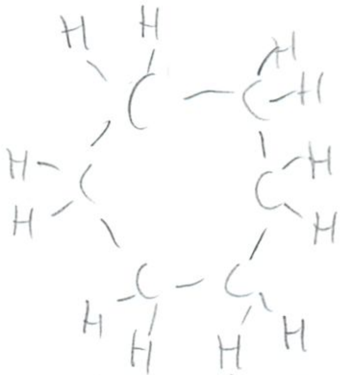


Yes there is more then one answer because you are allowed to have more then 2 carbon-carbon double bonds. ? Would it matter if the double bonded carbons were some were else?

5. Draw Cetylene (C_2H_2)

$H-C\equiv C-H$

6. Draw Hydrocarbon w/ 6 carbons + forms a ring



7. Some things are chemical free. Is this good? Some people think that chemicals in food such as steriods or preserivites are bad for them. Some of this has been proven and disproven. I see how it could matter but aren't we still living? Chemistry can help prove/disprove but its hard to tell

Michael Plasmeyer

Activity 6 - PIVNERT

20/20

What Do You Think?

You have seen how the behavior of gases can be influenced by changes in temperature and pressure. What about the amount of gas? In designing a toy that involves the use of a gas, it will be important to have an understanding of how all of these factors affect one another.

A toy dart gun uses a CO_2 cartridge to shoot darts. How does the toy company know how many times a cartridge can be used to shoot the darts before it runs out?

Record you ideas to this question in your *Active Chemistry* log.

Just try it. That's the best way - accounting for leaks and irregularities,

Goals

1. To determine the volume of one mole of a gas
2. To calculate the gas law constant "R"
3. To derive the Ideal Gas Law equation

20 mL acedic acid (cider vinegar)

Investigate

In this activity you will be reacting a piece of magnesium metal with ~~hydrochloric~~ acid to produce hydrogen gas. The quantity of hydrogen gas collected will be measured and the conditions of temperature and pressure will be recorded. You will then use the volume of hydrogen collected and the number of moles of magnesium consumed to calculate the volume of one mole of hydrogen gas under the present conditions of pressure and temperature. Further calculations will allow you to determine the volume of one mole of hydrogen under **Standard Temperature and Pressure (STP)** conditions (STP conditions imply 273 K and 1.0 atm.). Then you will be able to derive an equation that includes all of the factors that influence the behavior of gases – the **Ideal Gas Law**.

0°C

Make a table similar to the one below in your *Active Chemistry* log. Record your data in the appropriate places.

Mass of 100 cm Mg ribbon	Length of my Mg ribbon, cm	Mass of my Mg ribbon, g	Moles of my Mg Ribbon	Room Pressure, mmHg	Room Temperature, °C and K	Volume of Hydrogen Produced, mL and L
		.04	.0016	760	22°C	42.7 mL .0427 L

©2003 It's About Time

Your teacher will provide you with a gas collecting tube fitted with a one-hole stopper. The stopper has a piece of copper wire threaded through it. This is to hold the piece of magnesium in place while the reaction occurs in the gas tube.

Record the mass of 100. cm of magnesium ribbon. Then obtain a piece of magnesium ribbon and carefully measure its length to 0.01 cm. Using the mass of 100. cm, convert the length to mass in grams (to three significant figures).

_______ cm (X g / 100. cm) = __________ g

Record the room temperature and pressure. Convert the temperature to Kelvin (add 273 to Celsius).

You are now ready to begin the reaction.

1. Fill a 1000 mL beaker to about 2 cm from the top with water.

2. Carefully pour about 10 mL of the hydrochloric acid into the gas collecting tube. Then fill the rest of the tube with water. (The hydrochloric acid is denser than water and will remain at the bottom of the tube.)

done For us

3. Wrap the copper wire around the piece of magnesium and place the stopper in the gas tube. Place your finger over the hole in the stopper and invert the gas tube into the 1000 mL beaker of water. Remove your finger and watch as the hydrochloric acid falls toward the magnesium. As the reaction concludes, tap the gas tube in the beaker to dislodge any bubbles.

4. Record the volume of gas collected.

5. Rinse out the gas tube and dispose of the water in the beaker as directed by your teacher.

✓ # of moles

Calculations

$Mg_{(s)} + 2HC_2H_3O_2 \rightarrow Mg(C_2H_3O_2)_2 + H_{2(G)}$

acetate

The equation for the reaction is:

acetic acid

magnesium + ~~hydrogen chloride~~ → hydrogen + magnesium ~~chloride~~

~~$Mg_{(s)} + 2HCl_{(aq)} \rightarrow H_{2(g)} + MgCl_{2(aq)}$~~

The balanced equation shows that one mole of magnesium reacts with two moles of hydrogen chloride to produce one mole of hydrogen gas and one mole of magnesium chloride.

Grams in 1 mol of Mg = 24.3 g/mol

©2003 It's About Time For Evaluation Purposes Only– Do Not Duplicate

Since the ratios in the balanced equation are mole ratios, you need to convert the mass of magnesium to moles of magnesium in order to continue. One mole of magnesium has a mass of 24.3 g.

0's not signifiant here

Using this information, convert the mass of your piece of magnesium to moles.

source • target / source = target ? Conversion rate

.04 g (1 mole/ 24.3 g) = .00164 mole Mg or 1.6×10^{-3} .0016 <2 sig fig

Record this in your *Active Chemistry* log.

The balanced equation shows that one mole of magnesium will produce one mole of hydrogen gas, therefore, the number of moles of magnesium you used is also the number of moles of hydrogen gas that forms.

.0016 mole Mg (1mol H_2/1 mol Mg) = .0016 mole H_2 # of moles

?; Hg = Mg in terms of ~~molar~~ mass?

From the volume of gas collected and the number of moles of hydrogen gas, calculate the volume of one mole of hydrogen gas.

(.0427 liters of H_2) / (.0016 mole H_2) = 26.6875 → 27L Liters/mole

This is the number of liters in a mole of hydrogen gas at the conditions of temperature present in the room now. Record this as the Molar Volume of hydrogen in your *Active Chemistry* log.

What would the volume be at Standard Temperature (0 °C or 273 K) and Standard Pressure (1 atm or 760 mmHg) conditions? You can use a combination Boyle's and Charles' laws to determine this.

You learned in earlier investigations that pressure and volume of gases are inversely related and volume and temperature are directly related (as long as the number of particles remains constant). By combining the two equations you can get:

$$\frac{(\text{Pressure})\ (\text{Volume})}{(\text{Temperature})} = \text{constant}$$

For the same gas, if temperature and pressure conditions are changed, you can calculate the expected change in volume using this **Combined Gas Law**.

Convert to STP

$$\frac{P_1\ V_1}{T_1} = \frac{P_2\ V_2}{T_2}$$

STP

$$\frac{(1atm)(27L)}{295\ k} = \frac{1atm \cdot xL}{273\ k}$$

Units $\frac{L}{k} \cdot \frac{atm/k}{1}$

$\frac{x}{273} \cdot \frac{.0036}{1}$

.0915 atm•L / k = .0036 atm/k xL / 273 k

what is this = what to do w/ it

.0036 atmk

.0036 atm/k

25.15669 L = .0000139 Latms

For Evaluation Purposes Only– Do Not Duplicate ©2003 It's About Time

1 mol = 6.02×10^{23} representative particles in a substance

SI measurement for amt of a substance

Using this equation we can calculate the molar volume of hydrogen under the conditions of STP:

$$V_2 = \frac{P_1V_1T_2}{P_2T_1}$$

Where P_1 is room pressure, T_1 is room temperature, V_1 is your calculated volume of hydrogen, P_2 is Standard Pressure (1.0 atm) and T_2 is Standard Temperature (0 °C).

Remember to change Celsius temperatures to Kelvin. Record the molar volume at STP in your *Active Chemistry* log.

atom
Molecule
formula
Units

$$\frac{|exp - accp|}{accp}$$

The accepted value for the molar volume of *any* gas at STP is 22.4 liters/mole. How close was your value? Find your percent error by dividing the difference between your value and 22.4 and then dividing by 22.4. Multiply by 100 to get a percent.

-at 22k ↑ then STP

The Ideal Gas Law

12% error

You now have all the information you need to compare changes in pressure, temperature and volume for gases. The combined gas law allows you to look at changes in these three variables that affect gases. The fourth variable is the number of particles – the number of moles of gas involved. How does it fit in to the equation?

You know the combined gas law: $\frac{(\text{Pressure})\ (\text{Volume})}{(\text{Temperature})} = \text{constant}$

or more simply: $\frac{P\,V}{T} = k$

Using the conditions of STP, you can add the number of particles to the equation. The unit for molar volume is liters per mole, so the number of moles will go into the denominator in the equation. Using "n" for the number of moles you get:

$$\frac{P\,V}{T\,n} = k$$

Substituting the conditions for STP into the equation you get:

at $\left(\frac{(1\ \text{atm})\ (22.4\ \text{L})}{(273\ \text{K})\ (1\ \text{mole})}\right) = R$ in a

By convention, "R" is called the **Ideal Gas Law constant**. What is the value and unit of "R"? Record this in your *Active Chemistry* log.

What would be the value and unit for "R" if the pressure were 760 mmHg? Calculate and record this in your *Active Chemistry* log.

$$\frac{(760\ \text{mmHg})\ (22.4\ \text{L})}{(273\ \text{K})\ (1\ \text{mole})} = 62.L\ \text{mmHg/mol} \cdot \text{K}$$

©2003 It's About Time For Evaluation Purposes Only– Do Not Duplicate

Chem Talk

An **ideal gas** is defined as one in which all collisions between atoms or molecules are perfectly elastic and in which there are no intermolecular attractive forces. One can visualize it as a collection of perfectly hard spheres which collide but which otherwise do not interact with each other. Many common gases exhibit behavior very close to that of an ideal gas at low temperature and pressure.

↑↑ ↓temp + ↑pressure make it not ideal

The ideal gas law relates the pressure, temperature, and volume of an ideal gas. The ideal gas law was originally derived from the experimentally measured Charles' Law, Boyle's Law and Avogadro's Law. Let *P* be the pressure of a gas, *V* the volume it occupies, and *T* its temperature (which must be in absolute temperature units, i.e., in Kelvin).

The ideal gas law states

$$PV = nRT$$

Where P=pressure in atm
n=the number of moles of gas
R=the universal gas constant (62.4 L ·· mmHg/mol · K, or 0.0821 L · atm/ mol · K)
T=temperature in Kelvins

The ideal gas law can be used to make predictions about the pressure, temperature, volume and number of particles of a gas. For example, you might want to find the pressure of 2.5 moles of hydrogen gas at 25.0 °C if its volume is 8.55 L.

Using PV = nRT, rearrange the equation to solve for pressure:

P = nRT/V

Substituting values into the equation, we see:

P = (2.5 mol) (0.0821 L · atm/ mol · K) (298 K) / (8.55 L)

(remember to change the temperature to Kelvin)

Solving for P, we get P = 7.2 atm

Checking Up

1) What is the volume of 6.5 moles of oxygen gas if the temperature is 40.0 °C and the pressure is 6.2 atm?

V = nRT/P

V=(6,5 mol) (,0821 L • atm/mol • k) (313 k) / (6,2)

26,84 → 3sig fig (27 L)

For Evaluation Purposes Only– Do Not Duplicate ©2003 It's About Time

1.7 L air

2) A 2.0 L soda bottle is used as a water rocket. If 0.30 L of water is in the bottle and it is pumped with air to a pressure of 3.8 atm at a temperature of 25 °C, how many moles of air are in the rocket?

$n = \frac{PV}{RT} = (0.0821 \text{ L atm/mol K})(298 \text{ K}) / (3.8 \text{ atm})(1.7 \text{L})$

$n = 3.7877 \rightarrow \boxed{3.8 \text{ mol}}$

Reflecting on the Activity and the Chapter Challenge

You have now discovered a way to relate the pressure, volume, amount and temperature of a gas. If your toy was designed to employ a CO_2 gas cartridge to inflate a balloon, propel a car, or blast a rocket, you should include information telling the consumer how many times the toy can be used with one CO_2 cylinder.

Chem to Go

1. A sample of dry gas weighing 1.05 g is found to occupy 1.43 L at 23.5 °C and 0.951 atm. How many moles of the gas are present?

2. What is the mass of one mole of the gas in question #1?

3. Lets say that you are designing a toy that requires the generation of 1.0 L of oxygen gas to operate it. What reagents would you use, and how many moles would be required if the gas was being produced at 1.0 atm and 20 °C? (see activity 5 for help).

4. Many gases are stored in their compressed form (under pressure). Calculate the weight of N_2 that could be stored at 22 °C and 125 atm in a cylinder with a volume of 45.0 liters. The molecular mass of N_2 is 28.0 g/mole.

5. Calculate the weight in grams of the air in a hot air balloon that has a volume of 4.0×10^5 L when the temperature of the gas is 90.0 °C and the pressure is 750 mm Hg. Assume that the average molecular mass of air is 30.0 g/mole.

Preparing for the Chapter Challenge

Discuss with your group how you might employ a gas in your toy. Try to decide how much gas will need to be generated, how you will generate this gas, and the quantity of reactants will be required to produce the desired amount of gas. Keep in mind that you will probably be forming the gas under normal atmospheric conditions and room temperatures.

©2003 It's About Time For Evaluation Purposes Only– Do Not Duplicate

Pivnert

Chem to Go

3/7

PIVINERT Packet

1. $n = \frac{PV}{RT}$ $n = \frac{(.951\ atm)(1.43L)}{(.0821\ L\ atm/mol\ k)(296.5\ k)}$

$n = .05586$ mol = 1.05 g dry ice

2. $1g / .05586\ mol = 17.8990 \rightarrow 17.9\ g = 1\ mol$

3. $n = \frac{PV}{RT} = \frac{(1\ atm)(1L)}{(.0821\ L atm/mol\ k)(293\ k)} = .04157 \rightarrow \boxed{.0416\ mol}$

4 $n = \frac{PV}{RT} = \frac{(125\ atm)(45L)}{(.0821\ L atm/mol k)(295k)} = \frac{232}{28} \leftarrow \boxed{8.3g}$ mol in g

had to think ↗

5 $n = \frac{PV}{RT}$ $\frac{(.987\ atm)(400000\ L)}{(.0821\ L\ atm/mol\ k)(295\ k)} = \frac{16300\ mol}{28.0} = \boxed{582\ g}$

$750 \cdot \frac{1 atm}{760} =$ .987 atm

Don't really get why works

Chap Challange

Could release compressed gas to shoot a projectile, The gas would be precompressed and when it is punctured it is released. Use the ideal gas laws to find out how much gas (moles or g) can fit in a certain volume at 1 atm / 295 k)

Name ______________________ Date ________________ Class ______________

13 STATES OF MATTER

SECTION 13.1 THE NATURE OF GASES (pages 385–389)

This section introduces the kinetic theory and describes how it applies to gases. It defines gas pressure and explains how temperature is related to the kinetic energy of the particles of a substance.

equalibrium =
rate△ / direction =
rate△ other direction

▶ Kinetic Theory and a Model for Gases (pages 385–386)

1. The energy an object has because of its motion is called kinetic energy.

2. Circle the letter of each sentence that is true about the assumptions of the kinetic theory concerning gases.

[✓ **a.**] A gas is composed of particles with insignificant volume that are relatively far apart from each other.

b. Strong attractive forces exist between particles of a gas.

c. Gases tend to collect near the bottom of a container.

[✓ **d.**] The paths of uninterrupted travel of particles in a gas are relatively short because the particles are constantly colliding with each other or with other objects.

3. Is the following statement true or false? According to the kinetic theory, collisions between particles in a gas are perfectly elastic because kinetic energy is transferred without loss from one particle to another, and the total kinetic energy remains constant. true

▶ Gas Pressure (pages 386–387)

4. Gas pressure results from the force exerted by a gas per unit surface area of an object.

5. Simultaneous collisions of billions of particles in a gas with an object result in a mesurable force.

6. What force holds the particles of air in Earth's atmosphere? atmospheric presure

7. What kind of pressure is measured with a barometer?

athmospheric pressure in mm Hg

? barometric presure

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________________ Date ________________ Class ________________

8. Look at Figure 13.2 on page 386. What accounts for the difference in height of the two mercury columns shown in the figure?

A different atmospleric pressure which is exerted at different heights above sea level

9. Circle the letter next to every name of a unit of pressure.

[✓ **a.**] mm Hg
[✓ **b.**] standard
c. pascal
[✓ **d.**] kPa
[✓ **e.**] atm
f. degree

10. Standard temperature and pressure (STP) are defined as a temprature of 0°C and a pressure at 1 atm (101.3 kPa)

▶ Kinetic Energy and Temperature (pages 388–389)

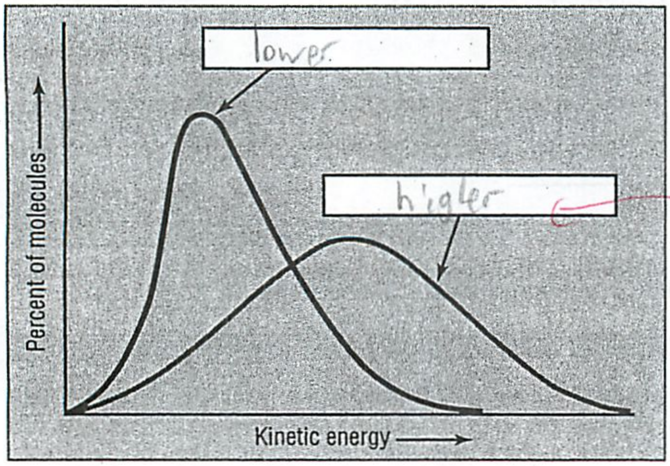
11. What happens to the temperature of a substance when the average kinetic energy of its particles increases?

it moves faster, exerts greater pressure

12. Is the following statement true or false. All the particles in a substance at a given temperature have the same kinetic energy. false

13. The temperature 0 K, or −273.15°C, is called absolute zero. Theoretically, particles of matter at this temperature would have no kinetic motion/energy.

14. On the graph below, write the labels *lower temperature* and *higher temperature* to identify the curve that depicts the kinetic energy distribution of particles in a liquid at a lower temperature and at a higher temperature.



not what you think
energy increase
avg ↑
spreads out
total same

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________________ Date ________________ Class ________________

15. Circle the letter of the temperature scale that correctly completes this sentence. Temperature on the _______ scale is directly proportional to the average kinetic energy of the particles of a substance.

a. Celsius

[✓ **b.**] Kelvin

c. Fahrenheit

d. Centigrade

SECTION 13.2 THE NATURE OF LIQUIDS (pages 390–395)

This section describes a model for liquids in terms of kinetic energy and the attractive forces between the particles in a liquid. It also uses kinetic theory to distinguish evaporation from boiling.

▶ A Model for Liquids (page 390)

1. Is the following sentence true or false? The kinetic theory states that there are no attractions between the particles of a liquid.

Halse — true for ~~[illegible]~~ gasses

sort-of

2. Circle the letter next to each sentence that is true about the particles of a liquid.

A few

✗ **a.** Most of the particles in a liquid have enough kinetic energy to escape into a gaseous state.

[✓ **b.**] Liquids are much denser than gases because intermolecular forces reduce the amount of space between the particles in a liquid.

[✓ **c.**] Increasing pressure on a liquid has hardly any effect on its volume.

? **d.** Liquid particles are free to slide past one another.

▶ Evaporation (page 391)

3. The conversion of a liquid to a gas or vapor is called vaporization.

4. When vaporization occurs at the surface of a liquid that is not boiling, the process is called Evaporization.

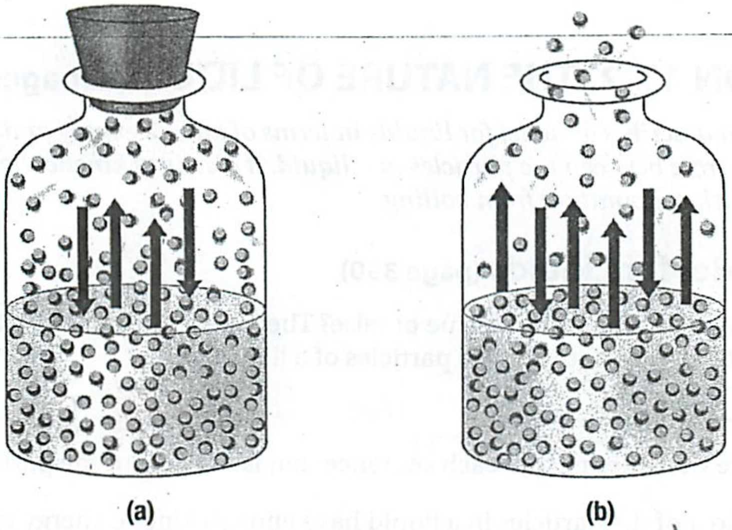
5. As a liquid evaporates, why do only some of the particles break away from the surface of the liquid? Why does the liquid evaporate faster if the temperature is increased?

Only the molecules with a certain min. kinetic energy can escape. At ↑ temp, the ↑ avg kinetic motion and allows more particles to overcome the attractive forces locking them together.

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

6. Is the following sentence true or false? Evaporation is a cooling process because the particles in a liquid with the highest kinetic energy tend to escape first, leaving the remaining particles with a lower average kinetic energy and, thus, a lower temperature. true

Questions 7, 8, 9, and 10 refer to either container A or container B below. Think of each container as a system involving both liquid water and water vapor.



(a) (b)

7. From which of the containers are water molecules able to escape? b

8. In which container can a dynamic equilibrium between water molecules in the liquid state and water molecules in the vapor state be established? a

9. In which container will the water level remain constant? a

10. From which container is it possible for all of the liquid water to disappear through evaporation? b

11. What causes the chill you may feel after stepping out of a swimming pool on a warm, windy day?

The water leaving your body takes some heat from your body.

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________________ Date ________________ Class ________________

▶ Vapor Pressure (pages 392–393)

12. Circle the letter next to each sentence that is true about vapor pressure.

[✓ **a.**] Vapor pressure arises when particles of a liquid in a closed, partly filled container vaporize and collide with the walls of the container.

[✓ **b.**] After a time in a closed, partly filled container, a liquid will evaporate and its vapor will condense at equal rates.

c. Look at Figure 13.6b on page 391. Condensation on the inside of the terrarium indicates that there is not a liquid-vapor equilibrium in the sealed terrarium.

[✓ **d.**] When the temperature of a contained liquid increases, its vapor pressure increases. ✓ more evaps then more condeses

13. Look at Figure 13.7 on page 393. How does the vapor pressure of the ethanol in the manometer change when the temperature is increased from 0°C to 20°C? Circle the letter of the correct answer.

a. The vapor pressure decreases by more than 4 kPa.

b. The vapor pressure remains constant.

[✓ **c.**] The vapor pressure increases by more than 4 kPa. $\uparrow T = \uparrow P$ (Gay Lussics)

d. There is no way to detect a change in vapor pressure with a manometer.

▶ Boiling Point (pages 393–395)

14. The boiling point of a liquid is the temperature at which the vapor pressure of the liquid is just equal to the external pressure.

15. Look at Figure 13.8 on page 394. Why does the boiling point decrease as altitude increases?

The ↑ altitude means ↓ pressure. This ↓ the external pressure which prevents the molecules from vaporizing, so more can escape

16. Use Figure 13.9 on page 394. At approximately what temperature would ethanol boil atop Mount Everest, where the atmospheric pressure is 34 kPa? Circle the letter next to the best estimate.

[✓ **a.** 50°C] **b.** 100°C **c.** 0°C **d.** 85°C

17. Is the following sentence true or false? After a liquid reaches its boiling point, its temperature continues to rise until all the liquid vaporizes. false

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________________ Date ________________ Class ________________

Reading Skill Practice

Writing a summary can help you remember what you have read. When you write a summary, include only the most important points. Write a summary of the discussion of boiling point on pages 393–395. Do your work on a separate sheet of paper.

SECTION 13.3 THE NATURE OF SOLIDS (pages 396–399)

This section describes the highly organized structures of solids, distinguishes between a crystal lattice and a unit cell, and explains how allotropes of an element differ.

▶ A Model for Solids (page 396)

1. Is the following sentence true or false? Although particles in solids have kinetic energy, the motion of particles in solids is restricted to vibrations about fixed points. true

2. A solid melts when the orginizations of particles break down as speed ↑

3. Is the following sentence true or false? The temperature at which the liquid and solid states of a substance are in equilibrium is the same as the melting point *and* the freezing point of the substance. False

~~?~~ ~~aren't the Δ~~

At that temp they both exist w/o Δ

▶ Crystal Structure and Unit Cells (pages 396–399)

4. How are particles arranged in a crystal?

Orderly, repeating, 3D pattern – "crystal lattice"

5. What type of solid has a relatively low melting point?

Molecular solids

6. Do all solids melt when heated? Explain.

No some turn to ash (decompose) like wood

cane sugar

why ?

energy req to decompose < energy to melt

7. Circle the letter next to each sentence that is true about solids.

a. Most solid substances are not crystalline. "amorphous" like glass, plastic

[✓ **b.**] All crystals have sides, or faces, that intersect at angles that are characteristic for a given substance.

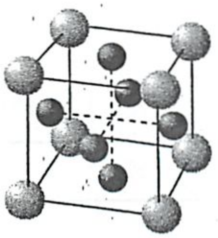
[✓ **c.**] There are seven groups, or crystal systems, into which all crystals may be classified.

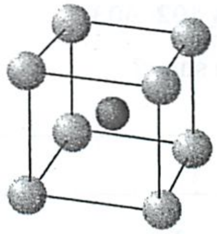
[✓ **d.**] The orderly array of sodium ions and chloride ions gives crystals of table salt their regular shape. fig 13.10

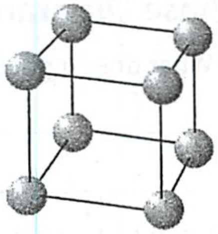
© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ____________ Date ____________ Class ____________

Identify the unit cell in each figure below as simple cubic, body-centered cubic, or face-centered cubic.







8. face centered **9.** body centered **10.** Simple

11. Is the following sentence true or false? Some solid substances can exist in more than one form. Give an example to support your answer.

True; allotropes (diff molecular form of same element in same physical shape) - carbon is - diamonds
- carbon
fullerenes - buckeyball

12. Two or more different molecular forms of the same element in the same physical state are called allotropes.

13. What is an amorphous solid?

lacks an ordered physical structure

14. Circle the letter next to each solid that is an amorphous solid.

a. table salt [✓ **c.** plastic]

[✓ **b.** rubber] [✓ **d.** glass]

15. How are glasses different from crystalline solids?

like supercooled liquid

no crystalized structure

soften when heated

when shattered - sharp jagged edges

SECTION 13.4 CHANGES OF STATE (pages 401–404)

This section describes the process of sublimation. It also explains phase changes between solid, liquid, and vapor states and how to interpret a phase diagram.

▶ Sublimation (page 401)

1. The process by which wet laundry dries on an outdoor clothesline in winter is called Sublimation.

2. Is the following sentence true or false? Solids have vapor pressure because some particles near the surface of a solid substance have enough kinetic energy to escape directly into the vapor phase. true

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Sublimation; vapor pressure > atm pressure

Name ______________________ Date ________________ Class ________________

CHAPTER 13, States of Matter *(continued)*

▶ Phase Diagrams (pages 402–403)

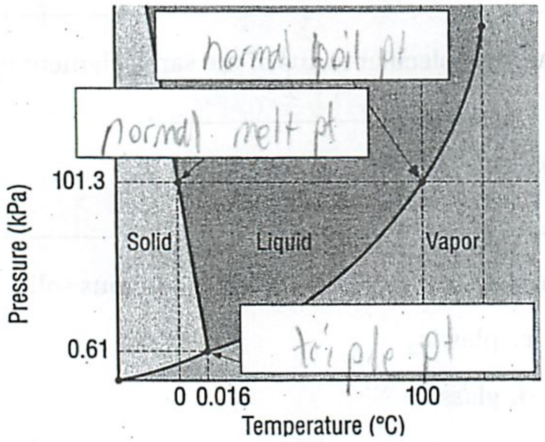
3. What does a phase diagram show?

the conditions of pressure + temprature at which 2 phases exist in equalibrium

4. What is the triple point of a substance?

only pt. where all 3 phases can exist in equalbrium

5. In the phase diagram for water shown below, label the melting point and boiling point at normal atmospheric pressure, and the triple point.



6. Use the phase diagram above to answer the following question. Why is a laboratory required to produce the conditions necessary for observing water at the triple point?

A pressure of .61 kPa and a temp or .016 °C is needed. Those exact measurements are hard to find in the real world

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________ Date ______________ Class ______________

GUIDED PRACTICE PROBLEM

GUIDED PRACTICE PROBLEM 2 (page 387)

2. The pressure at the top of Mount Everest is 33.7 kPa. Is that pressure greater than or less than 0.25 atm?

Analyze

Step 1. To convert kPa to atm, what conversion factor do you need to use?

1 atm = 101.3 kPa

Step 2. Why can you use an estimate to solve this problem?

1 atm = 100 kPa

.01 atm = 1 kPa or .1 = 10 kPa

.25 atm = 25 kPa

Calculate

just multiply

Step 3. Write the expression needed to find the answer.

$\frac{33.7\text{ kPa}}{1} \times \left(\frac{1\text{ atm}}{101.3\text{ kPa}}\right)$ $33.7 = 101.3 / x$ calc: .33

x / x $x\ x$

Step 4. Which common fraction is this number close to?

$\frac{1}{3}$ $\frac{33.7x}{33.7} = \frac{101.3}{33.7} =$

Step 5. What is this fraction written as a decimal? Is this number greater than or less than 0.25?

.33 > .25

Evaluate

Step 6. Are you confident your estimate gave a correct answer to this problem?

Yes

EXTRA PRACTICE (similar to Practice Problem 1, page 387)

1. What pressure, in atmospheres, does a gas exert at 152 mm Hg?

$20.3\text{ kPa} \cdot \frac{1\text{ atm}}{101.3\text{ kPa}} = .200\text{ atm}$

What is this pressure in kilopascals?

$152\text{ mm Hg} \cdot \frac{101.3\text{ kPa}}{760\text{ mm Hg}} = 20.3\text{ kPa}$

Conversion rate

origin $\cdot \frac{\text{target}}{\text{origin}} =$ target output

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________ Date ______________ Class ______________

14 THE BEHAVIOR OF GASES

SECTION 14.1 PROPERTIES OF GASES (pages 413–417)

This section uses kinetic theory to explain the properties of gases. This section also explains how gas pressure is affected by the amount of gas, its volume, and its temperature.

▶ Compressibility (pages 413–414)

1. Look at Figure 14.1 on page 413. How does an automobile air bag protect the crash dummy from being broken as a result of impact?

Gas is easily compressed while solids + liquids are not. This means that the gas compressing takes energy from the collison instead of you

2. What theory explains the behavior of gases? kinetic

3. Circle the letter next to each sentence that is true concerning the compressibility of gases.

* [✓ **a.**] The large relative distances between particles in a gas means that there is considerable empty space between the particles.

[✓ **b.**] The assumption that particles in a gas are relatively far apart explains gas compressibility.

c. Compressibility is a measure of how much the volume of matter decreases under pressure.

d. Energy is released by a gas when it is compressed.

▶ Factors Affecting Gas Pressure (pages 414–417)

4. List the name, the symbol, and a common unit for the four variables that are generally used to describe the characteristics of a gas.

a. Pressure (P) - kPa

b. volume (V) - L or cm³ solid only

c. temprature (T) T

d. num moles (n) - n

Mole - weight of atomic weight in grams -or- # of particles

5. What keeps the raft in Figure 14.3 inflated?

The increased pressure of the compressed gas inside of it. > pressure outside

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________________ Date ________________ Class ________________

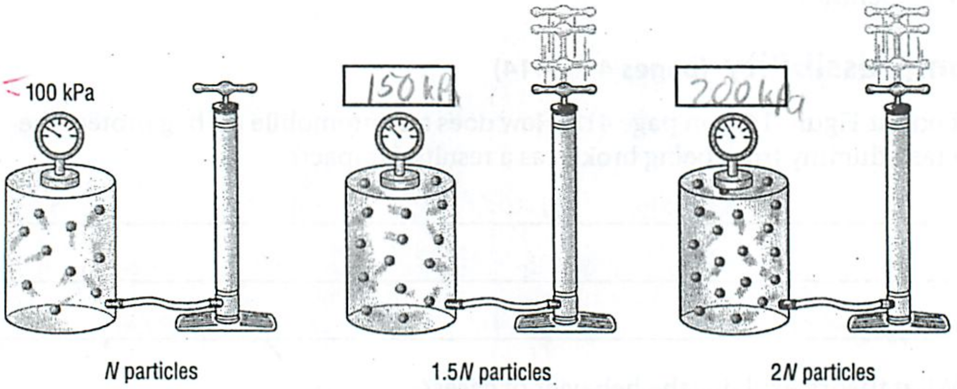
CHAPTER 14, The Behavior of Gases *(continued)*

6. How do conditions change inside a rigid container when you use a pump to add gas to the container?

The pressure againts the walls from the outside is greater - more particles inside

7. The diagrams below show a sealed container at three pressures. Complete the labels showing the gas pressure in each container.

2x particles = 2x pressure (at constant volume)

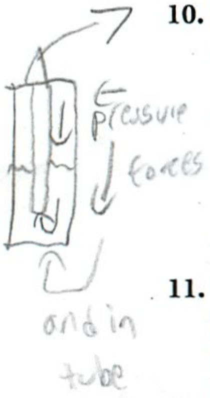


8. What can happen if too much gas is pumped into a sealed, rigid container?

The pressure inside is too great and the container bursts

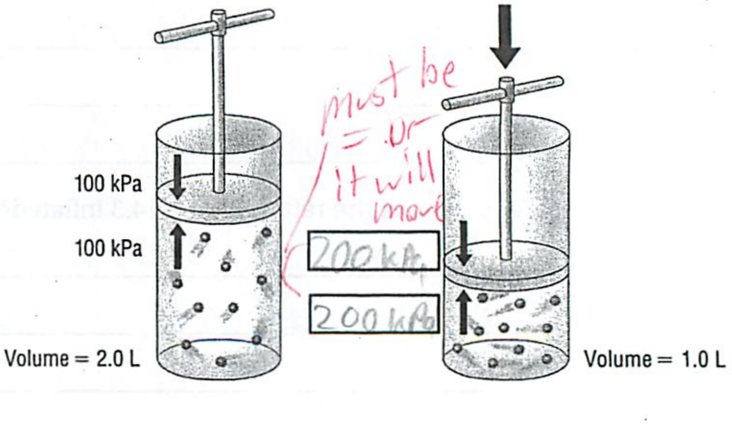
9. Is the following sentence true or false? When a sealed container of gas is opened, gas will flow from the region of lower pressure to the region of higher pressure. False

10. Look at Figure 14.5 on page 416. What happens when the push button on an aerosol spray can is pressed?



The pressure is greater inside the can, so gas forces the liquid down, up, and out of the can into the lower pressure region

11. In the diagram, complete the labels showing the pressure on the piston and the gas pressure inside the container.



must be = or it will move

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ____________ Date ____________ Class ____________

12. When the volume of a gas is reduced by one half, what happens to its pressure?
It doubles if the temp is also constant

13. Is the following sentence true or false? Raising the temperature of a contained gas causes its pressure to decrease. False

14. Circle the letter next to each sentence that correctly describes how gases behave when the temperature increases.

[✓ **a.**] The average kinetic energy of the particles in the gas increases as the particles absorb energy. but not total

[✓ **b.**] Faster-moving particles impact the walls of their container with more force, exerting greater pressure.

c. When the average kinetic energy of the enclosed particles doubles, temperature doubles and the pressure is cut in half.

15. Explain why it is dangerous to throw aerosol cans into a fire.
There is always high pressure in there ; heat ↑ pressure and it might be too strong for the container

16. Decide whether the following sentence is true or false, and explain your reasoning. When the temperature of a sample of steam increases from 100°C to 200°C, the average kinetic energy of its particles doubles.
False - the doubling law only works in kelvin its like 417 K → 517 K - that dosn't double but 100 K → 200 K does 2x the pressure

SECTION 14.2 THE GAS LAWS (pages 418–425)

This section explains the relationships among the volume, pressure, and temperature of gases as described by Boyle's law, Charles's law, Gay-Lussac's law, and the combined gas law.

▶ Boyle's Law: Pressure and Volume (pages 418–419)

1. Circle the letter of each sentence that is true about the relationship between the volume and the pressure of a contained gas at constant temperature.

[✓ **a.**] When the pressure increases, the volume decreases.

[✓ **b.**] When the pressure decreases, the volume increases.

c. When the pressure increases, the volume increases.

d. When the pressure decreases, the volume decreases.

inverse relationship

2. Boyle's law states that for a given mass of gas at constant temperature, the volume of the gas varies inversely with pressure.

P α V

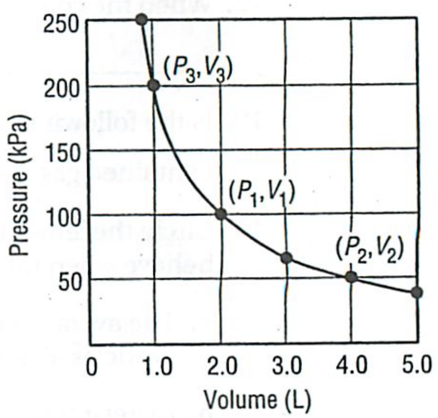
P↑ = V↓

P↓ = V↑

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________ Date ______________ Class ______________

Questions 3, 4, 5, and 6 refer to the graph. This graph represents the relationship between pressure and volume for a sample of gas in a container at a constant temperature.



3. $P_1 \times V_1 =$ 200 kPa·L

4. $P_2 \times V_2 =$ 200 kPa·L

5. $P_3 \times V_3 =$ 200 kPa·L

$\text{kPa} = \frac{200}{V}$

6. What do you notice about the product of pressure times volume at constant temperature?

its the same

▶ Charles's Law: Temperature and Volume (pages 420–421)

7. Look at the graph in Figure 14.10 on page 420. What two observations did Jacques Charles make about the behavior of gases from similar data?

$\frac{V}{T}$

Volume vs temp graph produces a line

or

extrapoliting the graph takes him to -317°C which is 0 k

just

8. What does it mean to say that two variables are directly proportional?

when one ↑, so does the other

" " ↓ " " " "

9. Is the following sentence true or false? Charles's law states that when the pressure of a fixed mass of gas is held constant, the volume of the gas is directly proportional to its Kelvin temperature. True

10. Charles's law may be written $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ at constant pressure if the temperatures are measured on what scale? kelvin

▶ Gay-Lussac's Law: Pressure and Temperature (pages 422–423)

$\frac{P}{T}$

11. Complete the following sentence. Gay-Lussac's law states that the pressure of a gas is proportional to the temprature (constant volume)

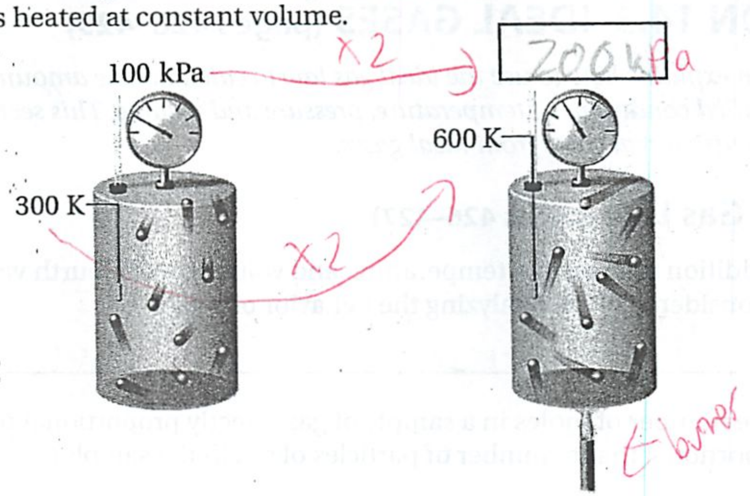
kelvin

12. Gay-Lussac's law may be written $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ if the volume is constant and if the temperatures are measured on what scale? kelvin

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________________ Date ______________ Class ______________

13. Complete the missing label in the diagram below showing the pressure change when a gas is heated at constant volume.



▶ The Combined Gas Law (pages 424–425)

14. Is the following sentence true or false? The gas laws of Boyle, Charles, and Gay-Lussac can be combined into a single mathematical expression. true

$\frac{PV}{T}$

Questions 15, 16, 17, and 18 refer to the following equation

$$\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$$

15. What law does this mathematical equation represent? Combined gas law

16. Which gas law does this equation represent if temperature is held constant so that $T_1 = T_2$? Boyle

17. Which gas law does this equation represent if pressure is held constant so that $P_1 = P_2$? Charles

18. Which gas law does this equation represent if volume is held constant so that $V_1 = V_2$? Gay-Lussacs

19. In which situations does the combined gas law enable you to do calculations when the other gas laws do not apply?
when only te amount of gas is constant

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

$\frac{P|V}{T}$ e flipped circle

Name ______________________ Date ________________ Class ________________

SECTION 14.3 IDEAL GASES (pages 426–429)

This section explains how to use the ideal gas law to calculate the amount of gas at specified conditions of temperature, pressure and volume. This section also distinguishes real gases from ideal gases.

▶ Ideal Gas Law (pages 426–427)

1. In addition to pressure, temperature, and volume, what fourth variable must be considered when analyzing the behavior of a gas?

\# of moles (n) = amt of gas

2. Is the number of moles in a sample of gas directly proportional or inversely proportional to the number of particles of gas in the sample?

proportional (directly)

3. At a specified temperature and pressure, is the number of moles of gas in a sample directly proportional or inversely proportional to the volume of the sample? proportional (directly)

4. Circle the letter next to the correct description of how the combined gas law must be modified to introduce the number of moles.

a. Multiply each side of the equation by the number of moles.

b. Add the number of moles to each side of the equation.

[✓ **c.**] Divide each side of the equation by the number of moles.

5. For what kind of gas is $(P \times V) / (T \times n)$ a constant for all values of pressure, temperature, and volume under which the gas can exist? ideal gas

6. What constant can you calculate when you know the volume occupied by one mole of gas at standard temperature and pressure?

ideal gas constant

$8.31 (L \cdot kPa)/(K \cdot mol)$

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ____________________ Date ______________ Class ______________

7. Complete the table about the ideal gas law. Write what each symbol in the ideal gas law represents, the unit in which it is measured, and the abbreviation of the unit.

Symbol	Quantity	Unit	Abbreviation for Unit
P	pressure	kilo Pascales	kPa
V	volume	Liters	L
n	moles	~~molecules~~ moles	mol
R	ideal gas constant	—	—
T	temprature	kelvin	k

-no inital of final conditions

need to figure w/ moles

8. When would you use the ideal gas law instead of the combined gas law?

When its not used

~~to calculate the # of moles in a gas at a know volume, pressure, temp~~ + fixed

Ideal Gases and Real Gases (pages 428–429)

9. Circle the letter of each sentence that is true about ideal gases and real gases.

a. An ideal gas does not follow the gas laws at all temperatures and pressures.

b. An ideal gas does not conform to the assumptions of the kinetic theory.

[✓ c.] There is no real gas that conforms to the kinetic theory under all conditions of temperature and pressure.

[✓ d.] At many conditions of temperature and pressure, real gases behave very much like ideal gases.

10. Is the following sentence true or false? If a gas were truly an ideal gas, it would be impossible to liquefy or solidify it by cooling or by applying pressure.

true   no attraction allowed

11. Red gases differ most from an ideal gas at low temperatures and high pressures.

12. Look at Figure 14.14 on page 428. What substance is shown? What change of state is occurring? How do you know this substance is not an ideal gas?

liquid nitrogen is vaporizing (liquid → gas)

not an ideal gas b/c its a liquid

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________________ Date ______________ Class ______________

CHAPTER 14, The Behavior of Gases *(continued)*

SECTION 14.4 GASES: MIXTURES AND MOVEMENTS

(pages 432–436)

This section explains Dalton's law of partial pressures, and Graham's law of effusion.

