AP Chem Lab Notebook

An (almost) verbatim digital version of my AP Chem Lab Notebook.

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1. Physical and Chemical Changes Lab.

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Partner: Jeff Somvorachith

Lab 1. Physical and Chemical Changes Lab

Purpose: To identify physical and chemical changes through different experiments.

Procedure:

Experiment A

1. Dissolve a small amount of copper sulfate pentahydrate in water(no more than 1.2cm deep).
2. Holding the test tube with a test tube holder, strongly heat the mixture over a Bunsen burner flame. Make sure that the mouth is directed at a wall, away from other people.
3. Continue heating until a residue remains and the color changes.
4. Allow the substance to cool.
5. Add approximately 10ml of water to the test tube.
6. Pour the solution into the waste container provided.

Experiment B

1. Place a test tube into a test tube rack.
2. Place HCl into the test tube(~2cm deep of HCl).
3. place a 1.5cm piece of magnesium ribbon into the HCl.
4. Pour Mg/HCl waste into the container.

Experiment C

1. Dissolve a scoop of Epsom salts into 50ml of water in a 250ml beaker and stir to dissolve.
2. Add 3 droppers full of household ammonia. Do not drink.
3. Pour the solution down the drain. clean/dry the equipment.

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Experiment D

1. Place 5 drops of silver nitrate into a test tube.
2. Add 5 drops of sodium chloride into the test tube.
3. Rinse test tube in the sink.

Experiment E

1. Put about 50ml of water into an Erlenmeyer flask and take the temperature of the water.
2. Add half a scoop of sodium bicarbonate to the flask and swirl to mix.
3. After 1 minute, measure the temperature ~~in the sink.~~ of the mixture.
4. Rinse flask in the sink.

Experiment F

1. Measure one small scoop of sodium polyacrylate into a beaker.
2. Add 10ml of water into the beaker.
3. Invert the beaker over the sink.
4. Add 10ml of water into the beaker.
5. Invert the beaker over the sink.
6. Put contents of the beaker into the appropriate waste container. Rinse out the beaker with lots of water.

Observations:

Experiment A

- The water is clear and the copper sulfate is blue.

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- There are some popping noises when heating and a smell is produced(water boiling smell).

- When mixed, the mixture is a transparent blue.

- It mixes as the water boils and becomes slightly darker and less transparent.

- After the liquid evaporates, there is a blue deposit left in the test tube. Some of it seemed to turn white. I think it is left over copper sulfate.

Experiment B

- HCI was clear, magnesium was metallic.

- When combined, it was exothermic, let off vapor, and bubbled.

Experiment C

- Water is clear, ammonia is clear(smells bad), Epsom salt is white.

- mixing Epson salt has no reaction.

- Mixing in ammonia makes a milky, cloudy mixture with things floating inside.

Experiment D

- Silver nitrate and sodium chloride are clear.

- When mixed, it turns white and begins to deposit a white substance. Cloudy.

Experiment E

- Water is ~~white~~ clear and 32^{o}C.

- Sodium bicarbonate is white powder.

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- End mixture is clear.

- End mixture is 29^{o}C.

~~Experiment~~ - It cooled by 3^{0}C.

Experiment F

- Water is clear, sodium polyacrylate is white.

- The mixture is a solid, opaque substance that sticks to the beaker.

- The sodium polyacrylate absorbed all of the water.

- It is a jelly like substance hat looks like ice.

Conclusion:

The purpose of the lab was to identify physical and chemical changes through different experiments. We achieved this purpose by performing 6 different experiments and determining weather ~~or not~~ it was physical or chemical. Experiment A was a physical change because mixing water and copper sulfate gave a blue liquid, boiling away that water left a white powder, and adding water back turned it blue again. Experiment B was a chemical change because adding a magnesium strip to HCl caused it to dissolve the strip and had an exothermic reaction. Experiment C was a chemical reaction because mixing water, Epsom salts and ammonia created a milky, cloudy substance with chunks floating in it. Experiment D was a chemical reaction because mixing the silver nitrate and sodium chloride yielded a white and cloudy liquid. Experiment E was a chemical changes because it was an endothermic reaction. It cooled by about 3^{o}C. Experiment F was a physical change because the sodium polycarbonate absorbed the water, which can be evaporated back out.

2. Mole Concept II

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Partner: Jeff Somvorachith

Lab 2. Mole Concept II

Purpose:

To practice using the mole and making mole calculations.

Procedure:

A. Mass a penny. Assuming it is pure copper, calculate:

1. The number of moles of copper.
2. The number of copper atoms present in the penny.

B. Mass 20 drops of water from your eyedropper, calculate:

1. The number of moles of water present in only one drop of water.
2. The number of molecules of water present in that one drop of water.

C. Mass one of the small pieces of the polyvinyl chloride(PVC) pipe, C_2H_3Cl, calculate:

1. The number of moles of PVC in your sample.
2. The number of molecules of PVC in your sample.
3. The number of carbon atoms contained in your sample.
4. The total number of all the types of atoms in your sample.

D. Mass one of the Styrofoam cups which are made of polystyrene, C_8H_8, calculate:

1. The number of moles of polystyrene in your cup.
2. The total number of all the types of atoms in your sample.

Data:

A.	3.07g	penny

B.	0.72g	20 drops of water
C.	6.49g	PVC
D.	3.61g	Styrofoam cup

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Analysis:

A. 3.07g/63.546 amu =0.0483 mol <-1.

0.0483 mol *6.022*10^{23}=2.91*10^{22} atoms Cu. <-2.

B. 0.72g/20=0.036g 0.036g/(1.00794+15.9994)=0.0021 mol <-1.

0.0021 mol *6.022*10^{23}=1.26*10^{21} molecules <-2.

C. 6.49/[2(12.0107)+3(1.00794)]=0.104 mol PVC <-1.

0.104*6.022*10^{23}=6.26*10^{22} molecules <-2.

6.26*10^{22}*2=1.25*10^{23} carbon atoms <-3.

6.26*10^{22}*3=1.88*10^{23} hydrogen atoms <-4.

D. 3.61g/(8*12.0107+8*1.00794)=0.0347 mol <-1.

0.0347*6.022*10^{23}=2.090 molecules

2.090*(8+8)=3.34 atoms

Conclusion:

Avogadro's number links the amount of atoms or molecules in a substance to the mole. the molar mass is used to convert between mass of a substance and the mole. {mole}={number of atoms or molecules}/{Avogadro's number} {mole}={mass(g)}/{molar mass(amu)} {atoms}={molecules}*{atoms per molecules} {molecule}={moles}*{Avogadro's number}. Molecules are made of atoms.

3. Hydrated Compounds

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Lab 3. Hydrated Compounds

Purpose:

To use hydrated compounds to practice calculating how much water molecules are attached to other molecules and compounds.

Procedure:

1. Clean the crucible with water, dry it with a paper towel, and heat it for 2-3 minutes using the Bunsen burner.
2. Let the crucible cool for 3-5 minutes. Use the crucible tongs to carry the crucible to a balance and mass the empty crucible.
3. Mass out 3g of the assigned hydrated compound into the crucible.
4. Heat the crucible and it's contents for about 10 minutes.
5. Let the crucible cool and re-mass it.
6. Heat the crucible and it's contents for 2 more minutes.
7. Let the crucible cool and re-mass it.
8. If the last mass does not agree within 0.02g with the last mass reading, reheat the crucible and re-mass it until the last to measurements are within 0.02g of each other.
9. After final massing, place the crystals in the waste container provided. Clean the crucible, and clean the lab area.
10. Determine the percentage of the water in the assigned hydrated compound.

Data:

Mass of crucible: 13.08g

Mass of hydrated compound: 3g $MgSO_4 \cdot H_2O$

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Mass of crucible and hydrated compound: 16.08g

Mass of crucible and anhydrous compound(first heating): 14.55g

Mass of crucible and anhydrous compound(second heating): 14.55g

Observations:

The $MgSO_4 \bullet xH_2O$ turned a more white color and seemed to melt into a single mass. It started off as clear granules(like table salt), and ended up as a single white mass. Mass went down after heating.

Analysis:

Mass of anhydrous compound: 14.55g-13.08g=1.47g $MgSO_4$

Mass of water in hydrated compound: 3g-1.47g=1.53g H_2O

Moles of anhydrous compound: 1.47g/(24.30+32.06+4*16.00)=$12.2*10^{-3}$ mol $MgSO_4$

Moles of water in hydrated compound: 1.53g/(2*1.008+16.00)=$84.9*10^{-3}$ mol H_2O

Ratio of moles of anhydrous compound to moles of water: $12.2*10^{-3}/12.2*10^{-3}$ mol MgSO4:$84.9*10^{-3}/12.2*10^{-3}$ mol H_2O=1:6.959 ~1:7 mol. $MgSO_4 \bullet 7H_2O$.

Conclusion:

The purpose of this lab was to use hydrated compounds to practice calculating how many water molecules are attached to hydrated compounds. We achieved this purpose by heating and weighing hydrated compounds until there was no water left and calculating the difference between the hydrated compound and anhydrous compound to find how much water was in the hydrated compound. Potential sources of error may be inaccuracies in weighing, contamination, or water that didn't get evaporated. Potential improvements may include a moisture-free environment, and a better controlled heating colution and weighing system. From the analysis, there was about 1.53g of water in the 3g of $MgSO_4 \bullet xH_2O$, leaving about 1.47g of $MgSO_4$, ~51% H_2O.

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I get a ratio of 1:9.6 mol $MgSO_4$:H_2O, the correct ratio is 1:7. My results are pretty accurate. The water must have been pretty close to completely evaporated, assuming the scale was 100% accurate. The error most likely comes from an inaccurate scale, given how close the answer is to the actual value.

4. Empirical Formula Determination

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Partner: Connor Engels

Lab 4. Empirical Formula Determination

Purpose:

To determine the empirical formula for a tin-oxygenated product.

Procedure:

1. Clean and dry an evaporating dish and a watch glass cover. To dry them, heat strongly for 2-3 minutes using a burner. Us forceps or tongs to handle the dish and cover throughout the experiment.
2. Place about 2g of granulated tin in the dish, cover with the watch glass and mass.
3. In the fume hood, add 5ml of 8M nitric acid and replace the watch glass.
4. After the chemical reaction had stopped, heat the dish over a low flame. An excessive amount of popping and spattering indicates that you are heating too rapidly. Continue to heat slowly until the contents are nearly dry.
5. When the popping and spattering no longer occur, remove the evaporating dish from the heat source. Remove the watch glass, taking care not o lose any of the product. Do not clean the watch glass until all the measurements have been made. Break up the solid with a stir rod.
6. Place the dish onto wire gauze. Heat carefully with a hot flame until the solid becomes a pale yellow. Remove the dish from the heat source and let it cool.
7. After the dish has cooled, replace the watch glass and re-mass. Reheat the dish for 2-3 and allow to cool. Re-mass, if the mass does not agree within 0.02g, reheat and re-mass until the last 2 measurements agree.
8. Discard the slid material into the designated container.