▶ Dalton's Law (pages 432–434)

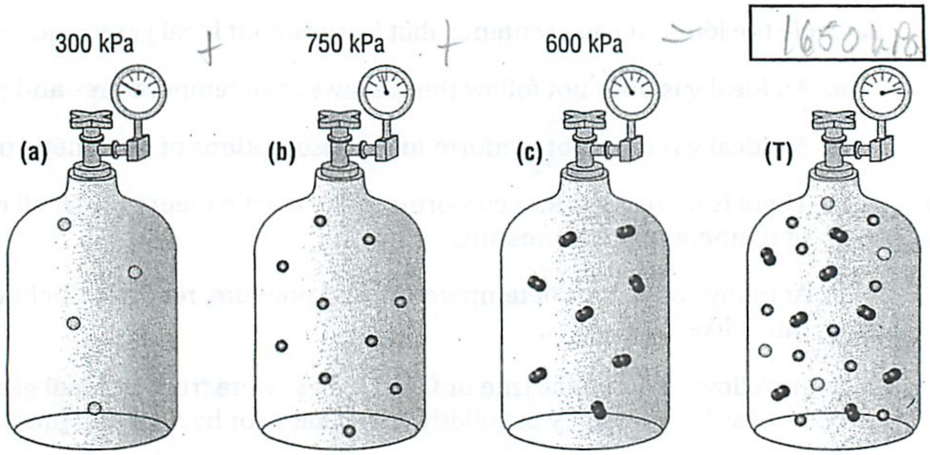
1. Is the following statement true or false? Gas pressure depends only on the number of particles in a given volume and on their average kinetic energy. The type of particle does not matter. true

2. The contribution of the pressure of each gas in a mixture to the total pressure is called the partial pressure exerted by that gas.

3. What is Dalton's law of partial pressures?

In a mixture of gasses, to total pressure is the sum of the partial pressure of gasses,

4. Container (T) in the figure below contains a mixture of the three different gases in (a), (b), and (c) at the pressures shown. Write in the pressure in container (T).



$P_{total} = P_1 + P_2 + P_3 \ldots$

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Graham's Law (pages 435–436)

5. The tendency of molecules in a gas to move from areas of higher concentration to areas of lower concentration is called diffusion.

6. What is Graham's law of effusion?

the rate of effusion is inversly proportional to square root of the gass molar mess

7. Is the following sentence true or false? If two objects with different masses have the same kinetic energy, the one with the greater mass must move faster.

false

lighter moves faster

Reading Skill Practice

You may sometimes forget the meaning of a vocabulary term that was introduced earlier in the textbook. When this happens, you can check its meaning in the Glossary on pages R107–R117. The Glossary lists all vocabulary terms in the textbook and their meanings. You'll find the terms listed in alphabetical order. Use the Glossary to review the meanings of all vocabulary terms introduced in Section 14.4. Write each term and its definition on a separate sheet of paper.

how fast particles spread out

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name

Date ______________ Class ______________

ıe Behavior of Gases *(continued)*

4 GASES: MIXTURES AND MOVEMENTS

(pages 432–436)

ıs Dalton's law of partial pressures, and Graham's law

e gas
olume?

W (pages 432–434)

ng statement true or false? Gas pressure depends only on of particles in a given volume and on their average kinetic type of particle does not matter. ______________

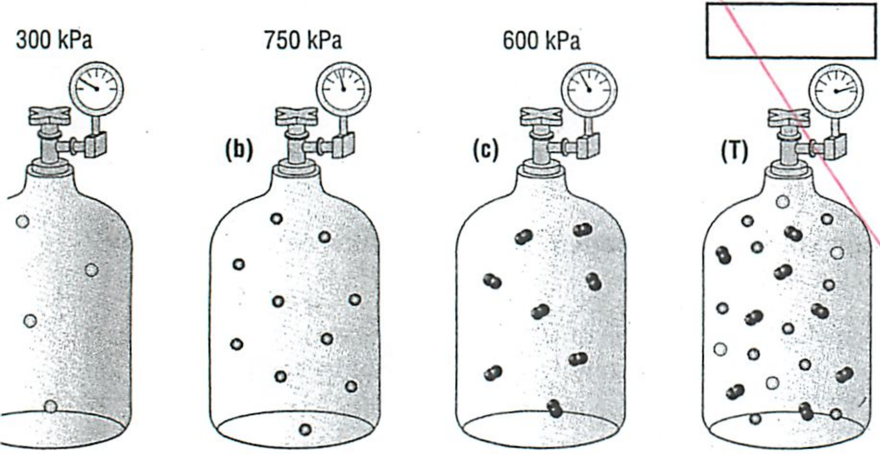
ıtion of the pressure of each gas in a mixture to the total pressure ______________ exerted by that gas.

on's law of partial pressures?

____

____

T) in the figure below contains a mixture of the three different , (b), and (c) at the pressures shown. Write in the pressure in T).



© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________________ Date ________________ Class ________________

EXTRA PRACTICE (similar to Practice Problem 11, page 423)

11. A gas has a pressure of 7.50 kPa at 420 K. What will the pressure be at 210 K if the volume does not change?

$? = \frac{7,50\,kPa \cdot 210\,k}{420\,k} = 3,75\,kPa$

looked y
up
now

GUIDED PRACTICE PROBLEM 31 (page 434)

31. Determine the total pressure of a gas mixture that contains oxygen, nitrogen, and helium if the partial pressures of the gases are as follows:

$P_{O_2} = 20.0$ kPa, $P_{N_2} = 46.7$ kPa, and $P_{He} = 26.7$ kPa.

93,4

Analyze

source • $\frac{target}{source}$ = target

a. What is the expression for Dalton's law of partial pressure?

$P_T = P_{O_2} + P_{N_2} + P_{He}$

b. What is the unknown in this problem? total → P_{total}

Calculate

c. Substitute the known quantities into the equation and solve.

93,4 kPa

Evaluate

d. Why is your answer reasonable?

yes - its the sum of all of them

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Review Graphs +
Grahman's Law

Review of States of Matter and Gases

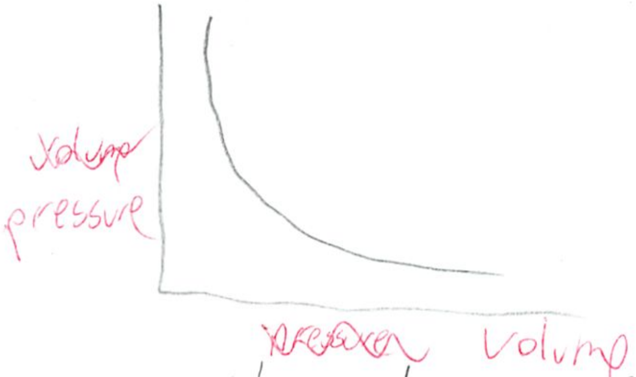
This information can be found in your notes, labs and Chapter 14 Big Book

1.State Boyle's Law (formula)

P • V   Pressure + volume are inversly related

2.Boyle's Law relates P and V when Temperature is constant

3. Sketch a graph of Boyle's Law.



4.Pressure = volume / temp. ← Charles

5. Units of Pressure are kPa, mm Hg and atm.

didn't learn, but same

6. 101.3KPa = 1 atm = 766 mmHg = 760 torr

7.What is meant by compressibility?

Same amount of particles in a smaller volume
Space between the particles is removed

8.In your own words, explain why gases are compressible.

The space b/w the particles is so large that particles can be forced to fit into those spaces

9.Factors which affect gas pressure are:

temp
pressure
volume

10.State Charles' Law.(formula)

gay lussacs

~~pressure~~ Volume
temp

11. Charles' Law relates volume and temp when Pressure is constant.

12.Sketch a graph of Charles' Law.

13.What is the Kelvin scale?How is it different from Celsius?

units of temprature where 0k is equal to absolute zero - its C -273

14. K= °C + 273

15.What is Gay-Lussac's Law?

$$\frac{\text{pressure}}{\text{temp}}$$

16. Gay-Lussac's Law relates P and T at constant volume.

17.What is the formula for the combined gas law?

$$\frac{P_1 \cdot V_1}{T_1} = \frac{P_2 \cdot V_2}{T_2}$$

18.Gases that obey all gas laws will be considered ideal.

19. What is the Ideal gas Law?

$$\frac{P_1 \cdot V_1}{T_1 \cdot n_1} = \frac{P_2 \cdot V_2}{T_2 \cdot n_2}$$

words?

20. How do Ideal gases differ from real gases?

assumes particles have no volume + no attraction forces b/w them

* only gasses - can't be solids or liquides

21. What is Dalton's law?

particles in each

Sum of ~~component of gas mixture =~~
~~the # of particles~~ in a mixture

sum of partial pressures = total pressure

related to pressure

22.What information does Graham's Law give us?

the * lighter gas the faster it travels

gasses w/ ↓ molecular mass diffuse faster ←

↗

less dense ← not really

rate of effusion is inversely proportional to $\sqrt{\text{molar mass}}$

23.What is effusion?

gas escaping thragh a tiny hole in a container

24. What is diffusion?

tendency of molecules to move toward areas of lower concentration until concentration is uniform throughout

25.Can you think of an application of Graham's law?(look up Uranium)

ballon deflatting through effusion

$$\frac{Rate_A}{Rate_B} = \sqrt{\frac{molar\ mass_B}{molar\ mass\ a}}$$

nitrogen = 28.0 g molar mass

helium = 4.0 g molar mass

$$\frac{Rate_{He}}{Rate_N} = \sqrt{\frac{28.0\ g}{4.0\ g}} = \sqrt{7.0} =$$

2.7x faster helium effuses then nitrogen

Michael Plasmeier

Laboratory Activity: Synthesis of an Artificial Flavor

$\frac{18}{20}$

NOTE: Do not taste or directly inhale your product;

Wear Safety Goggles during the entire laboratory activity

Dispose of solid product in the trash

Purpose: To identify functional groups in organic compounds

To synthesize an artificial flavor

To complete an organic reaction with structural formulas

Procedure;

1. In a small test tube add a small amount to salicylic acid. (Salicylic acid is a solid)
2. To that test tube, add a dropper full of methanol.
3. Have your instructor add a few drops of concentrated sulfuric acid. CAUTION! Sulfuric acid is extremely hazardous. You must wear eye protection.
4. Carefully stir your test tube.
5. Heat test tube in a water bath on a hot plate.
6. Note any odor by CAREFULLY WAFTING YOUR HAND ACROSS THE OPENING OF THE TEST TUBE. - smells like peppermint

Data: Describe the odor noted. (It will be pleasant)

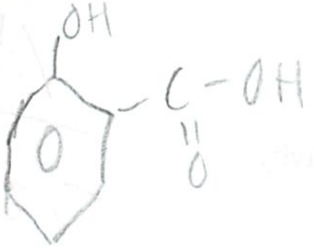
Calculations: Write out the reaction using structural formulas. The acid you used is salicylic acid.Look up structural formula in text book. The alcohol you used is methanol..

What is the role of the sulfuric acid in this reaction? Hint is it [✓ a reactant] or a product. Catalyst!

Reference: Wilbraham et al (Big Red Book) Ch. 23 p.742-3.

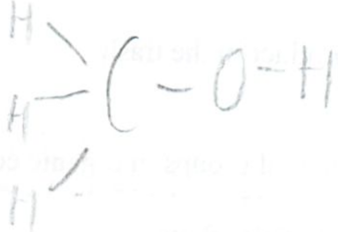
$C_6H_6(OH)_2CO$

Salicylic acid



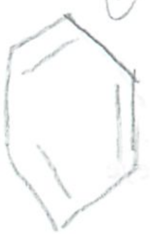
Methonal (methyl alcohol)

CH_3-OH -hydroxyl group
-primary alcohol



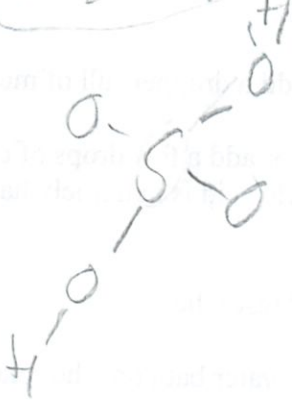
↗

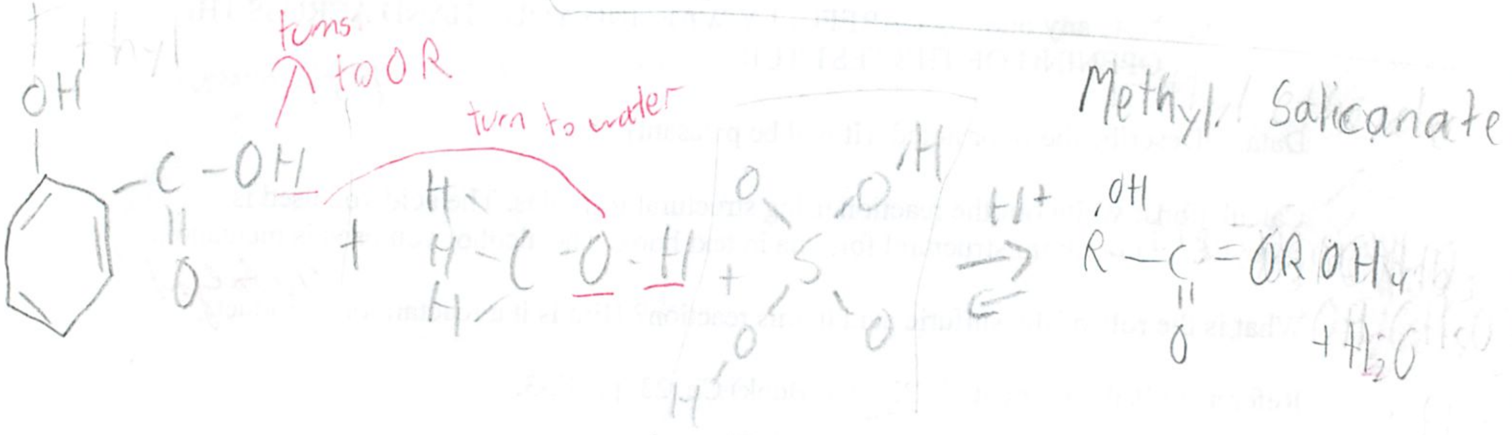
benzeen with ring in it means



Sulfuric Acid

H_2SO_4





salicyic acid + methonal + sulferic acid

corbonyl corbon

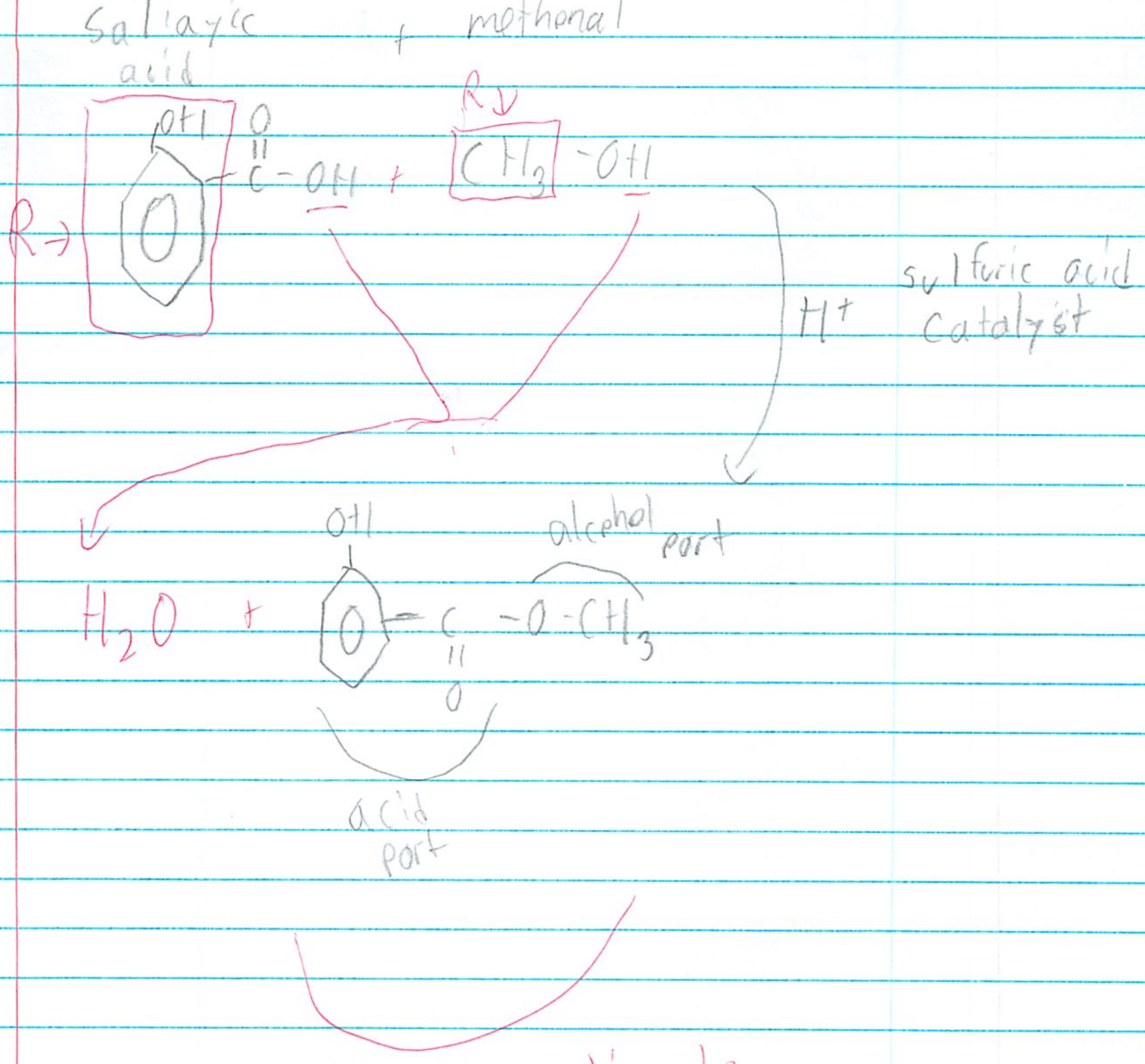
acid catylist

esters

Lab Formula

Answers

3/21



methyl salicyate

- commoni oid of wintergreen

Michael Plasmeier

3/12

✓ Makeup

MSE Act 7 p.53-54

1.What is a polymer? a substance that is a macromolecule consisting of many smaller molecules (monomers) linked together in long chains

2.What is meant by "viscosity"?

a property related to the resistance of a (Non-Newtonian) liquid to flow

3.What is meant by a non-Newtonian liquid?

A liquid which does not flow easily (in accordance w/ Newton's laws for liquid)

4.What is a monomer?

Similar smaller molecules which make up polymers

5. starch and cellulose are examples of natural polymers.

6.Humans can digest starch but not cellulose.

7. plexiglass ①, polyethylene, poly styrene, polyvinyl chloride are exaples.

Teflon ② manufacture from crude oil

Michael Plasneer   Organic Chem

READING GUIDE FOR CHAPTERS 22 AND 23 IN THE BIG RED BOOK

- study of molecular compounds of carbon

Read pages692-755

1.Simple organic compounds are called hydrocarbon.

C made up covalently bonded Carbon + Hydrogen atoms

2. Each carbon atom can form $\boxed{4}$ covalent bonds.

↑ one or more shared electron pairs

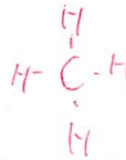
3.The shape of the methane molecule is due to the electrical repulsion of bonded electrons, this results in a tetrahedral shape. This is why 3D formulas are important.

ball + stick + space filling models (structural)

4.Alkanes contain only single ~~covalent~~ carbon carbon bonds. They can form straight chains or branched chains. or rings

Molecular formula - CH_4

Structural formula



5. Look at table 22.1 on p.695. Learn the names and molecular formulas for the first ten.

methane propane pentane heptane nonane
ethane butane hexane octane decane

memorize

6. What is the general formula for an alkane?

ends in -ane

C_nH_{2n+2}

7.Look at table 22.2. Identify each hydrocarbon:

a. CH_3-$CH_2(CH_2)_5CH_3$ octane

b. CH_3-$CH(CH_3)CH_2$-CH_2-CH_2CH_3

↑ branch

2methylhexane

8.How many carbon atoms are there in heptane? How many Hydorgens?

C_7H_{16}

9.An atom or group of atoms which takes the place of a hydrogen atom is called a substituent.

10. What is an alkyl group? Name the first ten.

A Hydrocarbon substitent to make up a branched chain alkane

-methyl
-ethyl
-propyl
-butyl
-pentyl
-hexyl
-octyl
-nonyl
-decyl

fator chains can be curved

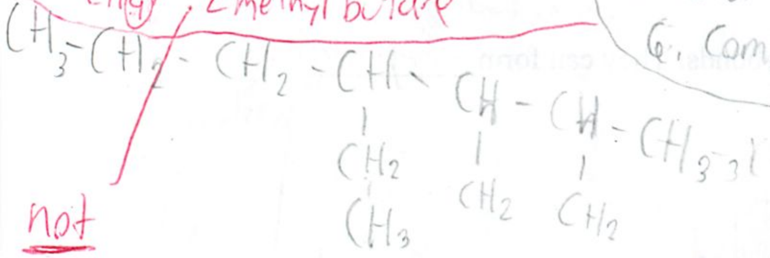
CH_3 CH_2 CH_3 ② ③
CH_2 ③ ①
CH ④ , 2 methyl butane

not 2-ethyl propane

11.In your own words, explain how to name a branched chain hydrocarbon.

1. Find name of longest carbon chain - its the parent
2. Number the carbons starting at the edge with the most branches
3. Add name of substituent branches with positions
4. Use prefixes if branch lenght occure more then once
5. List substitcent in alpha order w/o prefix
6. Commas seperate #s, Hyphens seperate #s and words

12.Name this molecule :



$CH_3-CH_2-CH_2-CH-CH-CH-CH_{3\,3}$ (branches below: CH_2–CH_3; CH_2; CH_2)

4-ethyl-2,3 dimethyl heptate

don't count to alphabetize

13.How are alkanes and alkenes different?

Sat.

alkane - hydrocarbon w/ 1 covalent bond

alkene - 1 or more carbon-carbon double covalent bond

Unsat.

(flat)

14.What is the general formula for an alkene?

ends in -ene

C_nH_{2n}

double

$$H\!>\!C=C\!<\!H$$
(H, H)

15. How do alkynes differ from alkanes and alkenes?

alkynes contain 1 or more carbon-carbon triple covalent bonds

both are Unsat.

C_nH_{2n-2}

triple bonds

✓ not plentiful

16. What is the general formula for an alkyne?

-yne

$H-C\equiv C-$

17. What is an isomer?

compounds w/ same molecular formula but diff. structure

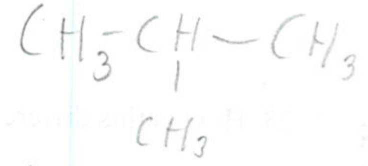
r are branched = boil point

18. Explain the three kinds of isomers.

Structure - same formula but diff order
- diff boiling + melting pt - diff chem properties

Stereoisomers - same formula, order - diff position
- geometric isomers - 2x bond prevents rotation of carbon atoms among each other
trans - methyl groups on opposet side of 2x bond
cis - methyl groups on same side
diff phy + chem properties
- optical isomers - carbon must be asymmetric
diff only in where 4 are arranged

$CH_3-CH_2-CH_2-CH_3$



$CH_3-CH(CH_3)-CH_3$

both C_4H_{10}

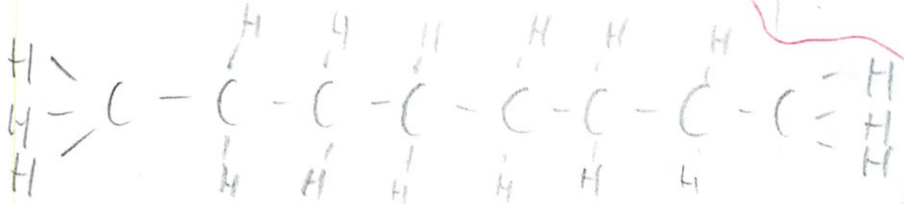
Functional - same formula - diff functional groups

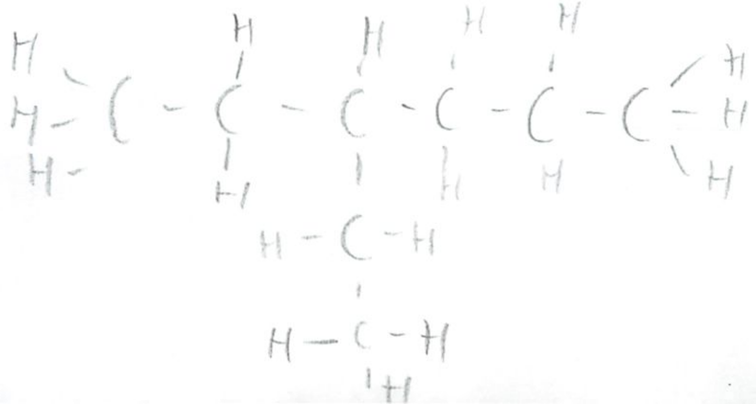
19. What is an asymmetric carbon?

Carbon has 4 different atoms or groups attached

aka 2hiro carbon esp

20. Draw the structural isomers of octane.

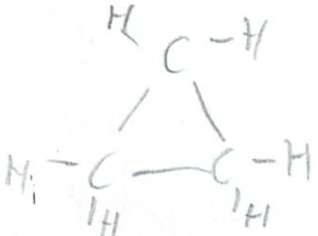




21.What is a cyclic hydrocarbon? Give some examples.

compounds which contain a hydrocarbon ring

cyclopropane



22.What structure does an aromatic hydrocarbon contain?

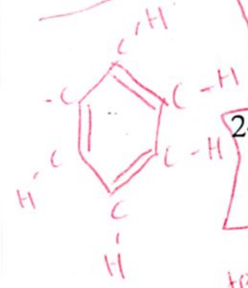
arene

Organic compound which contain benzene ring or other ring in which the bonding is like that of benzene

23. How is this different from a cyclic hydrocarbon?

benzene or benzene like ring which exibit resonance - more stable *alt. double bonds

benzene



(represented by

24.What is the major source of hydrocarbons?

the earth -natural gas -coal -petroleum } fossil fuels

25. Explain how hydrocarbons are separated from each other in crude oil.

by fractional distillation in a furnace each type boils at a different temp some is taken and condensed back into oil

26.How is coal classified?

hardness + carbon content

water content } some thing

27. Coal is a major source of condense aromatic compounds with a high molar mass, High proportion of carbons compared to hydrogen.

% is ocate which never knocks

and not % of heptane which knocks

28. Explain what is meant by the octane rating of gasoline. What is "engine knock"?

Octane raiting - measure of ability of resistance to engine knock (sound when gas ignites too soon b/c of poor timing) fuel ignites by itself

-premature ignition

29. What is a functional group?

specific arrangement of atoms in an organic compound caple of characteristic chem reactions - and classified by that

reactive group on molecule containing only O, N and S atoms

30.Triflouromethane is an example of a __________.

-dry cleaner

halide CHF_3

31. What functional group does an alcohol contain?

hydroxyl R-OH

32. What does the R stand for?

any carbon chains or rings attached to functional groups alkyl any group or aryl group

33.What is the difference between an amine and an amide?

amine - carbon single bond amide - O bonded to carbon

34. What is the difference between a carboxylic acid and an alcohol?

carboxylic acid ads an oxygen-carbon double bond

alcohol - doesn't have it

alcohol - OH group

35.What is a substitution reaction?

a atom/group of atoms replaces another atom/group of atoms

36.Explain hoe to form an alcohol from a halide.

Add an "-OH" group -react in water

halide has R + F, Cl, Br, or I

37.What is an addition reaction.? Which kind of hydrocarbons undergo addition reactions?

a substance (hydrogen usually) is added at the double or triple bond of an alkene or alkyne - introduce new functional groups

break it

38. List all of the functional groups containing the carbonyl structure.

Carbonyl - C=O

2x bond w/ oxygen

aldehyde - carbon always joined w/ one H (RCHO) -e → -al

ketone - carbon always joined w/ two carbons (RCOR) -e → -one

see table 23.1 big red book

39. What is a fatty acid?

Continuous-chain carboxylic acids

on Final exam

40. The reaction of a acid and a alcohol yields an ester and water.

esterification (

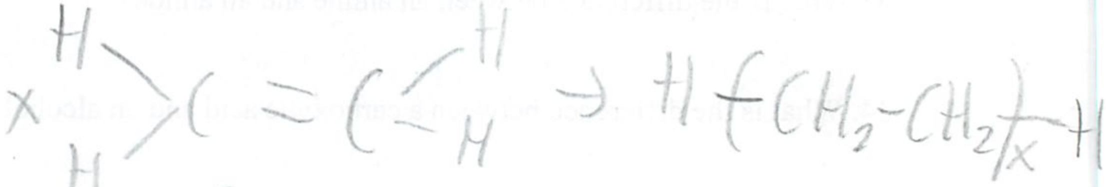
mineral acid as cataylist

41. What is polymerization? Write out the reactions for addition and condensation polymers.

Polymer - large molecule formed by covalent bonds of repeating smaller molecules (monomers)

polymerization - reaction which joins same-kind monomers to form polymers (w/ catylist)

addition - unsat. monomers react



$$\rightarrow H\{CH_2-CH_2\}_x H$$

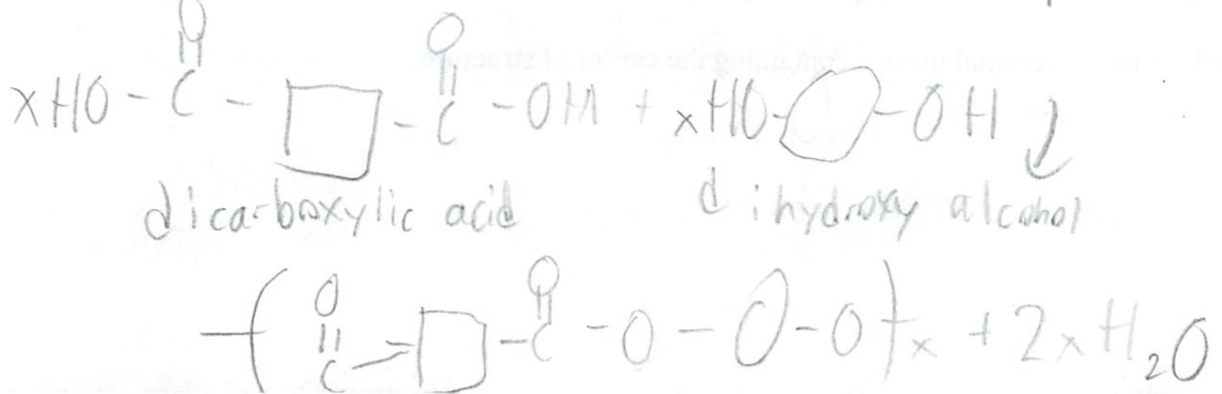
(X=# of ethylene units)

x=100 wax

x=1000 rigid

Condensation - head to tail joining of monomer units

- forms water
- circle + squares = unreactive portions



Michael Plasmeier

Elements + Their Properties

20/20

3/21

White
74

To test and classify materials and classify by properties

What do you think an element is?

A chemical element is the purest part of chemistry. It can not be seperated any furter.

1, Make list of 4 substances you think is an element. in everyday life

- oxygen
- iron
- nickel
- oxygen

2, Are these thing elements?

Greek: earth, air, fire, and water
Middle Ages: mercury, sulfur, and salt

The people of the middle ages were a lot closer. Air can be split up further and so can water. Hg, S, Na are all today still considered elements

3. Get + Record properties of metals to test

Element	Oberve	Conductive	Reactive	Metal?
glass	clear, brittle, hard	N	N	N
plastic	clear-ish, soft	N	N	N
silicon	shiny, rocklike, irreg shape, non maluble eaisly breakable	Y	Y	Y

20/20

copper	red orange color thin wire	Y	Y	Y
lead	non shiny silva color flat chips, maluble	Y	Y	Y
tin	silver-like color, shiney maluable, flat chips	Y	Y	Y
sulfer	yellow power, solid maluble	N	N	N
aluminum	shiney silver color chips thin - very maluable maintains creases	Y	Y	Y
magensium	matte silver color in a flat maluble ring	Y	Y	Y
emery cloth	rough, dark color, rust spots	N	N	N
iron	rusty, nonmaluble	Y	N	Y
steel wool	thin maluble wire mess dull silver color	Y	N?	Y
wood	brown color, nonmaluable hard	N	N	N

Michael Plasmeier

Reading Guide Activity 6 MSE

Read Chem Talk p.47-48

1.List and describe properties of metals.

have luster (reflection of light)
conductivity (transmits heat +electricity)
malleable + ductile (bends w/o breaking)
reactive (interact readily w/ other materials)
loose electrons readily (conductive)

2.List and describe some properties of non metals.

surfaces dull + brittle (no luster)
nonconductive
brittle - break easily
non ductile
don't oxidize

3,What is luster?

reflection of light from surface of material described by quality + intensity

4. What is elasticity?

property of a material to resist deformation + return to its normal size or shape after force has been applied to it (+removed

5. What property can be observed due to the "electron sea" model of metals?

conductivity of electricy

6.Another term for rusting is oxidation.

7.What is an alloy?

a substance that has metal characteristics and consits of two or more different elements

^can be non metal element

8.Why are alloys made?

- Some elements can't be worked with directly

9. Give some examples of processes which can help prevent oxidation.

metal surfaces can be painted, coated or combined w/ another metal to make them less reactive

10.Give some examples of alloys.

brass
solder
gold in jewlery

Do Chem to Go p,49

Michael Plasmeier

Metals + Non Metals

Chem to Go

3/22

White Book p49

1. 3 metals we know + 2 uses for each
   - steel
     - building
     - cars
   - copper
     - pennies
     - wires
   - aluminum
     - aircraft frame
     - foil to cover stuff
2. 3 non metals we know + 2 uses
   - wood
     - constiction
     - playgrounds
   - plastic
     - toys
     - chairs
   - glass
     - windows
     - jars
3. Characteristics of these + their componets
   - backpack
     - plastic
       - strong
       - waterproof
       - malvable
     - metal
       - zippers
       - nonmalvable in bulk
       - noncorrosive
       - durable
   - bicycle
     - aluminum
       - light
       - strong
       - nonmalvable when together
     - leather
       - seat cusion
       - soft
       - water proff
       - malvable
   - car
     - steel
       - strong
       - heat resistant
       - somewhat light
       - nonmalvable when togeter
     - plastic
       - see above

4. a 2 properties which are qualitative
- Scent - color (could be machine) - luster

b 2 properties quantative
- height - temp

5. Metal or non metal?
aluminum (Al) - metal
iron (Fe) - metal
Oxygen (O) - non metal
carbon (C) - non metal
mercury (Hg) - non metal

6. Factors to decide to change material
- saftey - weight - infrastucture
- strenght - cost - recycble?
- printabilty - how its made - env. Impact?

Michael Plasneier

Identifying Matter Lab

18/20

3/22

to identify metal ions present in materals by producing colored flames

White
Bech

What do you think - How to test for poisonous substances?

You could note the properties of poisonous substances and then check for those. Based on this lab's intro + picture, color when burned may play a big part.

1. How can distinguius one stick from another (each stick soaked in different chemical)
   You could expose it to a flame too see the color of the flame.

2. Light Bunsen burner and note its color.
   - blue with ocasional streeks of red
   - darker blue at bottom

3. Take splint and hold over fire. Note color and then extinguish in water BEFORE WOOD BURNS

Clocide X

Ion	color
Na	bright orange
K	red - blue (purple)
Ca	reddish - orange
Sr	orange
Cu	green
Li	red - fuscia
Ba	yellow
?	red - blue - Pink (k)

MSE Activity 8 Reading Guide

1. Did you observe a chemical reaction in this activity? WHY or WHY NOT?

No - the metal ions' electrons absorb energy from the Bunson burner, when they lose the energy again they fall back down giving off a unique amt, of energy represented by color.

2. What is an ion?

a electrically charge atom or grap of atoms that has acquired a net charge (+ or -)

3. What is an electron?

a subatomic particle that occurs outside of the nucleus and has a charge of -1 and a mass of $9,109 \times 10^{-28}$ g

4. Describe where electrons are located in atoms.

floating around the nucelus

changing levels

5.Explain in your own words how flame tests are produced.

The difference in electrons in each ion type produced a different energy let offs producing diff. colors allowing a flame test.

6.What is the nucleus of an atom?

The very dense core of the atom that contains neutrons and protons

7.The colors observed are produced by the changes in energy levels.

8. Is the color produced when the electron first changes level?

No, when they fall back down they give off energy as light

9. What is meant by energy is conserved?

energy can only be transfered- it can not be created or distroyed

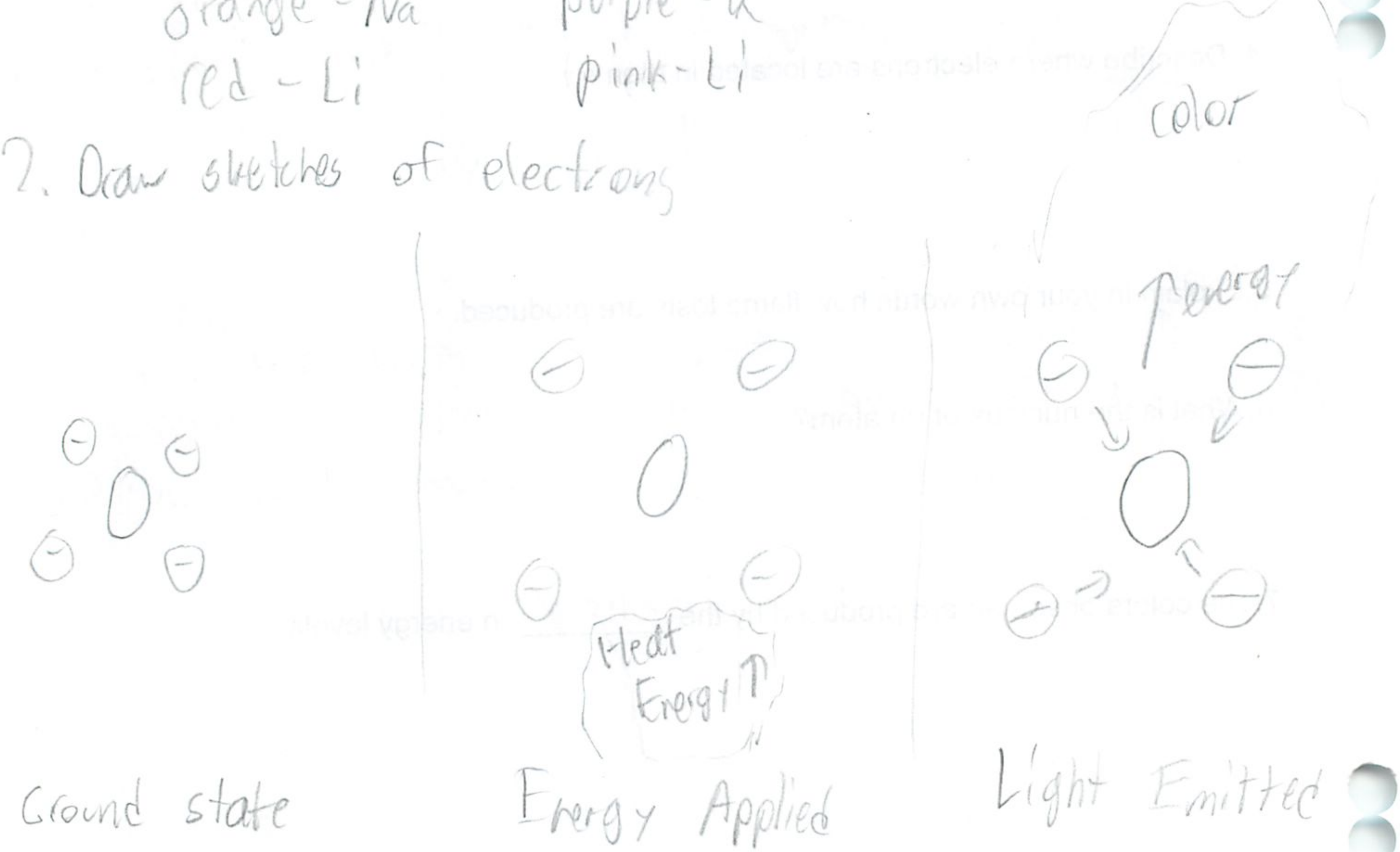
Do all Chem to Go. Remember to write out questions and/or incorporate the question in your answer.

1, What metals are used in fireworks,

green-Cu
orange-Na
red-Li

yellow-Ba
purple-K
pink-Li

2. Draw sketches of electrons



Michael Plasmeier

Atomic Structure
Questions

3/22

Read p100 Chap 4

1. How did Democritus characterize atoms?
Democritus thought atoms were indivisable + indestructable, However his ideas were not developed using the scientific method.

2. How did Dalton adv. the atom theroy,
Dalton used exprimental methods expanded it into scientific theroy.

3. How can atoms be observed?
Individual atoms can be seen with a scanning electron microscope.

4. State Dalton's theroy
All elements are made by indivisable particles called atoms. All of the atoms in one element are identical, and different from those of other atoms, These atoms are able to combine in simple whole number ratios to produce compounds, When atoms join, split or are reorranged, a chemical reaction occurs

5. Can you convert one element's atoms into another,
No, atoms of one element can't be changed into another element, Sorry alchmestist,

6. Describe the radii of most atoms in nm
5×10^{-11} to 2×10^{-10}

7.

7, A sample of Copper wire with a mass of 63,5 g has $6,02 \times 10^{23}$ atoms.

$$63,5 / 6,02 \times 10^{23} = 1,05 \times 10^{-22}$$

g atoms g/atom

4,2 p104 8, 3 types of sub atomic particles:

Proton
Neutron
Electron

9, Rutherford's model of atom

Protons + neutrons in the nucleus.
The electrons are distributed around the nuclus and occupy almost all of the volume of the atom

10, Charges + Relative mass (proton = 1) of 3 subatomic particles

Protons	⊕1	1
Neutrons	0	1
Electrons	⊖1	1/1840

11, Thomson's + Milkan's contribution to atomic Theory

J.J. Thomson discovered the electron when he made cathod-ray tubes. He saw that the gas was affected by a magnet.
Robert Millikan calculated the mass and charge of an atom very close to today's accepted values.

12. Rutherford's Gold Experiment

Ruthersford Gold Experiment showed how a small percentage of light bounced back hitting gold foil. He was expecting that most of light would have passed through with only minor deflection. However he had discovered the existance of the nucleus of an atom.

13. What experimental evidence told him that an atom was mostly open space.

The fact that most of the light passed straight through the gold sheet without deflection

14. How is Rutherford's model different from Thomson's

Rutherford was able to add the protron and neutron to Thomson's electron as subatomic particles.

17. # of neutrons in each atom

$\frac{\text{mass}}{\text{atomic}}$

neutrons = mass # - atomic

a. $^{16}_{8}O$ 16-8=8

b. $^{32}_{16}S$ 32-16=16

c. $^{108}_{47}Ag$ 108-47=61

d. $^{80}_{35}Br$ 80-35=45

e. $^{207}_{82}Pb$ 207-82=125

19. Write symbols of isotopes oxygen-16, oxygen-17, oxygen-18

$^{16}_{8}O$

$^{17}_{8}O$

$^{18}_{8}O$

20. How many neutrons in each isotope when atomic # = 24.

23-33

chromium-50 = 50-24 = 26
chromium-52 = 52-24 = 28
chromium-53 = 53-24 = 29

23. Calculate avg. atomic mass

copper-63 = 69.2% = 62.93 amu
copper-65 = 30.8% 64.93 amu
62.93 • .692 = 43.5 amu
64.93 • .308 = 20.0 amu
63.5 amu averge Cu

24. Avg. atomic mass of Br

Br-?₁ = 78.92 amu 50.69%
Br-?₂ = 80.92 amu 49.31%
78.92 • .5069 = 40.0 amu
80.92 • .4931 = 39.9 amu
79.9 amu avg Br

25. Whats different from one atom of an element to the atom of another element.

Elements are diff. because they contain different numbers of protons.

26. What equation tells you how to calculate the # of neutrons in an atom?

The number of neutrons in an atom is the diff. b/w mass number and atomic number

27. How do isotopes of a given element differ from one another?

They have the same # of protons but diff # neutrons.

28. How is avg. atomic mass calculated?
To calculate the atomic mass of an element, multiply the mass of each isotope by its natural abandence expressed as a decimal, then add the products

29. Why is the porodic table useful?
The periodic table is useful because it allows you to compare the properties of one element (or group) to anoter element or group. Each row of the table is a period and each vertical column is a group, with similar properties but amoung differen periods

30. What is platinum-194?
194. is the mass number (# of proton + neutron) of this isotope of platinum, Platinum has an atomic number of 94. This is also the number of protons (which happens to be = to the number of elections, There are 100 neutrons.

31. Why are AMUs not whole #s?
Atomic masses are not whole units because it is the mass compared to $\frac{1}{12}$ a carbon atom. Because most atoms are not based off that size atoms the sizes are not whole numbers.

32. List # protons, neutrons, + electrons.
a. $^{6}_{3}Li$ - pro 3 - ele 3 - nev 3
$^{7}_{3}Li$ - pro 3 - ele 3 - nev 4
b $^{42}_{20}Ca$ - pro 20 - ele 20 - nev 22
$^{44}_{20}Ca$ - pro 20 - ele 20 - nev 24
c $^{78}_{34}Sc$ - pro 34 - ele 34 - nev 44
$^{80}_{34}Se$ - pro 34 - ele 34 - nev 46

33. Name 2 elements w/ properties similar to Ca
The 2A group (Alka lines and Earth Metals)
including Be, Mg, Sr, Ba, Ra

Michael Plasmeier

Organizing a Store

20/21

3/26

p72
White

to plan the arrangement of a store in order to relate to the organization of a store

What do you think - How many items does a store have?

I would say around 10,000 skUs,

1. Brainstorm some item to sell at a supermarket

apples
bannas
Oranges
plums
tomatoes
peppers

hamburger
hot dogs
hot dog buns
turkey
hamburgers buns
frozen hambburgers

Coke
Sprite
pepsi
serra mist
root beer
mug

8 fl oz
12 fl oz
20 fl oz
2 lts
case cans
12pk cans

bread
- sliced
- nonsliced
- poppy
- whole wheat

bagels

1a Make a map of the store
- consider the raceway, cash register areas
- arrangement left to right

Chem to Go
p73

1. What pattern did you use?
We placed the single-serve stuff to go on the outside and the larger bulk items on the inside. Like items are sort-of together but not all drinks are in the same place.

2. Choose 1 asile and go front-to-back.
We placed are items sort-of around the circle with little reguard for front-to-back placement. The good foods are in the begining and the bad food toward the end. The shelves are placed to restrict side to side movement.

3. Where put chocolate covered peanuts?
I would put them in the candy near the cash register

4. Thanksgiving paper products.
We would placed them in with the paper products and the high profit items in islands on the raceway.

5. High profit items, where?
We have high margin items around the raceway. there might be place to put more in "islands" in the middle of the raceway. Actully, most of the store is a raceway.

6. Any blank spaces.
Not really, we can always re-arrange the store to get space for items. Mendeleev however cald not rearrange his table after it was published. Profit now is better then short term loss.

Michael Plasmeier

Atoms + Their Masses

20/20

3/27

to explore the ideas of atoms by trying to isolate a single atom and compare Dalton's experment results to the ones known today

When did you hear of atoms. What do they mean to you?

I first heard about them in elementry school. They mean the same to me now as they did then. Atoms are one of the smallest components of materials.

Atoms

Know	Think	Want
electron	more then 3	how broken up further
proton	sub-particles?	how gain + lose electrons
neutron		about isotopes

1. One way to think about atoms is imagining isolating one atom, keep cutting a piece of aluminm foil in half until you can't any more. How many cuts? How big final piece compared to orignal? Does cut piece have have same properties?

21 cuts

has same properties

2. You still haven't gotten to an atom. Imagine continuing to cut 10 x, 100 x, 1000 x. Eventually you will get to one atom.

3. Chemists combine elements to form new substances. By measuring the amts of elements used + produced, they draw conclusions about elements involved.

4+5. Measure mass of 50ml beaker, .20 g of aluminum foil, mass of paper - put foil into beaker chopped

6,7. Place 2.00g of $CuCl_2$ onto paper + then in beaker
What's happening in beaker? Color of foil? Why? +add water

8,9,10. Find mass of filter paper + make funnel. Pour beaker contense into filter. Place filter onto paper towel + let dry 12 hrs.

11. When paper is dry, measure contense. What elements?

12. This was a single-displacement reaction. Aluminum replaced Cu in $CuCl_2$. The Cu becomes its own free element. The Al joins w/ the Cl.
How much Al did you start with. How much Cu did you end with. Ratio of mass Cl : mass Al

12c. The mass of Cu and Al should be the same. How much more massive is a Cu atom then Al atom

13. Your results compared to class?

Data

Mass of empty 50ml beaker	29.69	g
Mass of beaker + foil	29.89	g
Mass of aluminum foil	.20	g
Mass of paper	.23	g
Mass of paper + $CuCl_2$	2.23	g
Mass of $CuCl_2$	2.00	g
Mass of filter	1.11	g
Mass of filter + product	2.75	g
Mass of product	1.64	g

Al + $CuCl_2$ in beaker
- Al sheets turning blackish red
- pieces statically stick together
- $CuCl_2$ turns brighter green
- beaker starts to feel warm

\+ Water
- sizzed immediaty
- condensation on side
- steam comming out
- bubbles in liquid
- foil immed. turing Cu red - like rust
- container warm
- liquid dark green colloid
- Al pieces disolving
- when mixing light gray rises w/ Al pieces
- after ~5 min dark red solids are being produced
- adding more water - purple color

sample
↓

filter

"ring" left on filter paper

~~Clear water~~ coming out of filter

- or light blue (Cl?)

after drying

- $ca blue powder at bottom - caked
- half of filter is blue
- bight yellow is top of filter
- fiter hard - keeps shape of contaier

Michael Plasmeier

Chem to Go

Atoms + Their Masses 3/28

p88 1. a. What conclusions could Dalton draw about relative masses? How many times more massive is O then H?

Could Dalton say that an oxygen molecule is 8 times more massive then a H hydrogen?

b. How do the current atomic masses of O + H compare to Daltons?

These compare to the values. Atomic mass = mass - # of neutrons which is what we found above. # Protons = # Neutrons

c. How many times more massive is an O then a H atom?

An oxygen is still 8x more massive then a hydrogen atom. - dosn't matter about compound - still the same ???

d. Are the revised masses more accurate?

Yes, duh, otherwise they would not be revised. They are right on now.

2. If Dalton thought ammonia formed with one atom of each element, he would have thought that their relative masses were the same. ??

b. Ammonia = H_3N If 3lbs H + 14 lbs N is b/c there are 3 H for each N, how many times more massive is the H atom then the N atom?

; would it matter about compound - no

H is 4,666 x more massive then a N

3. A student uses magnesium not aluminum and get the data listed on p 89 white book.

a. .20 g of magnesium used here

b. .52 g copper in filter

c. .52 : .20 → 2.6 : 1 ratio copper : magnesium

d. The masses above should have same # of atoms in each substance because of equal conversion. The relative size of a copper atom is 2.6 times the size of a magnesium atom, and a copper atom is 2.6 times as massive as a magnesium atom

4. Add data to element cards from chem talk. Does this new info (relative masses) change/help arrangement?

I don't know any thing about element cards. However the relative masses should help better charitize atoms. The perodic table does this I believe.

Michael Plasmeier

Atoms
Question

3/28

p84-87 Chem Talk 1. What's an atom?
An atom is the smallest representative part of an element.

Qu from Board 2 What did Aristotle and Democritus believe?
Aristotle did not believe in atoms because that would mean that there was empty space between the atoms. He did not believe that this was possible. Democritus, however did think that there were small particles called atoms which made everything up

3. What did "atomos" mean in Greek?
indivisable

4. How many atoms are there in 27g of Al? How many atoms are there in 63.5g of Cu?
6.02×10^{23} atoms - one mole) relative
6.02×10^{23} atoms - one mole again) atomic mass

5. How do compounds form?
Elements undergo chemical changes to have there atoms combine

6. Who was John Dalton? His ideas?
John Dalton was a early 19th century chemist who advanced the belief in atoms with his proportional combonations studies. He found that 5lbs nitrogen reacted w/ 1lbs H to make ammnomia.

7. What is an "atomic mass unit"?

An atomic mass unit is one-twelfth of the mass of a C-12 atom. The masses can be compared to find relative masses. 1 AMU = H

8. What is atomic mass?

Its the relative mass of an atom, determined by the mass of the protons + neutrons of an atom.

9. What is the law of defenite proportion?

The law states that the composition of a pure substance is always the same or the elements of the compound always combine in the same proportion by mass.

10. How can the existance of atoms help explain the Law of Definate Proportions?

The atoms provide the means for the law of definate proportions to work. It does not prove the existance of atoms, but makes it very convincing.

Michael Placmelor

Are Atoms Indivisble?

Activity 4

20/20 4/10

to observe the behavior of cathod rays and simulate two classic experments

Q. How are an atom of gold similor and different?

A. All atoms have nuclues, protons, and electrons spining orround them. The atoms of different elements have different numbers of these. (sometimes a same element has atoms with diff # of neutrons - an isotope)

1. Watch a demo of a cathod ray tube. What happens when a magnet is placed near the tube? What happens when magnet poles are reversed?

2. Magnets exert a force on moving electrons (as in a cathod ray tube) requardless of type of metal. This showed that the atom is not indivisable. The atom is made up of smaller parts, such as electrons. Magnets affect these electrons when exibited in cathod ray tubes.

3. Think about playing "Battleship". How hard would it be to hit a ship on a 1000 • 1000 board? This is what Lord Ernest Rutherford did in 1911 when he bombarded gold atoms with enorgetic particles. He needed a large number of "missles" to score a hit.

mass - pro + neu Michael Plasmeier
atomic # - pro/ele

au - mass

Reading Guide for Chem Talk p.94-96

1.What did JJThomson discover?

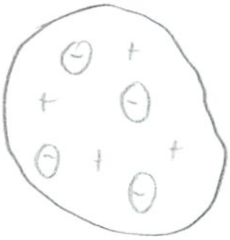
That there are negitily charged particles which Could be removed from atoms.

2.What did he call the particles he removed from atoms?

electrons

3.Discribe Thomson's model of the atom. (sketch)

a.



4.What did Rutherford discover?

That all of the mass + pos. charge of an atoms is concentrated at an extremly small part in the center (1/100,000 of the area)

5. Discribe how Rutherford made his discovery.

He sent alpha particles to a thin sheet of gold. Most of the particles passed right thragh but a few "hit" the sheet and bonced back

6.What was the role of alpha particles in Rutherford's experiment.

The "missiles" which tried to pass through the gold sheets

7.What happened that was unexpected?

Some particles bounced back

8.What term did he give the center of an atom?

nucleus

9.What did he call the smallest unit of positive charge?

proton

10.How did Rutherford discribe his results?

It's like firing a 15" shell at tissue paper and having it come back to hit you

11.What is Coulomb's force law?

positive charges repel each other

$$F = \frac{kq_1q_2}{d_2}$$

k = Coulomb's constant ($k = 9.0 \cdot 10^9 W\ m^2/C^2$)

q = charge in Coulombs

d = distance b/w charges

12. What does q1 and q2 represent?

charge in coulombs of two positive particles which are being compared

13. What is the relationship between the deflection of the alpha particle and the force of repulsion?

Its the same according to coulomb's law

Do Checking up p.96 and Chem to go p.97

ecking up .96

1, What is an electron?

An electron is a subatomic particle outside of the nucleus w/ a charge of -1 and mass of $9.109 \cdot 10^{-28}$ g

2, Why is thomson's model a "plum pudding model?

Thomson's model of the atom is like a plum pudding because the solid, small electrons are right inside a clump of light positive matter.

3, Why was Rutherford suppised w/ some atoms bounced back?

He was supprised that w/ the probability of hitting a small nuclus so large he somehow mannage to hit a few.

4, What is the nucls?

The very dense core of the atom which conteins neutrons, protons,

Chem to
60
p97

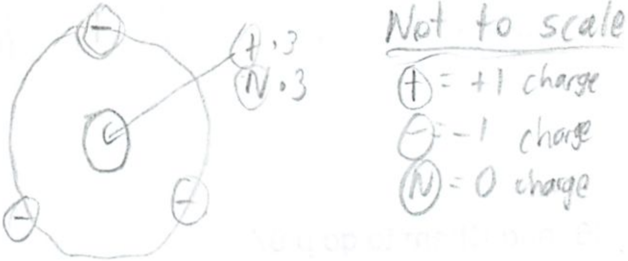
1. Since electrons are ⊖ and nucleus is ⊕ where would you expect it?

Ah I see that now, Shouldn't the electron be attracted to the nucleus.

2. Are atoms indivisible?

No there are smaller sub atomic particles called protons electrons and neutrons. They can be seperated like in the cathod ray tube or by seeing that most alpha particles pass through the gold.

3. Make a diagram of an atom.



4. How does atomic number compare to charges of nucleus?

The atomic number is equal to the number protrons in the nucleus of an atom, Because each protrons has a charge of ⊕1 the chare of the nucleus = the atomic number.

5. Chlorine has atomic number of 17. How many electrons does it have?

The number of elecctons is equal to the number of protons which is the definition of the mass number,

6. Whats probility of these situations

a 50,000 / 100,000 = ½ charch
b 25,000 / 100,000 = ¼ charch
c 1 / 100,000 = 1/100,000 charch

Situation C is similar to Rutherstord experiment beacuse the neutron is so small compored to most of the atom which is empty space

Michael Plasmeier Chem Behavior of Atoms

Activity 5

20/20

4/10

white
p99

to view the spectrum of hydrogen and interpeiet and predict with Bohrs model to identify the colors and regions of the electromagnetic spectrum

Q. How is color produced in a neon sign?
A. A neon sign uses electicity to activate a gas (neon) which glows. Different gasses (such as Ar, He, and Na are added for color.

1. To observe behavior of atoms, look at spectrum of light given off when gas excited by a high-voltage electric power supply (like a neon sign)

1. Set up a tube of H gas with high voltage power supply. View it through a spectroscope or a diffraction grating lens. Look at the colors of the spectrum to the side. What colors do you see? Draw with colors, orders, and spacing

2. The color of light is determined by its freq. The > the enrigy the greater the freq.

$f = \frac{c}{\lambda}$

f = freq. in waves, cycles, hz
c = speed of light
λ = wave lenght in nm

$E = hf$

E = energy
h = Planks constant
f = freq. (calc above)

Calculate freq. for other colors

80/80

Colors ~4-6

↑not suppose to see

The formula is in m so convert λ nm → m by $\times 10^{-9}$

	Wavelenghts λ (nm)	Freq (Hz)	E
purple	410.3	$7{,}307 \cdot 10^{\cancel{5}\,14}$	$4{,}84 \cdot 10^{-\cancel{28}\,19}$ J
blue	434.2	$6{,}905 \cdot 10^{\cancel{5}\,14}$	$4{,}58 \cdot 10^{-\cancel{28}\,19}$ J
green	486.3	$6{,}165 \cdot 10^{\cancel{5}\,14}$	$4{,}09 \cdot 10^{-\cancel{28}\,19}$ J
red	656.5	$4{,}567 \cdot 10^{14}$	$3{,}03 \cdot 10^{-19}$ J

3. In 1913, Niels Bohr, tried to explain the line Spectrum by hypotheeizing that the electron in the atom is only allowed to have a certain energy. Atoms would give off energy + light when they fall to a lower level. Like the Solor System electrons orbit around the nuclus. Thus electrons could only be a certain distance from the nuclues depending on the energy levels. The lights were a result of these energy jumps. Bohr said the lowest energy level to be $-2{,}18 \cdot 10^{-18}$ J.

$$E_n = \frac{1}{n^2} E_1$$

n = energy level

$E_1 = -2{,}18 \cdot 10^{-18}$ J

$$E_2 = \frac{1}{2^2}\left(-2{,}18 \cdot 10^{-18}\right) = -5{,}45 \cdot 10^{-19} \text{ J}$$

$$E_3 = \frac{1}{3^2}\left(-2.18 \cdot 10^{-18}\right) = -2{,}42 \cdot 10^{-19} \text{ J}$$

$$E_4 = \frac{1}{4^2}\left(-2{,}18 \cdot 10^{-18}\right) = -1{,}3625 \cdot 10^{-19} \text{ J}$$

$$E_5 = \frac{1}{5^2}\left(-2{,}18 \cdot 10^{-18}\right) = -8{,}72 \cdot 10^{-20} \text{ J}$$

$$E_6 = \frac{1}{6^2}\left(-2{,}18 \cdot 10^{-18}\right) = -6{,}05\overline{5} \cdot 10^{-20} \text{ J}$$

4. In Bohr's model to jump from $E_3 \to E_2 = 3{,}03 \cdot 10^{-19}$ J or the freq. of red light. Calculate other jumps

$E_4 \to E_2 = 4.08 \cdot 10^{-19}$ - green

$E_5 \to E_2 = 4{,}68 \cdot 10^{-19}$ - blue

$E_6 \to E_2 = 4{,}84 \cdot 10^{-18}$ - purple

Yes these numbers corrispond

$E_2 \rightarrow E_1 = 1,64 \cdot 10^{-18}$

$\rightarrow$ ultraviolet

Reading Guide for Bohr's Model p.104-106. PT

1. How are visible light and the electromagnetic spectrum related?

Visible light is a small part of the electromagnetic spectrum

2.List some forms of electromagnetic radiation.

radio waves, microrays, infared, ultra violet, x rays

3.Light is given off(emitted) by atoms when the electron level changes.

4.Which part of the electromagnetic spectrum do you hear?

You could recieve radio waves with special equip. + listen to that

5. What is frequency?

the numbar of waves per second or cycles per second
or hertz (H2)

6. What are the units of frequency?

H2 - hertz =1 cycle = $1 s^{-1}$

$f = \frac{c}{\lambda} \quad \frac{\cancel{m}/s}{\cancel{m}} \quad s^{-1}$

7.What does 1 s^{-1} mean?

an inverse second

8.What does the Greek letter lambda stand for?

wavelenght (the distance measured from crest to crest of one complete wave or cycle

9. All light travels at ~~the~~ Constan speed. This speed is:

The speed of light (c)

$2,998 \times 10^8$ m/s

10. What is the equation for the frequency of light?

$f = \frac{c}{\lambda} \leftarrow$ wavelenght

11. What does "c" represent?

speed of light

$2{,}998 \times 10^8$ m/s

12.Look at the example of the calculation on p. 105. What happens to powers of 10 when numbers are divided?

the exponents are subtracted

13. Look at the sample calculation on p. 105 (E=hf) What happens to powers of 10 when numbers are multiplied?

they are added

14.What is the value of "h", Planck's constant?

$6{,}63 \cdot 10^{-34}$ Js

15. How is energy of the electron calculated? (give equation)

$E = hf$

$\nwarrow$ freq.

16.Red light has $<$ energy than blue light.

17.What type of model did Bohr propose?

planetary model

18. What revolutionary idea did Bohr propose?

electrons travel in nearly circular orbits around the nuclus

19.What happens when electrons absorb energy?

They "jump" from a a lower energy level to a higher energy level

20.When are electrons most stable?

The one closer to the center

Now begin Activity 5. When you complete activity 5, do Chem to go p.107-8

Chem to go p107

1, What color in visable spectrum of H has > energy? < energy?

purple has the > energy > Freq

red has the < energy < Freq

2, How is $E_3 \to E_1$ fall compared to $E_3 \to E_2$ fall?

The larger fall will give off more energy in the fall..

3, What is the energy losss from $E_3 \to E_1$? How many times is the loss from $E_3 \to E_2$?

$E_3 \to E_1 = 1{,}94 \cdot 10^{-18}$

$E_3 \to E_2 = 3{,}03 \cdot 10^{-19}$ } 6,40 greater

4, $\lambda = 389.0$ nm $f = 7.707 \cdot 10^{14}$ Hz. Calculate this freq.

$f = \frac{2{,}998 \times 10^{8} \text{ m/s}}{389.0 \cdot 10^{-9} \text{ m}}$ ← -9 to convert nm → m

$f = 7{,}71 \cdot 10^{14}$

→

5. $\lambda = 389,0 \text{ nm}$ $E = 5,11 \cdot 10^{-19}$ Find E

$$\underrightarrow{E =} 7,71 \cdot 10^{14} \cdot 6,63 \cdot 10^{-34}$$

previous
problem

$$E = 5,11 \cdot 10^{-19}$$

6, Microwave $\lambda = 10 \text{cm}$

$$F = \frac{2,998 \cdot 10^{8} \text{ m/s}}{.1 \quad \text{m}} \qquad F = 2,998 \cdot 10^{9}$$

$$E = 2,998 \cdot 10^{9} \cdot 6,63 \cdot 10^{-34}$$

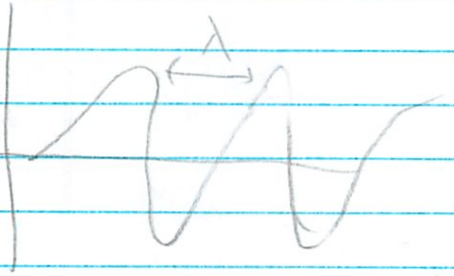
$$E = 1,99 \times 10^{-24}$$

6. How many x greater is Energy of red light (3.02×10^{-19})

~ 151759 times

Wave lenght

4/10



λ = wave lenght is nm
↳ $1 \cdot 10^{-9}$ m

f (ν) vibatrions per seconc

$f = \frac{c}{\lambda}$

c = speed of light
$2,998 \cdot 10^{8}$ m/s

$f = \frac{2,998 \cdot 10^{8} \text{ m/s}}{410 \cdot 10^{-9} \text{ m}}$ dividing # in exponent form → − exponents

$f = 1,78 \cdot 10^{8-(9)}$

$f = .00178 \cdot 10^{14}$

energy
↓
E = hf
↑
$6,63 \cdot 10^{-34}$ Js (f)
↑
Planks constant

↑ wave lenght = ↓ freq. = ↓ E

↓ " " = ↑ freq = ↑ E

purple ← red
↑E ↓E

Michael Plasmeier

20/20

Investigation of Rutherford's Model

Goal: To compare the results of a dart experiment to the prediction of the electron location around the nucleus of the atom

Procedure:

1. Obtain a target, tape and newspaper. Tape the target to the paper and place it on the floor. Standing straight, extend your arm over the target so that your arm is about 2 meters above the target. Drop the dart (pencil) to try to hit the bullseye. Repeat 100 times, dividing equally among the members of your lab group.
2. Count the number of hits in each circle and record in a data table.

Data Table:

Numbered Ring	Distance of Ring from Bullseye	# of Hits/ring

Calculations and Questions:

1. Plot the number of hits in each ring against the average distance from the bullseye. Draw a smooth curve through the plotted points.
2. What does this curve represent?
3. What is the probability that a dart hit would be found in Ring 4? Ring 8? Ring 1?
4. From your graph what ring has the most probability of being hit?
5. State in words what the graph tells us about the likelihood of locating an electron in the atom.
6. Would you expect the radius of maximum probability for a dart hit to be the same for each student?
7. Would you expect the most probable radius of the location of an electron in Li to be the same as that of an electron for hydrogen? Explain.
8. Is there a directional effect on your target? Divide your target into quadrants and count the hits in each quadrant for each ring.

Reflection:

This activity was mostly pointless. We did Battleship already and that even was unnessary. I know what the probability of 1/100,000 means. This probably confused people more.

Rutherford's Model of the Atom

PRELAB

1. How did Rutherford describe the atom? small center

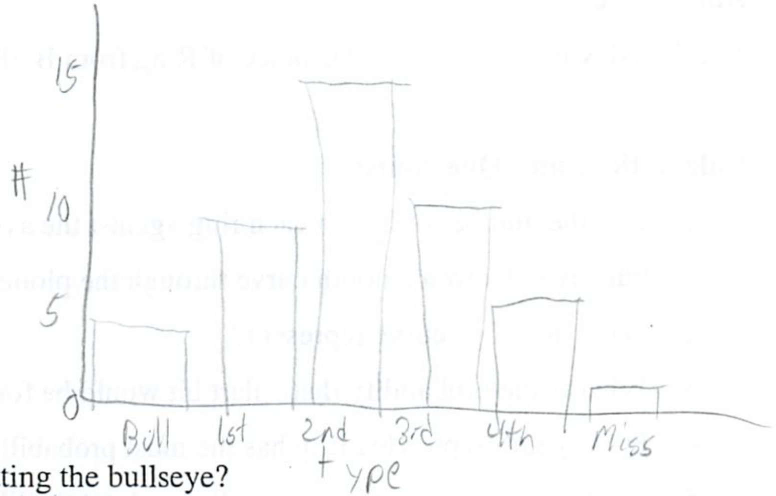
   A nuclus + proton at center

2. How did Rutherford come to these conclusions?

   With his alpha particles through gold experment

3. Draw a probability curve given the following data for a dart game. Plot the type of hit along the X-axis and the # of hits along the why axis.

Type of Hit	#
Bullseye	4
1st ring	10
2nd ring	15
3rd ring	11
4th ring	8
miss	2
Total Hits	50



4. What is the probability of hitting the bullseye?

   4/50 = 2/25

5. What is the probability of hitting ring 3?

   11/50

6. Which ring has the greatest probability of being hit?

   2nd ring

7. Why do we use probability to describe the position of an electron?

   Electrons are located in atom rings. they are very small compored to the rest of the empty space in the atom. Small probability of hitting them

Bullseye Questions

4/11

2. Why is the graph curved?
The graph is curved to show that larger areas are eaiser to hit. Also areas nearer to the target are hit because we were aiming for our goal.

3. Probability of Dart Hitting?
Ring 1 = 0/100 = 0
Ring 4 = 18/100 = 9/50
Ring 8 = 8/100 = 2/25

4. Which ring most likly to hit?
The fourth ring was most likly to be hit.

5. How hard is it to hit an atom?
It is very hard to hit an atom and very improbable with only a few shots

6. Would the most likly ring be = for all students?
No, depending on how they carried at the experiment things will differ, However the average should be close to many people.

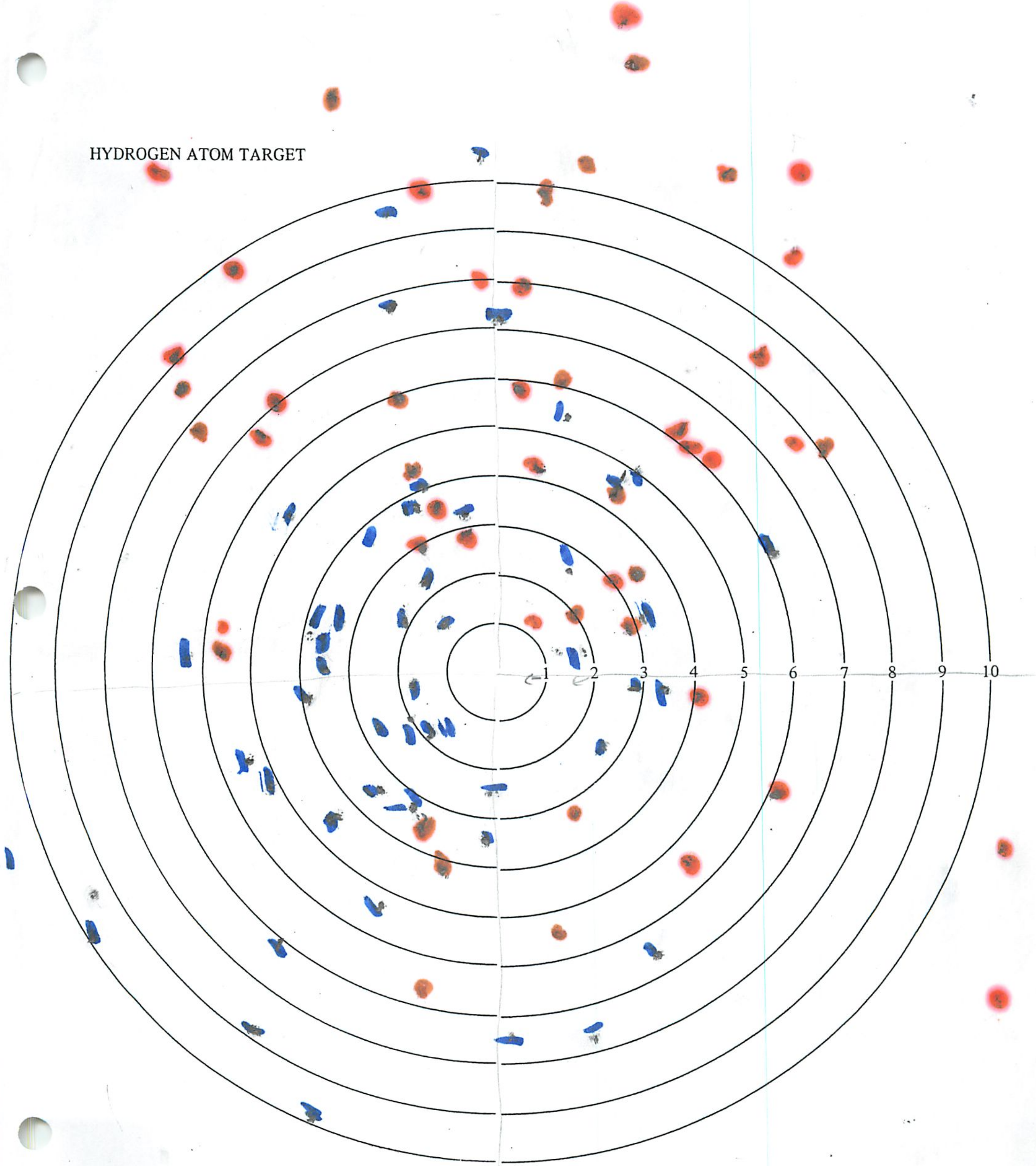
7. Would most probable ring be same in Li + H?
It would depend where the electron is. Each is the same size so individually they are equally likly to be hit, but Li has more electrons > the chanch of hitting one.

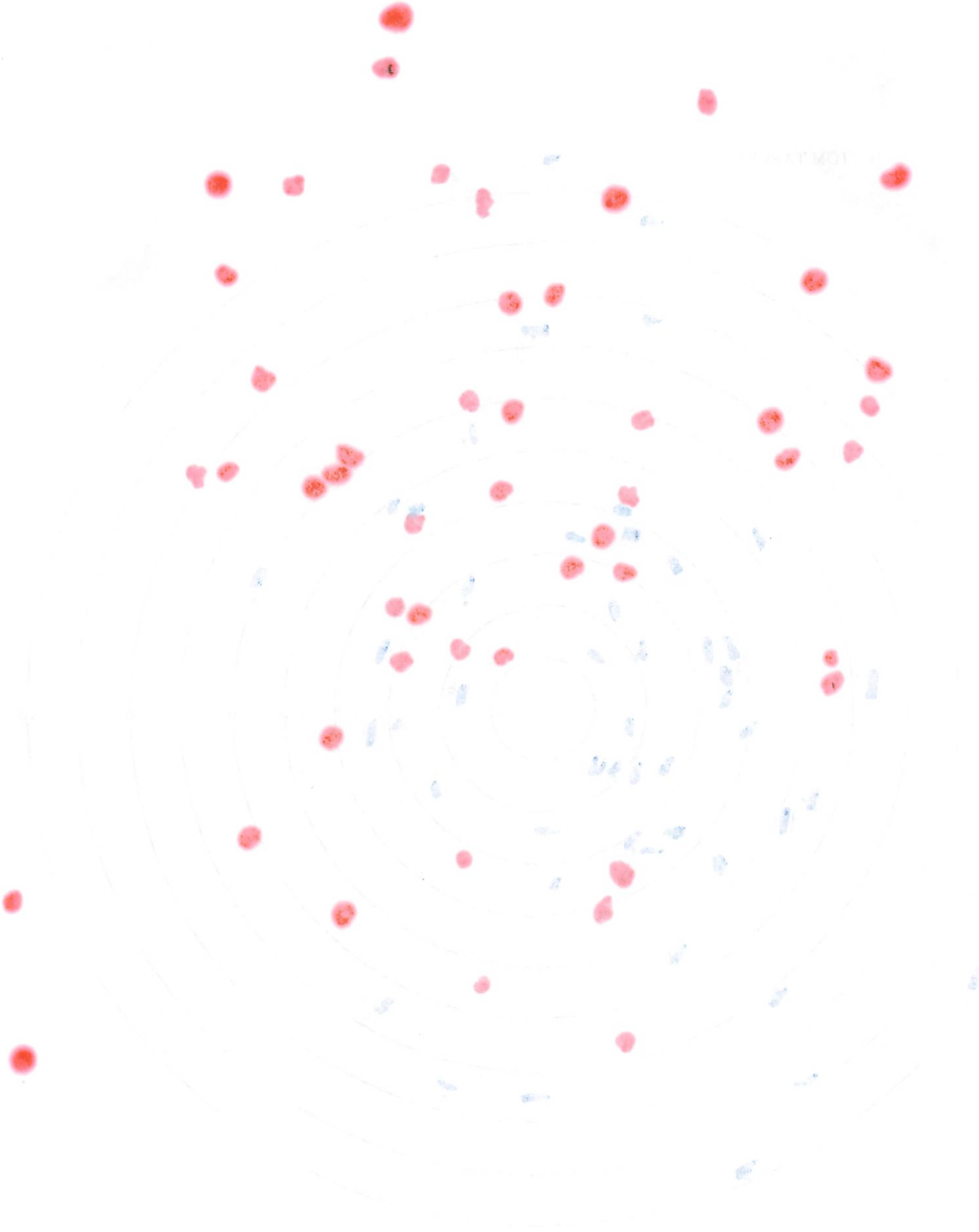
8. Is there a probability related to the four quadrants

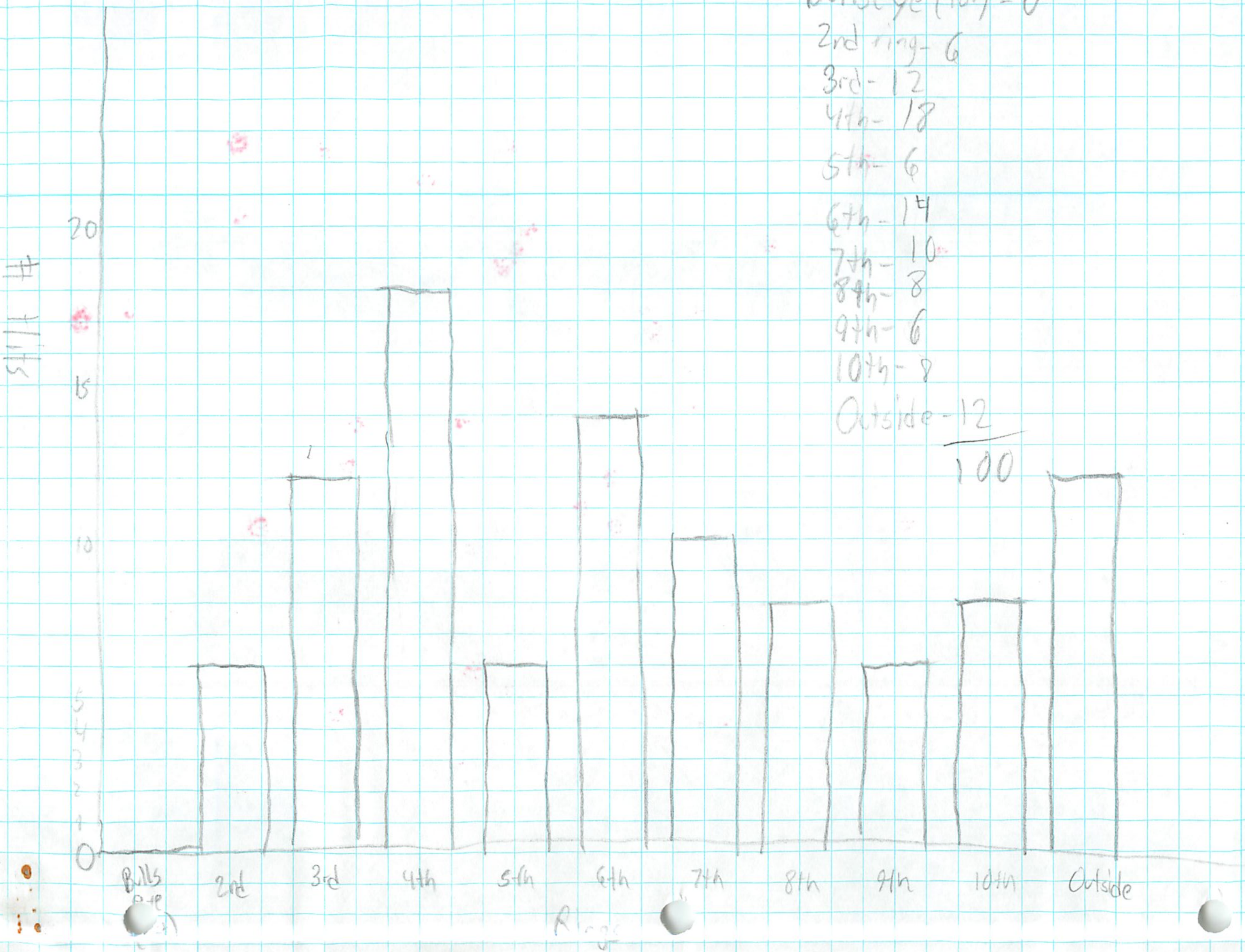
34	31
22	13

I guess there seems to be even though we were standing on opposit sides of the board.

HYDROGEN ATOM TARGET







Bullseye (1st) - 0
2nd ring - 6
3rd - 12
4th - 18
5th - 6
6th - 14
7th - 10
8th - 8
9th - 6
10th - 8
Outside - 12

100

Michael Plasmeier

Atoms with > 1 electron

Activity 6

20/20

4/12

to view spectra and graph energy required to remove electrons from diff. atoms to learn about the orginization of the periodic table.

Q How would >1 electron impact atoms' spectrum?

?? make it harder to remove atoms

(jumping)

Q What mods to Bohr's model would be needed to take into account >1 electron.

? They would jump like the same but would they be affected by others of those particles

1. Look at other gasses in a spectroscope. How are the colors different?

2. Bohr's law doesn't work w/ >1 atom though the basic idea of changing energy levels is still true. When ↑ electrons present some are easier to remove.

2a Make a graph of the ionizing energy for the first two ionizing energies. Include a title and a ledgend.

3. What patterns do you see. How would you describe it?
b. What are the $>$ and $<$ ionization energies?
c. What happens to ionization energies as atomic # ↑?
d. Group atomic energies into consecutive "periods"

4. How are lines 1+2 different?

+4.

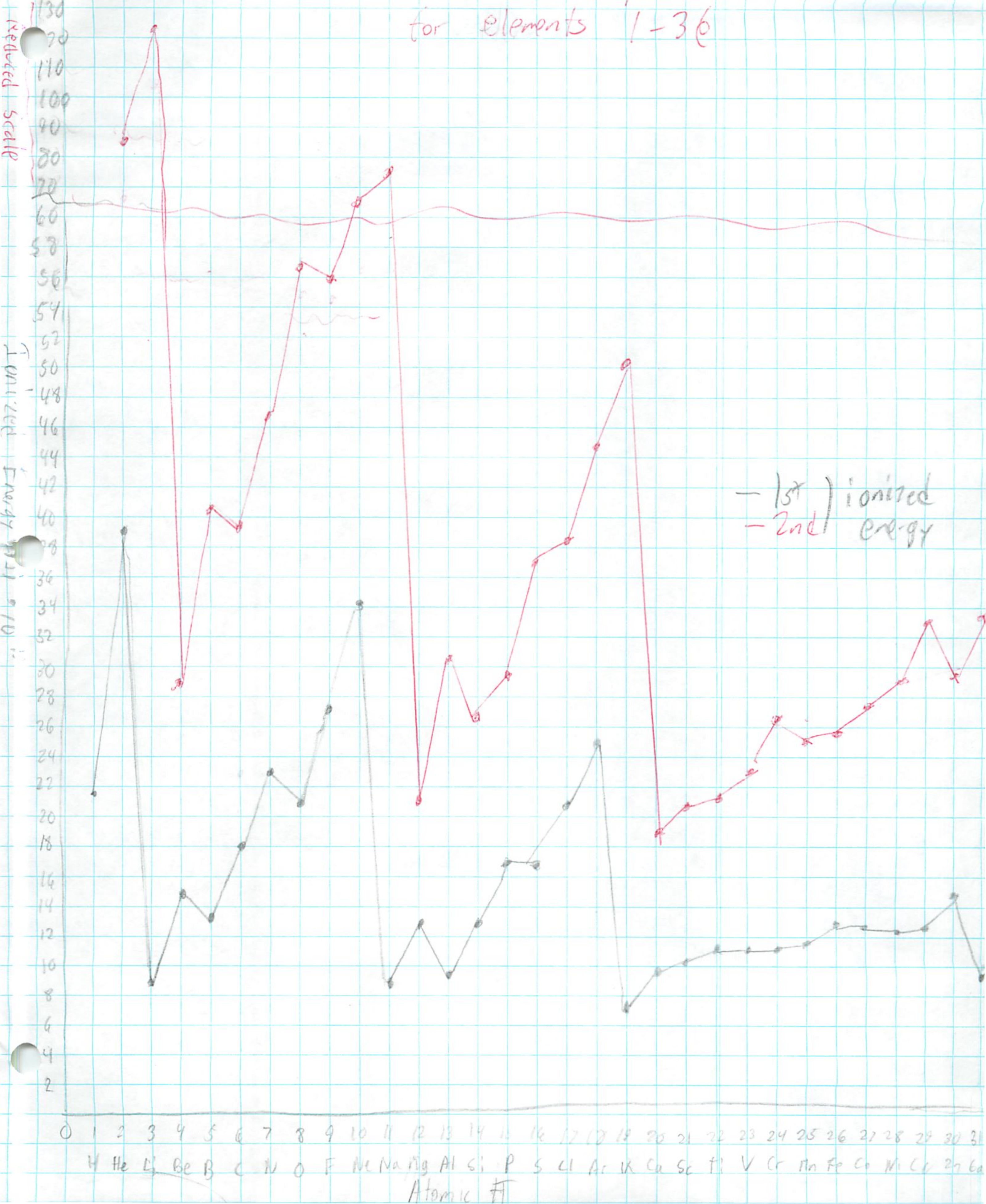
3. The lines points seem to increase mostly steadly until a certain point where it plummets and starts the climb over. Peaks in the 1st ionized energy are Li, Ne, Ar, Zn, Kr. Vallies are one place afterwords. Looking on the periodic table the peaks are the elements near the ends of the rows. The lows in the 1st ionizing energy happen to be the start of the rows in the periodic table. The drops from the 2nd ionizing energy are greater and seem to be shifted one electron to the left. In fact H, with one electron, dosn't have a second ionizing energy. Energies generally increase as they go accross the row but come back to a common low in group 1 (start of row)

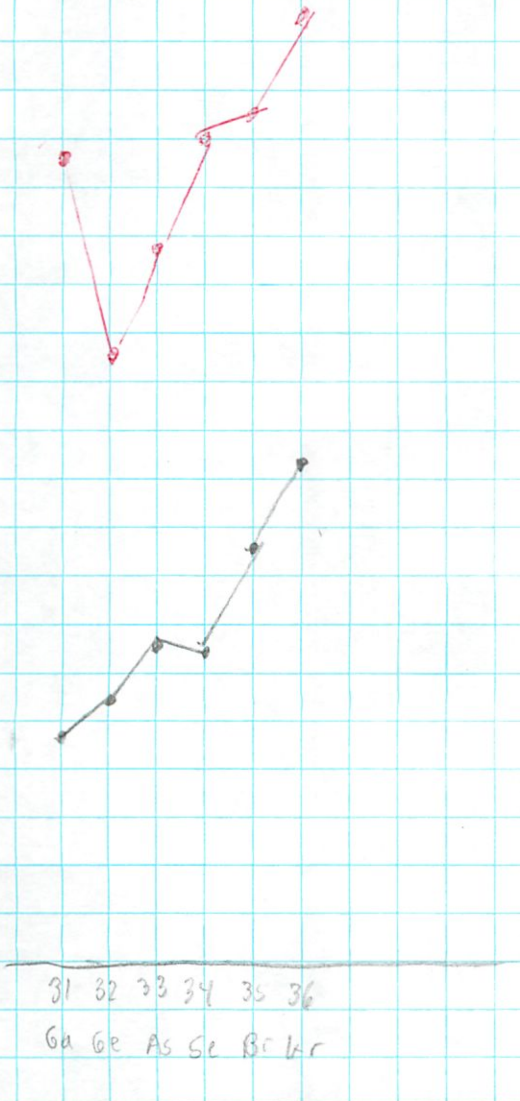
After demo explaining table

This is all related to the orbital patterns described in class. This is how the table is arranged and why jumps are like this. The different orbital patterns affect how hard it is to remove an electron (volatility) and the absence of a complete group makes the substance magnetigly charged

n isp

1st + 2nd Ionization Energy for elements 1-36




31 32 33 34 35 36
Ga Ge As Se Br Kr

5. If a large amount of energy is needed to remove an electron from an atom, the atom's considered stable. Thus high 1st ionization energy is needed to remove electron.