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Data:

Mass of tin: 2g

Mass of tin, evaporating dish, and watching glass: 39.21g

Mass after 1st heating: 39.76g

Mass after 2nd heating: 39.71g

Mass after 3rd heating: 39.71g

Observations:

It started to bubble, turn orange, and let off orange vapor. Bad, sulfur kind of smell. Orange vapor started to turn white. Dumped in a little more that 5ml(nitric acid). Left a residue on the watching glass. Tin dissolved into the acid. Acid kind of disappeared. Eventually stopped bubbling and letting off vapor. Took about 8 minutes to stop reacting. Liquid on top of the watching glass. ~~Starts to turn from white to yellow when heating. Yellow to white when heating.~~ Lets off more vapor when hated. A little popping. Some product fell out as we heated it. Turned orange, to white, to yellow. Has an indescribable smell, kind of like cleaning product. Let off an orange vapor. Pretty dry after heating. Some product stuck to the stir rod, some fell off the watch glass. Began to turn a darker yellow with more intense heating. Much stronger smell after stronger heating was finished. Some got more heated than others. It turns a more uniform color with more heating and stirring. Mass was 0.5g off on the first and second heating. Exact on for second and third heating.

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Analysis:

Mass of Sn 2.00 g/118.7 g/mol=0.0168 mol Sn

39.71g-39.21g=0.5g O

mass O 0.5 g/16.00 g/mol=0.0313 mol O

0.0313 mol O/0.0168 mol Sn=1.86≈2 O

0.0168 Sn/0.0168 Sn=1 Sn

Empirical Formula=SnO_2

Conclusion:

The purpose of the lab was to determine the empirical formula for a tin-oxygenated product. We achieved this by combining granulated tin and nitric acid, burning out the excess nitric acid and water, and weighing the tin and product to find the moles of tin and oxygen to derive the empirical formula. We got the formula $SnO_{1.86}$, which we rounded to SnO_2, which is the correct formula. We got this formula by calculating the moles of tin, subtracting the final mass from the initial mass, and using that to calculate the moles of oxygen. Then, we divided both answers by the moles of tin to get the empirical formula. If I were to do it again, I would use closer to 5mL of nitric acid, and try to lose less product. Sources of error likely include losing the product from it falling out while heating, and sticking to the stir rod. This would affect the final weight, thus skewing the how much oxygen is calculated, such as with the mole calculations. Not losing any product would have yielded a better result. Other than losing product, everything else went really smoothly.

5. Reactions in Solution

Precipitation

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Partner: Jeff Somvorachith

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Lab 5. Reactions in Solution Precipitation

Purpose:

> To practice identifying reactions that result in precipitation and practice recognizing and writing equations for double replacement reactions by testing pairs of reagents and looking for signs of chemical change and writing molecular and ionic equations for the reactions.

Prelab:

1. a. Zinc Sulfide: ZnS
   b. Chromium(III) Hydroxide: $Cr(OH)_3$
   c. Lead(II) Phosphate: $Pb_3(PO_4)_2$
2. a. Sodium Chloride: $NaCl \to Na^+ + Cl^-$
   b. Copper(II) Chloride: $CuCl_2 \to Cu^{2+} + 2Cl^-$
   c. Iron(III) Sulfate: $Fe_2(SO_4)_3 \to 2Fe^{3+} + 3SO_4^{2-}$
3. $Fe(NO_3)_3 + K_2CO_3 \to Fe_2(CO_3)_3$ (s) $+ KNO_3$
4. $0.2M*1L=M_1*2L$ $M_1=\underline{0.1M}$

Procedure:

1. Use 4-5 drops of each reagent for each pair of your tests. Note and record any sign of reaction. Test all possible pair of combinations of one group A reagent with one group B reagent. Record your observations in the data table in your notebook. Recall that not all pairs will react, and that sometimes the evidence of reaction, especially formation of a precipitate, take a few moments to appear.
2. If there are any combinations about which you doubt, repeat the tests.
3. Return pipettes containing unused portions of the reagent

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solutions to the proper location. Some of the solutions contain transition metal ions, so should not be washed down the drain. Instead, dump the contents onto a paper towel in the tray labeled "transition metal waste". Wash some items with water and clean others with cotton swabs.

Safety: Barium is toxic, don't touch. wash hands thoroughly.

Data Table:

	$CoCl_2$	**$CuCl_2$**	**$AlCl_3$**	**$BaCl_2$**	**$NiCl_2$**
NaI	No Reaction	Turned Brown, Precipitated	No Reaction	No Reaction	No Reaction
Na_2CO_3	Precipitated	Precipitated	Precipitated	Precipitated	Precipitated
Na_3PO_4	Tuned Purple, Precipitated	Precipitated	Precipitated	Precipitated	Precipitated
Na_2SO_4	No Reaction	No Reaction	No Reaction	Precipitated	No Reaction
NaOH	Turned Purple, Precipitated	Precipitated	Precipitated	Precipitated	Precipitated