[higest ↓peaks]

a. Which element in each period has most stable arrangment?
The element in the last group (last in the period, right of the table) [He, Ne, Ar, Kr]

6. Bohr can't describe when 71 electron, So we use sublevels s, p, d, and f. # of electrons = atomic #. The sublevels show how electrons are arranged.

a. Which sublevel are electrons for H and He?
1s

- As you move to the 2nd row, more electrons start filling more sublevels. Each sublevel must be filled before electrons start filling the next sublevel.

b-e. Which region starts adding electrons to the ↓ level.
s, p, d, f
When is it filled?
The s level is the first row. (period)
The p level starts with 5 - after the gap in group 15, and goes to 10, the end of the row
The d level starts the transition metals in period 4 and last for the end of the period
The f level starts at 57. (group 3, period 6) and lasts for the entire period.

→

f. Select a group and see in which sublevel the last electron was added. Is it the same for the entire gro

Group 17 always is one short of a filled p sublevel. The number of p groups there are increases as you move downward

g. How does the # and location of electrons in the outermost sublevel relate to chemical properties?

Elements which are in the same group are related. They share the same chemical and physical properties. They also have the same location and number of electrons in that location as atoms in the same groups

Reading Guide for Chem Talk-A Periodic Table Revealed p 114-115

1.What does the ionization energy represent?

the energy req. to remove an electron from a gaseas atom at ground state

2.All atoms are neutral. This means that the number of protons and electrons are equal.

¿except if ?

3. Atoms that have gained or lost electrons are called ions.

4. What is the “first ionization energy” ?

energy req. to remove a single electron from the highest occupied energy level.

5. How are the first and second ionization energies different?

the 2nd one is to remove a 2nd atom after the first has already been removed

6. What is meant by the term “electron configuration” ?

distribution of electrons in an atom's energy level

7.Energy levels are broken down into sublevels.

8. If all the sublevels are filled, the atom is ? stable.

9. What do the letters s, p, d, and f represent?

sub levels of wave patterns of electrons

10. What so these letters stand for?

s - sharp   d diffuse

p - principal   f - fundamental

11.What type of orbital is found when E=1?

1st   1s

12. What does the "1" identify and what does the "s" identify?

1 - energy level

s - type orbital

; # of

13.Which types of orbitals are found in the second energy level?

2nd   2s2p

14. How many and what kind of orbitals are in the third and fourth energy levels?

3rd   3s3p3d | 4th   4s 4f 4d4f

15.The maximum number of electrons in any s orbital is 2, in any p orbital

6.

16. What does $3s^2$ mean?

3rd level of s orbital filled (2 electrons)

?

?

17. What does $4p^4$ mean?

4th level w/ 4 electrons in the p

18. What does the term "excited" indicate?

Electrons that have been promoted to a higher energy level + then falls back

19. Particles arranged in an unstable way will.......

give off electrons to become more stable

20. What happens when elements are arranged by atomic number?

patterns emerge

21. The horizontal rows are called periods.

Chem to Go

1. Write seq. of energy levels from 1s to 4f

$1s^1$
$1s^2$
$1s^2 2s$
$1s^2 2s^2$
$1s^2 2s^2 2p^1$
$1s^2 2s^2 2p^2$
$1s^2 2s^2 2p^3$
$1s^2 2s^2 2p^4$
$1s^2 2s^2 2p^5$
$1s^2 2s^2 2p^6$
[Ne]

[Ne] $3s^1$
[Ne] $3s^2$
[Ne] $3s^2 3p^1$
[Ne] $3s^2 3p^2$
[Ne] $3s^2 3p^3$
[Ne] $3s^2 3p^4$
[Ne] $3s^2 3p^5$
[Ne] $3s^2 3p^6$
[Ar]

[Ar] $4s^1$
[Ar] $4s^2$
[Ar] $4s^2 3d^1$
[Ar] $4s 3d^2$
[Ar] $4s^2 3d^3$
[Ar] $4s^2 3d^4$
[Ar] $4s^2 3d5$
[Ar] $4s^2 3d^6$
[Ar] $4s^2 3d^7$

[Ar] $4s^2 3d^8$
[Ar] $4s^2 3d^9$
[Ar] $4s^2 3d^{10}$
[Ar] $4s^2 3d^{10} 4p^1$
[Ar] $4s^2 3d^{10} 4p^2$
[Ar] $4s^2 3d^{10} 4p^3$
[Ar] $4s^2 3d^{10} 4p^4$
[Ar] $4s^2 3d^{10} 4p^5$
[Ar] $4s^2 3d^{10} 4p^6$
[Kr]

Actually just

1s
2s
2p
3s
3p
4s
3d
4p
5s
4d
5p
6s
4f

22. The vertical columns are called groups.

23. Elements in chemical groups share........ similar electron configs + similar physical + chemical properties

24. They form similar kinds of compounds.

25. This behavior is due to......... elements in one chemical group have the same # of electrons in outer energy levels + tend to form ions by gaining or losing the same # of electrons

do chef TO go P.116-117

← #1

#2 What is Boron (B)'s atomic #? How many electrons? What is the complete sequence?

Boron = #5 atomic # = 5 electrons - $1s^2 2s^2 2p^1$

#3 What is zinc (Zn)'s atomic #? How many electrons? Electron sequence? Last sublevel? Where on table? What other elements have similar properties

Zn = #30 = 30 electrons = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

last sublevel = $3d^{10}$ Others in IIB / 2B (Group 12) like #48 - Cadmium (Cd) #80 - Mercury (Hg) or other transition metals

#4 What is calcium (Ca)'s atomic #? How many electrons? Complete sequence? Last subleve? Where on table? What other elements are similar?

Chem to Go
cont

4/16

Ca - #20 atomic - 20 electrons - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
last sublevel = $4s^2$ period 4, group 2
others in group 2 = #4 Be, #12 Magniseum
#38 Sr, #56 Ba - Alkaline Earth Metals

5. Element: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^8$ How many electrons? Atomic #? What can you predict? #110, 110 electrons Uun Ununnilium Part of Group 10, transition metal, similar to #28 - Ni, #46 - Pd, #78 Pt

6. What is this element? $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
#26 Iron (Fe)

spinning diff ways
poles cancle out
not magnetic

energy level

		1s	2s	2p	3s	3p	
+1 -1	H	↑					1 ← # electrons 1s ← orbital
+2 -2	He	↑↓					$1s^2$
+3 -3	Li	↑↓	↑				$1s^2 2s^1$
+4 -4	Be	↑↓	↑↓				$1s^2 2s^2$
+5 -5	B	↑↓	↑↓	↑ | |			$1s^2 2s^2 2p^1$
+6 -6	C	↑↓	↑↓	↑ | ↑ |			$1s^2 2s^2 2p^2$
+7 -7	N	↑↓	↑↓	↑ | ↑ | ↑			$1s^2 2s^2 2p^3$
+8 -8	O	↑↓	↑↓	↑↓ | ↑ | ↑			$1s^2 2s^2 2p^4$
+9 -9	F	↑↓	↑↓	↑↓ | ↑↓ | ↑			$1s^2 2s^2 2p^5$
+10 -10	Ne	↑↓	↑↓	↑↓ | ↑↓ | ↑↓			$1s^2 2s^2 2p^6$ very ← hard to remove electrons
+11 -11	Na	↑↓	↑↓	↑↓ | ↑↓ | ↑↓	↑		← magnetic
+12 -12	Mg	↑↓	↑↓	↑↓ | ↑↓ | ↑↓	↑↓		
+13 -13	Al	[NE] ← placeholder			↑↓	↑ | |	

etc

Michael Plasmeier

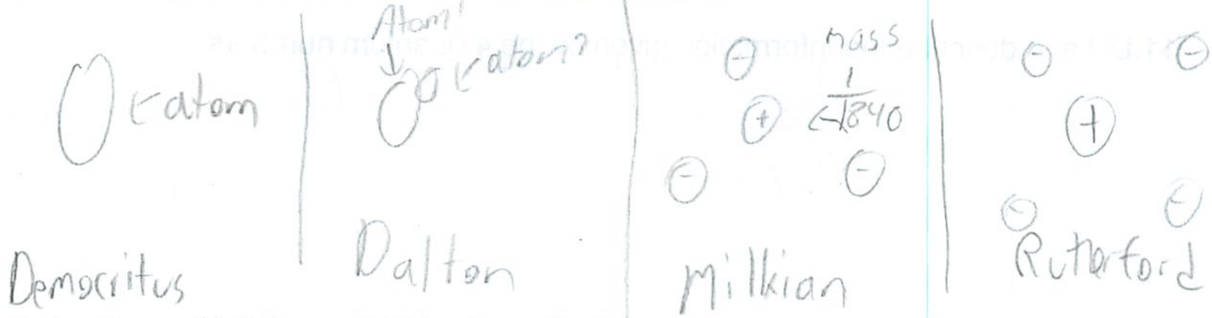
Reading Guide for Ch.4,5,6 Electrons in Atoms

1.Which idea was changed by the discovery of sub atomic particles?

Democrituss idea that the atom was indestructible

2. Sketch the various models of the atom and label all parts.

see table red book p128+129



3.What structures did Bohr model his atom after?

ladders of energy levels

4. How did he propose that the electrons did not fall into the nucleus?

they

5.What is an energy level? The information given by the energy level is:

fixed levels in space where electrons can be

electrons in certain energy levels define it

6.For an electron to change energy levels the quantum amount of energy must be gained or lost.

7.What is a"quantum"?

abrupt change

8.Why is a ladder a poor analogy for energy levels?

because the rungs in a ladder are equally spaces - really higher energy levels (farther away from nucleus)

9.What information came from solutions to Schrodinger's equation.

the quantum mechanical model - determines the allowed energies an electron can have and how likely it is to find the electron in various locations around the nucleus

10.The quantum mechanical model of the atom describes the probability of finding an electron in space.

11.List and describe the information given by the 4 quantum numbers.

each sublevel corresponds to an orbital of a diffrent shape, which describes where the electron is likely to be found

12.What are the shapes of the orbitals?

s - spherical
p - dumbbell shape
d - clover leaf shapes
f - complicated

13.The maximum number of electrons in any energy level is found by :

$2n^2$

1 - 2
2 - 8
3 - 18
4 - 32

14.Rules for electron configurations:

1. number of electrons in the atom

2. order of filling of orbitals and sub orbitals

3. size and shape of orbital

4.lowest energy level first (Aufbau)

5.2 electrons per box(Pauli)

6.Fill each sub orbital with one electron, then go back and double (Hund's)

15. Write the electron configurations for the first 22 elements.

H - $1s^1 2s^0$
He - $1s^2 2s^2$
Li - $1s^2 2s^2 2p^1$
Be - $1s^2 2s^2 2p^2$

B - $1s^2 2s^2 2p^3$
C - $1s^2 2s^2 2p^4$
N - $1s^2 2s^2 2p^5$
O - $1s^2 2s^2 2p^6$

F - $1s^2 2s^2 2p^6 3s^1$
Ne - $1s^2 2s^2 2p^6 3s^2$
Na - $1s^2 2s^2 2p^6 3s^2 3p^1$
Mg - $1s^2 2s^2 2p^6 3s^2 3p^2$

P - $1s^2 2s^2 2p^6 3s^1 3p^3$
S - $1s^2 2s^2 2p^6 3s^2 3p^4$
Cl - $1s^2 2s^2 2p^6 3s^2 3p^5$
Ar - $1s^2 2s^2 2p^6 3s^2 3p^6$

K - $1s^2 2s^2 2p^6 3s^1$ $3p^3 4s$
Ca - $1s^2 2s^2 2p^6$ $3s^1 3p^3 4s^2$
Sc - $1s^2 2s^2 2p^6 3s^1 3p^3 4s^2 3d^1$
Ti - $1s^2 2s^2 2p^6 3s^1 3p^3 4s^2 3d$

16. Describe the experiments which support the quantum mechanical model

Rutherford - gold model - discovered electrons
Bohr - proposed quantum levels
Schrödinger - math equation to determi levels

17.Define atomic number

# of protons

(which usually = electrons)

memorize

18. Define atomic mass

# of protons + neutrons

19. In your own words, explain how to determine the number of neutrons if you know the atomic mass and the atomic number.

atomic mass - atomic #

If an element has the same atomic number as another element , but has a different , atomic

mass the elements are isotopes /

p123 red

Do p.123 problems 61, 63, 65 p.149 probs. 22-29, 30-39

61. Whats wrong with Dalton's model today?
Atoms are no longer viewed as indivisible. We now know that atoms are made of protons, electrons, and neutrons. There is space between electrons

63. How is # of electrons related to atomic # in neutral atom?
They are the same.

→

65. Calculate approx. atomic mass of lead (Pb)

$\frac{\text{mass}}{\text{atomic\#}}$ $^{204}_{82}Pb$ • 1.37% | $^{206}_{82}Pb$ • 26.25% | $^{207}_{82}Pb$ • 20.82% | $^{208}_{82}Pb$ • 51.55%

2,7948
54,075
43,0974
+ 107,224
207,1912 atomic mass

p149 red

22. Whats wrong w/ Rutherford's model? What sub atomic particles in Thomson's plum-pudding model?
Rutherford did not talk about electron energy levels.
Thomson had electrons as sub-atomic particles

hard to ask

23. What did Bohr assume about electrons?
Bohr found that electrons rotate around te nucelus at fixed quantum energy levels.

24. Compare Rutherford + Bohr's model of te atom.
Bohr built on Rutherford's model. He proposed that an electron is only found in specific, circular paths, or orbits, around te nucels.

25. What is Te significance of te boundary of an electron cloud?
Te electron cloud is a visual model of te probable location of electrons in an atom. Te probability of finding an electron is higher in te denser regions of a cloud.

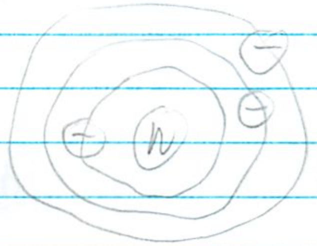
26. What is an atomic orbital?
An atomic orbital is often thought as a region of space in which tere is a high probability of finding an electron.
Te orbitals are in certain shapes s, p, d, f

27. How many orbitals are tere in te 2p sublevel?
Te 2p sublevel consists of 3 p orbitals of equal energy (6 electrons)

The Electron

4/13

Electrons are located outside the nucleus of the atom in energy levels. Energy levels are distinct regions in space at some distance from the nucleus. (Activity 5)

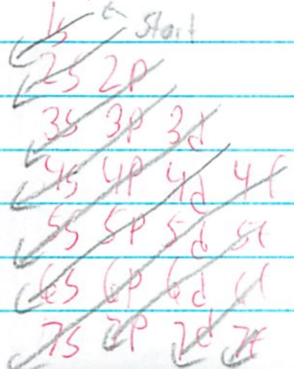


Within each of these energy levels are cloud-like regions of probabilty called orbitals. These orbitals are defined around axis and in planes as sub-orbitals.

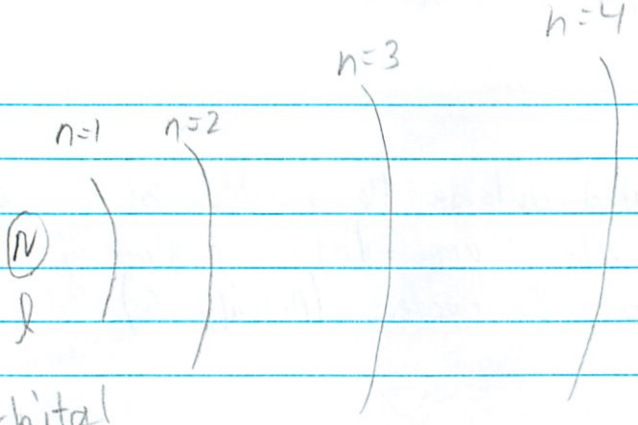
Evidence
flame tests
emission spectra
photo electric effect

Rules for Electron Config
1. know atomic # (# electrons)
2. know order of filling
3. know size of orbital (s=1, p=3, d=5, f=7)
4. Each orbital has 1 electron, Then start doubling up
5. Max two ⊖ in each suborbital
6: Spins are alternated. (per box)

Poly's exclusion principle
?sp



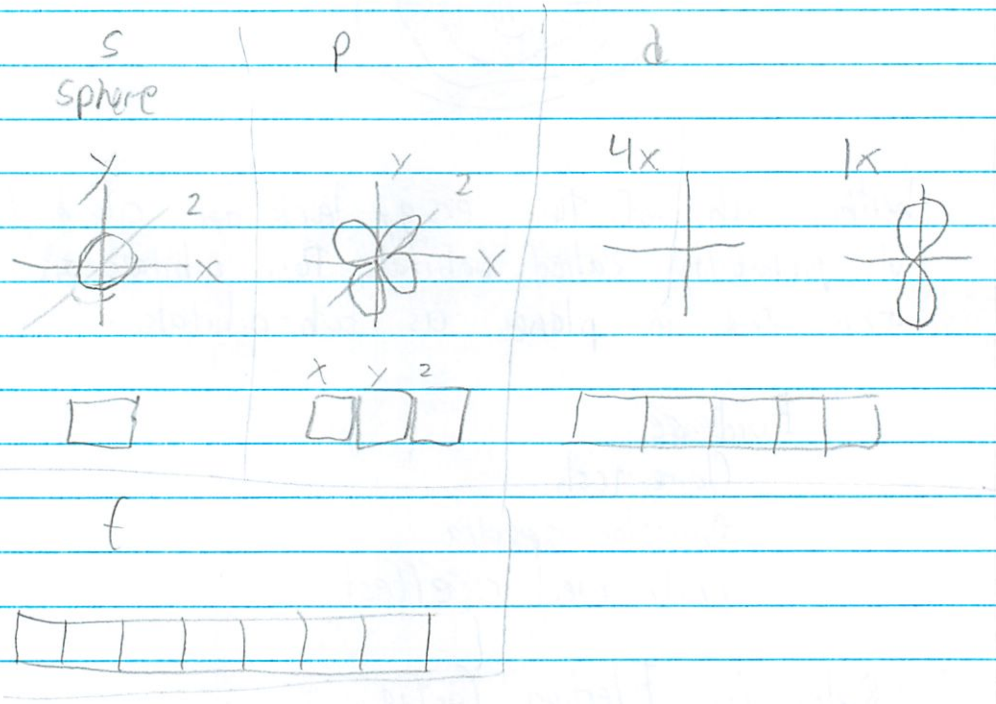
→



Quantam numbers
(really letters)

n= how far from nucleus.

Orbital



Orentation around axis

each box has max 2 electrons

l_m = suborbital

ms = spin

28. Sketch 1s, 2s, and 2p sublevels w/ same scale.


1s


2s


2p

29. How many sublevels in these principal energy levels

n=1 = 1 (1s)
n=2 = 2 (2s 2p)
n=3 = 3 (3s 3p 3d)
n=4 = 4 (4s 4p 4d 4f)

p149, red 30. How many electrons in the highest occupied energy levels?

a. Barium - $6s^2$ - 2
b. sodium - $3s^1$ - 1
c. aluminum - $3p^1$ - 1
d. oxygen - $2p^4$ - 4

31. What are rules which govern filling of atomic orbitals by electrons?

1. Aufbau principle - fill lowest 1st
2. Pauli exclusion principle - can by two electrons per orbital (box)
3. Hund's rule - fill boxes with 1 electron before adding another

32. Write electrons config.

a 15 - P - $1s^2 2s^2 2p^6 3s^2 3p^3$
b 12 - Mg - $1s^2 2s^2 2p^6 3s^2$
c 9 - F - $1s^2 2s^2 2p^5$
d 18 - Ar - $1s^2 2s^2 2p^6 3s^2 3p^6$

33. What is meant by $3p^3$

Its the 3rd energy level's p orbital half full.
(One electron in each orbital)

34. Give electron's config.

a- Na : $1s^2 2s^2 2p^6 3s^1$
b - S - $1s^2 2s^2 3p^6 3s^2 3p^4$
c - Mg - $1s^2 2s^2 3p^6 3s^2$
d - Ne - $1s^2 2s^2 3p^6$
e - K - $1s^2 2s^2 3p^6 3s^2 3p^6 4s^1$

35. What of tese orbital designaties are invalid

a 4s
b 3f ← 3f can't be done because f doesn't start
c 2d until the 4th energy levels
d 3d

36. What max # of electrons in each sublevel

a. 2s - 2
b. 3p - 6
c. 4s - 2
d. 3d - 10
e 4p - 6
f 5s - 2
g 4f - 14
h 5p - 6

37. Arrange sublevels in order of ↑ energy

3d, 2s, 4s, 3p → 2s, 3p, 4s, 3d

38. How many electrons in 2nd energy level

a. chlorine - 8 ↑ 2s + 2p
b. phosphorus - 8
c. potassium - 8

39. Write electron config?

a. selenium - $1s^2 2s^2 3p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$
b. vanadium - $1s^2 2s^2 3p^6 3s^2 3p^6 4s^2 3d^3$
c. nickel - $1s^2 2s^2 3p^6 3s^2 3p^6 4s^2 3d^8$
d. calcium - $1s^2 2s^2 3p^6 3s^2 3p^6 4s^2$

Michael Plasmeier

Atoms Interacting with each Other Lab
Activity 8

20/20

4/19

to relate patterns in ionization energies and valence electrons to study the arrangment of the periodic table and convalent / ionic bonds

The chem behavior is determined by the arrangment of valence electrons. How do they interact in NaCl to create a bond?
Some of the valence electrons are shared through either covalant or ionic bonds. They try to become noble gasses, each with 8 valence electrons they are connected to.

1. In activity 3, John Dalton thought that elements combined with one atom for each electron. In 6, you saw that atom's behavior was basel on excess or dificiency of electrons relative to closest noble gas. a) Which atoms have the lowest 1st ionization energy?

   Li, Na, K, Rb, Cs - all in group 1

   b) What do you observe of these atom's 2nd ionization energy?

   They have the higest 2nd ionization energy.

   c) Why?

   Because after you have removed 1 electron, its isoelectric (like a noble gas). It takes a lot of energy to break apart a noble gas because they are so tightly bonded.

   d) What atoms' have the lowest 2nd ionization energy?

   Elements in the 2nd group like Be, Mg, Ca, Sr, Ba have the lowest 2nd ionization energy.

   e) Why?

   Elements try to become like noble gasses. Elements in group 2 need to lose 2 electrons to become stable noble gasse chem for Compounds

2. You can predict formulas if you know the position of electrons (esp. valence electrons) Na has one valence electron in 3s, When it loses an atom it becomes sodium ion like Ne. What is the electric charge of the ion?
-1

a). Consider chlorine. How many valence electrons?
7

b) How many electrons to gain to be like Ar?
+1

c) When Cl gets +1 electron to become chlorine ion, what is the charge?
+1

d) Each Cl can gain an electron. How is NaCl formed?
NaCl is formed when Na loses an electron to Cl.

Zinc Cloride

3. Consider reaction b/w Al + ↑ Cl. The Zn atoms have 2 valence electrons in the 4s level. In order to be like Ar, the Zn give up 2 electrons and becomes +2 charge.

a) Each Cl atom can accept 1 electron. How many Cl needed to take the 2 electrons from the Zn?
2

b) Write formula for Zinc Cloride.
$ZnCl_2$

4. In reaction b/w Al + $ZnCl_2$, Al replaces the Zn in zinc chloride, forming aluminum cloride + zinc.

a) How many valence electrons in Al? 3

b) How many electrons Δ to be like Ne? ~~+3 or -5~~

c) What are Al called after give up valence electrons? Charge
ions +3 ~~or +5~~ 3 protons w/o electrons

d) How many Cl atoms needed to accept Al electrons? ~~5~~ 3

e) How write formula $AlCl_3$

Michael Plosmski

4/19

Reading Guide "The Octet Rule"

READ p.128-9

1.Which two ways allow the formation of compounds?

ionic bond - attraction between oppositely charged ions

Covalent bond - when 2 atoms combine + share their paired electrons w/ each other

2.What is the octet rule?

atoms tend to gain or lose electrons during chemical reactions so that the atoms have an outer shell config of 8 electrons

3.Which elements are exceptions to this rule?

transition elements (groups 3-12)

4. Which elements are "halogens"? What happens when a halogen gains an electron? group 17

F, Cl, Br, I, At

when gains electron -1 charge

like noble gas

5.How are chemical names organized?

Metal → nonmetal

6.What does "binary" mean?

formed from combining of 2 diff chem elements

7.The names of binary compounds end in -ide.

except water + ammonia

8. Water and ammonia are compounds with common names.

binary compounds

9.How many valence electrons does Carbon have?

4

10.Atoms with a Small number of valence electrons give up electrons.

Atoms with a large number of valence electrons will gain electrons to have the same electron arrangement as the nearest noble gas. (Your answers will be gain, lose, few or many)

11.Atoms which form compounds by giving or taking electrons form ionic bonds.

12.Atoms like carbon share electrons to form convalent bonds. (ionic or covalent)

13.Why are electron dot diagrams used?

to indicate convalent bonds

14. How are valence electrons represented?

dots on the chart

15.Look at the examples on p.129. How many dots are around each atom other than Hydrogen?

8

How many dots are around Hydrogen? What is the nearest noble gas to Hydrogen?

2 - He

Do ~~Checking Up~~ and Chem to Go p.129-132.

1, Write chem formula

a) sodium oxide = Na(+1) O(-2) Na_2O

b) magnesium chloride - Mg(+2) Cl(-1) $MgCl_2$

c) aluminum oxide - Al(+3) O(-2) Al_2O_3

3, If you know sodium chloride is NaCl, how could know the formulas for potassium chloride, cesium bromide, lithium iodide and Sodium fluoride.

I know the formulas are similar because all are made up of a combonation of group 1 and group 17 on the Periodic table.

4, Calcium chloride = $CaCl_2$ - Can you write formulas for Group 2 + Group 17 combining?

Magnesium bromide - Mg(+2) Br(-1) $MgBr_2$

Strontium iodide - SrI_2

beryllium fluoride - BeF_2

barium chloride - $BaCl_2$

Calcium iodine - CaI_2

You can see that Group 2 (+2) needs 2 of group 17 (-1) to balance out into like a noble gas

5. Why are elements grouped together and called a family?
Write formulas: Elements try to combine (w/ electrons) to be stable like a noble gas

boron fluoride - B(+3) F(-1) → BF_3
aluminum bromide - $AlBr_3$
gallium iodide - GaI_3
indium(III) chloride - $InCl_3$
thallium (III) bromide - $TlBr_3$

6. How does Group 1, 2, 13 combine with Group 16?
potassium sulfide - K(+1) S(-2) → K_2S
lithium telluride - Li_2Te
cessium oxide - Cs_2O
rubidium selenide - Rb_2Se
Sodium sulfide - Na_2S

6b. How about Group 2 + Group 16?
magnesium sulfide - Mg(+2) S(-2) → MgS
barium oxide - BaO
calcium sulfide - CaS
strontium selenide - $SrSe$
beryllium telluride - $BeTe$

6c. Group 3 + Group 16?
boron sulfide - B(+3) S(-2) → B_2S_3 drop sign + switch
gallium telluride - Ga_2Te_3
thallium(III) sulfide - Tl_2S_3
aluminum selenide - Al_2Se_3
indium(III) oxide - In_2O_3

Ion Notes

4/18

⊕ protons
⊖ electron

all atoms are electrically neutral (thay have = # ⊕ and ⊖)

noble gasses have all electrons filled - don't combine well

when an electron is removed atom Δ → ion
(#⊕ ≠ #⊖)
ions are charged atoms (or group of atoms)
by forming ions, the element gets an electric charge
but when ion has the same config as a noble gas
= isoelectronic
-example Na + Cl are isoelectronic
example: Na^+Cl^- called ionic compound

Li | +1
F | +7(-1)

Li^-F^+

ionic bonds when electrons are transfered from metals to nonmetals

The magic # is 8
tries to be a noble gas

Mg | +2
O | +6(-2)

$Mg^{-2}O^{+2}$
↳ MgO

still ionic compound

Na | +1
O | +2(-2)

→ Na_2O
need two to = 8

Ca | +2
Cl | +7(-1)

→ $CaCl_2$ needs to Cl - one for each Ca

→

Mg | +2 Mg_3N_2
N | +5 (-3)

* write one closest to left 1st
(metals 1st - then non metals)

Simple way

Rb | +1 $\overset{+1}{Rb}_1 \overset{-1}{I}$ drop sign, write as oppset subscript
I | +7 (-1)

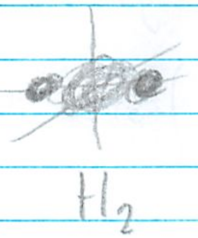
$\overset{+2}{Ca}\overset{-1}{Br} \rightarrow Ca_1Br_2$

Sharing + Bonding

H_2 → H• H• give electrons doesn't do anything

H••H ← can share - each has 2 electrons
covalent bond

← 2 s orbitals



← distinct combined s orbital

$F_2 \rightarrow F-F$

↑ indicate sharing
↗ acts like Ne

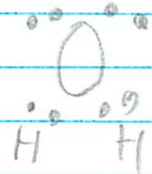
→

N_2 → N - N

O_2 → O - O

← unpored for complicated reasons

H_2O → H - H - O



← covalent bond

Michael Plasmeier

What Determins
+ Limits an Atoms
Mass - Lab 9

20/20

4/26

to investigate the composition of an atoms nucelus and learn about isotopes.

What makes up the nucelus of an atom where most of an atoms mass is located.

Protons + neutrons are in the nucelus of an atom

1, Atomic mass is the avg. mass of the atoms (isotopes) of each element. a) How many protons in H?

1 = atomic #

b) To nearest whole # - atomic mass of H.

1

c) How many protons in He?

2

d) What would you expect the atomic mass of He to be?

about 2 (# of protons)

e) What is it?

4 - because its the # of protons + neutrons.

2, Yes He's atomic mass is 4 × H b/c neutrons also count, B has atomic #5. Avg. mass = 10.811. Most B has 11 amu and some has 10. B/c atomic mass = # protons + # neutrons - B can have 5 or 6 protons.

a) How many protons do you expect to find in nucleus of H?

1 - b/c # protons = # electrons

b) How many neutrons in H?

0 - because none shown in mass

c) How many protons + neutrons?

atomic #

mass # - atomic #

			Protons	neutrons
Li - 3	3.941	Be -	4	5.012187
B - 5	5.811	C -	6	6.011
N - 7	7.00674	O -	8	7.9994
F - 9	9.998	Ne -	10	10.1791

3.a Atomic mass Mg + Cl? Na + F? Which closer to whole #?

Mg/Cl = 24,3050/35,4527

Na/F = 22,98977/18,998403 ← closer, but to 2x+1
(where x = atomic #)

b/ We can only have whole #s of protons + neutrons.
What should atomic masses be?
If #protons = #neutrons then atomic mass shold be 2x atomic #.

4. The fact that atomic masses are not close to whole #s is because the # neutrons is not the same in all atoms of a given element. Only the # protons (atomic #) is the same in all atoms in a given element. Atoms of same elements w/ diff # neutrons are isotopes (means same # protons).
Li has atomic # 3 and avg. atomic mass of 6,941.
Most Li has 4 neutrons, but some has 3 neutrons,
(Li-7 and Li-6 respectivly)

a) What isotopes do you think acount for these atomic masses?

- C (12,011) = C-6 + C-7
- H (1,00794) = H-0, H-1, H-2
- B (10,811) = B-5 + B-6
- Be (9,012) - Be-4, Be-5, Be-6
- Na (22,98) - Na-11, Na-12, Na-13
- Mg (24,30) - Mg-12, Mg-13

b/ $\frac{\text{mass}}{\text{atomic \#}}$ - How many neutrons + protons

	nev	pro		nev	pro
$^{3}_{2}He$	1	2	$^{4}_{2}He$	2	2
$^{6}_{3}Li$	3	3	$^{7}_{3}Li$	4	3
$^{12}_{6}C$	6	6	$^{13}_{6}C$	7	6
$^{14}_{7}N$	7	7	$^{15}_{7}N$	8	7

Part B

1. There are 2 different forces action on electrons, protons + neutrons. Cut 2 strips of tape ~12 cm long. Form tabs on each by bending a corner over. Place one sticky side down on table + label "B." Place other one sticky side " " top + " "A". peel off top strip w/ 1 hand and then the other. Hold the strips near each other. Do they attract or repel? What is the force.

The top one is attracted towards the bottom one. Is the force static electicty

2. Make a second strip of tapes. Peel the top ones off from each group. Hang them next to each other again. Do they attract or repel?

They repel from each other.

b) What do you think will happen if you do this with the 2 bottom strips?

Again they repel from each other.

3. The two diff. strips of tape in a set have different charges. The top strip lost some electrons and has a positive charge. The bottom strip gained some electrons and has a negitive charge.

a) Is the force between 2 top strips attractive or repulsive?
It is like a magnet, two similar charged strips repel each other.

b) Is the force between 2 bottom strips attractie or repulsive?
It is like a magnet, two similer charged strips repel each other.

c) What about a positive + negitive strip?
Again, like a magnet unlike polarities attract.

4. a) What kind of forces exist b/w the nucelus + the electrons,
They attract each other because they are unlike charges.
b) What kind of forces exist b/w protons in the nucelus,
They would try to repel each other,

5. If the protons are repeling each other, there must be a stronger attractive force which keeps them together, called the nuclear or strong force, This acts between pairs of protons, neutrons, and protons + neutrons, The electron is not affected by it

a)

Particles	coulmb/electrostatic force	strong nuclear force
⊖ ⊕	attractive	
⊖ ⓝ		
⊕ ⊕	repslsive	✓
⊕ ⊖	attractice	
⊖ ⊖	repuslive	

6, If the nucleus were too large, the protons on one side of the nucleus are too far away to attract the protons on the other side of the nucleus. The protons could still repel one another since the calomb electrostatic force works long-range, The repuslive force wins and the nucelus breaks apart with a lot of energy called nuclear fission, It occurs in uranium when an added neutron brings instability,

$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{94}_{36}Kr + ^{139}_{56}Ba + 3\,^{1}_{0}n$$

a+b) Is atomic mass + number conserved on both sides?
Yes to both, Atomic mass = 236 on both sides + atomic number = 92 (# of electrons)
c) Why does the neutron have a mass of 1 + atomic # of 0?
A neutron dosn't have an protron/electron but still has mass
Small nuclei can combine + release energy @ fusion

Activity 9

Reading Guide : The Nucleus of the Atom p138-142

1. How does atomic mass relate to atomic number?

The atomic mass = # of protons + neutrons
The atomic # is the # of protons
So atomic mass = atomic # + # neutrons

2.What is wrong with organizing the periodic table by atomic mass?

Some atoms do not follow the rules of atomic mass because of isotopes
so sorting by mass did not always match the propertys

3.When were neutrons discovered?

Lord Rutherford suggested it
Chadwick actually discovered it

4.What does Helium's nucleus have to do with the discovery of neutrons?

It confirmed their existance because why is the mass of He 4x the mass of H, when the charge is only 2x. It's because the mass of neutrons $\approx$ mass of protons

5.What is a nucleon?

protons + neutrons
(reside in the nucelus)

6.Why are atomic masses not whole numbers?

it is the average of all atoms (diff isotopes) of an element

7.What is an isotope?

atoms of the same element but diff atomic mass because diff # of neutrons

~~7.What is an isotope?~~

8. How are different isotopes of the same element indicated?

Chlorine-20 or ^{37}Cl

↑ # neutrons

↑ atomic mass (# pro + # neutrons)

9.How are the forces between two protons in the nucleus related?

The protons are repeling each other $\left(F = \frac{kq_1q_2}{d}\right)$
But strong nucleor force attracts them
- very short range

10. Why is the force between pairs of protons very large? (What does the equation on page 139 mean?)

Because they are very close to each other

$F = \frac{kq_1q_2}{d^2}$

F = force
k = Coulomb's force constant
$q_1 + q_2$ = Charges
d = distance b/w charg

11.What holds the nucleus together if electrical forces repel?

a Strong nuclear force

12.Why aren't electrons affected by nuclear forces?

They are a different class of particle

13.What are the characteristics of the strong nuclear force?

- very, very strong at small distances
- actons only in neutron
- always attractive
- very short range ($< 10^{-14}$ m)

coulomb

14.What are the characteristics of the electrostatic force?

- strong at small distances, weak at large distances
- acts only b/w charged particles
- attractive or repuslive
- long range (but force gets weaker at longer range)

15.Why does the nucleus stay together?

nuclear force overwhelms repulsive electrostatic force

16.The highest atomic number to occur in nature is 92. The highest atomic number created in the laboratory is 117.

17.What do "unstable" isotopes do?

try to convert to a more stable combo of neutrons + protons - according to a systematic pattern (reactivity)

18.What is meant by "radioactive"?

an atom that has an unstable nuclei and will emit alpha, positron, or beta particles to achieve more stable nuclei.

19.The most stable nucleus belongs to which element?

$^{56}_{26}Fe$

20.Light elements become more stable when they gain mass.

21. What is meant by the term "nuclear fusion"?

nuclei of lighter atoms combine to form nuclei with greater mass + release a large amants of energy

22.What is nuclear fission?

The process of breaking apart nuclei into smaller nuclei into smaller nuclei + release a large amount of energy

23. Where does nuclear fusion occur?

occurs in the Sun +stars

24.What is the process of fusion good for?

Supplies lots of safe energy

25.How do we produce nuclear energy? (process for submarines and electric power)

fission

26.What are some major problems associated with the use of nuclear fission?

Security
radiation
removal of spent rods
disposal of waste

Chem to Go - (White p 143)

1. If Li becomes Li^+, what is the avg. atomic mass,

It wouldn't change, because electrons have almost no mass

2. H has isotopes 1, 2, 3, Write chem symbol,

$^{1}_{1}H$ $^{2}_{1}H$ $^{3}_{1}H$ ← Don't account for protons

27.Can you think of a problem with trying to use nuclear fusion to produce energy?

We have not yet gotton it to work industrally.

3. Give chem sysbols for 16 protons, 16 electrons, 17 neutrons

$^{33}_{16}S$

Do ~~checking up~~ and Chem to Go.
Read Prentice Hall Text book (Big BooK) Ch 25 do all questions #1-24

4.

chem symbol	$^{39}_{19}K$	$^{19}_{9}F$	$^{31}_{15}P$	$^{127}_{53}I$
atomic #	19	9	15	53
# protons	19	9	15	53
# electrons	19	9	15	53
# neutrons	20	10	16	74
atomic mass	39	19	31	127

5. Neutrons can be used to bombard the nucleus of U. Why is it harder to inject a proton?

Would a proton resist being added because of the repulsive electrostatic force,

6. Complete the reaction: $^{215}_{92}U + ^{1}_{0}n \rightarrow ^{94}_{38}Sr + \_\_\_ + 2\,^{1}_{0}n$

$^{119}_{54}Xe$

7. Radon is dangerous. Complete: $^{222}_{86}Rn \rightarrow ^{218}_{84}Po + \_\_\_$

$^{4}_{2}He \rightarrow \alpha$ (alpha particle)

8. Explain why He exists. What keeps 2 electrons, 2 pro, + 2 neu together?

The 2 protons + 2 neutrons are attracted to the nucleus by the nuclear force,

The electrons float around by the electrostatic force

Chap 25 Questions (Red Book)

1. How does an unstable nucleus release energy?

It emitts radiation during process of radioactive decay

2. What are 3 main types of nuceler radiation?

Alpha (α), gamma (γ), and beta (B)

3. What part of atom undergoes changes?

The nucelus decays

4. How is atomic # changed by 3 types of decay?

Alpha = -2

Beta = +1

Gamma = No Change

5. How is atomic mass changed by 3 types of decay?

Alpha = -4

Beta = No Change

Gamma = No Change

6. Which type is most penetrating?

Gamma Rays

25.2

7. Mangarese-56 is B emitter w/ half life of 2.6 hours. What is mass of 1 mg sample after 10.4 hrs

1mg = 0 hrs

.5 mg = 2.6

.25 mg = 5.2

.125 = 7.8 hrs

$\frac{1}{16}$mg = 10.4 hrs

8. Thorium-234 has $\frac{1}{2}$ life of 24.1 days. Will it be gone in 48.2 days?

No. It will keep getting smaller and smaller. Only $\frac{1}{4}$ would be left after 48.2 days.

9. What determins type of radioisotope decay?

The neutron to proton ratio.

→ seperate sheet

Chap 25 cont.

4/30

10. How much remains after 1 $\frac{1}{2}$ life? 2 $\frac{1}{2}$ lives?
    $\frac{1}{2}$ of original
    $\frac{1}{4}$ of orignal

11. What are 2 ways transmutation occurs?
    Transmutation can occur by radioactive decay or when particles bombard the nucelus of an atom.

12. Balance the equation
    a) $^{27}_{13}Al + ^{4}_{2}He \rightarrow ^{30}_{14}Si + ^{1}_{1}H$
    b) $^{214}_{83}Bi \longrightarrow ^{4}_{2}He + ^{210}_{81}Ti$
    c) $^{27}_{14}Si \rightarrow ^{0}_{-1}e + ^{27}_{15}P$
    d) $^{66}_{29}Cu \rightarrow ^{66}_{30}Zn + ^{0}_{-1}e$

13. Sample has $\frac{1}{2}$ life of 4 days. How much after
    a) 4 days = $\frac{1}{2}$
    b) 8 days = $\frac{1}{4}$) orignal

14. Mass of cobalt-60 went from .8 to .2 g in 10.5 years.
    Calculate half life.
    2.1 years

15. Explain what happens in a nuclear chain reaction.
    Some of the neutrons produced react with other fissionable atoms producing more neutrons which react with still more fissionable atoms. The process starts when a neutron is fired into uranium-235 or plutonium-239 atoms

16. Why are spent fuel rods kept in water.
    To cool + shield the rods + prevent radiation from escaping

17. How are fusion reactions diffgent then fision?
Atoms with small nuclei combine to form larger atoms and release a larger amount of energy. However fusion only occures > 40,000,000°C

18. What does nuclear moderation accomplish in a reactor.
Without it the neutrons would move too fast through the nuclei with out being absorbed. This (water + graphite) slows it down.

19. What is the source of radioactive nuclei in spent fuel rods?
Both remainers of the orignal fuel and products of the reaction.

20. What are advatages to using fusion?
The materials are easy to find + it seems not to produce hazardous waste.

21 3 methods to detect radiation
Geiger counter - metal tube where gas becomes ionized during a reaction, making curent flow
Scintillation counter - phosphor coated surfaces produces light which can be recorded
Film badge - film darkens with exposure to radiation + can be saved to keep a record

22. Describe 2 applications in medicne.
Radioisotopes can be used to dectect thyroid + other medical problems. It can also be used to treat certain diseases, such as cancer, where radiation kills unhealthy cells.

23. Why use film badge over Geiger counter.
The film badge provides a lasting record.

24. What is advatage of using a radioactive seed (radioisotope in small gold tube) to treate a cancerous tumor.
The radioisotopes can kill surronding bad cells. (tumor)
The container keeps it enclosed so it dosn't spread.

Cations

s

anions

p

sp^6

0

Noble Gas

Periodic Table of the Elements

Department of Chemistry
Lebanon Valley College
Annville, Pennsylvania 17003

+1

+2

valance electrons

+3 ±4 +5 -3 +6 -2 +7 -1

d

1 H 1.0079																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 -2 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.086	15 P 30.974	16 S 32.066	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.88	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.847	27 Co 58.933	28 Ni 58.693	29 Cu 63.546	30 +2 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.922	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.224	41 Nb 92.906	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 +1 Ag 107.87	48 +2 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (210)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Ha (262)	106 Sg (263)	107 Ns (261)	108 Hs (265)	109 Mt (266)									

metals

nonmetals

f

		58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (147)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97		
		90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (239)	95 Am (241)	96 Cm (244)	97 Bk (249)	98 Cf (252)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)		

IUPAC Standard Atomic Weights Abridged to Five Significant Figures
Pure and Applied Chemistry **1991**, *63(7)*, 987-988
See *Chem. Eng. News* May 2, 1994, p. 26 for details of elements 104-109.