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Conclusion:

a. Formula of Precipitates

	$CoCl_2$	**$CuCl_2$**	**$AlCl_3$**	**$BaCl_2$**	**$NiCl_2$**
NaI		CuI_2			
Na_2CO_3	$CoCO_3$	$CuCO_3$	$Al_2(CO_3)_3$	$BaCO_3$	$NiCO_3$
Na_3PO_4	$Co_3(PO_4)_2$	$Cu_3(PO_4)_2$	$AlPO_4$	$Ba_3(PO_4)_2$	$Ni_3(PO_4)_2$
Na_2SO_4				$BaSO_4$	
NaOH	$Co(OH)_2$	$Cu(OH)_2$	$Al(OH)_3$	$Ba(OH)_2$	$Ni(OH)_2$

b. Molecular Formulas

$CuCl_2(aq)+2NaI(aq)\text{->}2CuI_2(s)+2NaCl(aq)$

$CoCl_2(aq)+Na_2CO_3(aq)\text{->}CoCO_3(s)+2NaCl(aq)$

$CuCl_2(aq)+Na_2CO_3(aq)\text{->}CuCO_3(s)+2NaCl(aq)$

$2AlCl_3(aq)+3Na_2CO_3(aq)\text{->}Al_2(CO_3)_3(s)+6NaCl(aq)$

$BaCl_2(aq)+Na_2CO_3(aq)\text{->}BaCO_3(s)+2NaCl(aq)$

$NiCl_2(aq)+Na_2CO_3(aq)\text{->}NiCO_3(s)+2NaCl(aq)$

$3CoCl_2(aq)+2Na_3PO_4(aq)\text{->}Co_3(PO_4)_2(s)+6NaCl(aq)$

$3CuCl_2(aq)+2Na_3PO_4(aq)\text{->}Cu_3(PO_4)_2(s)+6NaCl(aq)$

$AlCl_3(aq)+Na_3PO_4(aq)\text{->}AlPO_4(s)+3NaCl(aq)$

$3BaCl_2(aq)+2Na_3PO_4(aq)\text{->}Ba_3(PO_4)_2(s)+6NaCl(aq)$

$3NiCl_2(aq)+2Na_3PO_4(aq)\text{->}Ni_3(PO_4)_2(s)+6NaCl(aq)$

$BaCl_2(aq)+Na_2SO_4(aq)\text{->}BaSO_4(s)+2NaCl(aq)$

$CoCl_2(aq)+2NaOH(aq)\text{->}Co(OH)_2(s)+2NaCl(aq)$

$CuCl_2(aq)+2NaOH(aq)\text{->}Cu(OH)_2(s)+2NaCl(aq)$

$AlCl_3(aq)+3NaOH(aq)\text{->}Al(OH)_3(s)+3NaCl(aq)$

$BaCl_2(aq)+2NaOH(aq)\text{->}Ba(OH)_2(s)+2NaCl(aq)$

$NiCl_2(aq)+2NaOH(aq)\text{->}Ni(OH)_2(s)+2NaCl(aq)$

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c. Net ionic formulas

$\mathrm{Cu^{2+}+2I^{+}\text{->}CuI_2}$

$\mathrm{Co^{2+}+Co_3^{2-}\text{->}CoCO_3}$

$\mathrm{Cu^{2+}+Co_3^{2+}\text{->}CoCO_3}$

$\mathrm{2Al^{3+}+3CO_3^{2-}\text{->}Al_2(CO_3)_3}$

$\mathrm{Ba^{2+}+CO_3^{2-}\text{->}BaCO_3}$

$\mathrm{Ni^{2+}+CO_3^{2-}\text{->}NiCO_3}$

$\mathrm{3Co^{2+}+2PO_4^{3-}\text{->}Co_3(PO_4)_2}$

$\mathrm{3Cu^{2+}+2PO_4^{3-}\text{->}Cu_3(PO_4)_2}$

$\mathrm{Al^{3+}+PO_4^{3-}\text{->}AlPO_4}$

$\mathrm{3Ba^{2+}+2PO_4^{3-}\text{->}Ba_3(PO_4)_2}$

$\mathrm{3Ni^{2+}+2PO_4^{3-}\text{->}Ni_3(PO_4)_2}$

$\mathrm{Ba^{2+}+SO_4^{2-}\text{->}BaSO_4}$

$\mathrm{Co^{2+}+2OH^{-}\text{->}Co(OH)_2}$

$\mathrm{Cu^{2+}+2OH^{-}\text{->}Cu(OH)_2}$

$\mathrm{Al^{3+}+3OH^{-}\text{->}Al(OH)_3}$

$\mathrm{Ba^{2+}+2OH^{-}\text{->}Ba(OH)_2}$

$\mathrm{Ni^{2+}+2OH^{-}\text{->}Ni(OH)_2}$

6. Stoichiometry and Limiting Reagent

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Partner: Jeff Somvorachith

11/4/24

6. Stoichiometry and Limiting Reagent

Purpose:

> To practice using stoichiometry and identifying and calculating the limiting reagent by reacting copper(II) sulfate with iron, determining the limiting reagent, and observing the reaction for precipitates and changes and calculating the percent yield and writing formulas using stoichiometry.

Procedure:

1. Place about 7.00g of copper(II) sulfate in a beaker.
2. Add about 50mL of water to the beaker.
3. Arrange the beaker and stand.
4. Carefully heat and stir the mixture in the beaker. The solution should be hot but not boiling. After all the crystals have dissolved, remove the beaker from the heat.
5. Add about 2.00g of iron fillings slowly to the hot $CuSO_4$ solution while stirring. Record observations.
6. Allow the beaker to cool for 10-15 minutes.
7. Pour off(decant) the solution into a different beaker. Pouring the solution down a stirring rod is recommended. Make sure not to disturb the solid product.
8. Add a small amount of water(at least 10mL) to the copper and stir.
9. Let the copper settle to the bottom of the beaker and decant again.
10. Dry the copper and mass it.

Data:

Mass empty beaker w/ tape	**Mass $CuSO_4$**	**Mass Fe**	**Mass Beaker+Copper**
117.25g	7.00g	2.00g	119.42g

Observations:

Water was exactly 50mL, all other measurements were exact.