6/1/c

revised 5/6/94

Michael Plosmeier

50

50

READING GUIDE FOR THE PERIODIC TABLE AND BONDING

Read Chapters 6,7 and 8 in the Big Red Book

1.The elements are sorted according to ↑ atomic number.

2.The first attempt at classifying elements was done by ~~Mendeleev~~, who grouped the elements into ~~periods~~. triads (groups sort-of) Doberelner

3.He did this by taking the comparing of the atomic masses of two similar elements.

4.Mendeleev arranged the elements in his periodic table according to increasing atomic mass.

5.Mendeleev's periodic table was used to ~~discover~~ predict the existence of new elements.

6.Mendeleev's periodic table was arranged in increasing order of atomic mass. Find as many exceptions to this in the modern periodic table as you can.

he forgot group 18 (noble gasses)

↑ Ar + K Co + Ni

7. The modern periodic table is arranged in order on increasing atomic number.

8.State periodic law:

When elements are arranged in the order of increasing atomic number, there is a periodic repetition of their physical + chemical property

9.What does the stair step line indicate?

The metalloids seperating the metals + non metals

10.Explain the numbering of the groups.

3 systems ↙# valance electron

US - 1A, 2A, 3B, 4B, 5B, 6B, 7B, 8B, 1B, 2B, 3A, 4A, 5A, 6A, 7A, 8A

Europe - IA, IIA, IIIA, IVA, VA, VIA, VIIA, VIIIA, IB, IIB, IIIB, IVB, VB, IVB, VIB, VIIB

IUPAC - 1-18

11. The three classes of elements are:

metals, nonmetals, metalloids

12. Most elements are solid metals (classification and state)

13. Summarize properties of metals and nonmetals.

most (good conductors, solids (STP) green, ductile, malleable shiny | greater variation most gas (STP) poor conducters brittle dull

14. What is a metalloid?

properties of metals + nonmetals depending on condition

15. What kinds of information can be found in the periodic table?

atomic #, energy level, symbol, name, avg. atomic mas

\# Vallance electron, state @ STP, metal/nonmetal

16. Describe each of the representative families in Groups 1A, 2A, and 7A.

all poisonous

1A - alkali metals - silver gray, soft solids - low melting + boiling pt

reactive, "burn" in air, form bases when react w/ water

2A - alkaline earth metals - relatively soft, harder then 1, gray white luster when fresh cut, hardens to thin, tough oxide coating, strong lightweight

7a - Halogens - nonmetals - very reactive, salt former when react

17. What is a noble gas and what is significant about noble gases?

Group 8A or 18 - very stable, inert, unreactive atoms combine to be like noble gasses - uncombined in nature

18. What is a representative element ? Look at p. 165. What is this page trying to explain?

groups 1A - 7A (1, 2 + 13-17) everything other than noble gasses metals + nonmetals display a wide range of properties

19. What is a transition metal? How do the inner transition elements differ from the transition elements. (Hint: Look at the order of filling orbitals)

Group 3-12 - higest s filled + higest d filling up transion metal → nometal

inner - higest s filled + higest f filling up (normaly shown under the table)

20. Do problems10-15 on p. 167.

10. What info in periodic table?

The periodic table displays the symbols + names of the elements along w/ info about the structure of their atoms,

11. What 4 classes can atoms be seperated in based on electron config?

Elements can be seperated into noble gasses, representative elements, transition metals, and inner transition metals.

→

12. Why do K + Na have similor properties?

They are both in the same group (Group 1, alkali metals)

only 1 valence electron

13. Classify as representative, transition, or noble.

a $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$ - noblegas
b $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$ - ~~representative element~~ transition metal
c $1s^2 2s^2 2p^6 3s^2 2p^2$ - representative

14. Which are transition metals?

1. Cu - Yes
2. Sr - No
3. Cd - Yes
4. Au - Yes
5. Al - No
6. Ge - No
7. Co - Yes

15. How many electrons in highest energy level of Group 5A (15)

~~np^3~~ 5 (valence electron)

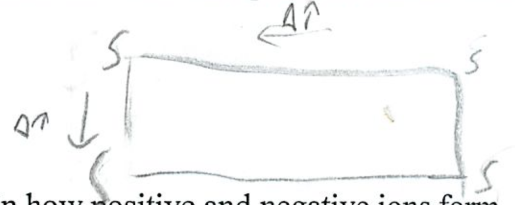
↑
what?

21.In your own words, explain how paint is made.

pigment (color), binder (to stick), + liquid mixed
↗
organic or inorganic
different elements = different colors,

22.Summarize the following periodic trends: Atomic Size,

atomic radius
½ distance
b/w nuceuli
when joined



23. Explain how positive and negative ions form. electrons gained

⊖ lost
electrons are transfered between atoms, so that the combined form can be more like a stable noble gas

24. Define: anion define cation.

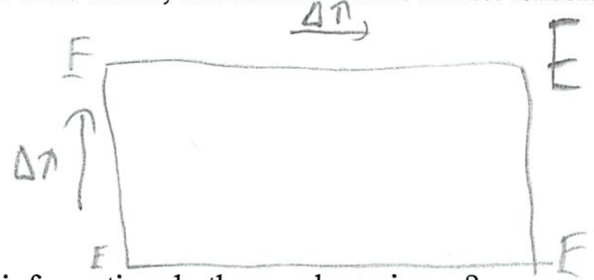
ion w/ negitive charge ⊖
gain electrons
(Groups 14-17)

electrons > protons

ion w/ positive charge ⊕
(Groups 1-13) electrons < protons
lose electrons

(after getting electrons and thus are ions)

25,In your own words, summarize trends in first ionization energies.



26. What information do these values give us? stability of atom + ion formed

How easy to remove an atom. Atoms with ↓ 1st ionization energy easily lose an electron to become like noble gases.

27. Look at the graph on page 174, How do the peaks and valleys relate to the periodic table?

See # 25
peaks - noble gasses (Group 18)
valleys - Group 1 - 13 - active metals

28.Do the Quick Lab on p.175.

See other sheet (graph)

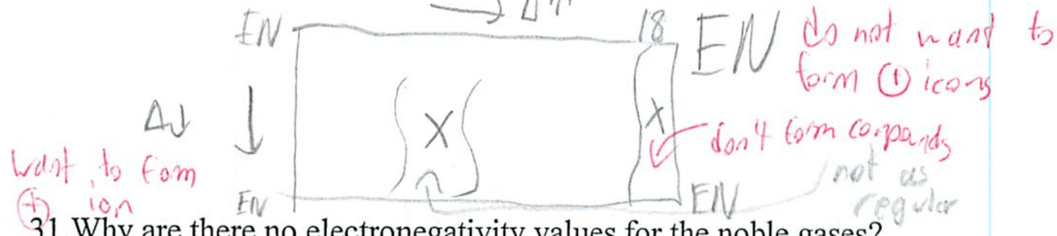
29.What is electronegativity?

The ability of an atom to attract electrons when the atom is in a compound

determines which type of bond

calculated w/ ionization energies

30.In your own words, summarize the trends in electronegativity.

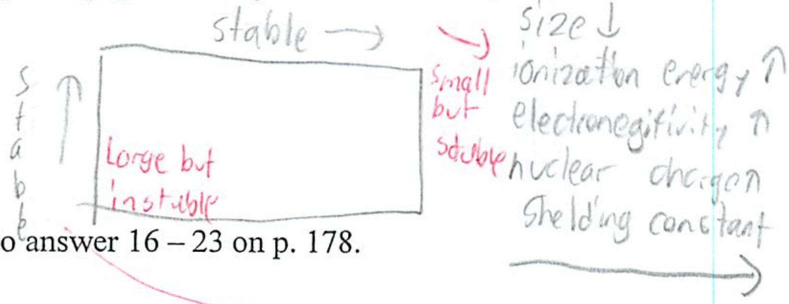


31.Why are there no electronegativity values for the noble gases?

They tend not to attract electrons

32. Look at the graph on page 178. In your own words summarize the information given by it.

trends explained by atomic propertis



size ↑ (atomic + Ionic)
ionization energy ↓
Electronegitivy ↓
nuldar charge ↑
Shelding increase

33. Use the graph to answer 16 – 23 on p. 178.

Back

Read Ch 7 and answer the following questions:

1. What is a valence electron?(s)

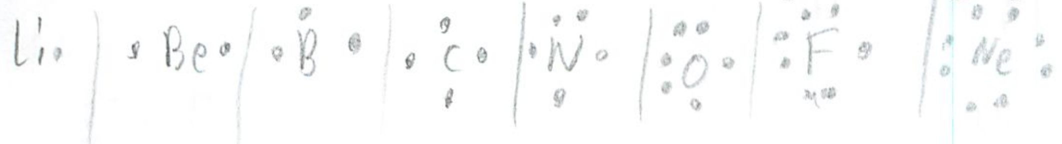
electrons in the higest occupied energy level of the atom

2. How can the number of valence electrons be found easily?

look at its group number (US system)

3.Draw electron dot structures for the first element in each representative group.

Pod
2



Only shows valence electrons

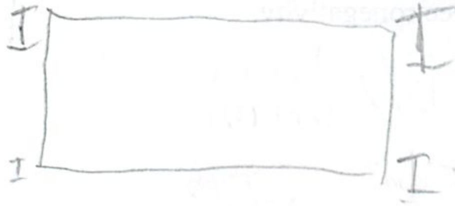
16. How does atomic size Δ within groups & periods?

In general, atomic size ↑ from top to bottom + decreases left to right.

17. When do ions form?

Positive + negitive ions form when electrons are transferred between atoms.

18. What happens to 1st ionization energy?

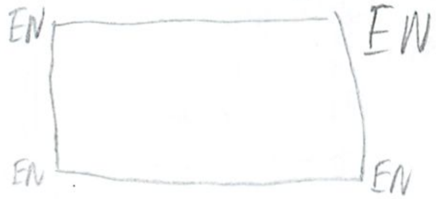


19. Compare size of ions to size of atoms they formed from?

Cations (1-13) always are < size atoms) formed from

anions always > size atoms

20. How does electronegativity vary within groups & periods?



21. In general how can periodic trends be explained?

The periodic trends are explained by variations in atoms structure.

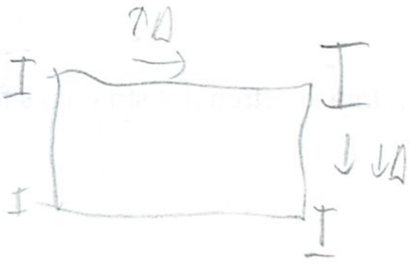
22. Arrange atoms in Δ↓ atomic size. Periodic trend or group trend?

Cl, S, Al, Na - Periodic trend

23. Which element has the larger 1st ionization energy?

a. Na, K ← higher period

b. P, Mg ← higher group



Oh within the pair opps.

to get 8

4. What is the “octet rule?”

Atoms of the metallic elements tend to lose their valence electrons, leaving a complete octet in the next-lowest energy level. Atoms of some non-metallic elements tend to gain electrons or share electrons with another non metallic element to achieve a complete octet.

5. Explain how a + ion (cation) forms.

Atoms lose valence electrons which can be easily removed to form octets

6. Write out the electron configuration for the potassium ion.

relate to protons

$K \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

$K^+ \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6$ octet $\rightarrow Ar$

anions - gain electrons - charge

7. What is the charge on this ion?

~~$1e^-$~~ $+1$

8. What would the charge be on the Calcium ion.?

Ca^{+2} thats its charge not that it gains 2 electrons

9. Explain in your own words how an anion (-) ion, forms.

gains electrons to have an octet

10. Write out the electronic configuration for the formation of the fluoride ion.

$F \rightarrow 1s^2 2s^2 2p^5$

$F^- \rightarrow 1s^2 2s^2 2p^6$

11. Write out the configuration of the oxide ion.

$:\ddot{\underset{\cdot\cdot}{O}}:$ ~~2-~~ $1s^2 2s^2 2p^6$ O^{2-}

12. All halide ions have a __-1__ charge (number and sign)

13. Learn the names , formulas and charges for the common anions on p. 192.

p193.

3. How determin # valence electrons in representative element?
To find the # of valence electrons in an atom is to look at its USA group #.

4. Which atoms tend to gain elections? lose electrons?
Atoms of the metallic elements tend to lose their valence electrons, leaving a complete octet in the next lowest energy level. Atoms of some nonmetallic elements tend to gain electrons or to share electrons with another nonmetalic element to achieve a complete octet.

5. How do cations form?
Atoms lose valence electrons to form octets

6. How do anions form?
Atoms gain electrons to form octets

7+8. How many valence electrons? Draw dots.
a) k - 1 $k\cdot$ c) Mg - 2 $\cdot Mg \cdot$
b) C - 4 $\cdot\overset{\cdot}{\underset{\cdot}{C}}\cdot$ d) O - 6 $:\overset{\cdot\cdot}{\underset{\cdot}{O}}\cdot$

9. How many atoms gain/lose to form ion?
a) Ca - lose 2 b) F - gain 1
c) Al - lose 3 d) O - gain 2

10. Write name + symbol now has (+) charge
a) k loses 1 electron → k^+ potassium ion
b) Z loses 2 electrons → Z^{2+} ← weird order zinc ion
c) F gains an electron → F^- floride

11. Write electron config of Cd^{2+}. → short 2 electrons
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$ ~~$5s^2 4d^8$~~
the s is empty $5s^0 4d^{10}$ ← higest level

14. Do. 3-11 on p. 193.

← Back

15.Compounds are electrically neutral this means the sum of ionic charges adds to 0.

16.What is an ionic bond? Explain how it forms and give one example using electron configurations.

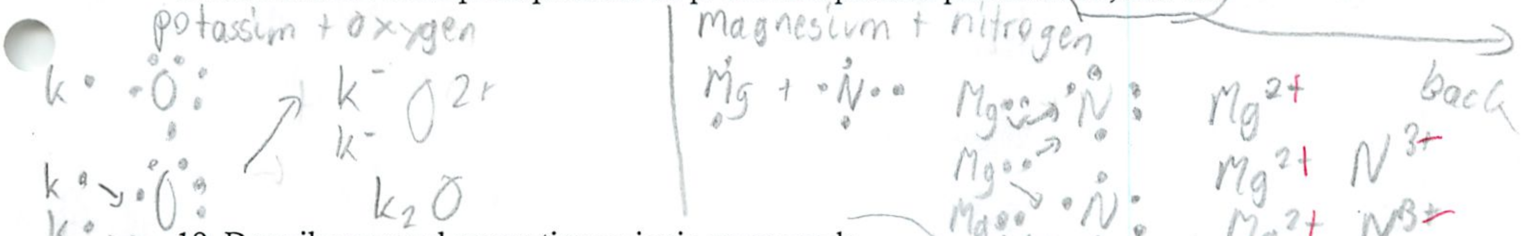
When electrostatic forces hold anions + cations togeter b/c they have oppset charge
cations can edisly lose electrons while anions can edisly gain electrons. They bond when they both form octets

17.What information is given by a chemical formula? What is a formula unit?

chem formula kinds + # of atoms in the smallest representative unit of a substance

formula unit - lowest whole # ratio of ions in an ionic compand

18. Look at the conceptual problem on p. 196. Do practice problems 12, and 13. → back



↑ charges not electrons

Mg_3N_2

19. Describe general properties on ionic compounds.

most are crystalline solids at room temp
3D patterns
Strong attraction to neiobocs → stable
high melting points (>800°C)
conduct electricity when disolved + melted

20. What information is given by the coordination number of a crystal?

the # of ions of opposite charge that surrounds the ions in a crystal
ions

21.What enables a dissolved ionic solid to conduct electricity?

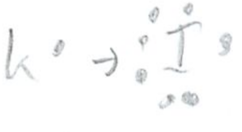
When melted, the orderly crystal structure breaks down cations can migrate freely to 1 electrode + anions migrate to the other - movement allows electricity to conduct
pass charge

22.Do problems 14-23 on p. 199. ← back

p 196

12. Show dots to form ionic compounds

a) k reacts w/ I



KI

k^+ I^-

b) Al reacts w/ O

→ Al^{3+} O^{2-}
Al^{3+} O^{2-}
O^{2-}

Al_2O_3

13. What is formula of calcium + chloride?

Using cross trick

~~Ca_2Cl~~ $CaCl_2$

---

p 199

14. How can you describe te electrical charge of an ionic compound?
They ore electically nectral.

15. What properties characterize ionic compounds?
Ionic compounds are crystalline solids at STP + have a high melting point.

16. Define an Ionic Bond.
An ionic bond is when an electrostatic charge holds together an anion and a cation as tey try to form an octet.

17. How can you represent te composition of an ionic bond?
You can represent te bonds with electron dot structures.
or writing formula

18. Write te chem. formulas:

a) K^+S^{2-} - K_2S

b) $Ca^{2+}O^{2-}$ - CaO

c) Na^+, O^{2-} - Na_2O

d) $Al^{3+}N^{3-}$ _ AlN ← reduce w/ cross method

~~Cross method~~ ~~doesn't always~~ ~~work~~

continued next back

7.3

in metals - nuclels remans in place

23. What is the "electron sea model"?

Valence electrons can move fairly freely + be aranged

why metals hold together (

24. What is a "metallic bond"?

attraction of free-floating valence electrons for positivly charge metal ions (cations)

25. Why are alloys important?

Their properties are often superior to those of their component alloys -combos of metals

26. Read p. 204-205 and explain some advantages of building with alloys.

lets tall buildings be built quickly, taller
-not be all walls
can be used to enforce concrete (rebar)

8.1

READ Chapter 8.

1. What is a covalent bond and how is it different from an ionic bond?

atoms held together by sharing electrons not "stealing" electrons transferred

3. What is a diatomic molecule?

molecule consisting of 2 atoms

4. What is a molecular compound?

Compounds made up of molecules covalently bonded

5. Compare molecular compounds to ionic compounds.

have lower boiling + melting points
most comprised of atoms of 2 or more nonmetals
gas + liquid usually STP

6. What is a molecular formula?

Chem formula of molecular compound
shows how many atoms of each element a molecule contains ratio 1 atom to other

8.2

7. In your own words, explain how the octet rule applies to covalent bonding.

elements try and bond so that each atom has 2 or 8 electrons to be like a noble gas. Note the total # of electrons does not have to be ÷ by 2 or 8

8. Use the Chlorine molecule to describe the formation of a single covalent bond.

They share electrons to try + form an octet

9. What information is given by a structural formula?

Shows arrangment of covalently bonded atoms.

p199

19. Write chem formulas ?BaCl

a) barium chloride - ($Ba^{2+}Cl^{2-}$) b) magrasium oxide - ($Mg^{2+}O^{2-}$) MgO

c) lithium Oxide ($Li^{+}O^{2-}$) Li_2O d) calcium fluoride - ($Ca^{2+}F^{-}$) CaF_2

20. Which pair most likly to form ionic compound?

a) Cl, Br b) Li, Cl c) K, He d) I, Na

↑ equally likly ↑ farther down (↓) ~~most likly~~ ?

21. Describe arrangment of sodium ions and chloride ions in crystal of sodium chloride.

NaCl is very stable because of the alternating +even arrangments of Na + Cl. They have equal coordination numbers meaning they touch each other an equal amount. solid

22. Why do ionic compounds conduct electricty when disolved in water.

When they are disolved the ions are attracted to the polar electroids and can migrate ~~freely~~ carrying a charge.

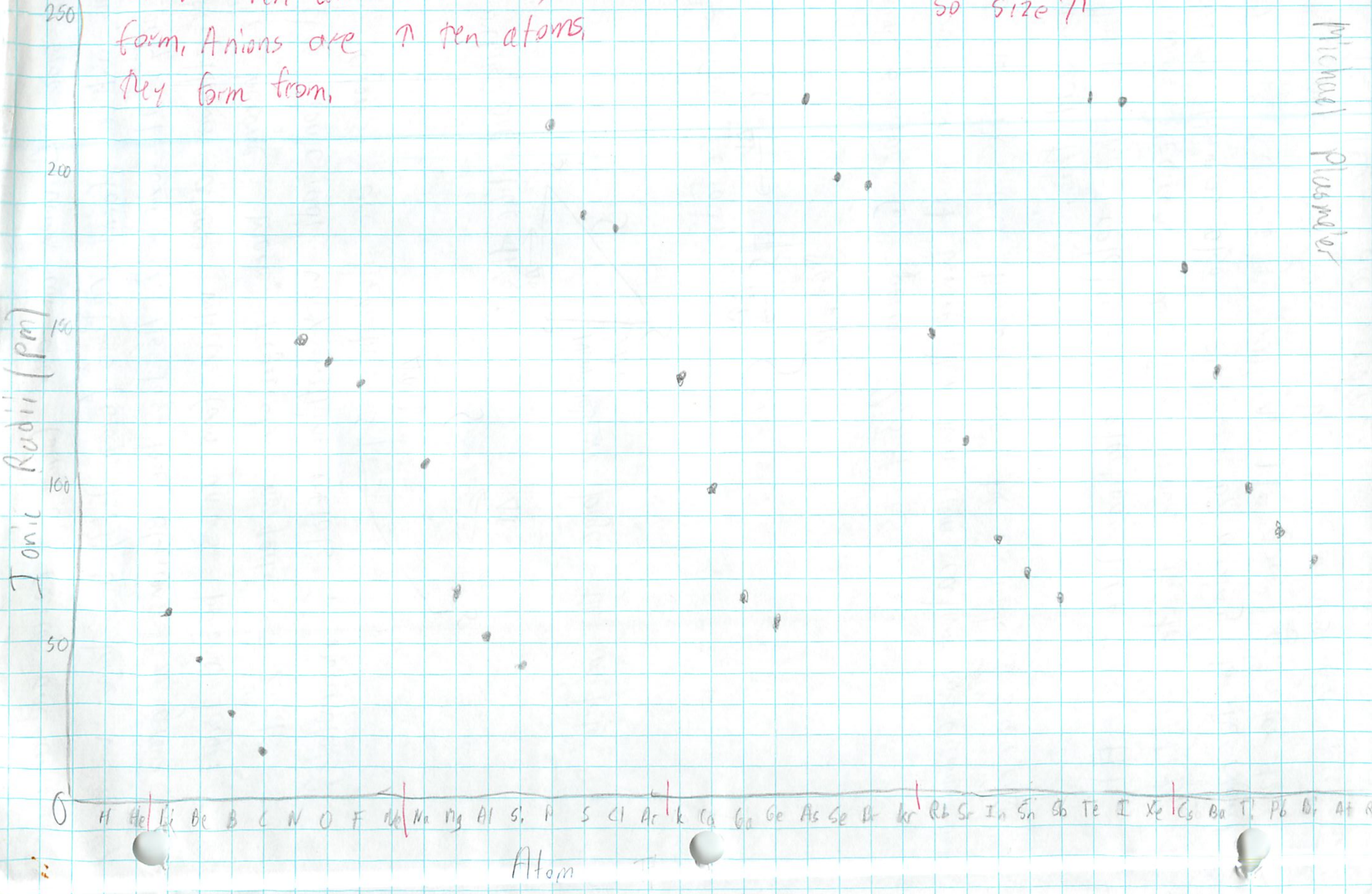
4/25

Periodic Trends in Ionic Radii

* Cations are always smaller than atoms which they form. Anions are ↑ than atoms they form from.

As # electrons ↑ attraction force for each ↓ so size ↑

Michael Plasmeier

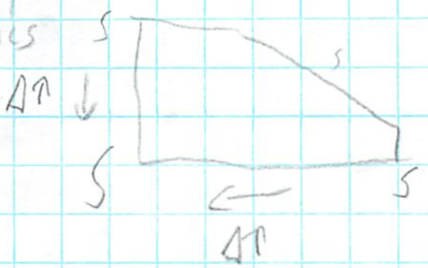


1. Describe size Δ when an atom forms a cation and anion?

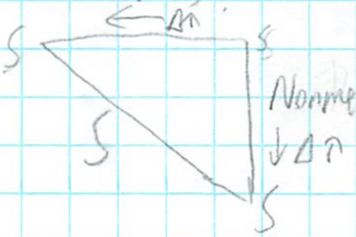
Atoms which form cations (Groups 1-13) are larger when they are only +1 over +2. As they lose electrons they grow smaller. Hower atoms which form anions become larger as they become more negitivly charged -2 > -1

2. How does ionic radii vary within metals? nonmetals?

Metals' atoms are like this

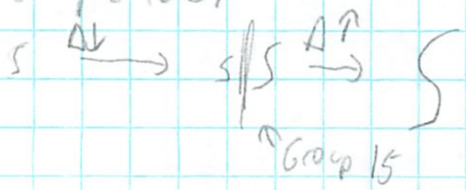


Nonmetals are like this:



Nonmetals are larger then metals

3. Describe one period.



4. Is trend similar for periods 2, 3, 4, + 5?

Yes, but the point where they turn from cations to anions moves to the left because of the metalloid "staircase"

5. Explain all this.

I think this relates to electrons like in #1. Cations lose electrons as they move to the right, thus get smaller. Anions also electrons by moving to the right explaining the decreasing pattern. It goes greater the displacement near the end? b/c the electrons are less stable as their are more.

○ RESCORE ○ MARK ✓ ○ TOTAL ONLY

USE NO. 2 PENCIL ONLY

Use a No. 2 Pencil

Ⓐ ● Ⓒ Ⓓ Ⓔ Fill circle completely

Ⓐ Ⓑ Ⓒ Ⓓ Ⓔ Erase cleanly

	T	F				
1	●	Ⓑ	Ⓒ	Ⓓ	Ⓔ	✓
2	Ⓐ	Ⓑ	●	Ⓓ	Ⓔ	
3	●	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
4	Ⓐ	●	Ⓒ	Ⓓ	Ⓔ	
5	Ⓐ	●	Ⓒ	Ⓓ	Ⓔ	
6	Ⓐ	●	Ⓒ	Ⓓ	Ⓔ	
7	Ⓐ	Ⓑ	Ⓒ	●	Ⓔ	
8	Ⓐ	Ⓑ	Ⓒ	●	Ⓔ	
9	Ⓐ	Ⓑ	●	Ⓓ	Ⓔ	
10	Ⓐ	●	Ⓒ	Ⓓ	Ⓔ	
11	Ⓐ	Ⓑ	Ⓒ	●	Ⓔ	
12	Ⓐ	Ⓑ	Ⓒ	●	Ⓔ	
13	Ⓐ	●	Ⓒ	Ⓓ	Ⓔ	
14	Ⓐ	●	Ⓒ	Ⓓ	Ⓔ	
15	Ⓐ	●	Ⓒ	Ⓓ	Ⓔ	
16	Ⓐ	●	Ⓒ	Ⓓ	Ⓔ	
17	Ⓐ	Ⓑ	Ⓒ	●	Ⓔ	
18	Ⓐ	●	Ⓒ	Ⓓ	Ⓔ	
19	Ⓐ	Ⓑ	Ⓒ	●	Ⓔ	
20	●	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
21	Ⓐ	●	Ⓒ	Ⓓ	Ⓔ	✓
22	Ⓐ	Ⓑ	Ⓒ	●	Ⓔ	
23	●	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
24	●	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
25	Ⓐ	●	Ⓒ	Ⓓ	Ⓔ	
26	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
27	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
28	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
29	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
30	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
31	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
32	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
33	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
34	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
35	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
36	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
37	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
38	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
39	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
40	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
41	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
42	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
43	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
44	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
45	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
46	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
47	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
48	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
49	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	
50	Ⓐ	Ⓑ	Ⓒ	Ⓓ	Ⓔ	

1. Incorrect?
A B C D
↑sort-of true

4. N^{3-} - gained 3 electrons like Ne.

6. anions - easy to gain electrons

7. ?? red ↑F = ↓λ = ↓E
infa-red [visable] ultra violet

9. B^{3+}

10 $Ca^{+2} F^{-1} \rightarrow CaF_2$

11. $K^{+1} S^{-2} \rightarrow K_2S$

12. $Al^{+3} O^{-2} \rightarrow Al_2O_3$

17. Noble Gas

18. $[Ar] 3d^{10} 4s^2 4p^3 = 5$

19. $Si^{+4} Br^{-1} \rightarrow SiBr_4$

23. $F = \frac{c}{\lambda}$

SCORE	23	# CORRECT
	92	% CORRECT
RESCORE		# CORRECT
		% CORRECT
ROSTER NUMBER	16	SCORE
		RESCORE

NAME Michael Plasmeier

COURSE Chem - Periodic table + Bonding

HOUR 06-07S2B3 DATE 5/2/07

TEST RECORD	
PART 1	
PART 2	
TOTAL	

TEST ANSWER SHEET J
Form No. 72512

ACCU-SCAN #26180-TR 10/00 70976 APPERSON PRINT MANAGEMENT SERVICES (800) 473-6761 05/05

DO NOT WRITE IN THIS AREA

5/3

Name Michael Plasmeier Block ____________

Active Chemistry Quizo #2

Chemical Dominoes–Activity #2 Quiz

Balance the following equations:

1. 2 Na + 1 Cl_2 = 2 NaCl ← can't change subscript

2. 2 Cu + 1 S = 1 Cu_2S

3. 2 HgO = 2 Hg + 1 O_2

4. 1 NH_4NO_3 = 1 N_2O + 2 H_2O H-4 O-3 N-2

5. 2 KI + 1 Cl_2 = 2 KCl + 1 I_2

6. ____ CH_4 + ____ Cl_2 = ____ $CHCl_3$ + ____ HCl I can't be done

7. 1 $CaCO_3$ + 2 HCl = 1 $CaCl_2$ + 1 H_2CO_3

8. 2 $AgNO_3$ + 1 H_2S = 1 Ag_2S + 2 HNO_3

9. 1 $CaCO_3$ = 1 CaO + 1 CO_2

10. 1 C_3H_8 + 5 O_2 = 3 CO_2 + 4 H_2O

Bonus Question:

11. 2 $Al(OH)_3$ + 3 H_2SO_4 = 1 $Al_2(SO_4)_3$ + 3 H_2O

Al-2 H-6
O-18 S-3

Michael Plasmeier

Chemical Reactions

20/20

5/2

to prove mass is conserveld and so chem equations will balance out

Get a ziplock bag, Na HOC$_3$ (sodium bicarbonate) and HCl (HydroCloric acid), and a balance.

Find the mass of each material sepertly, (including bag)

Combine the two compounds in the ziplock bag.

Find the mass of the bag + reactants

Subtract the mass of the bag from the mass of the reactants, Has mass been consered? Write out the chem formula of what happened. Calculate the % yield and error $\frac{\text{product}}{\text{reactants}}$

Chem Formula:

$$HCl + NaHCO_3 \rightarrow CO_2 + H_2O + NaCl$$

Mass Beaker = 30.00 g
Mass Dish = 2.22 g
Mass Bag = 1.59 g

Mass Beaker + HCl = 51.06 g
Mass Dish + $NaHCO_3$ = 3.66 g

Mass HCl = 21.06 g
Mass $NaHCO_3$ = 1.44 g

22.5 g

Mass Bag + Products = 23.75 g

Ma
Mass Products = 22.16 g

% Yield = $\frac{22.16}{22.5\text{ g}}$ = 98.5 %

% Error $\frac{.34}{22.5\text{ g}}$ = 1.5%

100%

Calculations

1. $\% \text{ yield} = \frac{\text{mass products}}{\text{mass reactants}} \cdot [100] = \%$

2. Look at Chem equation

$$\overset{\text{reactants}}{HCl + NaHCO_3} \xrightarrow{} \overset{\text{products}}{CO_2 + H_2O + NaCl}$$

3. Count the number + kind of atoms on both sides. Are they equal?

H_2 Na C
Cl O_2 } yes they are equal!

Look up atomic weights of each

H - 1,00794 Na - 22,98977 C - 12,011
Cl - 12,011 O - 15,9994

What is meant by molar mass?

The mass of 1 mole of the element,
↑ $6{,}02 \times 10^{23}$ particles

Find the molar mass (find # g in 1 mole - add masses)
molar mass = amu · 1 g

$HCl = 1 + 34 = 35\text{g}$
$NaHCO_3 = 22 + 1 + 12 + (3 \cdot 16) = 83\text{g}$ } 118g
$CO_2 = 12 + (2 \cdot 16) = 44\text{g}$
$H_2O = (1 \cdot 2) + 16 = 18\text{g}$
$NaCl = 22 + 34 = 56\text{g}$ } 118g

=

law of conservation of mass/matter - in any physical change or chemical reaction, mass is conserved, mass can be neither created nor destroyed; products = reactants

Symbols

+	seperates 2 reactants or 2 products	
$\rightarrow$	"yields" seperates reactants from product	
$\leftrightarrows$	reversable reaction	
(s)	solid	after formula
(l)	liquid	
(g)	gas	
(aq)	aqueos solution	
$\xrightarrow{\Delta}$ $\xrightarrow{heat}$	heat is added	
$\xrightarrow{Pt}$	catalyst used (above yield sign)	

Chap 11.1

Describing Chem Equations 5/3

1. Write a sentence that describes this:

$Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_2\ (g)$

The salt and water are the reactants. One H and an O join with Na, and leave 2 Hydigen bonded togeth.
Sodium hydroxide (in solution)

2. S burns in O to make sulfur dioxide. Write skeleton equations.

$S_{(s)} + O_{(g)} \rightarrow SO_{(g)}$

3. Balance the equation

a) $2Ag\cancel{_2}NO_3 + H_2S \rightarrow Ag_2S + 2H_2NO_3$
← don't change subscripts - use coefficent

b) $Zn(OH)_2 + H_3(PO_4) \rightarrow Zn_3(PO_4)_2 + H_2O$

4. Write balanced chem equations

a) $H_2 + S_2 \rightarrow H_2S_2$

b) $FeCl_3 + Ca(OH)_2 \rightarrow Fe(OH)_2 + CaCl_3$

5. Balance the equation

a) $FeCl_3\ 3NaOH \rightarrow Fe(OH)_3 \overset{3}{+} NaCl$

b) $CS_2 + 3Cl_2 \rightarrow CCl_4 + S_2Cl_2$

6. Write + Balance equation

$Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + H_2O + H_2$

Mole Notes

5/4

mole = $6.02 \cdot 10^{23}$ things

↑ of any substance ↑ atoms, molecles, ions formula units

molar mass = mass in grams of 1 mole of a substance

" " " " of $6.02 \cdot 10^{23}$ things

Counting Particles + Relative Masses Questions

5/4

1. There are $6{,}02 \times 10^{23}$ H atoms in one mole of H,

b. How many He molecules in a mole? $6{,}02 \times 10^{23}$
In any substance! $6{,}02 \times 10^{23}$ particles
↑ always

c. How many grams does 1 mole of C weigh 12g
Mole of H_2O - 18g
Mole of Fe - 55,87g
molar mass

2. How many moles make up $12{,}06 \times 10^{23}$ of He $\frac{2}{\text{mole}}$
6×10^{20} O = $\frac{1}{1000}$ mole
1000 H_2O particles = $\frac{1}{6.03 \cdot 10^{20}}$ mole

3. Calc # atoms or molecules
a) 15 moles CH_4 = $15 \cdot 6{,}02 \cdot 10^{23}$ = $9{,}03 \cdot 10^{24}$
b) 3 g H_2O $\frac{1}{6}$ mole = $1{,}003 \cdot 10^{23}$
c) 2 moles Fe = $2 \cdot 6{,}02 \cdot 10^{23}$ = $1{,}204 \cdot 10^{24}$
d) 2 g Fe = $\frac{1}{28}$ mole = $2{,}15 \cdot 10^{22}$
particles

4. What would be mass of 20 moles C?
12 • 20 = 240 g ←
1mole = $6{,}02 \cdot 10^{23}$
1mole = 12g
20 moles = $1{,}204 \cdot 10^{24}$
20 moles

5. One mole of beans =
$6{,}02 \cdot 10^{23}$ beans
1 bean = ,2478 g
1 mole = $1{,}49 \cdot 10^{23}$ g

$6 \cdot 10^{26} / 1{,}49 \cdot 10^{23}$ = The earth is 4022 × greater than 1 mole of beans.

Extension Qu

1. 100 g of the heavest bean = 263,64 beans
   263.64 of the ligest bean = 65,33 g

2. If a mole = 10 g H it is 10× current value or $6.02 \cdot 10^{24}$
   1 (new) mole carbon would be 120 g
   1 (new) mole CH_4 molecules = 160 g

Michael Plasmeier

5/7

Mole Calculations- Easy

1. Molar mass: To find the molar mass of a substance, add up the atomic masses of the elements in a compound. If the element is followed by a subscript, multiply that subscript by the atomic mass of the element;

(of 1 mole) 1 AMU = 1g

example: NaOH Na =23g O = 16g and H = 1g molar mass = 23+16+1 = 40g/mole

$CaCO_3$ Ca =40g C =12g O= 3 x 16g molar mass= 40+12+ 3(16) = 100g/mole

$(NH_4)_2SO_4$ = 2(14g) +8(1g) +32g =4(16g) =132g/mole

* remember: molar mass has units of g/mole

atomic mass

Do these problems: Find the molar mass of each compound:

a. KNO_3
19 + 7 + 3(8)
50
101

b. $Na_2C_2O_4$
2(11) + 2(6) + 4(8)
66
133

c. $Fe(OH)_3$
26 + 3(8 + 1)
53
149 (?)

d. $(NH_4)_3PO_4$
3(7 + 4(1)) + 15 + 4(8)
80
150 (?)

x2 b/c mass not atomic #

2.One mole contains the same number of objects. This number is 6.02×10^{23}. It does not matter what the substance is. Atoms, molecules, formula units are all objects (particles)

two factors can be prepared from this statement:

1 mole = 6.02×10^{23} particles

These factors are created by multiplying one side by the other. The numerical value of both sides is 1.

$$\frac{1 \text{ mole}}{6.02 \times 10^{23} \text{ particles}} \quad \text{and} \quad \frac{6.02 \times 10^{23}}{1 \text{ mole}}$$

When solving problems, follow these simple rules:

1. Read the problem
2. Underline what you are given and what you are being asked for.
3. Start with what you are given INCLUDE UNITS.
4. Use the factor which will cancel the units of the given value.

example: How many atoms are there in 4 moles of Boron.

given: 4 moles of Boron   asked: how many atoms

$$4\text{moles of Boron} \times \frac{6.02 \times 10^{23}\text{atoms}}{1\text{mole}} = 2.41 \times 10^{24}\text{ atoms}$$

* To multiply numbers in exponential form, multiply numbers and add exponents
* To divide numbers in exponential form, divide numbers, subtract exponents

3. When finding masses of moles, you need the molar mass of the substance. This is what you determined in 1. You can make two factors for this:

$$\frac{1\text{ mole}}{\text{molar mass}} \qquad \frac{\text{molar mass}}{1\text{ mole}}$$

Example: What is the mass of 3 moles of KOH?

$$3\text{ moles of KOH} \times \frac{56\text{g(molar mass of KOH)}}{1\text{ mole}} = 168\text{g}$$

Do the following problems:

a. How many molecules are there in 2.5 moles of Silver?

$2.5 \cdot 6{,}02 \cdot 10^{23} = 1{,}505 \cdot 10^{24}$ molecules

b. What is the mass in grams of 1.7 moles of $NaNO_3$?

$Na\ N\ O_3$

$23 + 14 + 3(8) = 61$

$1{,}7\text{ moles} \cdot \frac{61\text{ g (molar mass)}}{1\text{ mole}} = 103{,}7\text{ g}$

c. How many moles are there in 4.2×10^{24} atoms of Argon?

$4{,}2 \cdot 10^{24} / 6{,}02 \cdot 10^{23} \approx 6{,}97$ moles write $40g\ Mg(OH)_2 \cdot \frac{1\ mole\ Mg(OH)_2}{(24+32+2)}$

d. How many moles are there in 40g of $Mg(OH)_2$

$Mg + (OH)_2$ $40/58 = ,69$ moles

$24 + 2(16+1) = 58g$ (or Amu) = molar mass

e. What is the mass in grams of 1 molecule of CO_2 ? (you will need two factors to solve this) $1\ mole\ CO_2 = 12 + 2(16) = 44g / 6.02 \cdot 10^{23} = 7.31 \cdot 10^{-23}$

Molar Volumes: One mole of any gas will occupy 22.4L at STP(standard temperature and pressure) 0 degrees Celsius, 101.3KPa pressure.

example: $N_2 + 3H_2 = 2NH_3$ How many liters of ammonia will be produced by reacting 2.5 L of nitrogen?

$$2.5LN_2 \times \frac{1\ mole\ N_2}{22.4L} \times \frac{2\ mol\ NH_3}{1\ mol\ N_2} \times \boxed{\frac{224\ L}{1\ mol\ NH_3}} = 5L\ of\ NH_3$$

$N_2 + 3H_2 \rightarrow 2NH_3$

How many moles of H_2 are needed to react with .42 moles N_2?

$$\underset{\text{given}}{.42 \text{mole } N_2} \cdot \frac{3 \text{mole } H_2}{1 \text{ mole } N_2} = 1.26 \text{ mole } H_2$$

---

p 360 Red Book

11. $4Al + 3O_2 \rightarrow 2Al_2\ O_3$

a. $\frac{2 \text{ mole } Al_2}{4 \text{ mol } Al}$ $\frac{2 \text{ mole } Al_2}{3 \text{ mole } O_2}$ $\frac{2 \text{mole } O_3}{4 \text{ mole } Al}$

(?)

any combo

$\frac{2 \text{mole } O_3}{3 \text{ mole } O_2}$ $\frac{4 \text{ mole } Al}{2 \text{ mol } Al_2O_3}$ $\frac{2 \text{ mole } Al_2O_3}{4 \text{ mole } Al}$

$\frac{4Al}{3O_2}$ $\frac{4Al}{2Al_2O_3}$ $\frac{3O_2}{2Al_2O_3}$ $\frac{3O_2}{4Al}$

b. $3.7 \text{m } Al_2O_3 \cdot \frac{4 \text{mole } Al}{2 \text{ mol } Al_2O_3} = 7.4 \text{ mol } Al$

same cancle out

12a. $14.8 \text{ mole } Al \cdot \frac{3 \text{ mole } O_2}{4 \text{ mole } Al} = 11.1 \text{ mole } O_2$

b. $0.78 \text{ mol } O_2 \cdot \frac{2\ Al_2O_3}{3\ O_2} = .52 \text{ mole } Al_2O_3$

Naming Compounds

5/10

First count # of elements in compound

If 2 : Binary compound
- Representative metal + non metal
  - Name metal
  - Nare nonmetal
  - Drop ending
  - Add -ide
  - NaBr : Sodium bromide
- Transiton metal + non metal
  - Determine the charge on the metal ion,
  - Name metal using ( ) + roman numeral to indicate charge on metal
  - Name nonmetal
  - Drop ending
  - Add -ide
  - $Fe^{3+}Cl_3^-$ iron = 3 → iron(III) chloride
- 2 non metals
  - Name first element
  - Nare second element
  - Drop ending
  - Add -ide
  - Use prefixes to indicate atoms of each
  - P_2O_3 - di phosphorous - tri oxide
- Binary acids (H X) Hydrogen
  - hydro ____ ic acid

Binary Molecular
Mono - 1
Di - - 2
tri - 3
tetra - 4
Penta - 5
Hexa - 6
Hepta - 7
Octa - 8
Nona - 9
Deca - 10

Compounds with > 2 elements
- representative metal + polyatomic ion
  - Draw a box around all elements after the metal
  - Name metal
  - Look up ion + nare it
  - NaOH → sodium hydroxide

Compounds with >2 elements

Transition metal+ ion

Determine charge of metal

Use ( ) and roman numeral to indicate charge

Name ion

$Cu^{2+} SO_4^{2-}$ e Copper (II) sulfate

Oxy acids (Hydrogen+ polyatomic ion)

If ion ends in -ite

____ous acid

If ion ends in -ate

____ic acid

} no hydro-

Prefixs

ClO hypoclorite

ClO_2 clorite

ClO_3 clorate

ClO_4 perclorate

Michael Plasmeier

II. Names, Symbols, and Oxidation Numbers of Common Elements and Radicals

+1 Oxidation No.