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The tape on our beaker burned. Took about 1 minute 50 seconds to dissolve. Letting off steam. Seems to heat up more when adding the iron. Solution turned from blue to grey. Iron seemed to rust. Combined to form Fe_3O_3. Has an iron sort of smell. Precipitate settled to the bottom. Was a darker blue fluid with a red bubbling deposit on the bottom. Some filings stuck to the stirring rod and side of the beaker. Some of the tape burned off. After decanting, deposit was a wet, red powder. Decanted liquid was more of a teal color. There was a little bit of deposit that was decanted. Second decant was a more red color liquid. Definitely lost some of the precipitate. Third decant was more cloudy. Decanted liquid was a very clear, very light blue liquid with some precipitate. Final product was wet, red, some black, and powdery. After drying, it is a solid mass with a reddish-brown color. Very dry. Some iron left over. Some of the product was black.

Analysis:

Limiting Reactant	**Theoretical Yield**	**Percent Yield**
Iron	2.28g	95.2%

1. Iron is the limiting reactant. No more iron was left over, only copper.
   (2.00g Fe /55.85)*63.55=2.28g Cu
   (7.00g $CuSO_4$ /[63.55+32.06+4*16.00])*63.55=2.56g Cu
2. Iron was the limiting reactant so that there would be no iron left over with the copper. If copper(II) sulfate was the limiting reactant, there would be iron left in the copper, affecting the weight and calculation.

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3. Water was added to remove any underacted copper(II) sulfate, iron, and iron(II) sulfate. The solution was washed to isolate the copper.

4. The water added didn't matter if the measurement w for water was exact because it wasn't included in the reaction. Water was only used to wash the copper and help facilitate the reaction.

5. $Fe+CuCO_4$->$Cu+FeSO_4$ The formula for the iron-containing compound was $FeSO_4$. The observations support this as the color of the copper(II) sulfate solution got darker, so it is iron(II) that was used. So, it would be $Fe^{2+}+SO_4^{2-}$->$FeSO_4$, making the formula $FeSO_4$.

6. <u>95.2% Yield</u> 119.42g-117.25g=2.17g (2.17/2.28)*100=95.2%

Conclusion:

The purpose of this lab was to practice identifying and calculating limiting reactants. We did this by combining copper(II) sulfate with iron filings and weighing the mass of the empty beaker, beaker with copper, and both reactants to use the mass to calculate the theoretical yield, actual yield, and percent yield. Through achieving this purpose, we were able to determine the limiting reactant and percent yield using the mass, and used observations to determine the charge of the iron used, when the reaction was complete, and what

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was left over after and what was left over was the limiting reactant. The limiting reactant turned out to be the iron, both through math and observations, as copper was all that was left after. All of the iron would have reacted to form iron(II) sulfate, leaving behind copper. The calculated theoretical yield was 2.28g of copper. We ended up with a final weight of 2.12g, making for a 95.2% yield. Some possible errors that may have affected our results are some copper(II) sulfate that was left unreacted, interacted iron filings stuck on the beaker and stirring rod, the tape on the beaker partially burned off reducing the mass, and accidentally decanting off some of the reactants or copper. If I were to to the lab again, I would use an amount of reactant close to what is required so that there are less unreacted reagents left over, I wouldn’t use tape, and I would make sure everything gets reacted fully, not stuck to the beaker or stirring rod. The percent yield tells me that some of the product was lost or never created. There was less copper left in the beaker than there could have been. Our methods were not thorough or careful enough to get perfect results and our results are only as good as the scale we mass it it and how complete our reaction was.

7. Standardization of a NaOH Solution

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Partner: Jeff Somvorachith

7. Standardization of a NaOH Solution

Purpose:

> To determine the molarity of an NaOH solution by reacting a kniwn volume of each reagent, NaOH and KHP, and using the known molarity of the KHP solution to determine the molarity of the NaOH solution when the stoichiometric endpoint is detected.

Procedure:

1. Flush a buret several times with distilled water. Then, rinse the buret with 5mL portions of the NaOH solution. Drain the buret each rinse through the buret tip. Discard each rinse in the "water bases" container.
2. Using a clean funnel, fill the buret with the NaOH solution. Wait a few seconds, then record the initial buret volume of NaOH.
3. Place approximately 2 grams of KHP in an Erlenmeyer flask and dissolve it in distilled water. Add 2 drops of phenolphthalein indicator. Make certain that all the solid has dissolved.
4. Slowly add the NaOH to the KHP solution in 1-2mL increments. As the endpoint nears, the color change of the indicator slows. Occasionally rinse the walls of the flask with distilled water. Continue to add NaOH until the endppint is reached. The color should persist for 30 seconds. Record the final volume of NaOH in the buret.
5. Refill the buret if necessary and repeat the titration process at least 2 more times.
6. Calculate the molarity of the NaOH solutions for each trial.

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7. Calculate the percent deviation between the trials.