Hydrogen	H
Lithium	Li
Sodium	Na
Potassium	K
Rubidium	Rb
Cesium	Cs
Silver	Ag
Copper(I)	Cu
Mercury(I)	Hg
Ammonium	NH_4

+2 Oxidation No.

Beryllium	Be
Magnesium	Mg
Calcium	Ca
Strontium	Sr
Barium	Ba
Chromium(II)	Cr
Manganese(II)	Mn
Iron(II)	Fe
Cobalt(II)	Co
Nickel(II)	Ni
Copper(II)	Cu
Zinc	Zn
Cadmium	Cd
Mercury(II)	Hg
Tin(II)	Sn
Lead(II)	Pb

+3 Oxidation No.

Boron	B
Aluminum	Al
Nitrogen	N
Phosphorus	P
Arsenic(III)	As
Antimony(III)	Sb
Bismuth(III)	Bi
Chromium(III)	Cr
Iron(III)	Fe
Cobalt(III)	Co
Nickel(III)	Ni

+4 Oxidation No.

Carbon	C
Silicon	Si
Tin(IV)	Sn
Lead(IV)	Pb
Manganese(IV)	Mn

+5 Oxidation No.

Nitrogen	N
Phosphorus	P
Arsenic(V)	As
Antimony(V)	Sb
Bismuth(V)	Bi

NOTE: Certain elements such as copper, iron, tin, and others exhibit more than one oxidation number. In these cases, the name of the element is followed by the oxidation number in roman numerals in parenthesis. The case of nitrogen and phosphorus are discussed in Section IV.

+1
NH_4

-1 Oxidation No.

Fluoride	F
Chloride	Cl
Bromide	Br
Iodide	I
Hypochlorite	ClO
Chlorite	ClO_2
Chlorate	ClO_3
Perchlorate (also Br and I)	ClO_4
Bicarbonate	HCO_3
Bisulfite	HSO_3
Bisulfate	HSO_4
Nitrite	NO_2
Nitrate	NO_3
Cyanide	CN
Cyanate	CNO
Thiocyanate	CNS
Hydroxide	OH
Permanganate	MnO_4
Acetate	$C_2H_3O_2$

-2 Oxidation No.

Oxide	O
Sulfide	S
Peroxide	O_2
Sulfite	SO_3
Sulfate	SO_4
Thiosulfate	S_2O_3
Dithionite	S_2O_4
Dithionate	S_2O_6
Persulfate	S_2O_8
Carbonate	CO_3
Oxalate	C_2O_4
Chromate	CrO_4
Dichromate	Cr_2O_7
Silicate	SiO_3
Stannate	SnO_3
Tetraborate	B_4O_7

-3 Oxidation No.

Nitride	N
Phosphide	P
Phosphite	PO_3
Phosphate	PO_4
Arsenate	AsO_4
Ferricyanide	$Fe(CN)_6$

-4 Oxidation No.

Perphosphate	P_2O_6
Pyrophosphate	P_2O_7
Ferrocyanide	$Fe(CN)_6$

1A	2A	3B		4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	No Ga
1 H +1 -1																		2 He
3 Li +1	4 Be +2												5 B +3	6 C +4 +2 -4	7 N +5 +4 +3 +2 -3	8 O -1 -2	9 F -1	10 Ne
11 Na +1	12 Mg +2												13 Al +3	14 Si +4 -4	15 P +5 +3 -3	16 S +6 +4 +2 -2	17 Cl +7 +5 +3 +1 -1	18 Ar
19 K +1	20 Ca +2	21 Sc +3		22 Ti +4 +3 +2	23 V +5 +4 +3 +2	24 Cr +6 +3 +2	25 Mn +7 +6 +4 +3 +2	26 Fe +3 +2	27 Co +3 +2	28 Ni +2 +1	29 Cu +2 +1	30 Zn +2	31 Ga +3	32 Ge +4 +2	33 As +5 +3 -3	34 Se +6 +4 -2	35 Br +7 +5 +1 -1	36 Kr +4
37 Rb +1	38 Sr +2	39 Y +3		40 Zr +4	41 Nb +5 +4	42 Mo +6 +4 +3	43 Tc +7 +6 +4	44 Ru +8 +6 +4 +3	45 Rh +4 +3 +2	46 Pd +4 +2	47 Ag +1	48 Cd +2	49 In +3	50 Sn +4 +2	51 Sb +5 +3 -3	52 Te +6 +4 -2	53 I +7 +5 +1 -1	54 Xe +8 +6 +4 +2
55 Cs +1	56 Ba +2	57 La +3	58 - 71 Ce - Lu +3	72 Hf +4	73 Ta +5	74 W +6 +4	75 Re +7 +6 +4	76 Os +8 +4	77 Ir +4 +3	78 Pt +4 +2	79 Au +3 +1	80 Hg +2 +1	81 Tl +3 +1	82 Pb +4 +2	83 Bi +5 +3	84 Po +4	85 At -1	86 Rn
87 Fr +1	88 Ra +2	89 Ac +3	90 - 103 Th - Lr															

Problems

1. Determine the oxidation number of S in each of the following compounds.

a) $\overset{+}{Na_2}S_2O_3^{2-}$ $2(1)+2S+3(-2)=$ $2S=4$ ans. a) 2

b) $\overset{+}{H_2}SO_3^{-2}$ b) 4

c) SO_2^{-3} c) 6

d) $\overset{+}{K_2}S_2O_4^{-3}$ d) 5

e) $\overset{3+}{Al_2}S_3$ e) -2

f) $\overset{2+}{Ba}S_2O_8^{-3}$ f) 11

2. Name the following compounds.

a) $\overset{+2}{Pb}I_2^{-}$ ans. a) lead (II) iodide

b) $Fe(SO_4)^{2-}$ b) iron (II) sulf~~ide~~ ate ←

c) $\overset{2+}{Ag}(CO_3)^{2-}$ c) silver (II) carbon~~ide~~ ate ←

d) $\overset{+}{Na}(CN)^{-}$ d) sodium cynide

e) $\overset{2+}{Ca}(C_2H_3O_2)_2^{2-}$ e) calcium actate ← don't Δ ending w/ poly-atomic ion

f) $\overset{2+}{Cu}(NO_3)_2^{2-}$ f) copper (II) nitr~~ide~~ ate ←

g) $\overset{2+}{K_2}(C_2O_4)^{2-}$ g) potassium oxal~~ide~~ ate

h) $HgCl^{-}$ h) mercury (I) cloride

3. Write formulas for the following compounds.

NH_4^{+} S^{2-}

a) ammonium sulfide ans. a) $(NH_4)_2S$

Mg^{2+} PO_4^{3-}

b) magnesium phosphate b) $Mg_3(PO_4)_2$

Hg^{2+} CNS^{-}

c) mercury(II) thiocyanate c) $Hg(CNS)_2$

Na^{+} I^{-}

d) sodium iodate d) NaI

Cr^{3+} Cl^{-}

e) chromium(III) chloride e) $CrCl_3$

K^{+} MnO_4^{-}

f) potassium permanganate f) $K(MnO_4)$

Zn^{2+} Br^{-}

g) zinc bromide g) $ZnBr$

Co^{+2} ClO_4^{-}

h) cobalt(II) perchlorate h) $Co(ClO_4)_2$

4. Determine the oxidation number of Cr in each of the following compounds.

a) $\overset{2+}{Ca}Cr\overset{2-}{O_4}$ ans. a) 6

b) $Cr\overset{-}{Br_2}$ b) 2

c) $\overset{+}{Ag_2}Cr_2\overset{2-}{O_7}$ c) 6

d) $Cr_2(\overset{2-}{SO_4})_3$ d) 3

e) $\overset{+}{Li_2}Cr\overset{2-}{O_4}$ e) 6

5. Name the following acids and bases.

a) H_2SO_3 ans. a) sulfous acid

b) $\overset{4+}{Sn}\overset{-}{(OH)_4}$ b) tin (IV) hydroxide

c) HNO_3 c) nitric acid

d) KOH d) potassium hydroxide

e) HIO_4 e) periodic acid

f) HF f) hydrofloric acid

g) $\overset{3+}{Fe}\overset{-}{(OH)_3}$ g) iron (III) hydroxide

h) $H_2\overset{2-}{SO_4}$ h) sulfic acid

i) $H_3\overset{3-}{PO_3}$ i) phosphous acid

6. Write formulas for the following acids and bases.

a) nitrous acid ans. a) $H(NO_2)$

b) phosphoric acid b) $H(PO_4)$

c) sodium hydroxide c) $Na(OH)$

d) bromic acid d) $H(BrO_3)$

e) tin(II) hydroxide e) $Sn(OH)$

f) hydroiodic acid f) $H(I)$

g) hypobromic acid g) $H(BrO)$

h) aluminum hydroxide h) $Al(OH)$

i) zinc hydroxide i) $Zn(OH)$

7. Name the following compounds.

a) $BaCrO_4$ ans. a) barium per Cchromate
b) $\overset{+2}{Ni_2}Fe\overset{-4}{(CN)_6}$ b) nickel ferrocyanide
c) HBr c) hydro-bromic acid
d) $KCNO$ d) pot assium cynate
e) H_2O_2 e) hydro-oxyic acid
f) $AlPO_4$ f) aluminum per potassiate
g) CuO g) hypocoppite
h) $\overset{+2}{Pb}\overset{-1}{(C_2H_3O_2)_2}$ h) lead acetate
i) PCl_3 2 non metals i) mono-phosphorus tri-chloride
j) $\overset{+1}{NH_4}\overset{-1}{CN}$ ← 2 ions straight name → j) ammonium cyanide
k) $\overset{+2}{Ni}C_2O_4$ k) nickel oxalate
l) $\overset{+1}{Na_2}\overset{-2}{SiO_3}$ l) sodium silicate
m) $\overset{+2}{Ca}(BrO_4)_2$ m) calcium per-bromate
n) $\overset{+1}{Ag}MnO_4$ n) silver permanganate
o) $\overset{+2}{Sn}\overset{-1}{F_2}$ o) tin (II) floride
p) As_2S_3 2 non metals p) di-arsenic tri-sulfide
q) $\overset{-1}{Na_2}\overset{+2}{O}$ q) sodium oxide
r) $\overset{+2}{Mg}\overset{-1}{(IO_3)_2}$ r) magnesium iodate
s) $\overset{+2}{Hg_2}\overset{-1}{SO_4}$ s) mercury (II) per sulfate
t) H_3AsO_4 t) per arsenic acid
u) $\overset{+2}{Co}\overset{-1}{Cl_2}$ u) cobalt (II) cloride
v) $NaClO$ v) sodium hypochlorite
w) $\overset{+4}{C}\overset{-2}{S_2}$ w) carbon sulfide
x) $\overset{+1}{(NH_4)_2}\overset{-1}{SO_3}$ x) ammonium sulfate
window → y) $\overset{+3}{Bi}\overset{-1}{(OH)_3}$ y) bismuth hydroxide
z) $\overset{+2}{Fe}\overset{-2}{S_2O_3}$ z) iron (II) thiosulfate

8. Write formulas for the following compounds.

a) $\overset{+3}{\text{chromium(III)}}$ $\overset{-1}{\text{nitrate}}$ ans. a) $Cr(NO_3)_3$

b) $\overset{+2}{\text{manganese(II)}}$ $\overset{-1}{\text{hydroxide}}$ b) $Mg(OH)_2$

c) $\overset{\times\ +3}{\text{nitrogen}}$ triiodide binary molecular c) NI_3

d) $\overset{+1}{\text{sodium}}$ $\overset{-1}{\text{tetraborate}}$? d) $NaBrO_3$

e) $\overset{+2}{\text{zinc}}$ $\overset{-4}{\text{carbonate}}$ e) Zn_2C

f) $\overset{+1}{\text{ammonium}}$ $\overset{-1}{\text{nitrite}}$ f) NH_4NO_3

g) $\overset{+2}{\text{magnesium}}$ $\overset{-2}{\text{oxalate}}$ g) MgC_2O_4

h) $\overset{+2}{\text{copper(II)}}$ sulfite h) Cu_2SO_3

i) $\overset{+4}{\text{silicon}}$ $\overset{-4}{\text{dioxide}}$ i) SiO_2

j) $\overset{+2}{\text{lead(II)}}$ $\overset{-2}{\text{chromate}}$ j) $PbCr$

k) $\overset{+1}{\text{silver}}$ $\overset{-1}{\text{cyanide}}$ k) $AgCN$

l) $\overset{+}{\text{sodium}}$ bicarbonate l) $NaHCO_3$

m) $\overset{+2}{\text{calcium}}$ $\overset{-2}{\text{dithionate}}$ m) CaS_2O_6

n) $\overset{+3}{\text{antimony(III)}}$ $\overset{-2}{\text{sulfide}}$ n) Sb_2S_3

o) $\overset{+}{\text{potassium}}$ $\overset{-2}{\text{oxide}}$ o) K_2O

p) $\overset{3+}{\text{boron}}$ $\overset{3-}{\text{trifluoride}}$ p) BrF_3

q) $\overset{+4}{\text{tin(IV)}}$ $\overset{-1}{\text{nitrate}}$ q) $Sn(NO_3)_4$

r) $\overset{+2}{\text{barium}}$ $\overset{-1}{\text{chloride}}$ r) BaC_2

s) $\overset{+3}{\text{aluminum}}$ $\overset{-1}{\text{acetate}}$ s) $Al(C_2H_3O_2)_3$

t) $\overset{+1}{\text{copper(I)}}$ $\overset{-2}{\text{oxide}}$ t) Cu_2O

u) $\overset{+2}{\text{manganese(II)}}$ $\overset{-4}{\text{pyrophosphate}}$ u) $Mg_2(P_2O_7)$

v) $\overset{+3}{\text{chromium(III)}}$ $\overset{-2}{\text{sulfate}}$ v) $Cr_2(SO_4)_3$

w) $\overset{+}{\text{lithium}}$ $\overset{-}{\text{hydride}}$? w) LiH

x) $\overset{+2}{\text{iron(II)}}$ $\overset{-3}{\text{phosphate}}$ x) Fe_3P_2

y) $\overset{+}{\text{ammonium}}$ $\overset{-2}{\text{oxalate}}$ y) $(NH_4)_2C_2O_4$

z) $\overset{+2}{\text{mercury(II)}}$ $\overset{-1}{\text{iodate}}$ z) $Hg(IO_3)_2$

9. Name the following compounds.

a) $\overset{+2}{K_2}\overset{-2}{S_2O_8}$	ans. a)	potassium persulfate
b) $\overset{+2}{Mg_3}\overset{-3}{N_2}$	b)	magnesium nitride
c) HIO	c)	hypochlorous acid
d) $Sr(OH)_2$	d)	strontium hydroxide
e) Na_3PO_3	e)	sodium phosphite
f) $Ag_2Cr_2O_7$	f)	silver dichromate
g) $\overset{+2}{Cd}CO_3$ -2	g)	cadmium (II) carbonate
h) $HC_2H_3O_2$	h)	acetic acid
i) $LiHSO_4$	i)	lithium bisulfate
j) $\overset{+2}{Ni_2}P_2O_7$ -4	j)	nickel (II) pyrophosphate
k) $\overset{+3}{As}P$ 3 ?need mono →	k)	mono-arsenic $\overset{mono}{\text{phosphoide}}$
l) $\overset{+5}{Sb_2}\overset{-2}{S_3}$	l)	di-antimony tri-sulfer
m) $HBrO_4$	m)	perbromic acid
n) $\overset{+2}{Mn}C_2O_4$ -2	n)	maganese (IV) oxalate
o) $Co(ClO_4)_2$ -1	o)	cobalt (II) perchlorate
p) H_2S	p)	hydrosulfuric acid
q) BF_3	q)	boron floride
r) $NaClO_2$	r)	sodium chlorite
s) $PbSO_4$	s)	lead persodiate
t) $H_2C_2O_4$	t)	oxalic acid
u) $\overset{+}{Cu}Cl$ 1	u)	copper (I) cloride
v) $\overset{+2}{Ba}\overset{-2\,-2}{O_2}$	v)	barium oxyide
w) $HClO$	w)	hypochlorous acid
x) $\overset{+}{Rb}\overset{-}{OH}$	x)	rubidium hydroxide
y) $\overset{+2}{C}\overset{-2}{O}$	y)	dicarbon di oxide
z) $\overset{+3}{P}\overset{-3}{I_3}$	z)	mono-phosphorus tri-iodide

Types of Chem Reactions Notes

5/17

1. Synthesis → 2 reactants → 1 product

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$

2. Decomposition → 1 reactant → 2+ products

$$2H_2O_{(l)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

3. Single Displacement → element replaces an element in a compound

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

4. Double Displacement → ⊕ and ⊖ ions switch

$$\overset{+}{A}\overset{-}{B}_{(aq)} + \overset{+}{C}\overset{-}{D}_{(aq)} \rightarrow \overset{+}{A}\overset{-}{D}_{(aq)} + \overset{+}{C}\overset{-}{B}_{(aq)}$$

* you must form water, a precipitate or a gas

Michael Plasmeier

Electrons Determine Chem Behavior Lab Activity 7

20/20

4/18

to investigate and relate atom's electron patterns according to chemical + physical properties to find periodic table arrangment

How does the arrangment in electrons determain its chem behavior?

The arrangment of electrons in subleves determains its chemical properties because ~~depending on~~ how many atoms are in the last filled sublevel deteimis how easily at atom gains, loses, attracts, and combines with other electrons/atoms. It also deteimis the elements volatilty (inc. how ealsly it explodes.)

1. In Act. 6 you saw that atoms with relativly high ionization energies are stable + have stable arrangments of electrons. These are noble gasses.

	"A"	"B"	"C"	"D"
	energy level last electron	sublevel last electron	# electrons in last sublevel (# valence electrons)	total # electrons of "A" energy level
He	1	s	2	2
Ne	2	p	6	8
Ar	3	p	6	8
Kr	4	p	6	8
Xe	5	p	6	8
Rn	6	p	6	8

a. How "A" related to respective row

"A" (energy level) is equal to the row or period of each respective element.

→

b. What pattern in "B" and "C"?
Except for the first period (He), there are always 6 electrons in the last sublevel "p".

c. What pattern in "D"?
Except for the 1st period (He), there are always 8 electrons in each respective energy level.

2. The chem behavior can be seen by arrangment of atoms related to other nearby noble gasses. Chem inactivity of noble gasses is because stable arrangment of electrons.

	# electrons > (He)	energy level last electron	energy sublevel last electron	period on table	group on table
Li	1	2	s	2	1
Be	2	2	s	2	2
B	3	2	p	2	13

	# electrons > (Ne)	energy level last electron	energy sublevel last electron	period on table	group on table
Na	1	3	s	3	1
Mg	2	3	s	3	2
Al	3	3	p	3	13

d) Any patterns you notice?
As you move from left to right in a period, you add electrons

"more"

	# electrons > Ne	energy level last electron	energy sublevel last electron	period on table	group on table
N	-3	2	p	2	15
O	-2	2	p	2	16
F	-1	2	p	2	17

	# electrons > Ar	energy level last electron	energy sublevel last electron	period on table	group on table
P	-3	3	p	3	15
S	-2	3	p	3	16
Cl	-1	3	p	3	17

f) Any patterns you notice?
The next period down has a higher energy level (1 exactly higher)

3. Draw simple period table w/ 1st 3 periods

	1	2	13	14	15	16	17	18
1	H							He
2	Li +1	Be +2	B +3	C ±4	N -3	O -2	F -1	Ne
3	Na	Mg	Al	Si	P	S	Cl	Ar

d) Carbon is farthest from a stable noble gas.

4. Electrons are added to s + p sublevels of the period they're in. The # of the energy level is significant because the ↑ the #, the farther from the nucleus. The electrons on the outermost sublevel determain how close it is to a noble gas and are responsible for the atom's chem behavior. They are often called valence electrons. You can have max 8.

a. How many valence electrons are in:?

Group			
18	Noble Gasses	He - 2	Kr - 8
		Ne - 8	Xe - 8
		Ar - 8	Rn - 8
1	alkali	Li - 1	Rb - 1
		Na - 1	Cs -
		K - 1	
2	alkaline earth	Be - 2	Sr - 2
		Mg - 2	Ba - 2
		Ca - 2	
13		B - 3	Al - 3
14		C - 4	Si - 4
15		N - 5	P - 5
16	chalcogens	O - 6	S - 6
17	halogens	F - 7	Br - 7
		Cl - 7	I - 7

Reading Guide for Chem Talk "The Noble Gases"

READ p.123-4

1.In your own words, explain how Argon was discovered.

Lord Rayleigh made 1L N two different ways but they weiged different amounts. Ramsay looked at it the gasses in a spectrometer and saw that some ports where different. This was Argon.

2. What role did mass play in this discovery?

They showed that 2 quantites were not the same. Atoms of different elements have different masses.

3. What role did the inactivity of Argon play?

The inactivity of Argon had it escape notice

4.When were most of the other noble gases discovered?

Ar - 1894
He - 1895
Ne - 1898
Kr - 1898

5. What is different about the elements we call noble gases?

They don't react, or react only in unusual circumstances.

4, Mg: $1s^2 2s^2 2p^6 3s^2$ -2 valence electrons

Ba: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$

-2 valence electrons ← same

The # valance electrons are = for all elements in group 2 (alkaline earth metals)

5, Co: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$

9 valance electrons

6. Why are they called rare?

Not eaisly found

7. What does inert mean?

Not eaisly chemically reactive

Do Chem to Go p124-125

1, Indicate excess/dificiency of electrons compared to nearest noble gas

a - Ca +2

b - As -3

c - K +1

d - I -1

2, Group the — two similar electrons

a C, N, Si - C+Si same group (15)

b F, Cl, Ne - F+Cl same group (17)

c S, Br, O - S+O same group (16)

d Na, Mg, S - Na+Mg - both metals

e He, Ne, H, He+Ne - same group (18)

3, From each pair which more stable?

a - He+Li - He The ones with more electrons in highest

b - Li+Be - Be energy level (↑ group #) are more stable

c - Mg+Cl - Cl

d - Mg+Ar - Ar

e - Ne+Kr - Ne less electrons = higher marginal 1st ionization energy

Michael Plasmeier

Weight Relationship in Chem Changes Lab

5/8

to determine % of an element in a compound

1. Weigh a crucible 9,86 g

2. Add approx 20 cm Mg to crucible 10,04 g

3. Heat until reaction is complete. Cool. 10,15 g

Mass of metal	,18	g	(#2 - #1)
Mass of product	,29	g	(#3 - #1)
Mass of reactant	,11	g	(#3 - #2)

Moles of metal = $\frac{\text{g metal}}{24,3}$,007407

Moles of oxygen $\frac{\text{g oxygen}}{16}$,006875

Formula Units subscripts $\frac{\text{moles non metal}}{\text{moles metal}} = x$

round (x) to nearest whole #

Mg_xO_x

$Mg_1O_1 \leftarrow \frac{,006875}{,007407} = ,928125$

$\boxed{MgO}$

20/20

Questions

1. State the formula for the compound formed.
MgO

2. What is the % Mg? % O?
$Mg = \frac{.18}{.29} = 62\%$ $O = \frac{.11}{.29} = 38\%$

3. Why did we round to whole # for formula subscripts?
Atoms combine in whole # ratios.

Chemical Changes
Activity 2

20/20

5/9

to observe seveal chemical changes and draw conclusions about chemical reactions

There are 2 types of reactions
A baking mix in an oven
$Na_2CO_2 + NaHSO_4$
Which results in a chemical reaction? - Both?
The two chemicals will combine and change
Similarity + Differences between 2 reactions
One (the oven) requires heat as a catylst.
They both have a finished product whos properties are different from every reactant.

1. There are 8 solutions. Mix each one (3 drops) togeter sepertly and record changes like color, precipate, gas, Δ temp.

See chart next page
- barium nitrate ($Ba(NO_3)_2$)
- sodium hydroxide ($NaOH$)
(Na_2CO_3)
- copper (II) sulfate ($CuSO_4$)
- potassium iodine (KI)
- silver nitrate ($AgNO_3$)
($FeCl_3$)
- hydrochloric acid (HCl)

3g. What are evidences of chem changes?
- heat
- bubbles (gas)
- color change (?)
- precipitate formation

3. a) Which combo seemed to produce no reaction

$Ba(NO_3)_2$ + $AgNO_3$
" + HCl
$NaOH$ + Na_2CO_3
" + $AgNO_3$
+ $FeCl_3$
+ HCl
Na_2CO_3 + KI
+ $AgNO_3$
+ HCl
$AgNO_3$ + HCl

(I forgot to record if colored becomes clear - so listed as "no reacting")

b) Which formed a gas
I believe none formed a gas

c) Which produced a color change
Most did, All of the Ones not listed in #1

d) Which formed precipitates?
KI + $CuSO_4$ did, I did not keep track of the time period to do this. Also see below

e) Yellow / brown precipitates
$CuSO_4$ + KI - brown KI + $FeCl_3$ - brown
$CuSO_4$ + $FeCl_3$ - yellow
Blue precipitates
$Ba(NO_3)_2$ + $CuSO_4$
Na_2CO_3 + $CuSO_4$

f) Which produced heat?
None which I could tell. →

$Ba(NO_3)_2$	clearish	white cloudy	blue cloudy liquid	slightly yellow liquid	clear	light yellow	clear
clear liquid bluish	NaOH	clear	dilutes blue liquid	dilutes existing yellow liquid	clear	darker yellow liquid	clear
clear liquid with white clouds	clear	Na_2CO_3	dark blue oily liquid	clear liquid	clear	clear liquid	clear
cloudy blue liquid	light blue	dark blue oily	$CuSO_4$	dark brown oily	light blue	yellow blue liquid turns dark brown	light blue
Yellowish clear	light yellow	mid yellow	dark brown oily liquid redish	kI	light yellow	darker yellow liquid	light yellow
clear	clear	clear	light blue	light yellow	$AgNO_3$	very light yellow	clear
light yellow liquid	darker yellow liquid	clearish	dark yellow swirls	yellow	light yellow	$FeCl_3$	light yellow
clear	clear	clear	light blue	light yellow	clear	light yellow	HCl

Activity 2

Chem Talk +
Chem to Go

5/9

Checking Up

p161 1. What is a chem test?
A chem test is a physical procedure or chemical reaction used to identify a substance.

2. Describe how you can use a burning or glowing splint to test for H or O?
When a burning splint is exposed to a concentrated amount of H, the splint is relit with a loud pop. In the presence of significant amounts of O, the glowing splint will burst into flames.

3. Why does this not work with CO_2?
CO_2 does not burn or support the burning of the splint. CO_2 extinguises the fire if exposed for a significant amount of time.

4. What can you do to detect CO_2?
You blow bubbles (CO_2) into limewater (Calcium hydroxide) If the water turns cloudy, a precipitate has formed and you know that the gas is CO_2

5. What is a precipitate?
Solids which form in liquids during a chemical change.

6. What are acid-base indicators and how are they useful?
A die which has different colors depending on if it is mixed with an acid or a base.

Chem to Go

p162 1, When you mix something and the color changes, that might not always represent a chemical change. For example kool-Aid mix in water. Are these chem changes

a) acid disolves in water releasing heat
Yes

b) burning match
Yes

c) crystal in saline solution
?? No

d) CO_2 released when soda opened
No

e) Glowing filiment in lightbulb
No

f) Metal in acid releasing H
Yes

g) Mixing $NaOH + CuSO_4$ which forms a precipitate
Yes

2, Anhydrous copper (II) sulfate ($CuSO_4$) is a white solid. When it disolves in water, it becomes blue. Is this a chem change?

Yes, the chemicals are reacting in water producing a change in state and color.

? how to tell more

3. When soda becomes flat, is this chem or physical change?

It's a physical change, the gas is escaping ? No other compounds are being produced.

Michael Plasroki

Chemical Names and Formules Lab 3

20/20

5/10

to find out how to name ionic compounds

What information (and significance thereof) does the periodic table give for every element?

It gives the name and abbriviation first of all for every element. Also listed is the electron configeration, atomic # (# protons [electrons]) and average atomic mass (of each isotype) [protons + neutrons]

1, The periodic table lists elements in order of increasing atomic numbers. (# protons [electrons]) write the atomic number of these elements

a) Cu - 29
b) S - 16
c) Zn - 30
d) Au - 79
e) O - 8
f) C - 6
g) Ag - 47
h) Cl - 17
i) N - 7
j) H - 1
k) Mg - 12
l) I - 53
m) Fe - 26
n) Ca - 20
o) Al - 13
p) Na - 11
q) k - 19
r) Pb - 82

2. Elements can combine to form compounds, Some compounds are made of ions of different elements, which form by gaining or losing electrons,

a) Group 1 (+1) + Group 17 (-1) ions have names like kBr, RbCl, LiF, kI, and CsBr

b) Group 16 (+2) + Group 16 (-2) ions also have simple names like MgO, CaS, CaO, SrSe, and BeTe

3, If an ion loses a certain number of electrons and another atom gains those electrons, However multiple ions may be required to give off / absorb an equal number of electrons.

3. Write the formulas when these elements are combined

a) calcium + oxygen → $Ca^{2+}O^{2-}$ → CaO

b) aluminum + florine → $Al^{3+}F^{-}$ → AlF_3

c) boron + oxygen → $B^{3+}O^{2-}$ → O_2B_3

d) stontium + nitrogen → $Sr^{2+}N^{3-}$ → Sr_3N_2

e) barium + selenium → $Ba^{2+}Se^{2-}$ → $BaSe$

4. Some compands like sodium hydrogen carbonate (baking soda) [$NaHCO_3$] have polyatomic ions. These are made up of several atoms joined togeter. [The HCO_3 has a charge of -1]

a) potassium nitrate → $K^{+}NO_3^{-}$ → KNO_3

b) barium sulfate → $Ba^{2+}SO_4^{2-}$ → $BaSO_4$

c) potassium sulfate → $K^{+}SO_4^{2-}$ → K_2SO_4

d) sodium acetate → $Na^{+}C_2H_3O_2^{-}$ → $NaC_2H_3O_2$

e) $(NH_4)_2SO_4$ → ammonium sulfate

f) $Al_2(CO_3)_3$ → aluminum carbonate

g) $LiHCO_3$ → lithium hydrogen carbonate

h) HNO_3 → hydrogen nitrate

5. In Molecules electrons are shared through convalent bonds. You can imagine the charges called oxidation numbers

a) Carbon Dioxide is CO_2 What is the oxidation number?

$C^{4-}O_2^{2+}$

b) Oxidation #s for carbon monoxide (CO)

$C^{4-}O^{2+}$

c) Explain how I found this.

Basically I assumed it was an ionic compound and looked up its charges. There seem to be much more complicated rules on p 639 red book (Section 20.2)

6. Do chemical changes occur each time reactants are mixed?
Place equal amounts of baking soda, crushed Alka-Seltzer tablets and baking power in 3 test tubes. Add water. Place a glowing splint at the top of each tube. What happens?
Carbon Dioxide (CO_2) is produced each time extinguing the glowing splint.

7. Place to tubes into boling water. What happens?
The same reaction as before happens, but faster.

8. Try with vinegar and ammonia.
↑water ↑ammonia asotate

Lab 3 Chem Talk +
Chem to Go

5/11

Chem Talk: Checking up

p166

1. Why are there so many materials in the world if there are only ~100 elements?
   Elements can combine to form compounds. There are many different combonations.

2. What is an ion?
   An electrically atom/group (atoms) that has aqcired a net charge (⊕ or ⊖)

3. How are ionic compounds formed?
   Anions easily lose electrons which cations easily gain. These ions bond in ionic bonds to form ionic compounds.

4. What's a polyatomic ion?
   An ion which consists of 2 or more atoms which are convalently bonded and have either a negitive or positive charge. For example: sulfate SO_4^{2-}

5. How are molecular compounds formed?
   They are formed when electrons are attracted to two nuclei and are shared between the two atoms.

6. Distinguish b/w an ionic + convalent bond?
   Convalent bonds are normaly between 2 non metels while ionic bonds are between a metal + non metal.

Chem to Go

p168 1. Write Chem formula

a) $Na^+ + Br^- \to NaBr$
b) $K^+ + S^{2-} \to K_2S$
c) $Mg^{2+} + Cl^{2-} \to MgCl_2$
d) $Cs^+ + I^- \to CsI$
e) $Al^{3+} + O^{2-} \to Al_2O_3$

2. Write name:

a) nitric acid → HN
b) ammonium hydroxide → $NH_4^+(OH)^-$
d) cesium + iodine → CsI
c) magnesium^{2+} + chlorine$^-$ → MgC_2
e) aluminum^{3+} + oxygen^{2-} → Al_2O_3

3. a) copper^{2+}(II) sulfate^{2-} → CuS
b) iron^{2+}(II) oxide^{2-} → FeO
iron^{3+}(III) oxide^{2-} → Fe_2O_3

4. What is the charge for Group 2: +2

5. Sodium phosphate = Na_3PO_4. What is charge on polyatomic
3−, I know from my sheet. Also charge atom; must balance: $Na^+_3(PO_4)^{3-}$

6. Sodium hydroxide - $Na^+ OH^-$
aluminum hydroxide - $Al^{3+}(OH)^-_3$ ← Parenthese needed because you need 3 groups to balance the equation

7. Formula:

a) iron^{3+}(III) nitrate^{1-} → $Fe(NO_3)_3$
b) lead^{2+}(II) sulfate^{2-} → $PbSO_4$
c) silver^{1+} nitrate$^-$ → $AgNO_3$
g) iron(III) nitrate $Fe(NO_3)_3$
h) hydrocloric acid HCl

8. a) Barium nitrate → $Ba(NO_3)_2$
b) sodium hydroxide → $NaOH$
c) sodium hydrogen carbonate → $NaHCO_3$
d) copper (II) sulfate → $CuSO_4$
e) potassium iodine → KI
f) silver nitrate → $AgNO_3$

Michael Plesner

Chemical Equations
Activity 4

20/20

5/16

to represent different classes of chemical reactions using chemical equations

How cald you communicate about the reactants and products using chemical equations
You could write the reactants on one side and products on the other side with an arrow seperating them.

1. Your teacher added 45mL of 2.0M ammonium hydroxide solution to A-20 drops phenolphthalein solution
B-15 drops magnesium sulfate solution
C-15 drops 1M copper (II) sulfate
A-turns pink
B-turns cloudy white
C-turns oily purple / blue
a) Yes a chemical change took place in every reaction, I could see this by noticing a color change.

3. There are 4 main types of reaction
- synthesis reactions
- decomposition reaction
- single - displacement reaction
- double - displacement reaction

a) Synthesis means a false change ? combination.
Decomposition is taking things apart

4. A synthesis reaction is when 2 or more chemicals combine to form a compand
$A + B \rightarrow AB$
Magnesium + Oxygen → magnesium oxide

a) What are the advantages of writing a reaction using chemical symbols?

It is shorter and you can see how many atoms are in each molecle.

b-g) The number of atoms must be = before and after the reaction. What about $Mg + O \rightarrow MgO$?

$2Mg^{+} + O_2^{-2} \rightarrow 2MgO$ ← 2 magnesium + 2 oxygen

↑2 magnesium ← 2 oxygen

h) The equation is now balanced. There are 4 atoms on each side. You can see that matter (and mass) are conserved, because no matter (mass) is created or distroyed.

g) You should write it using letters in subscript to indicate state. Rewrite our equation

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

5. Write a word + balanced chem equation

a) carbon + oxygen → carbon dioxide

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

b) hydrogen reacts with oxygen to form water

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

c) iron reacts with oxygen to form iron (III) oxide

$2Fe_{(s)} + O_{2(g)} \rightarrow 2FeO_{(s)}$

+2·2 -2·2 ↑1 oxygen

d) sodium solid + chlorine gas produces sodium chloride soli

$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$

6. Water can be seperated into its forms with energy

$2H_2O_{(l)} + \text{energy} \rightarrow 2H_{2(g)} + O_{2(g)}$

↑4 hydrogen + 2 oxygen ↑4 hydrogen ↑ 2 oxygen

a) Yes the equation is balanced. You can tell by counting the # of atoms of each element.

When a substances breaks down into its parts it is called decomposition.

$AB \rightarrow A + B$

Write word + balanced formula:

b) sodium chloride → sodium + clorine

$\underset{+}{Na}\underset{-2}{Cl_{2(s)}} \rightarrow Na_{(s)} + Cl_{2(g)}$

c) potassium iodide → potassium + iodine

$kI_{(s)} \rightarrow k_{(s)} + I_{(s)}$

d) magnesium bromide → magnesium + bromine

$MgBr_{2(s)} \rightarrow Mg_{(s)} + Br_{2(l)}$

7. The reactions mentioned above involve compands forming or breaking up. Other types of reactions involve elements reacting with compands to form products.

$Zn_{(s)} + 2HCl_{(aq)} \rightarrow H_{2(g)} + ZnCl_{2(aq)}$

1 Zn
2 H
2 CL

before + after - so balanced

This is a single displacement reaction

$A + BC \rightarrow B + AC$

Write word + chemical formulas

b) iron is added to copper (II) sulfate to produce iron sulfate + copper

$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$

c) lead is added to silver nitrate to produce lead nitrate + silver

$Pb_{(s)} + AgNO_{3(aq)} \rightarrow Pb(NO_3)_2 + Ag$

d) aluminum foil in copper (II) hydroxide to produce aluminum hydroxide + copper

$Al_{(s)} + Cu(OH)_{2(aq)} \rightarrow Al(OH)_{3(aq)} + Cu_{(s)}$

8. Another type of reaction is a double displacement reaction. Some solutions have been put out. Mix 3 drops of each solution as indicated by the chart.

initial

kI - light yellow liquid
Na_2CO_3 - clear liquid
$FeCl_3$ - orange-yellow liquid
$NaOH$ - clear liquid
$AgNO_3$ - dark, dirtly looking gray liquid
$CuSO_4$ - light blue liquid

Table	Silver Nitrate $AgNO_3$	Copper (II) Sulfate $CuSO_4$	Magnesium Sulfate $MgSO_4$	Sodium Hydroxide $NaOH$
Sodium Carbonate Na_2CO_3	① dilutes	② blue cloudy	③ dilutes	④ dilutes
Sodium Hydroxide $NaOH$	⑤	blue cloudy ↑ light		X
Potassium Iodide kI	⑨	dark dirty brown		dilutes
Iron (III) Chloride $FeCl_3$	⑬	turns darker yellow		turns darker orange cloudy

11. Look at your data. Do you think a chemical change took place each time?
No, it seemed that some times (about half) only diluted the colored liquid. Only when it changed colors could I really tell a chemical change took place. I couldn't tell temp changes.
Pretty much every time

b) Could you predit what products would be formed?
Not by looking at it but I may have been able to rememba results from past labs. However if you know the rules of Chemistry you could write out the reactions and predict the product.

c) Write the word + chemical reaction for each reaction
1. silver nitrate + Sodium carbonate →
silver carbonate + sodium nitrate
$AgNO_3 + Na_2CO_3 \rightarrow Ag_2CO_3 + NaNO_3$
+1 −1 +2 −2 +2 −2 +1 −1

2. copper (II) sulfate + sodium carbonate →
copper (II) carbonate + sodium sulfate
$CuSO_4 + Na_2CO_3 \rightarrow CuCO_3 + Na_2SO_4$
+2 −2 +2 −2 +2 −2 +2 −2

3. magnesium sulfate + sodium carbonate →
magnesium carbonate + sodium sulfate
$MgSO_4 + Na_2CO_3 \rightarrow MgCO_3 + Na_2SO_4$
+2 −2 +2 −2 +2 −2 +2 −2

4. sodium hydroxide + sodium carbonate →
no reaction - both soluble
$NaOH + Na_2CO_3 \rightarrow Na_2\cancel{OH_2 + Na_2CO_3}$
+ − +2 −2 +2 −2 −1 −2

5. Silver nitrate + sodium hydroxide →
silver hydroxide + sodium nitrate
$AgNO_3 + NaOH \rightarrow AgOH + NaNO_3$
\+ − + − + − + 1

6. Copper (II) sulfate + sodium hydroxide →
sodium sulfate + copper (II) hydroxide
$CuSO_4 + NaOH \rightarrow Na_2SO_4 + Cu(OH)_2$
+2 −2 + − + −2 +2 −2

7. magresium sulfate + sodium hydroxide →
sodium hydroxide + sodium sulfate
$MgSO_4 + NaOH \rightarrow Mg(OH)_2 + Na_2SO_4$
+2 −2 + − +2 − +2 −2

8. Same thing

9. silver nitrate + potassium iodide →
silver iodide + potassium nitrate
$AgNO_3 + kI \rightarrow AgI + kNO_3$
\+ − + − + − + −

10. copper (II) sulfate + potassium iodide →
copper (II) iodide +
$CuSO_4 + kI \rightarrow CuI_2 + k_2SO_4$
+2 −2 + − +2 −2 +2 −2

11. magresium sulfate + potassium iodide →
magnesium iodie + potassium sulfate
$MgSO_4 + kI \rightarrow MgI_2 + k_2SO_4$
+2 −2 + − +2 −2 +2 −2

12. Sodium hydroxide + potassium iodide →
Sodium iodide + potassium hydroxide
$NaOH + kI \rightarrow NaI + kOH$
\+ − + − + − + −

→

13. silver nitrate + iron (III) chloride →
- silver chloride + iron (III) nitrate
$AgNO_3 + FeCl_3 \rightarrow AgCl + Fe(NO_3)_3$
+ − +3 −1·3 + − +3 −1·3

14. copper (II) sulfate + iron (III) chloride →
copper (II) chloride + iron (III) sulfate
$CuSO_4 + FeCl_3 \rightarrow CuCl_2 + Fe_2(SO_4)_3$
+2 −2 +3 −3 +2 −2 +3·2 −2·3

15. magnesium sulfate + iron (III) chloride →
magnesium chloride + iron (III) sulfate
$MgSO_4 + FeCl_3 \rightarrow MgCl_2 + Fe_2(SO_4)_3$
+2 −2 +3 −3 +2 −2 +3·2 −2·3

16. sodium hydroxide + iron (III) chloride →
sodium chloride + iron (III) hydroxide
$NaOH + FeCl_3 \rightarrow NaCl + Fe(OH)_3$
+ − +3 −1·3 + − +3

Chem Talk + ChemtoGo

Activity 4

5/21

Checking Up p 178

1. What's a synthesis reaction?
A synthasis reaction is when two or more compounds combine to form a compound. For example $2Na + Cl_2 \rightarrow 2NaCl$

2. Whats a decomposition reaction?
It's when a compound breaks up into individual elements/products. For example, $KI \rightarrow K + I$

3. What's the difference between a single + double displacement?
In a single displacement reaction, one element leaves a compound and joins a single element to form a new compound. In a double, an element leaves 1 compound + goes and joins another compound. This displaces a different eletron which goes back to the orignal compound.

4. What's the evidence of a double displacement reaction?
A precipitate, a gas, or water will form.

5. Will hydrocloric acid strip clean copper?
No, hydrocloric acid is less active then hydrogen in a salt solution.

6. Is calcium sulfate soluable in water?
No, Rule 4 says that all sulfates with the exception of calcium are soluible, so this will form a precipitate.

Chem to Go p 179

1. $NaHCO_3$ when heated produces water vapor carbon dioxide and a solid product, What is it?

$2NaHCO_3 \rightarrow H_2O_{(g)} + 2CO_{2(g)} + Na_{(s)}$ ← sodium

2. Sodium hydroxide + potassium carbonate } no apparent reaction
Sodium hydroxide + potassium iodide

Potassium is more active, so the sodium plates the zinc, Apparently a reaction is going on here.

3. Sodium sulfate + barium nitrate = precipitate why?
A double displacement reaction occured, I know this because it was used as the example in the book, Also the book states that because of rule #4, barium sulfate is not soluble in water. This is why a precipitate forms.