    % deviation=(highest molarity - lowest molarity)/average molrity *100

8. If the % deviation is greater than 5%, repeat the titration process.

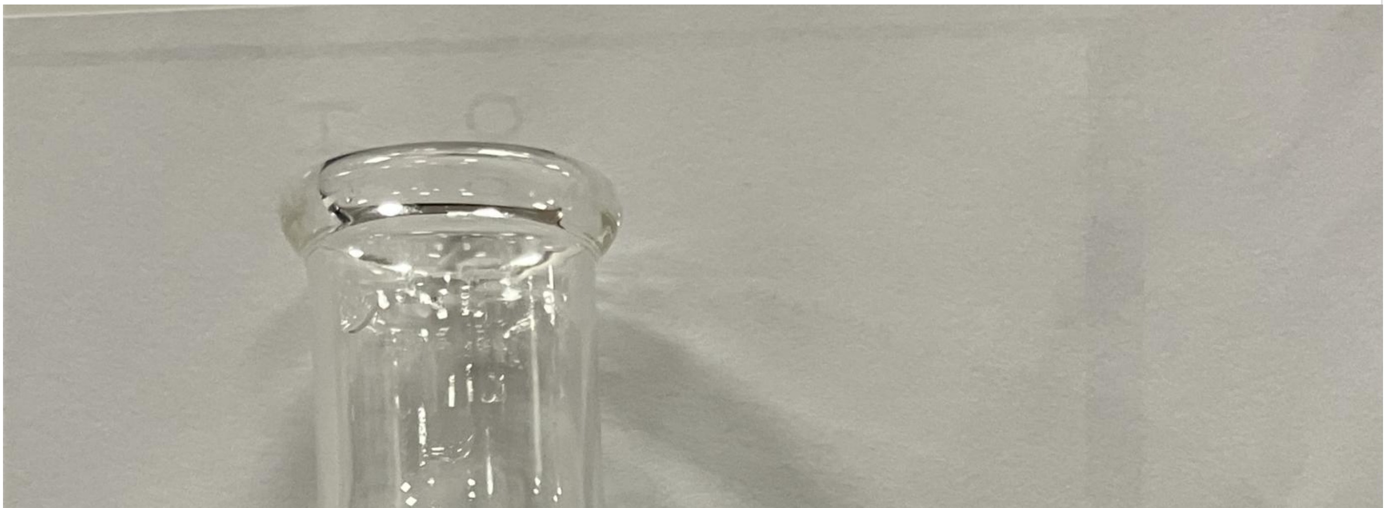
Data:

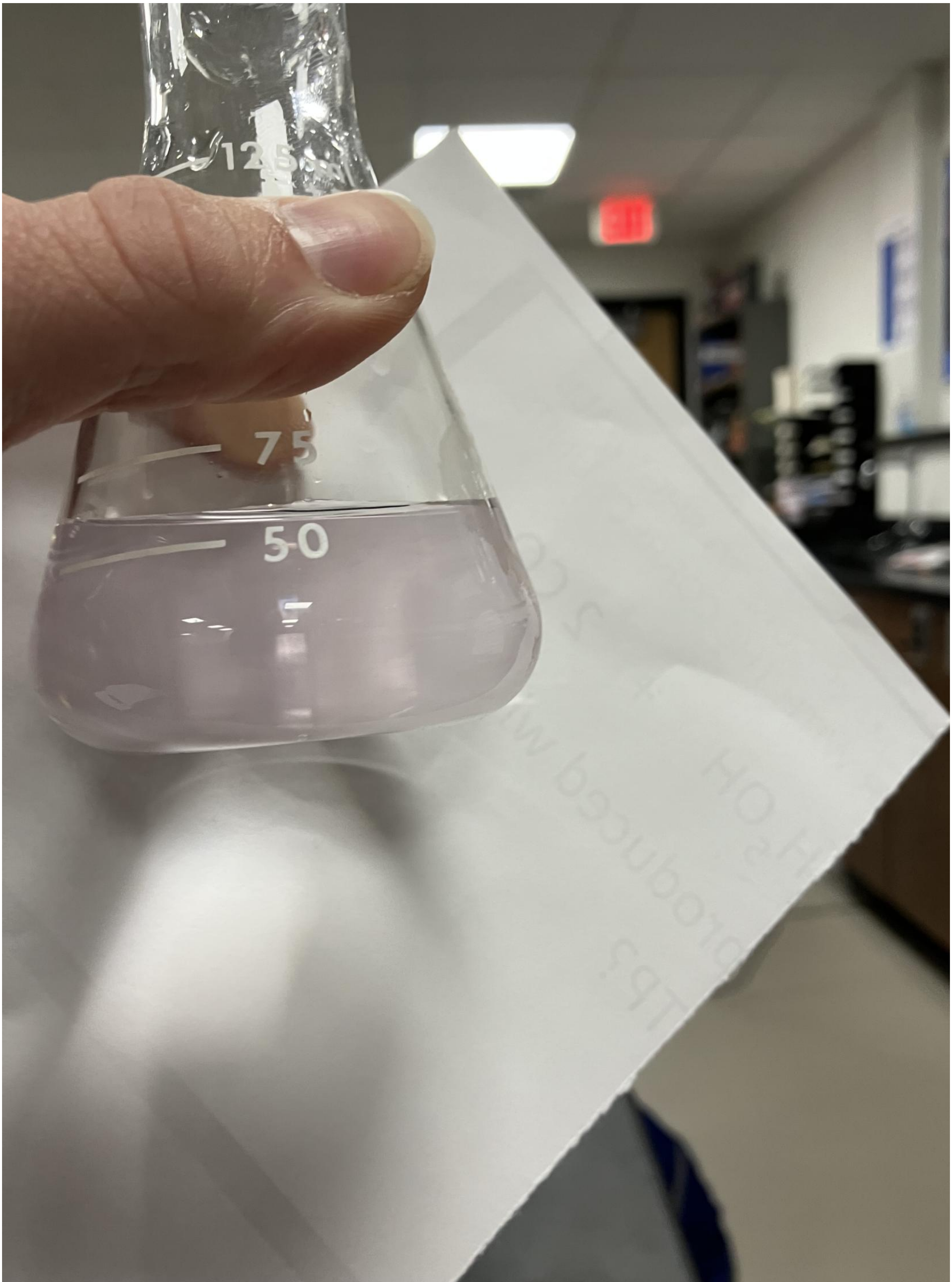
	Trial 1	**Trial 2**	**Trial 3**	**Trial 4**	**Trial 5**
KHP used	2.00g	2.00g	2.00g	2.00g	2.00g
Initial NaOH	0.20mL	10.00mL	19.70mL	29.4mL	39.15mL
Final NaOH	10.00mL	19.70mL	29.4mL	39.15mL	48.85mL