4+5 Combine + explain why it dosn't form a precipitate.

a) $K_2CO_{3(aq)} + NaOH_{(aq)} \rightarrow K_2(OH)_{2(aq)} + Na_2CO_3$
+2 -2 + − +2 -2 +2 -2

rule 1 - all salts soluble rule 5 - carbonates + alkali metals

b) $KI_{(aq)} + MgSO_{4(aq)} \rightarrow K_2SO_4 + MgI_2$
+ − +2 -2 +2 -2 +2 -2

rule 4 - most sulfates rule 1 ionic compound

c) $KI_{(aq)} + NaOH_{(aq)} \rightarrow KOH + NaI$
+ − + − + - + −

rule 1 - salts rule 1 salts

d) $FeCl_{3(aq)} + CuSO_{4(aq)} \rightarrow Fe_2(SO_4)_3 + Cu_3(Cl_3)_2$
+3 -3 +2 -2 +6 -6

rule 4 - sulfate rule 2 chlorides
+3 -3 +2 -2

e) $FeCl_{3(aq)} + MgSO_{4(aq)} \rightarrow Fe_2(SO_4)_3 + Mg_3(Cl_3)_2$

Mole Ratio
Notes

5/23

a How many mole O_2 when 2.6 mole Al produced

$$4Al + 3O_2 \rightarrow 2Al_2O_3$$

$$2.6 \text{ mole Al} = \frac{3 \text{ mole } O_2}{4 \text{ mole Al}} = \left(\frac{2.6 \cdot 3}{4}\right) \text{mole } O_2$$

1.95 mole O_2

sig fig ↳ 2.0 mole O_2

How to do
1. Write what is given

Are you given # moles? } yes
2. Set up factor (mole ratio) to cancle units given
3. Use coefficen from balanced equations to create mole ratio

b How many moles Al are formed when 100g Al_2O_3 decompose.

$$100g\,\cancel{Al_2O_3} = \frac{1 \text{ mole}\cancel{Al_2O_3}}{102g\,Al_2O_3} \cdot \frac{4 \text{ moles Al}}{2 \text{ moles } Al_2O_3} = \left(\frac{100 \cdot 4}{102 \cdot 2}\right) = \text{~~3.92~~ } 1.96 \text{ moles Al}$$

Use mass not atomic #

no ↳ 1. Convert grams to moles w/ molar mass

4. If you want ans in grams → convert ans to grams

c How many grams O_2 produced with production of 80g Al?

$$80g\,\cancel{Al_2} \cdot \frac{1\,\cancel{\text{mole }Al_2}}{54g\,\cancel{Al_2}} \cdot \frac{3\,\cancel{\text{moles }O_2}}{4\,\cancel{\text{moles }Al_2}} \cdot \frac{16g\,O_2}{1\,\cancel{\text{mole }O_2}} =$$

$$\frac{80 \cdot 3 \cdot 16}{54 \cdot 4} = 17.778g\,O_2$$

18 g O_2 ↲ } 2 sig fig

$$\text{given}(g) \times \overset{\text{convert to moles}}{\left(\frac{1\,\text{mole}}{g\,Fm}\right)} \cdot \text{ratio coefficon} \cdot \overset{\text{convert ans}}{\left(\frac{g\,Fm}{1\,\text{mole}}\right)} = \text{answer}$$

↑ opt if given in g

↑ opt if ans in g

or

$$\frac{1\text{ mole}}{6.02 \times 10^{23}\text{ particles}}$$

or

$$\frac{6.02 \cdot 10^{23}\text{ particles}}{1\text{ mole}}$$

gas at STP

0°c

1atm

$$\frac{1\text{ mole}}{22.4\text{ L}}$$

or

$$\frac{22.4\text{ L}}{1\text{ mole}}$$

or

Extra Problems
Mole Ratios

5/23

1. How many g SO_3 formed from 20g SO_2?

$SO_2 + O_2 \rightarrow SO_3$

$2SO_2 + O_2 \rightarrow 2SO_3$

$$20\cancel{g\,SO_2} \cdot \frac{1\cancel{mole\,SO_2}}{32+32\cancel{g\,SO_2}} \cdot \frac{2\cancel{mole\,SO_3}}{2\cancel{mole\,SO_2}} \cdot \frac{32+48\,g\,SO_3}{1\cancel{mole\,SO_3}} = \text{answer}$$

$$\frac{20\cdot 2\cdot 80}{64\cdot 2} = 25\,g\,SO_3$$

2. How many L O_2 needed to form 30g SO_3 (at STP)?

$2SO_2 + O_2 \rightarrow 2SO_3$

$$30\cancel{g\,SO_3} \cdot \frac{1\cancel{mole\,SO_3}}{80\cancel{g\,SO_3}} \cdot \frac{1\cancel{mole\,O_2}}{2\cancel{moles\,SO_3}} \cdot \frac{22.4L\,O_2}{1\cancel{mole\,O_2}} = \text{ans}$$

$$\frac{30\cdot 22.4}{80\cdot 2} = 4.2\,L\,O_2$$

↑ type + copy correctly!

3. How many ~~atoms~~ molecules N_2 needed to produce 2 L of NH_3 (at STP)

$N_2 + H_2 \rightarrow NH_3$

$N_2 + 3H_2 \rightarrow 2NH_3$

$$2\cancel{L\,NH_3} \cdot \frac{1\cancel{mole\,NH_3}}{22.4\cancel{L\,NH_3}} \cdot \frac{1\cancel{mole\,N_2}}{2\cancel{moles\,NH_3}} \cdot \frac{6.02e23\ \overset{\text{molecules}}{\cancel{\text{atoms}}}\ N_2}{1\cancel{mole\,N_2}}\ \text{ans}$$

$$\frac{2\cdot 6.02e23}{22.4\cdot 2} = 2.6875e22$$ ~~atoms~~, ↓ sig fig

3×10^{22} ~~atoms~~ molecules

4. How many g H_2 needed to produce 8×10^{20} molecules NH_3?

$N_2 + 3H_2 \rightarrow 2NH_3$

$$8e20 \cancel{\text{molecules } NH_3} \cdot \frac{1 \cancel{\text{mole } NH_3}}{6.02e23 \cancel{\text{molecules } NH_3}} \cdot \frac{3 \cancel{\text{moles } H_2}}{2 \cancel{\text{moles } NH_3}} \cdot \frac{2 \text{ g } H_2}{1 \cancel{\text{mole } H_2}}$$

$$\rightarrow \frac{(4{,}8e21)}{(1{,}2e24)} = ,003986 \text{ g}$$

in calc put parthenses around

5. How many g CO_2 produced by 100g C_4H_{10}

$C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$

$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$

$$100\text{g } \cancel{C_4H_{10}} \cdot \frac{1 \cancel{\text{mole } C_4H_{10}}}{12 \cdot 4 + 10 \text{g } \cancel{C_4H_{10}}} \cdot \frac{8 \cancel{\text{moles } CO_2}}{2 \cancel{\text{moles } C_4H_{10}}} \cdot \frac{12 + 16 \cdot 2 \text{ g } CO_2}{1 \cancel{\text{mole } CO_2}}$$

$$\frac{100 \cdot 8 \cdot 44}{58 \cdot 2} = 303.44 \text{ g } CO_2$$

Michael Plasmeier

Bean Counting

EXPERIMENT 4

COUNTING PARTICLES AND FINDING THEIR RELATIVE MASSES

18/20

OBJECTIVES

1. To develop and use a new system for counting small particles.
2. To determine the relative mass of different small particles by comparison with an arbitrarily chosen mass.
3. To better understand the mole system and the table of atomic masses by comparison with a model system.

DISCUSSION I: WAYS TO COUNT BEANS

Counting a large number of eggs, or sheets of paper, or beans, is a time-consuming, tedious task. One way to minimize the difficulty of such a job is to use units larger than "one" in the counting process. It's a lot easier to count 15 dozen eggs than it is to count 180 eggs. Two reams of paper mean the same to the stationery store manager as 1000 sheets, without counting. In chemistry, we have a system for counting atoms and molecules that is somewhat different from the dozen or the ream, but which has similar advantages. In order to help you become familiar with the chemical system we will do an experiment involving different kinds of beans. In the experiment, we will count beans in groups called "bunches." You will determine the number of beans in a bunch by relating it to an arbitrary standard of mass. As you will see, this procedure will allow you to avoid the burden of actually counting large numbers of beans. In the second part of the experiment, we will see how a bunch of beans is related to a mole of atoms or molecules.

PRELIMINARY STUDY

1. Review the concept of the mole and mole-gram conversions in Section 2.6 of your text.

PROCEDURE

1. (a) Weigh an empty paper cup (or any other suitable container). Count out exactly 100 beans of one type, discarding any beans which differ greatly from an average bean. Weigh the cup of beans.
   (b) Repeat this procedure for each type of bean provided.
2. (a) Calculate the average mass of one bean of each type and record it in the data table.
   (b) Determine the relative mass of each type of bean by comparison to the lightest type of bean. Record these values in the data table.

$$\text{Relative mass} = \frac{\text{Avg mass of bean}}{\text{Avg mass of lightest bean}}$$

+ YR

3. Weigh out the relative mass (in grams) of each kind of bean and count the beans weighed.

SAMPLE DATA TABLE **EXPERIMENT 4**

	Marble Bean 1	Orange Bean 2	White Bean 3	Dark Red Bean 4
Mass of beans and cup	51.90 g	44.00	38.75	49.69
Mass of cup	13.97 g			
Mass of beans	37.93 g	30.03 g	24.78 g	35.72 g
Average mass	.3793 g	.3003	.2478	.3572
Relative mass	1.53	1.21	1	1.44
No. of beans in relative mass	65.35	82.64	100	69.44

100/relatie mass

these #s should be approx =

CALCULATIONS AND QUESTIONS: I

1. (a) If a "bunch" is defined as the number of beans in 1.00 g of the lightest bean, how many beans are in a bunch? 4.0355
   (b) What statement can you make about the number of beans in a relative mass of each type of bean? its 100
   (c) What statement can you make about the mass of a bunch of beans whose relative mass is: 3.5? 2.6? 17? x heavyer then ligest bean
2. (a) Calculate the number of bunches in 1000 beans. 247.8 bunches
   (b) If you wanted 1000 of the heaviest bean used in this experiment, how could you measure out that approximate number without counting?

You would multiply 37.93 · 10 = 379.3 g

type 3

3. Calculate the number of beans in:
   (a) 15 bunches of garbanzo beans. 60.53
   (b) 3.0 g of a bean whose relative mass is 1.2. 3/1.2 = 2.5 bunches = 10,088 g
4. What would be the mass of 35 bunches of the heaviest bean?

35 · 4.0355 = 141.24 · 1.53 = 216.10

DISCUSSION II: WAYS TO COUNT ATOMS AND MOLECULES

Knowing the relative mass of a bean and the number of beans in a bunch, we have seen that we could use the bunch as a unit with which to measure out given amounts or numbers of beans. Chemists use a remarkably similar system for measuring amounts of atoms and molecules.

By methods which need not concern us here, it is possible to determine the relative masses of atoms and molecules. Roughly speaking, we can say that we base the relative masses on the average mass of the smallest atom, hydrogen, in the same way that we based the relative mass of a bean on the lightest bean. Using that system, we obtain the relative masses of the atoms and molecules listed below:

or AMU

PARTICLE	RELATIVE MASS	PARTICLE	RELATIVE MASS
H atom	1.0	O atom	16.0
He atom	4.0	H_2O molecule	18.0
C atom	12.0	CH_4 molecule	16.0

Given the relative masses, we can say that a helium atom is four times as heavy as a hydrogen atom, and that a CH_4 molecule is four times as heavy as a helium atom.

In counting atoms and molecules, and ions as well, we use a unit called the "mole." A mole is equal to a certain number of atoms or molecules, in the same way that a bunch is equal to a certain number of beans. We define the size of the mole in the same way as we did the size of the bunch, namely, as that number of atoms or molecules which has a mass in grams equal to their relative mass. A mole of hydrogen atoms weighs 1.0 gram. A mole of carbon atoms weighs 12.0 grams. A mole of water weighs 18.0 grams. The number of atoms or molecules in a mole is very large, since atoms and molecules are very small. That number is called Avogadro's number and is a constant equal to about 6×10^{23}. Clearly, it is much larger than the number of beans in a bunch! Knowing its value, we can use the mole, as we did the bunch, to count particles and to determine their masses.

Relative masses of atoms are not determined in quite the way we have indicated in the previous paragraph. If they were, then the relative mass of H would be 1.0000, rather than the actual value, 1.0079. In the system now used, the relative mass of the carbon-12 isotope is taken to be exactly 12, and the relative masses of the other atoms are based on it. For practical purposes the results are essentially the same as if we take the relative mass of H to be 1. At the beginning it's probably easier to see the principles if you base the system on hydrogen, so that is why we described it that way. Relative masses of atoms are the same as the atomic masses discussed in Chapter 2. You can find values of the atomic masses inside the back cover of the text.

CALCULATIONS AND QUESTIONS: II

The following questions involving atoms, molecules, and moles are very similar to those you answered involving beans and bunches.

1. (a) If a mole is defined as the number of H atoms in 1.0 g of hydrogen, how many H atoms are there in a mole?
   (b) How many He atoms are there in a mole? How many H_2O molecules? How many Na^+ ions? How many $C_{12}H_{22}O_{11}$ molecules?
   (c) How many grams will a mole of C atoms weigh? A mole of H_2O molecules? A mole of Fe atoms?
2. (a) How many moles are there in a sample containing 12×10^{23} atoms of He? 6×10^{20} O atoms? 1000 H_2O molecules?
   (b) If you wanted a sample of H_2O containing 6×10^{23} molecules, how could you measure out that approximate number, without counting?
3. Calculate the number of atoms or molecules in:
   (a) 15 moles of CH_4 molecules
   (b) 3.0 grams of H_2O molecules
   (c) 2.0 moles of Fe atoms
   (d) 4.0 grams of Fe atoms
4. What would be the mass of 20 moles of carbon atoms?
5. Calculate the mass of one mole of the lightest bean used in this experiment. Contrast that mass with the mass of the earth, 6×10^{27} grams.

EXTENSIONS

1. Assuming that the "bunch" is redefined as the number of beans in 100 g of the heaviest bean in this experiment, calculate the new value for the number of beans in a bunch. Calculate the new mass for 1 bunch of the lightest bean used in this experiment. Check the new values by mass measurements.
2. Determine the new value of Avogadro's number if the mole were to be redefined as "the number of atoms in exactly 10 grams of hydrogen." Using this definition, what would be the mass of (a) one mole of carbon atoms and (b) one mole of CH_4 molecules?

Name Michael Plameler Date ____________ Class ____________

5/25

SMALL-SCALE LAB: Making a Solution

Laboratory Recordsheet Use with Section 16.4

20/20

SAFETY

Wear safety glasses and follow the standard safety procedures outlined in the Small-Scale Lab Manual.

PURPOSE

To make a solution and use carefully measured data to calculate the solution's concentration.

MATERIALS

- solid NaCl
- water
- 50-mL volumetric flask
- balance

PROCEDURE

Measure the mass of a clean, dry volumetric flask. Add enough solid NaCl to approximately fill one tenth of the volume of the flask. Measure the mass of the flask again. Half fill the flask with water and shake it gently until all the NaCl dissolves. Fill the flask with water to the 50-mL mark and measure the mass again.

100 mL

ANALYSES AND CONCLUSIONS

Using your experimental data, record the answers to the following questions.

1. Percent by mass tells how many grams of solute are present in 100 g of solution.

salt

$$\% \text{ by mass} = \frac{\text{mass of solute}}{\text{mass of solute + solvent}} \times 100\%$$

mass flask = 25.56 g

a. Calculate the mass of the solute (NaCl).

37.21 g − 25.56 = 11.65 g

b. Calculate the mass of the solvent (water).

94.61 g

c. Calculate the percent by mass of NaCl in the solution.

$$\frac{11.65}{94.61 + 11.65} = 10.96\%$$

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________________ Date ________________ Class __________________

2. Mole fraction tells how many moles of solute are present for every 1 mol of total solution.

$$\text{Mole fraction} = \frac{\text{mol NaCl}}{\text{mol NaCl} + \text{mol } H_2O}$$

a. Calculate the moles of NaCl solute. (Molar mass NaCl = 58.5 g/mol)

$11,65g \circ \frac{1 \text{ mole NaCl}}{58,5g} = ,1991 \text{ moles NaCl}$

b. Calculate the moles of water. (Molar mass H_2O = 18 g/mol)

$94,61g\ H_2O \cdot \frac{1 \text{ mole } H_2O}{18g} = 5,26 \text{ moles}$

c. Calculate the mole fraction of your solution.

$\frac{,1991}{,1991 + 5,26} = ,0189 \text{ mole ratio}$

3. Molality (m) tells how many moles of solute are present in 1 kg of solvent.

$$m = \frac{\text{mol NaCl}}{\text{kg } H_2O}$$

Calculate the molality of your solution.

$\frac{,1991 \text{ mole}}{,09461 \text{ kg}} = 2,1044 \text{ molarity}$

4. Molarity (M) tells how many moles of solute are dissolved in 1 L of solution.

$$M = \frac{\text{mol NaCl}}{\text{L solution}}$$

a. Calculate the liters of solution. (1000 mL = 1 L)

100 ml = ,1 L

b. Calculate the molarity of the NaCl solution.

$\frac{,1991 \text{ mole NaCl}}{,1 \text{ L solution}} = 1,991 \text{ L solution}$

5. Density tells how many grams of solution are present in 1 mL of solution.

$$\text{Density} = \frac{\text{g solution}}{\text{mL solution}}$$

Calculate the density of the solution.

$\frac{106,26 \text{ g solution}}{100 \text{ mL solution}} = 1,06 \text{ g/mL density}$

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

Name ______________________ Date ________________ Class _______________

YOU'RE THE CHEMIST

Use the space below to record your observations of the small-scale activities in the *You're the Chemist* section.

Theoretical →

1. Mass small flask = 38.33g

Mass NaCl solution = 3.22 g

mass water = 46.61g

moles NaCl $= 3.22g \cdot \frac{1}{58.5} = .055$ moles NaCl

moles $H_2O = 46.61g \cdot \frac{1}{18} = 2.589$ moles H_2O

$$\text{mass/mass ratio} \quad \frac{3.22}{3.22 + 46.61} = 6.46\% \text{ m/m}$$

1.18m The mole ratio half as below, so

1.1M is the Molarity. This is because the concentration was diluted

└+ molality

2. Its basically the same as before. Measure the beaker's mass. Add sugar + measure mass. Add water + mix. Measure mass again. Use the same calculations to find the mole ratio of the sugar solution.

© Pearson Education, Inc., publishing as Pearson Prentice Hall. All rights reserved.

← volume

$V_c C_c = V_d C_d$ ← diluted

↑ concentrated ← concentration

$(5ml)(1.99M) = (50ml)(?)$

$C_d = .199\ M$

↑ concentration

Michael
Placerato

Chap 12

Stoichiometry
Questions

5/22

p358 5. How is a balanced equation similar to a recipe?
They are similar because both provide quantity information. All quantities can be doubled, and you get 2x the products.

6. How do chemists use chem equations?
Chemists use the balanced equations as a basis to calculate how much reactants are needed or how much product is generated in a reaction.

7. Chem reactions can be described in terms of what quantitie
They can be described and measured by number of atoms, molecules, moles, mass or volume.

8. What is conserved after a chemical reaction?
Mass and the number of atoms are conserved.

9. Interpert as particles, moles, and mass.

1 mole H = 1 g

$2K_{(s)} + 2H_2O_{(l)} \rightarrow 2KOH_{(aq)} + H_2$

2 atoms + 2 molecles H_2O → 2 molecules KOH + 2 atoms

2 moles + 2 moles → 2 moles + mole ← ≠ always

76 g + 2(2 + 16) → 2(38 + 16 +) + 2 g

76 + 36 → 110 + 2 g

112 g

10. Balance

$C_2H_5OH_{(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

not ✓

+ −

$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + \ \ + 3H_2O$

p360 11. $4Al + 3O_2 \rightarrow 2Al_2O_3$

a) $1 \text{ mole } \cancel{Al} = \frac{3 \text{ mole } O_2}{4 \text{ mole } \cancel{Al}} = \left(\frac{1 \cdot 3}{4}\right) \text{ mole } O^2 \rightarrow \frac{3}{4} \text{ mole } O^2$

$1 \text{ mole } O_2 = \frac{4 \text{ mole } Al}{3 \text{ mole } O_2} = \frac{4}{3} \text{ mole } Al$

$1 \text{ mole } Al_2O_3 = \frac{4 \text{ mole } Al}{2 \text{ mole } Al_2O_3} = 2 \text{ mole } Al$

$1 \text{ mole } Al_2O_3 = \frac{3 \text{ mole } O_2}{2 \text{ mole } Al_2O_3} = 1,5 \text{ mole } O_2$

$1 \text{ mole } Al = \frac{2 \text{ mole } Al_2O_3}{4 \text{ mole } Al} = \frac{1}{2} \text{ mole } Al_2O_3$

$1 \text{ mole } O_2 = \frac{2 \text{ mole } Al_2O_3}{3 \text{ mole } O_2} = \frac{2}{3} \text{ mole } O_2$

b) How many moles Al to form 3,7 mole Al_2O_3

$3,7 \text{ mole } Al_2O_3 = \frac{4 \text{ mole } Al}{2 \text{ mole } Al_2O_3} = 7,4 \text{ mole } Al$

12. a) How many mole O_2 to react with 14,8 mole Al?

$14,8 \text{ mole } Al = \frac{3 \text{ mole } O^2}{4 \text{ mole } Al} = 11,1 \text{ mole } O_2$

b) How many mole Al_2O_3 formed with ,78 mole O_2?

$,78 \text{ mole } O_2 = \frac{2 \text{ mole } Al_2O_3}{3 \text{ mole } O_2} = ,52 \text{ mole } Al_2O_3$

13. $CaC_2 + 2H_2O \rightarrow C_2H_2 + Ca(OH)_2$ How many grams of acetylene produced by adding water to 5g $CaCl_2$?

$5g\, CaC_2 \times \frac{\cancel{1 \text{ mole } CaC_2}}{40+68\, g\, CaC_2} \cdot \frac{1 \text{ mole } C_2H_2}{1 \text{ mole } CaC_2} \times \frac{24+2g}{1 \text{ mole } C_2H_2} = 71,25g$

b) How many moles CaC_2 needed to react with 48,0g H_2O?

$48g\, H_2O \cdot \frac{\cancel{1 \text{ mole } H_2O}}{18g\, H_2O} \cdot \frac{\cancel{1 \text{ mole } CaC_2}}{2 \cancel{\text{ mole } H_2O}} \cdot \frac{40+24g}{\cancel{1 \text{ mole } CaC_2}} = 87,11g$

Common Anions

Raticals

Polyatomic Ions

4/30

-1	-2	-3
F^- floride	O^{2-} oxide	N^{3-} nitride
Cl^- Chloride	S^{2-} sulfide	P^{3-} phosphide
Br^- bromide	SO_4^{2-} sulfate	PO_4^{3-} phosphate
I^- iodine	CO_3^{2-} carbonate	
OH^- hydroxide		
ClO^- hypochlorite		
NO_3^- nitrate		
$C_2H_3O_2^-$ acetate		
HCO_3 - hydrogen carbonate		

Polyatonic Atoms
to Memorize

5/23

nitrate	NO_3^{-1}
sulfate	SO_4^{-2}
hydroxide	OH^-
carbonate	CO_3^{2-}
hydrogen carbonate	HCO_3^-
acetate	$C_2H_3O_2^{-1}$
ammonium	NH_4^+

Review of Activities 2,3, and 4.

1.Write the formulas for each compound:

a. sodium bromide

$NaBr$

b. Iron(III)Nitrate

$Fe(NO_3)_3$

c. Calcium Hydroxide

$Ca^{+2}(HO)_2$

d. Hydrochloric acid

HCl

e.sulfur hexafluoride.

$\overset{-2}{S}F\cancel{l}_{\cancel{2}6}$ flurine = F ≠ Fl

2. Name each compound:

a. KOH

potassium hydroxide

b.$CrCl_3$

Chromium (III) chloide

c.$Cu\overset{-2}{S}O_4$

copper (II) sulfate ← charge

d.PCl_5

don't need mono → ~~mono~~ phosphorus - pent cloride

e.HNO_3

nitric acid

Memorize

NO_3^- nitrate

OH^- hydroxide

SO_4^{2-} sulfate

CO_3^{2-} carbonate

ratio of coefficent

3. How many moles of oxygen are needed to react with 3.5 moles of CH_4 ?

study → red book p 360

CH_4 + 2 O_2 = CO_2 + 2 H_2O ← balance 1st

$3.5 \text{ mole } CH_4 \cdot \frac{2 \text{ mole } O_2}{1 \text{ mole } CH_4} =$ 7 moles

4.Classify each reaction according to type:

a. Zn + 2 HCl = $ZnCl_2$ + H_2

Single displacement

b. 2 SO_2 + O_2 = 2 SO_3

~~Synergy~~ → Synthesis

c. 2 KOH + H_2SO_4 = K_2SO_4 + 2 H_2O

2 K
6 O
4 H
1 S

double displacement

5. Balance each equation in 4.

↑

6. Complete and write a balanced chemical equation for each:

Sodium carbonate + Silver nitrate = Sodium nitrate + silver carbonate

use up

$Na_2CO_3 + AgNO_3$ $2NaNO_3 + Ag_2CO_3$

+ −2 +1 −

Hydrochloric acid + potassium hydroxide = ~~hydroxic acid~~ + water + potassium cloride

$HCl + KOH \rightarrow$ ~~HOH~~ $H_2O + KCl$

+ − + −

Magnesium + acetic acid =

$Mg + HC_2H_3O_2 \rightarrow HMg(C_2H_3O_2)_3$

+2 + − + +2 −3

Michael Plasmeier

5/23

Test on Quantitave Calculations: Note YOU MUST SHOW ALL STEPS TO RECEIVE A GRADE HIGHER THAN 80%.

98

1. How many atoms are there in 3.5 moles of Silver?

$3.5 \cancel{\text{moles Ag}} \cdot \frac{6.02 \cdot 10^{23} \text{ atoms}}{1 \cancel{\text{mole Ag}}} = 2.107 \cdot 10^{24}$

$\boxed{2.1 \cdot 10^{24} \text{ atoms Ag}}$ ← 2 sig fig

+10

2.What is the mass in grams of 4 moles of CH 4?

$4 \text{ moles } CH_4 \cdot \frac{(12+4) \text{ g}}{1 \text{ mole } CH_4} = \boxed{64 \text{ g } CH_4}$

+10

3. .How many moles are there in 250 g of NaOH?

$250 \text{ g NaOH} \cdot \frac{1 \text{ mole NaOH}}{(23+16+1) \text{ g}} = \boxed{6.25 \text{ moles NaOH}}$

+10

4.
Write formulas, complete and balance these equations:

calcium nitrate + silver hydroxide

a. calcium hydroxide + silver nitrate =

$\overset{+2}{Ca}(OH)_2^{-1 \cdot 2} + \overset{+}{Ag}NO_3^{-} \rightarrow \overset{+2}{Co}(NO_3)_2^{-\cdot 2} + \overset{+}{Ag}\overset{-}{OH}$

b. hydrochloric acid and Iron (III) hydroxide =

$\overset{+}{H}\overset{-}{Cl} + \overset{+3}{Fe}(OH)_3^{-\cdot 3} \rightarrow H_2O + \overset{+3}{Fe}\overset{-1 \cdot 3}{Cl_3}$

water + iron(III) cloride

c. zinc + sulfuric acid =

alway + 2 right? → $\overset{+2}{Zn} + H_2SO_4^{-2} \rightarrow \boxed{\overset{+2}{H_2}\overset{-2}{Zn}} + SO_4^{-2}$

hydrozinc acid + sulfide

$2SO_2 + O_2 \rightarrow 2SO_3$

actully 'its this

d. sulfur dioxide + oxygen gas = ← that's why this doesn't work ?? sufide

$\overset{+4}{S}\overset{-2\cdot 2}{O_2} + O_{2(g)} \rightarrow SO_4^{-2}$

+8

e.Sodium carbonate + lead (II) acetate = Sodium acetate + lead(II) carbonate

$\overset{+\cdot 2}{Na_2}CO_3^{-2} + Pb^{+2}(C_2H_3O_2)_2^{-\cdot 2} \rightarrow \overset{+}{Na}\overset{-}{C_2H_3O_2} + \overset{+2}{Pb}\overset{-2}{CO_3}$

5.In your own words, explain what is meant by mass is conserved?

It is not lost or gained - it only changes form it is "recycled"

+10

6. How many moles of atoms are there in 4 x 10 ^{23} atoms of Ne?

$4 \cdot 10^{23}$ atoms Ne $\cdot \frac{1 \text{ mole}}{6.02 \cdot 10^{23} \text{ atoms Ne}} = .66445$ moles Ne

.7 moles Ne

1 sig fig

+10

7. What is the molar mass of (NH $_4$) $_3$PO $_4$? (extra point for correct name of this compound)

$3(14 + [1 \cdot 4]) + 31 + (16 \cdot 4)$

149 g/mol

+10

ammonium phosphate

8.How many atoms of hydrogen are there in each compound:

a. sodium hydroxide

NaOH → (1)

b.$H_2 C_2O_4$

(2)

c. nitric acid

$H^+NO_3^-$ (1)

d. KOH

(1)

e. ammonium sulfate $\overset{+1\cdot 2}{}$ $(NH_4)_2SO_3^{-2}$ ⑧ +10

9.If one mole of any gas occupies 22.4 L under standard conditions, What is the volume of 150 g of Ar at standard conditions?

$$150\,g\,Ar \cdot \frac{1\ mole\ Ar}{40\,g} \cdot \frac{22.4\,L}{1\ mole\ Ar} = \boxed{84\,L}$$

✓ +10

10.Given the following equation: 4Al + $3O_2$ = $2Al_2O_3$

How many moles of oxygen are needed to react with 10 g of Al?

$$10\,g\,Al \cdot \frac{1\ \cancel{mole\ Al}}{27\,g\,Al} \cdot \frac{3\ moles\ O_2}{4\ \cancel{moles\ Al}} = \boxed{\frac{5}{18}\ moles\ O_2}$$

+10

Michael Plasreier

5/29

Solutions Work Sheet/ Reading Guide

1 List and DESCRIBE the five factors which affect the rate and formation of solutions. (471-472)

Surface area

stirring/agitation = ↑SA exposure

temprature - ↑T = ↑KE = ↑SA exposure + collsions

SA particles

pressure (gasses mostly) P↑ = Solubility ↑

2.Read p. 445-452) Explain what is meant by solvated ions.

Ions from an ionic compand are attracted from there by solvent ions where they break free of the ionic compand

3.Solubility is defined in terms of solute per and 100g solvent.

4.Explain the formation of an unsaturated , saturated and supersaturated solution.

unsaturated = solute < saturated solution) STP

saturate = max solute for the solvent

supersaturate solute > then it can hold at temp - easy to add a seed crystal

✓ seed crystal

5. What would happen if 1 crystal of the salt dissolved was added to each type of solution?

Unsaturated - te solute will disolve
saturate - more solute will not disolve
supersaturated - will start to happen becase it's a seed crystal

6. What would happen if a crystal of a different salt were added?

It would disolve regardless of saturated state - saturation only relates to that substance

7.What do the terms miscible and immiscible indicate?

miscible - liquids which infinitely disolve in each other at eaual proportions

immiscible - liquids which are insoluble in each other

8.Look at the graph on the top of page 474. answer a, b, and c.

a - As temp ↑, solubility greatly increases
b - $Yb_2(SO_4)_3$ solubility decreases with temp ↑
NaCl's solubility doesn't Δ w/ temp
c - Te extra NaCl doesn't disolve in a solution which (solvent) is saturated

9.How does temperature affect the solubility of a gas in a liquid?

Opposit of solid
> in cold water ten hot

10.What is Henry's Law?

$$\frac{S_1}{P_1} = \frac{S_2}{P_n}$$

at a give temp - te solubility of a gas in a liquid is directly proportional to te pressure of te gas above a liquid

11. Where are nephrons and what do nephrons do?

blood filtering Units in humans which filter out waste

11. Where are nephrons and what do nephrons do?

hemodialysis

12. In your own words describe how renal dialysis works? (see p.478-479)

When someone's kidneys are broken, 3 times a week they are hooked up to a machine which circulates their blood through a machine. The machine has a tube of semi-pormeble membrane. The waste particles fit through this, but the red blood cells return to the body

13. What is the definition of Molarity?

number of $\frac{\text{Moles in solute}}{\text{liters of solution}}$

14. How is a molar solution abbriviated?

$\underline{X}$ M solution

↑ Molarity

15.P.499 Do problems 51,53,

51. Calculate Molarity

a. 1 mole KCl in 750 mL solution

$\frac{1\,\text{mole}}{.750\,\text{mL}} = 1.3\text{ M KCL}$

b) .50 mole $MgCl_2$ in 1.5 L solution

$.5\text{mole} \cdot \frac{1\,\text{mole}}{1.5\,L} = .33\text{ M } MgCl_2$

53. Back →

53. Calculate moles + grams Solute

a) 1.0 L of .50 M NaCl

$$\cancel{1\,L} \cdot \frac{.5\,\text{mole}}{\cancel{1\,L}} = \frac{1}{2}\,\text{mole} \cdot \frac{58.5\,g}{\cancel{1\,\text{mole}}} = 29.25\,g \xrightarrow{2\,\text{sig fig}} 29\,g$$

b) 500 mL of 2.0M kNO_3

$$\overset{\downarrow}{.5\,L} \cdot \frac{2\,\text{mole}}{\cancel{1\,L}} = 1\,\text{mole} \cdot \frac{101\,g}{\cancel{1\,\text{mole}}} = 101\,g$$

always → 1 L

c) 250 mL of .10M $CaCl_2$

$$.250\,L \cdot \frac{.10\,\text{moles}}{1\,L} = .025\,\text{moles} \cdot \frac{40+(35\cdot 2)}{1\,\text{mole}} = 2.75\,g$$

16.Explain how a solution is diluted from a stock or more concentrated solution. (p.483)

You add more solvent, The solute is more spaced out (less concentrated) but there is still the same amount of solute.

17.Describe the two ways of calculating % solutions.

$\frac{\text{volume solute}}{\text{volume solution}}$ $\frac{\text{mass solute}}{\text{mass solution}}$

18.Do p 499 problems 54 and 55.

54. Calculate grams solute needed to make these solutions

a) 2500 g saline solution @ 90% m/m

$2500 \text{ g NaCl} \times \frac{90 \text{ g solute}}{100 \text{g solution}} = 2250 \text{g solute}$ or (2500 · 90%)

b) .050 kg of 4.0% m/m $MgCl_2$

$50 \text{g } MgCl_2 \cdot \frac{4 \text{g}}{100 \text{g}} = 2 \text{g solute}$

19.What are colligiative properties?

Properties which depend only on # of solute particles (not identity) to determin what happens

20.List the main colligiative properties.

Vapor-pressure lowering
boiling-point elevation
freezing-point depression

21.What do colligiative properties depend on?

(not type) of particles

21. What do colligiative properties depend on?

22. Define molality. How is it abbreviated?

$$\frac{\text{moles solute}}{\text{kg solvent}} = \text{molal concentration} \quad m$$

23. What is mole fraction?

ratio of $\frac{\text{moles of solute}}{\text{moles solute + solution}}$

24. What are the formulas for Boiling point elevation and freezing point depression?

$\Delta t_f = k_f \cdot m$

? molal freezing-pt depression constant

$\Delta T_B = k_B \cdot m$

p490 25. Do p.499 61,63,64 65 (63,64,65 for honors classes only)

61. Diff between 1M and 1m?

$1M = \frac{\text{moles solute}}{\text{L solution}}$ q $\quad 1m = \frac{\text{moles solute}}{\text{kg solvent}}$ q

Note: M is solution while m is solvant

63. What is the boiling point?

a) .50 mol glucose in 1000 g H_2O

$\frac{.50 \text{ mole solute}}{1 \text{ kg solvant}} = .50\ m$

for water → $.512\ °C/m \cdot .50\ m = .256\ °C = \boxed{100.26\ °C}$

b) 1.50 mol NaCl in 1000 g H_2O

$\frac{1.56}{1 \text{ kg}} \cdot .512\ °C/m = .768\ °C = \boxed{100.77\ °C}$

64. What's the freezing point?

a) 1,40 mole Na_2SO_4 in 1750 g H_2O

$$\frac{1,40}{1,750} \cdot 1,86\ °C/m = 1,488 = \boxed{-1,49°C}$$

b) ,06 mole $MgSO_4$ in 100 g H_2O

$$\frac{,06}{,1} \cdot 1,86\ °C/m = 1,116 = \boxed{-1,12°C}$$

65. Detamie the freezing point of each ,20 m aqueous $\downarrow H_2O$ solution.

a) K_2SO_4 = ,20 · 1,86 =

b) $CsNO_3$,20 · 1,86 = $\boxed{-,93°C}$

c) $Al(NO_3)_0$,20 · 1,86 =

Dosn't matter what it is because freezing pt. is a Colligative property.

MOLARITY (M)

Name Michael Plosnder

mol
L | M

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liter of solution}}$$

Solve the problems below.

1. What is the molarity of a solution in which 58 g of NaCl are dissolved in 1.0 L of solution?

$\frac{1 \text{ mol solute}}{1.0 \text{ L solution}}$ already counted the 58g NaCl

1 ~~.95~~ M

2. What is the molarity of a solution in which 10.0 g of $AgNO_3$ is dissolved in 500. mL of solution?

$\frac{10\text{g}}{(108+14+48)}$ ↑ round

$\frac{.060 \text{ moles solute}}{.500 \text{ L solution}}$ ← .060

.12 M

3. How many grams of KNO_3 should be used to prepare 2.00 L of a 0.500 M solution?

.5 M · 2 L = 1 mol · (39+14+48)

101 g

4. To what volume should 5.0 g of KCl be diluted in order to prepare a 0.25 M solution?

$\frac{5\text{g}}{(39+35)}$

$\frac{.067 \text{ mol}}{.25 \text{ M}}$

≈ .27 L

5. How many grams of $CuSO_4 \bullet 5H_2O$ are needed to prepare 100. mL of a 0.10 M solution?

.1 M · .1 L = .01 mole · (64+32+64 + 10+80)

2.5 g

©Instructional Fair, Inc.

MOLARITY BY DILUTION

Name ____________________

Acids are usually acquired from chemical supply houses in concentrated form. These acids are diluted to the desired concentration by adding water. Since moles of acid before dilution = moles of acid after dilution, and moles of acid = M x V then, $M_1 \times V_1 = M_2 \times V_2$. Solve the following problems.

1. How much concentrated 18 M sulfuric acid is needed to prepare 250 mL of a 6.0 M solution?

18M · ? = 250mL · 6 M

$83\frac{1}{3}$ mL

2. How much concentrated 12 M hydrochloric acid is needed to prepare 100 mL of a 2.0 M solution?

12 M · ? = 100 mL · 2M

$16\frac{2}{3}$ mL

3. To what volume should 25 mL of 15 M nitric acid be diluted to prepare a 3.0 M solution?

15 M · 25 mL = 3 M · ?

125 mL

4. To how much water should 50. mL of 12 M hydrochloric acid be added to produce a 4.0 M solution?

12 M · 50 mL = ? · 4 M

1800 mL − 50 mL

1750 mL

5. To how much water should 100. mL of 18 M sulfuric acid be added to prepare a 1.5 M solution?

18 M · 100 mL = ? · 1.5 M

1200 mL − 100 mL

1100 mL

MOLALITY (m)

Name ____________________

moles
m | kg

$$\boxed{\text{Molality} = \frac{\text{moles of solute}}{\text{Kg of solvent}}}$$

Solve the problems below.

1. What is the molality of a solution in which 3.0 moles of NaCl is dissolved in 1.5 Kg of water?

$$\frac{3 \text{ moles solute}}{1.5 \text{ kg solvent}}$$

2 m

2. What is the molality of a solution in which 25 g of NaCl is dissolved in 2.0 Kg of water?

$$\frac{25 g}{58 g/mol}$$ ← rand molar mass

$$\frac{.43 \text{ moles}}{2 \text{ kg}}$$

molal

.22 m

3. What is the molality of a solution in which 15 g of I_2 is dissolved in 500. g of alcohol?

$$\frac{15 g}{(127 g \cdot 2)}$$

$$\frac{\overset{.06}{\cancel{.12}} \text{ moles}}{.5 \text{ kg}}$$

.12 molal

~~.24~~ m

4. How many grams of I_2 should be added to 750 g of CCl_4 to prepare a 0.020 m solution?

.02 m · 750 g = 15 moles · (127 · 2)

3810 g (-750 g don't have to to this step 3060 g)

5. How much water should be added to 5.00 g of KCl to prepare a 0.500 m solution?

$$\frac{5g}{(39+35)g}$$

$$\frac{.069 \text{ moles KCl}}{.5 m \text{ solution}} = .14 \text{ kg solvent}$$

139 mL

EFFECT OF A SOLUTE ON FREEZING AND BOILING POINTS

Name ____________________

We use the following formulas to calculate changes in freezing and boiling point due to the presence of a nonvolatile solute. Freezing point is always lowered, boiling point is always raised.

ΔT_F = m x d.f. x k_F

ΔT_B = m x d.f. x k_B

$k_B H_2O$ = 0.52° C/m

$k_F H_2O$ = 1.86° C/m

m = molality of solution

k_F and k_B = constants for particular solvent

d.f. = dissociation factor (how many particles solute breaks up into: for a nonelectrolyte, d.f. = 1)

(Theoretical Dissociation Factor is always greater than observed effect.)

mass

molar mass | moles

Solve the problems below.

1. What is the new boiling point if 25 g of NaCl is dissolved in 1.0 Kg of water?

$\frac{25g}{\cancel{18g}}$ 59g

$\frac{\cancel{1.38} \; .423 \text{ moles}}{1 kg}$ = ~~1.38m~~ · 1.86 = 2.5668

.423 · 1.86 = .7881 · 2 d.f. =

E count Na+Cl - stupid error

~~102.57°C~~

101.57°C

2. What is the freezing point of the solution in Problem 1?

~~1.38~~ mol · .52°C/m = ~~.7176~~

.423 · .52°C/m · 2 = -.44°C

~~-.72°C~~

3. What are the new freezing and boiling points of water if 50. g of ethylene glycol (molecular mass = 62 g/mol) is added to 50. g of water?

$\frac{50g}{62g}$ $\frac{.8065 \text{ moles}}{.05 kg}$ = 16.12 m · 1.86°C/m = -30°C

(df=1)

16.12m · .52°C/m = ~~+8.38°C~~, ~~.30°C~~ flipped the ans

108.38°C

4. When 5.0 g of a nonelectrolyte is added to 25 g of water, the new freezing point is -2.5° C. What is the molecular mass of the unknown compound?

2.5°C/1.86°C/m = ~~4.807~~ 1.344 m · .025kg = ~~.1201~~ .0336 moles

5/~~.12~~ .0336 = ~~41.6~~ 148.8

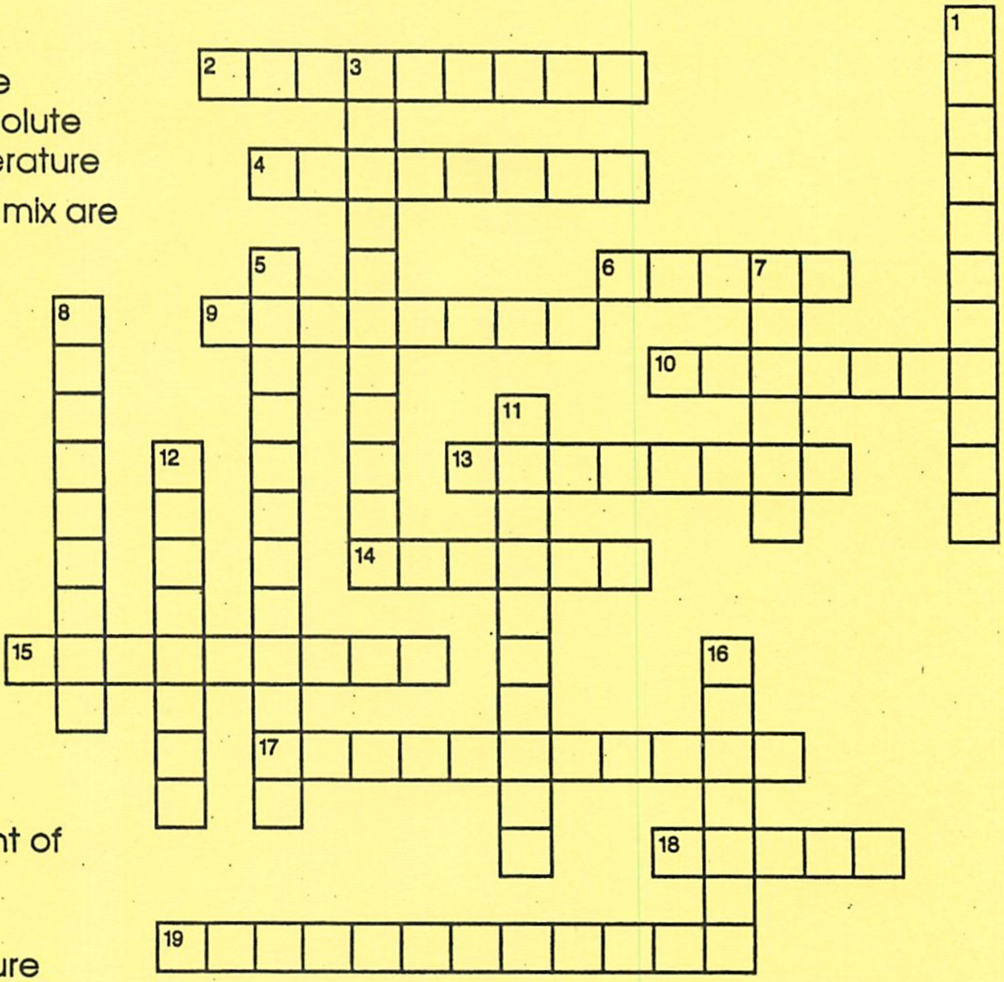
~~41.6~~ g

148.8g

SOLUTIONS CROSSWORD

Name ____________

Across



2. Solution containing the maximum amount of solute possible at that temperature
4. Two liquids which can mix are said to be ________ .
6. The presence of a nonvolatile solute will ________ the boiling point of a solvent.
9. A homogeneous mixture
10. Substance present in larger amount in a mixture
13. Moles of a solute per kilogram of solvent
14. Solution containing a relatively large amount of solvent
15. The solubility of gases ________ as temperature increases.
17. State in which the rate of dissolving is equal to the rate of precipitation
18. The presence of a nonvolatile solute will ________ the freezing point of a solvent.
19. These substances dissociate or ionize in water and are then able to conduct an electric current.