Observations:

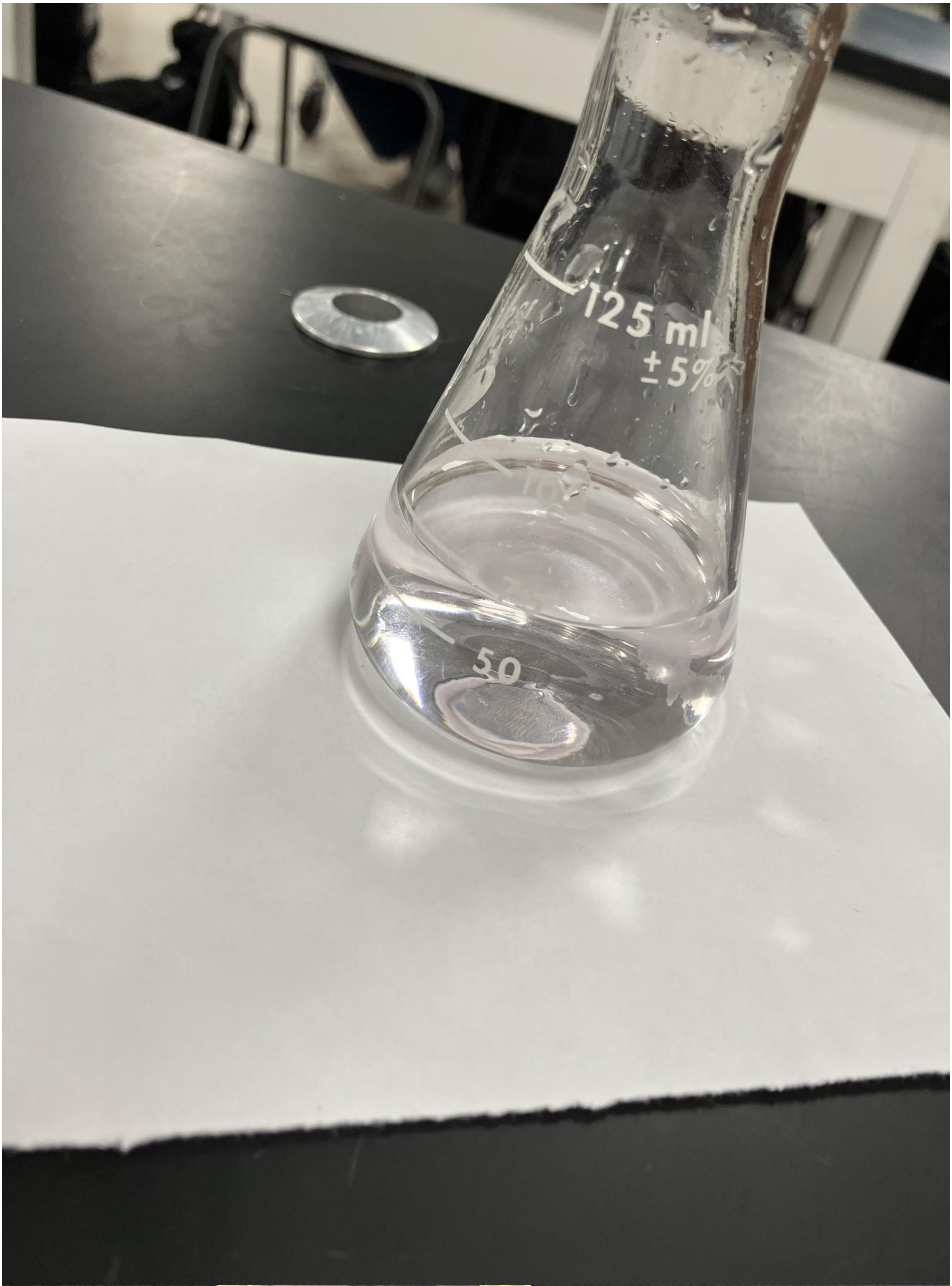
First trial turned a dark pink after about 10mL of NaOH. Way too dark. About a drop too much of NaOH. Trial 2 turned a dark pink, not as dark as trial 1. Trial 3 was similar to trial 2, same color and same amount of NaOH used. Magic numer seems to be 9.7mL when it turns pink. We lost a little bit of the KHP for the 4th trial. Trial 4 got a very very faint pink color. Even though we lost some KHP, it looks near perfect titration. Trial 4 looked clear until help against white paper. Trial 5 was a dark pink, though better than 1, 2, and 3. 4th trial was the best out of the 5.