Down

1. Properties that depend on the number of particles in solution
3. Solution in which more solute can be dissolved
5. Solution containing a relatively large amount of dissolved solute
7. Substance present in smaller amount in a mixture
8. The solubility of most solids ________ as temperature increases.
11. Maximum amount of solute that can dissolve in a stated amount of solute at a given temperature
12. Moles of solute per liter of solution
16. Solutions in which water is the solvent are called ________ .

Michael Plasmer

Chemical Energy

Activity 5

20/20

5/30

p180 to make hot + cold packs + observe chemical changes

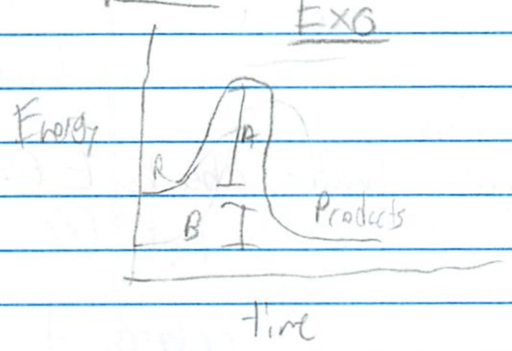
Two chemicals at room temp are mixed together + they cool drastically. How does this work
→ Some sort of chemical reaction

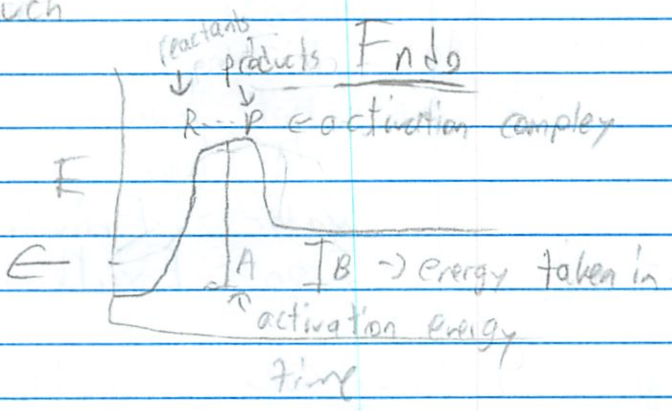
1. 10g $NaNO_2$ + 20mL $H_2O_{(l)}$ - mixed in a bag
   - It's cold to the touch
   - Gray milky liquid

gets cold

Endotermic - energy (heat) absorbed) apply to both
Exotermic - energy (heat) given off) chem + physical exersies

2. 20g $NaCO_3$ + 20mL $H_2O_{(l)}$ - mixed in a bag
   - hot to the touch





3. Take a flask with 16g ammonium thinocyanate + 32g barium hydroxide + stopper + shake. Put on a wet piece of wood + let sit.
   - Its strongly frozen to the board because of an endotermic reaction.

4. NaOH + $H_2O_{(l)}$ in test tube

It gets hot - as a exothermic reaction.

Chem to Go
p185

Identify endo or exo?

1. a) melting ice - endo
b) lighting a match - exo
c) dry ice subliming - end
d) frying an egg - endo
e) burning gas - endo
f) explosion H gas - endo - feels warm (exo) ?
but bonds break (endo) ?

How does it feel?

Bonds form or break?

2. Water in a teapot is heated. Endo or exo?
Exothermic change

3. Red-hot iron is dropped into water. Endo or exo - in the iron + water?
Water - Endothermic Water absorbs E from iron
Iron - Exothermic - E out - but feels colder

4. How does a cold pack work related to E flow?
The cold pack absorbs heat energy from the ankle. This heat ? comes from the blood cells which now move slower?

5. If -20°C ice is put on a hot plate - why does it not appear to melt?

p17 graph

The extra heat makes the particles move faster, but they do not escape + melt until they reach boiling point.

Reading Guide CCS Act 5

1.Explain in your OWN words the difference between an endothermic change and an exothermic change.

endo thermic Δ = Δ where energy is absorbed as heat makeing it feel cooler

exothermic Δ = Δ where energy is given off as heat making it feel warmer

2.Look at the diagrams in the chem talk section. What does E_a represent?

The activation energy
· min energy required for colliding particles to break existing bond

3. What is Δ H?

The actual heat Δ caused by the reactants → products

4. The terms endo and exothermic can be applied to both physical and chemical changes. Give examples of each.

physical - melting ice is an endotermic Δ (taking energy to make it melt)

Chemical - $NaNO_2 + H_2O_{(l)}$ = endo thermic Δ

5.Explain why energy is so important.

Energy is the great organizing principle of all sciences. Energy is always conserved - it changes from 1 form to another

You can track these Δ

6.In order for a chemical reaction to take place, colliding particles must ....

have enough energy to enable themselves to react with each other

7.The minimum energy required for a chemical reaction is the activation

___________.

energy

8.Bond breaking is an endo- process. Bond formation is a exo- process.

↑ add energy
gets cooler

↑ release energy
gets warmer

9.What is the formula for kinetic energy? What is moving here?

$KE = \frac{1}{2}mv^2$ Particles are moving

10.The cells in animals use light to get energy. This process is called Photosynthesis

_______.

endothermic (energy in)

11. What is the chemical formula for respiration?

$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$ + energy

glucose

12.What is photosynthesis? What is formed as a result of photosynthesis?

Plants use light as energy to make oxygen + glucose

13.The ______ amount of energy a plant stores from the sun, is ___?___ reverted by the

?

metabolism of glucose in an animal cell.

14.Explain the difference between heat and temperature.

heat - form of energy

temperature - number relative to avg. kinetic energy of atoms in a material

15. What happens when materials at different temperatures interact?

They exchange energy until they have the same temps.

16.Heat is the transfer of energy, which often results in a change of KE

of Particles and a change in the temp. of the system.

Chap 19

Acids, Bases, Salts 5/31

p593 3, Property of acids + bases

Acids, taste sour, be strong or weak electrolytes solution

Bases Taste bitter, feel slippery, be strong or weak electrolytes solution

4, How did Arrhenius describe acids + bases,

He said that acids are hydrogen-containing compounds which ionize to yield hydrogen (H^+) ions in aq solution

Bases are compounds that ionize to yield hydroxide ions (OH^-) in an aq solution.

5, How did Brønsted + Lowry expand this?

They changed it that an acid is an hydrogen ion donor + a base is a hydrogen ion acceptor

6. How about Lewis?

He said that an acid accepts a pair of electrons while a base donates a pair of electrons.

7, What is a conjugate acid-base pair?

A conjugate acid-base pair consists of 2 substances related by the loss or gain of a single hydrogen pair

b) What are the equations + pairs?

$$\underset{\text{acid}}{HNO_3} + \underset{\text{base}}{H_2O} \rightarrow \underset{\substack{\text{conj.}\\ \text{base}}}{NO_3^-} + \underset{\substack{\text{conj.}\\ \text{acid}}}{H_3O^+}$$

$CO_3^{2-} + H_2O \rightarrow H_2CO_3 + O$

base acid conj acid conj base

8, Identify as monoprotic, diprotic, or triprotic

a) H_2CO_3 - diprotic - 2 H

b) H_3PO_4 - triprotic - 3 H

c) HCl - monoprotic - 1 ionizable hydrogen

d) H_2SO_4 - diprotic - 2 H

p 596 9, Acid or base?

$[H^+] = 6,0 \times 10^{-10} M$ - base

$[OH^-] = 3,0 \cdot 10^{-2} M$ - base

$[H^+] = 2,0 \cdot 10^{-7} M$ neutral

$[OH^-] = 1,0 \cdot 10^{-7} M$ neutral

10, If $[OH^-] = 1 \cdot 10^{-3} M$ - Whats $[H^+]$? Acid or base?

$[H^+] = 1 \cdot 10^{-11} M$ = base

11, pH:

a) $[H^+] = 1 \cdot 10^{-4} M$ = 4,0 ph

b) $[H^+] = ,0015 M$ = 2,02 ph

12, a) $[H^+] = 1 \cdot 10^{-12} M$ = 12,0 ph

b) $[H^+] = ,045 M$ = 1,35 ph

13, $[H^+]$;

a) ph = 5,00 = $1 \cdot 10^{-5}$

b) 12,83 = $1,48 \cdot 10^{-13}$

14, a) 4,00 = $1 \cdot 10^{-4}$

b) 11,55 = $2,82 \cdot 10^{-12}$

15. Find ph:

a) $[OH^-] = 4{,}3 \cdot 10^{-5} M$ $-\log\left(\frac{1 \cdot 10^{-14}}{4{,}3 \cdot 10^{-5}}\right) = 9{,}63$

b) $[OH^-] = 4{,}5 \cdot 10^{-11} M$ $-\log\left(\frac{1 \cdot 10^{-10}}{4{,}5 \cdot 10^{-11}}\right) = 3{,}65$

16. a) $[H^+] = 5{,}0 \cdot 10^{-5} M$ $-\log(5{,}0 \cdot 10^{-5}) = 4{,}30$

b) $[H^+] = 8{,}3 \cdot 10^{-10} M$ $-\log(8{,}3 \cdot 10^{-10}) = 9{,}08$

Acids + Bases

Notes

6/1

Acids H^+... HCl, HBr, HI, HNO_3

Arrhenius Acids

- depends on water
- strongest acid is H_3O^+

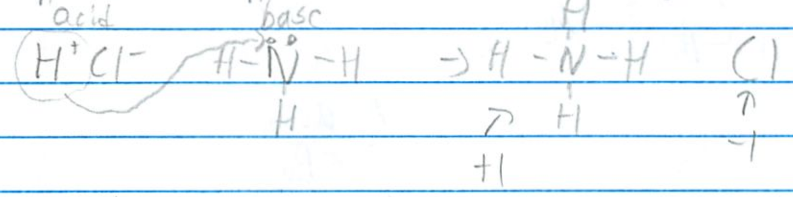
$$H_2O + H^+ \rightarrow H_3O^+$$ ← strongest acid in water

Arrhenius Bases

$$\overset{+}{Na}\overset{-}{OH} \rightarrow Na^+ + OH^3$$ ← strongest base in water

Bronsted-Lowry (gasses)

$$\underset{\uparrow\text{acid}}{HCl_{(g)}} + \underset{\uparrow\text{base}}{NH_{3(g)}} \rightarrow NH_4Cl_{(s)}$$



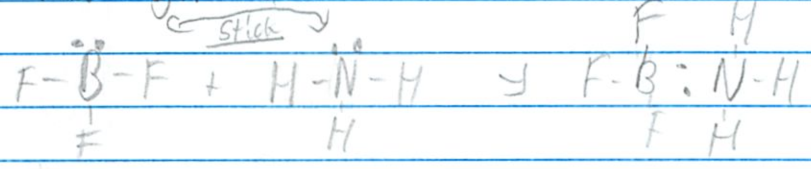
$$H^+Cl^- \quad H-\ddot{N}-H \rightarrow H-\overset{H}{\underset{H}{N}}-H \quad Cl$$

$$\underset{+1}{\uparrow} \qquad \underset{-1}{\uparrow}$$

acid is a proton donor
base is proton accepter

Lewis

most general



$$F-\overset{\cdot\cdot}{\underset{F}{B}}-F + H-\overset{\cdot\cdot}{\underset{H}{N}}-H \rightarrow F-\overset{F}{\underset{F}{B}}:\overset{H}{\underset{H}{N}}-H$$

stick

acid is electron pair accepter
base is electron pair donor

H_2O

molecler compound

[ ] ← molar concentration

$$H_2O \rightleftarrows \frac{[H^+]\cdot[OH^-]}{[H_2O]} \leftarrow k_{eq}$$

← the forward rate of reaction is equal to the reverse rate of the reaction

ionic product constant of water

$$* \quad k_w = [H^+]\cdot[OH^-] = 1.0 \cdot 10^{-14}$$

always equals this

$$* \quad pH = -\log[OH^-]$$

$$V_A C_A = V_B C_B$$

v = volume
c = concentration
A = acid
B = Base

Michael Plasmeier

Acid - Base
Concentration Lab

Try 1

6/1

At 0mL on both sides

2,15 mL   4,15mL
acid   bae

$V_A C_A = V_B C_B$

2,15mL ∘ ? = 4,15ml ∘ ,010M

$C_A = ,0193M$

$C_A = ,019M$ @ 3 sig fig

Michael Plasre

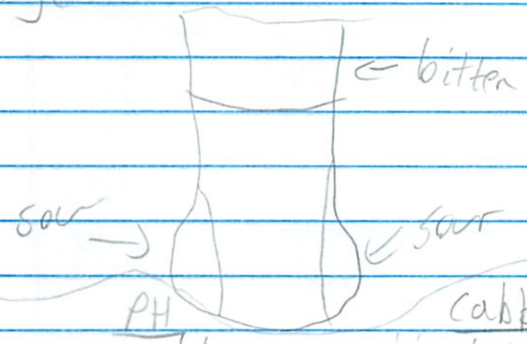
ph

Online Activity

6/3

20/20

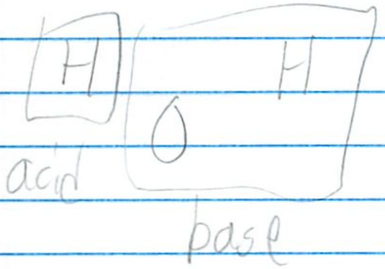
Excite Tounge



Explore

	PH	cabbage	strip
Lemon	4	pink	orange
Baking Soda	8	green	Yellow
Aspirin	5	purple	brown
Ammonia	10	green	green
Milk Magesia	8	light green	yellow-green
Lava Soap	10	yellow	orange
Borax	9	red	green
water	7	purple	Yellow
Lemon Soap	6	purble/blue	orange
vinigar	4	red/pink	orange
tonic water	4	red	orange
Liquid plumber	12	tan	tan

Explain Water = H_2O



6/1

white
1197

Activity 7

Acids and Bases CCS

1.How were acids and bases classified?

according to their characteristic properties

2.What is an indicator?

substances that change colors depending on the presence of acid or base

3. Give some examples of indicators.

cabbage juice/broth
rose peals
litmus paper

4, Acids react with metals, bases do not.

5.Bases feel bitter, taste slipery. Acids taste Sour.

6. Acids and bases are generall good conducters of electicty.

7.How did Arrhenius classify acids?

the ability to prodce hydrogen ions when disolved in water

8.All acid formulas begin with H.

9. When acids are added to water, what happens?

a H atom can be drawn off into the water solution. The H atom leaves an electron behind, forming a positive H^+ ion

$$HCl_{(g)} \xrightarrow{water} H^+_{(aq)} + Cl^-_{(aq)}$$

10. The hydrogen ion is a H^+ which attaches to a other element and forms

negitive ion

a ______________ion.

11.What is the formula for a hydronium ion?

$H_3O^+_{(aq)}$

12.According to Arrhenius, bases produce the OH ion.

13. Show what happens when KOH is dissolved in water.

$$KOH \xrightarrow{H_2O} K^+ + OH^-$$

14.Which other scientists proposed acid/base theories?

Bronsted
Lowry
Lewis

15.When acids react with bases, the H^+ ion and the OH^- ion react to produce

water on a 1 to 1 ratio.

16.The remaining ions join to form a salts.

17. This type of double displacement reaction is called a neutralization.

18.Complete and balance the following equations:

a. $KOH + HBr = KBr + H_2O$

b. $2NaOH + H_2SO_4 =$

$2H_2O + 2Na_2SO_4$

c. $\overset{+2}{Ca}(OH)_2 + 2HNO_3 =$

$2H_2O + Ca(NO_3)_2$

19.Why is an acid considered strong?

it ionizes completely in every solution

20.Give an example of a strong acid and a strong base.

HCl NaOH

21.What does the pH scale tell us?

How acidic or basic a solution is

22.The pH scale has a range of 0 to 7.

23.The acid range is:

$pH < 7$

24. The base range is :

$pH > 7$

25. If the pH of a solution is 7, the solution is neutral.

26. What is the definition of pH?

"power of hydrogen"

a quantity used to represent the acidity of a solution based on the concentration of hydrogen ions

$-\log[H^+]$

27. What do the [ ] stand for?

the concentration of

28. What is a log?

is the exponent to which 10 must be rased to produce that #

29. How many times stronger is a solution with a pH of 2 than a solution with a pH of 3?

10

30. If I know the pH of a solution, how do I find the pOH?

14 - pH

31. What is the pH of each solution?

a. H^+ =1 x10^{-3}

3

b. H^+ = 1 x 10^{-8}

8

c.pOH = 1 x 10^{-5}

9

Do Chem to Go p202 and 203

1, Acid or base?

a) taste sour -acid

b) release OH when disolved in water- base

c) feels slippery -base

d) release H when disolved in water - acid

e) turns pink w/ phenolphthalein - base

f) reacts with metal to produce $H_{(g)}$ - arrhenics acid

g) taste bitter - base

h) turns cabbage green - base

2. Write chem equation that shows these Arrhenius acids are acidic,

a) sulfuric acid $H_2SO_4 \rightarrow H^+_{2(aq)} + SO_4^-$

b) carbonic acid $H_2CO_3 \rightarrow H^+_{2(aq)} + CO_3^-$

More Chem to Go
Activity 7

6/4

3. Show these are bases w/ Arrhenius's method.

a) potassium hydroxide

$$KOH \xrightarrow{H_2O} K^+_{(aq)} + OH^-_{(aq)}$$

b) calcium hydroxide ($(Ca(OH)_2)$)

$$CaOH_2 \xrightarrow{H_2O} Ca^+(aq) + OH^-_{(aq)}$$

4. If you had 2 strong acids (sulfuric + hydrochloric) at the same concentration - how would the ph levels compare?

?

5. Distilled water should have a pH of 7 but tap water often has less (more acidic). Why?

There must be impurities in the water which lean acidic.

6. If you bubbled carbon dioxide through would the pH change?

? Yes - the chart lists that soft drinks are acidic

7. Lemon juice, curdled milk and vinegar taste sour, Why?

They are all acidic. Acids react with metals, produce stinging sensations and conduct electricty. All will turn litmus paper red.

Titration of an Unknown Acid

Final Lab Project 6/4

* to use the concentration of a known base to determine the concentration of an unknown acid

titate to endpoint a sodium hydroxide solution with an acetic acid solution of an unknown concentration
Use phenolphthalein as an indicator
Start with 10mL acid to start

Using Acid "B" .1012 M = C_B
Filled up Acid + Base to 0mL

Titates with: 10.20mL Acid "B"
10.11mL Base

$$V_A C_A = V_B C_B$$

10.20mL Acid · ? M Acid = 10.11 mL Base · .1017 M Base

$$\frac{10.20mL \cdot ?}{10.20mL} = \frac{1.023132 \text{ mL M}}{10.20mL}$$

$$\boxed{.1003 M = C_A}$$

Write out the chemical formula

$$\underset{\text{base}}{NaOH_{(aq)}} + \underset{\text{acid}}{CH_3COOH_{(aq)}} \rightarrow Na^+_{(aq)} + H_2O_{(l)} + CH_3COO_{(aq)}$$

Plaz

6/5/2007

Titration of an Unknown Acid

Final Project Lab for Chemistry

Purpose

To use the concentration of a known base to determine the concentration of an unknown acid

Directions

Titrate to endpoint a *sodium hydroxide solution* with a *acetic acid solution* of an unknown concentration. Use phenolphthalein as an indicator. Start with 10mL of acid.

Results

- Using Acid "B" at .1012M concentration
- Filled up both burettes to 0mL

10

Titrated with 10.20mL of Acid "B" and 10.11mL Base

- **$V_A*C_A=V_B*C_B$**
- 10.20ml Acid * ?M Acid = 10.11mL base * .1012M Base
- 10.20mL * ?M Acid = 1.023132 mlM base
- (divide both side by 10.20ml)
- **$.1003M = C_A$**

Formula

$NaOH_{(aq)}+CH_3COOH_{(aq)} \text{ -> } Na_{(aq)} + H_2O_{(l)}+CH_3COO_{(aq)}$

Michael Plasreiro

Reaction Rates
Lab 6

20/20 9/7

to discover conditions + moleculor happenings which make a readion proceed faster + slower

Why do reactions proceed at different rates. What factors influence this rate of reaction? How could you change the speed of this?
I guess the speed of the molecules moving and the strenght of the bonds influence this. Heating should make it take place faster, and cooling should slow a reaction.

Concentration

1, Put 20mL, 10mL, and 5mL of vinegar in 3 test tubes,
0ml 10mL and 15mL of water into those tubes
Mix,
Put 3 pieces Mg in each,
Time until the Mg strip disappears.

10 min, 25 min, never

2, Repeat the same reaction using a well plate,
I would put

Vinegar	1drop	2 drop	3 drop	4 drop	5 drop
water	5 drops	4 drop	3 drops	2 drop	1 drop
Mg	1 Piece	1 piece	1 piece	1 piece	1 piece

I would add all of it in 1 colum,
Observe

1 vinigor 5 water 1 strip	→ ← →			5 vinigor 1 water

← less bubbles — more bubbles →

← clear water — cloudy water →

Mg looks clearer →

3. Place 10 drops 1.0 M HCl into test tube.
Place " " 0.1 M HCL " another tube.
Drop a Mg strip into each

1.0 M → 14 min 26 sec ← faster
0.1 M → never

4. $2H_2O_2$ is sold as disinfectant. It decomposes slowly to form Oxygen + water

$$2\,H_2O_2(l) + \underset{\overset{\leftrightarrow}{\text{or}}}{\text{light energy}} \xrightarrow{\overset{\swarrow\,\text{catalyst}}{MnO_2}} 2\,H_2O + O_2$$

Pour 15 mL H_2O_2 into 2 test tubes. Add MgO_2 as a catalyst for the reaction. A catalyst is a material which speeds up a reaction w/o being permanently changed

5. Design a reaction to prove the MnO_2 was not used up.
Place a Mn strip into the H_2O_2. As you see it decompose to O_2 + H_2O - note that the Mg strip is still there. Try doing it again with 2 strips. Any different?

Temp

6. Pour 200 mL hot water in a beaker
200 mL cold water in a beaker
Add alka seltzer in each.
Obsarbe

The hot one bubbled more, faster
Entire beaker cloudy
Colder one - only a stream of bubbles
beaker mostly clear
reaction continues for longer

The hotter temperature makes the reaction go faster.

Surface Area

7. Get 2 beakers
Add crushed alka seltzer to one
Leave the other one whole
Obsarbe.

The crushed one went faster - more bubbles) afterwards
Whole one took longer - less bubbles

We compared surface area. Greater surface area made the reaction go faster

8. Describe how these factors affect reaction times
- Concentration - a higher concentration of acid - made the reaction go faster
- catalyst - a catalyst made the reaction go faster
- temp - ↑ temp = faster reaction
- surface area - ↑ SA = faster reactions

10. Reaction

A	5 drops	4 drops	3 drops	2 drops	1 drops	
B	5	6	7	8	9	
Results	clear	light blue	clear	clear	dark blue	after 10 min

Chem Talk +
Chem to Go

6/7

Checking Up p191 1. What are the 4 factors which influence a reaction rate?
Surface area, temperature, catalysts, and/concentration of a reaction. All of these influence the collision frequency of the molecular particles.

2. For each influence (above) describe how it influences the collision freqvency of the particles.
Surface area - More molecules are in contact + in contact more often
Concentration - ↑ concentration = ↑ # particles in reaction = ↑ collision freq.
Temperature - Particles move faster as temp ↑ - faster particles = ↑ chance they will collide.
Catalysts - lower the action energy needed for the reaction to take place.

3. How is the catalyst different from products / reactants?
The catalyst is not used up or permonetly changed.

Chem to Go p192 1. What will happen?
a. What cooks 1st? 50°C or 150°C?
The 150°C cooks faster because the particles move faster.
b) Animals hibrate during the winter. Their metabolism slows down. Why?
Particles move slower when they are cool.
c) Granulate sugar disolves faster then a sugar cube
A larger surface area allows particles to interact more, causing it to disolve faster

d

e) Burning a sugor cube is hard. But when ash is added - the sugor would burn
The ash must be acting like a catalyst, causing the sugor to burn.

e d) Why are 2 antiacids better then 1?
2 provide a higher concentration - meaning they act faster then just 1

f) Powdered aspirin disolves faster then whole aspirin
Surface area - see c

g) Sugar disolves faster in hot tea then ice tea.
Higher temp makes particles move faster, increasing contact, increasing the reaction speed.

2. As temp ↑, reaction rates ↑ Why?
See #1 g

3. If you purchased a glo-stick and wanted to make it last longer - what could you do?
You could put it in the freezer to make it last a bit longer by decreasing the reaction time.

4. Why did the Mg strip react slower in acid with a lower contration?
The less concentrated acid has less particles to interact with the Mg strip.

5. Why is it possible for someone stuck in colder water to survive longer in cool water then warm water.
Cool water slows the reactions in their bodies allowing your stored food to last longer. Also you don't lose water sweating. However in very cold water, you use up a lot of energy shivering.

6. Grain elevators have been known to explode because of the fine powder.
Yeah - there must be a lot of surface area for the reaction to happen. But there must be some sort of spark...

~~Study~~ Read all chem labs in white book

— I know
— I looked up
— She said

Active Chem Review 1

- same way every time
- formulas
- new properties

1.In your own words, explain how a compound and a mixture are different.

Compound - 2 or more elements in fixed proportions

mixture - physical combo - not chemical
- remain properties
- no def. proportions
- seperated physically

- not new substance w/ diff formula

asked 4-5 diff times

2. What is an element?

- purest form of matter
makes up everything
- can't be simplifies ~~w/ chem means~~ by any means
Unique properties

3.Explain how to write the symbols for the elements. a. If the element has only one symbol

b. if the element has two symbols.

H ← symbol cap

Pb ← small letter

4.How was each gas in activity 1 produced? Explain

Electrolsis caused a chem reaction

H - active metal (Zn) + acid (HCl)

O - decomposition of H_2O_2 w/ MnO_2 as catylist

? → CO_2 - $Na_2 + CO_3 + \Delta$

5.Compounds can be separated by electrolysis. This is a process which adds electricity / (electrons),

6.Compounds are represented by formulas. These are simple, whole number ratios.

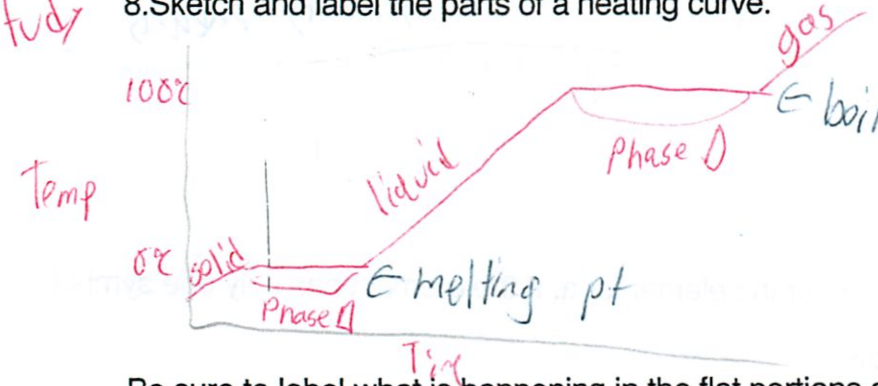
7. How many atoms of each element are there in this formula:

$(NH_4)_2SO_4$ ✓1 ←4

2 4·2

study

8.Sketch and label the parts of a heating curve.



absorbing KE when temp ↑

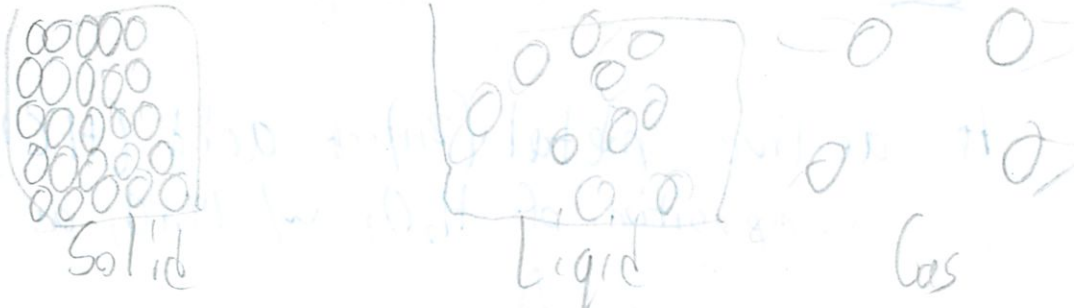
PE Δ when state Δ

Be sure to label what is happening in the flat portions as well as the diagonal portions.

endo

exo

9.Describe how particles are organized in the states of matter. Drawings will work.



Study

Boyle

$P_1 V_1 = P_2 V_2$

10. As temperature is increased, volume ↑.

in k

Charles

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

11. As pressure is increased, volume ↓

Gay-Lussac

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$

12. What is sublimation? Give an example of a substance which sublimes.

Solid → gas

mothballs

dry ice

iodine

13.. The normal boiling point of a substance depends on Atmospheric Pressure and Temp.

Study

14. In your own words explain how to distinguish among solutions, suspensions and colloids. Your answer should include an explanation of the Tyndall effect, settling and ions.

Suspensions - filters + hetero ← will settle Tyndal Yes

Colloid - Tyndal Yes doesn't settle homo

Solution - Tyndal no

↑clear, particles ion

15. What is a solution? How can it be distinguished from a pure substance?

Two elements/compounds link together - homogenus mixture

substances

through a physical change

like evaporating

16. In your own words, explain how to read a graduated cylinder. Your answer should include the terms meniscus and significant figures.

Read to line and fractionall division between them

Lowest point in meniscus

Look at eye level

density = g/cm 3

17. Density is the ratio of ~~grams~~ mass to volume. Create a memory circle for density.

18. A liquid weighs 31.64g and occupies a volume of 32.06 ml. Find its density.

Show all work

$$\frac{31,64g}{32,06ml} = ,9869\ g/mL$$

19. What is the mass of 25.0 ml of a liquid with a density of 0.79g/ml? Show all work

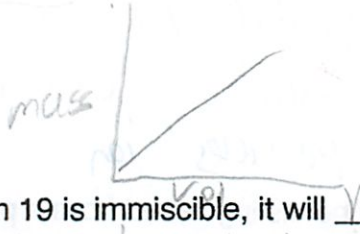
↓ 2 sig fig

25,0 mL • ,079 g/mL = 1,975 g → 2sig fig 2,0 g

20. A graph of mass (y) versus volume (x) will be a linear graph. The slope will be

equal to 1.

?



21. If the liquid in problem 19 is immiscible, it will ~~no~~ not dissolve in water. Why?

Because its immiscible / insoluble

Float in water (lower density)

21. When rounding numbers, the number of significant figures should.......

alwas be = to the inputs of least accurate input

Red
6.1

22. List and describe the properties on metals.

Conduct heat + electricy
sheen
Solid room temp (except Hg)
ductile

23. List and describe the properties of nonmetals.

Most gas (STP), Some solid, 1 liquid
Poor conductors
brittle
dull

24. What is oxidation?

20.1 rust - a process that involves complete or partial loss of electrons or a gain of oxygen - results in an ↑ of oxidation # of an atom

25. What is an alloy?

mixture of ≥2 elements
1 is a metal
(solution of metals)

26.Elements can be identified by flame tests. Why does each element have its own color?

the different electron levels it falls

Study

Unique electron config

Study

27. What is an organic substance? Why are organic substances represented by structural formulas?

made of hydrocarbons
↑ carbons
the position matter
types of bonds

28.The most common reaction of hydrocarbons is ~~Covalent bonding~~ burning.

29. Explain the difference between potential and kinetic energy?

potential - energy of position - stored

kinetic - energy of motion

30. Explain how to distinguish among hydrogen, oxygen and carbon dioxide.

put a splint in it

hydrogen - explodes
oxygen - relights
CO_2 - extinguises

glowing splint

- know
- book
- teacher review

Review For Final Part 2

1.The symbol for the element Flourine is:
F

2. The symbol for tin is:
~~Sn~~ Sn * * Study these

3. Given $2KClO_3$, How many atoms of each element are indicated? How many molecules? How many moles?
2 k
2 Cl
~~3O~~
6O
opps - looked like an I
2 molecules
~~2(39+12+127+16·3)~~
2 moles
↑
particles

4. Which particles contribute to mass in the atom?
protons + neutrons p^+ n^0

5.All acids contain the H^+ ion. All bases contain the OH^- ion.

6. If a substance has a pH of 12 it is (acid or base)
base

7. Balance:

a. $4Cu + O_2 \; 2 = Cu_2O$

b. $2 C_6H_{14} + 19 O_2 = 12 CO_2 + 14 H_2O$
12 C
22 H
36 O

c. $2 KOH + H_2SO_4 = 2 H_2O + K_2SO_4$
2 k
6 O
4 H
1 S

8. Evidence that a chemical change has taken place includes:

Formation of a *precipitate, formation of a gas, __________,changes in temp, color Δ smell
? ↑ sort of right

9. Metals are ductile, this means that they can be made into wires. Metals are ~~ductile~~ mallable, this means that they will flatten into sheets when hammered.

10.List the elements in the third period. List the members of the halogen family.

3rd pd: Na, Mg, Al, Si, P, S, Cl, Ar

Group 17: F, Cl, Br, I, At

11. Elements in the same family have what in common?

electron level / properties

vallance electrons

12. If a metal is active, this means:

exact

reacts a lot replaces H from an acid

13. Oxidation is ~~loss~~ loss of electrons.

* atom becomes +ly chaged

Redox

14. Reduction is gaining of electrons and takes place at which electrode?

~~negitive goes to anode~~ cathod ?

15. To separate metals from ores reduction takes place. (oxidation, reduction)

16. Iron can be distinguished from tungsten by which chemical property?

rusting / Oxidation ?

* Iron rusts, tungsten doesnt

17. Nitrogen can be distinguished from oxygen by which physical property.

reviewing

~~fire test~~ density

18. One mole of any substance contains ~~$6.9 \cdot 10^{24}$~~ particles. study

6.02×10^{23}

19. The molar mass of each compound is:

a. $NaOH$

$23+16+1 = \boxed{40}$

c. $CaCO_3$

$40+12+(16 \cdot 3) = 100$

Unit

g/mol

b. $(NH_4)_2SO_4$

$2[14+(1 \cdot 4)]+32+(16 \cdot 4)$

$\boxed{132}$

d. $KMnO_4$

$39+55+(16 \cdot 4)$

$\boxed{158}$

by what - mass
- particles

20. What is the percentage of oxygen in each compound?

a. $KClO_3$	b. NaOH	c. $K_2C_2O_4$
$\frac{16\cdot3}{39+35+16\cdot3}$	$\frac{16}{23+16+1}$	$\frac{16\cdot4}{39\cdot2+12\cdot2+16\cdot4}$
$\frac{48}{122} = 39\%$	$\frac{16}{40} = 40\%$	$\frac{64}{166} = 39\%$

21.Atoms with filled outer shells are:

this

a. stable	b. unstable	c. all metals	d. all gases
↑	↑	↑	↑
totally filled	nearly filled	need to lose electrons	need to gain electrons

22. List at least 3 ion that always form soluble salts.

Na^+, K^+, NH_4^+ - salts which disolve in water

23.How many dots should I draw around S?

6

16 th family

6 valance electrons

24. All noble gases have 8 (2) electrons in their outer shells.

25. If an atom loses 2 electrons, the resulting ion has a +2 charge.

Chemistry Review for Periodic Table and Bonding

1.What is a period? What do elements in periods have in common?

row - the familes repeat
don't share properties except filling same sub level

2. What is a group or family? What do elements in groups have in common?

Columns - same # valance electrons
Similar properties - mostly chemical
make similar compounds

3.If Nitrogen combines with 3 hydrogen atoms to form ammonia, What would the formula for similar compounds with Phosphorous and arsenic be?

H_3N ! → NH_3
H_3P → PH_3
H_3As → AsH_3
? H last (irregular

2/2

4.What is meant by the tern" representative elements"?

not a noble gas
filling "p+s" orbitals

not really

ion forms same type of charge

5.List and describe the representative families of elements. List the elements in each and explain something about each family (what kinds of compounds formed)

?

the 7 groups

very reactive

right

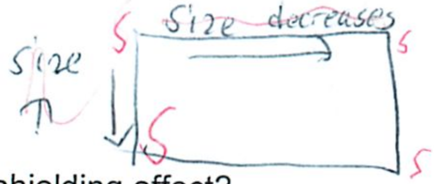
Alkali metals - soft silver solids - low melting + boiling pt
Alkaline earth metals - relativly soft - gray white luster
Halogens - non metals - very reactive
salt formers
poisinous
smell like bleach
Noble gasses — full electrons

6.Which type of metals are filling d orbitals? Which type are filling f orbitals?

transition inner transition

(+1) 9.Describe trends in atomic and ionic radii.

the radius is half the distance between 2 nuclei



① 10.What is the shielding effect?

* Big atom - w/ lots of electrons loses outside electrons eaiser

① 11. What information does ionization energy give us?

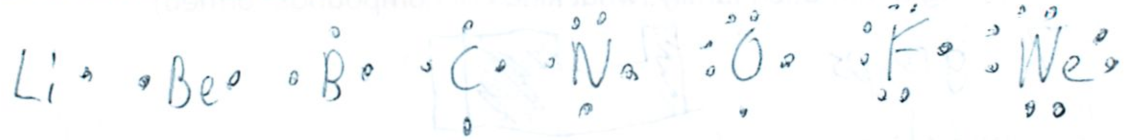
how easy (E required) to separate

12.What is electronegativity? What kind of information can we use these values to obtain? The ability of an atom of an electron to attract electrons when the atom is in a compound

13.What is meant by the term "valence electrons"?

Electrons in last sublevel "Outer shell electrons"

14.Draw electron dot structures for the elements in the second period.



15.What is the “octet rule”?

Want to have 8 valance electrons

lose, gain, or share electrons

16. How is a cation formed?

non metals get extra electrons

flip

17. How is an anion formed?

metals lose electrons

18. Give a detailed description of the formation of an ionic bond.

Anion + cations, which have oppset charges, attract each other w/ electrostatic forces

* electrons move from metals to non-metals

19. What is a crystal? What information is given by the coordination number?

ionic structures - orderly repeating pattern

Coordination # - # of ions of opposite charge that surrond that ion in a crystall

20.What causes ionic compounds in solution to conduct electricity?

electrons can move around

electrical partacles provide a pathway for the flow of charge

21. In your own words, describe a metallic bond.

The valance electrons are like a sea can be moved around the nuclei

- conduct
- malable
- ductile

22. Explain how a single covalent bond forms.

Sharing of electrons

7

Electron pair equally shared

2 non metalic atoms

23. Write a structural formula for the water molecule.

H – O – H

24. Why do nonmetallic atoms share electron pairs?

N want 8 electrons

like a noble gas

25. Give examples of multiple covalent bonds.

$O + O = O_2$ - $:O = O:$

$N + N = N_3$ - $:N \equiv N:$

26. List and draw electron dot structures for all the diatomic molecules.

molecule consisting of 2 atoms

F_2 $:F - F:$

Cl_2 ⟶ $:Cl - Cl:$

Br_2 $:Br - Br:$

I_2 ⟶ $:I - I:$

H_2 $H - H$

N_2 $:N \equiv N:$

O_2 Inadiquate

Review for Final Exam :Cool Chemistry Show

1.What are the two basic types of changes that matter can undergo?

Physcial
Chomical

2. Explain how to tell the difference between these two types.

Chem- ~~molecules rearrange~~ - ~~can't be undone~~
~~produced~~
- new substances
- diff properties
- Δ temp, color, smell
- not easy to unde

physcal
- changes in state
- easy to reverse

3. What is a solution? Use words like solute, solvent, concentration, saturated, unsaturated, and supersaturated in your explanation.

homogenous mixture
Solute disolved in solvent
dilluted Unsat - can add more solute
Sat - more solute won't disolve
Sup. sat - more solute then it can handdle

Molar
molal
study

4.A solution of barium nitrate is mixed with a solution of sodium hydroxide, a solid white product forms.Complete and balance this equation, using correct formulas.

$Ba(NO_3)_2 + 2NaOH \rightarrow$ + balance

study. diatomic again

$\rightarrow 2\overset{+1}{Na}(\overset{-1}{NO_3})_{aq} + \overset{+2}{Ba}(OH)_2$

5.Look at the chart on p.158 in the Active Chemistry book. Write the formulas for all of the products formed.(Remember to check solubility rules) tightly bound grap of atoms which have a ⊕ or ⊖ charge + behave as a unit

6.Write out the name, charge and formula for the common polyatomic ions..

Nitrate NO_3^{-1}

Sulfate SO_4^{-2}

Hydroxide OH^-

Carbonate CO_3^{2-}

hydrogen carbonate HCO_3^-

acetate $C_2H_2O_2^-$

ammonium NH_4^+

7.What is the “activity series”? What information does it give us.

metals in ↓ reactivity

which metals reacts with each other

8. Write out the solubility rules.

*

all Clorides soluble

Carbonates, Hydroxides Not soluble

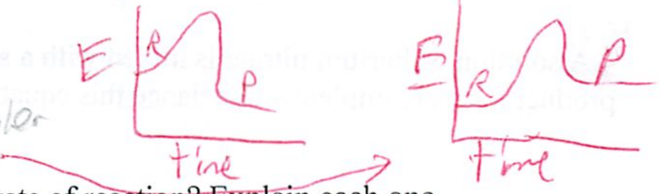
9. Define exothermic and endothermic .Label a diagram to explain each.

↑ heat out

feels warmer

↑ heat in

feels cooler



10. What are the factors which affect the rate of reaction? Explain each one.

~~Solubility~~ catalyst - ↓ reaction rate

Surface aree - more area for collosion

temp - particles move + clash faster

concentration - more particles able to react

11. What is a catalyst?

Δ speed (reaction) w/o being used up

↑ not reactant or product

OH

12. What is an acid? What is a base?

+proton H donor

H + proton acceptor

releases OH in water

H

13. An acid + a base yield water + salt.

14. Oxidation is:

losing electrons

15. Reduction is:

gaining electrons

16. Given this equation: $\overset{0}{Zn} + \overset{+2}{Cu}\overset{-1}{SO_4} = \overset{0}{Cu} + \overset{+2}{Zn}\overset{-2}{SO_4}$ Identify what is oxidized and what is reduced..

$Zn^0 \rightarrow +2$ - lossing - ox

$Cu^{+2} \rightarrow 0$ - gaining - reduction