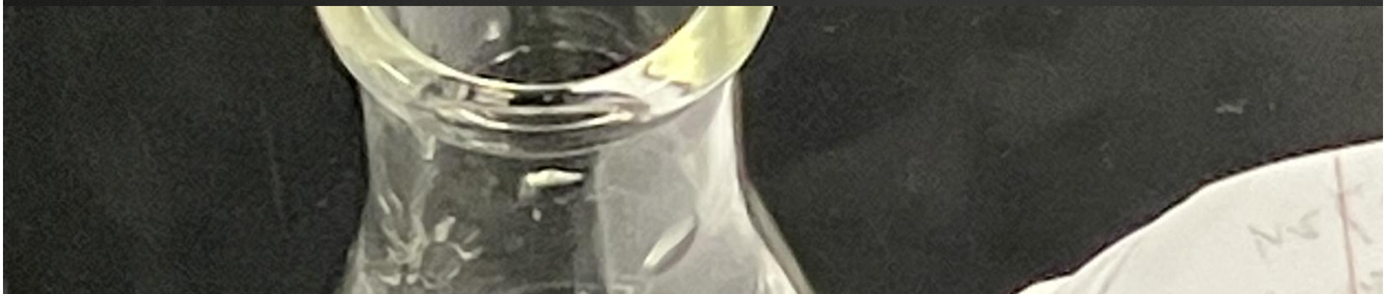
Trial 4




125
75
50


125 ml
±5%


125 ml
±5%
100
75
50



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Analysis: $KHC_8H_4O_4+NaOH\text{->}NaKC_8H_4O_4+H_2O$

Trial 1:

10.00mL-0.20mL=9.80mL NaOH used

2.00g/(39.10+1.008+8*12.01+4*1.008+4*10.00)*(1/1)*(9.80/1000)=0.999~1.0M

Trial 2:

19.70mL-10.00mL=9.70mL NaOH used

2.00g/(39.10+1.008+8*12.01+4*1.008+4*10.00)*(1/1)*(9.70/1000)=1.01M

Trial 3:

29.40mL-19.70mL=9.7mL NaOH used

2.00g/(39.10+1.008+8*12.01+4*1.008+4*10.00)*(1/1)*(9.70/1000)=1.01M

Trial 4:

39.15mL-29.40mL=9.75mL NaOH used

2.00g/(39.10+1.008+8*12.01+4*1.008+4*10.00)*(1/1)*(9.75/1000)=1.00M

Trial 5:

48.85mL-39.15mL=9.70mL NaOH used

2.00g/(39.10+1.008+8*12.01+4*1.008+4*10.00)*(1/1)*(9.70/1000)=1.01M

Average Molarity:

(1.00M+1.01M+1.01M+1.00M+1.01M)/5=1.006~1.01M on average

% Deviation:

(1.01-1.00)/1.01 *100=0.990% deviation

Conclusion:

The purpose of the lab was to determine the molarity of an NaOH solution by reacting KHP with NaOH and getting as close to the endpoint as possible. We achieved this purpose by takings the before and after volume of the NaOH to when the enpoint is reached or passed, the difference of which was used to calculate the molarity of the NaOH solution. On average, 9.7mL was too far, though 9.75mL was used to get a near perfect titration. We got an average molarity of 1.01M NaOH. The target was about 1.00M.

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Our % deviation was 0.990%, meaning our answers between each trial was pretty close to the average. Since we got one really good titration, the % deviation means we were really close to being exact and getting a light pink color, within a few hundreths of a milliliter. Some possible errors may include inaccurate recording of the NaOH solution in the buret, NaOH stick to the beaker or tip of the buret, we lost some KHP when transferring from the weight paper to the flask, and improperly cleaned flasks between trials. If I did it again, I would pay more attention to how long it takes for it to turn clear again to get a more accurate titration, be more careful so that everything measured gets mixed, and spending more time reading the buret and measuring out the reactants.

8. Vinegar Titration Lab

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8. Vinegar Titration Lab

Purpose:

Determine the concentration of acetic acid in vinegar.

Procedure:

1. Using a clean funnel, fill a buret with NaOH solution, which was determined to have a molarity of 1.00M. Record the initial volume of NaOH in the buret.
2. Measure out 10mL of vinegar using a graduated cylinder. Transfer the vinegar into an Erlenmeyer flask. Add 2 drops of phenolphthalein to the flask.
3. Slowly add the 1.00M NaOH to the vinegar in 1-2mL increments. As the endpoint nears, the color change of the indicator slows. If you feel that you need less than a drop of NaOH, wash the tip of the buret and walls of the flask with a very small amount of distilled water. Continue to add NaOH until the endpoint is reached. The endpoint should be a very faint pink color. Record the final volume of NaOH in the buret.
4. Refill the buret if necessary and repeat the titration process at least 2 more times.
5. Calculate the molarity of the NaOH solutions for each trial.
6. Calculate the percent deviation between the trials. (high-low)/avg*100)
7. If the % deviation is greater than 2%, repeat the titration process.

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Data:

	Trial 1	**Trial 2**	**Trial 3**	**Trial 4**	**Trial 5**	**Trial 6**
Amount of vinegar	10mL	10mL	10mL	10mL	10mL	10mL
Initial NaOH	1mL	9.9mL	18.0mL	26.2mL	34.4mL	36.65mL
Final NaOH	9.9mL	18.0mL	26.2mL	34.4mL	42.65mL	44.6mL

Observations:

First trial turned dark pink, too much NaOH. Accidentally let out a stream of NaOH near the endpoint. Trial 2 was a light pink, better than trial 1. Trial 3 was terrible. I accidentally added a stream of NaOH. Trial 4 was clear. half a drop turned it pink. Trial 5 was also a drop over from being clear and pink. Trial 6, you can barely tell that it is pink. If you look really close, you can see a very faint pink. It is really close to perfect. With some distilled water.

Analysis: $CH_3COOH+NaOH\text{->}CH_3COONa+H_2O$

Trial 1:

9.90mL-1.00mL=<u>8.90mL NaOH used</u>

1.00M*8.90mL/1000=0.0089 mol NaOH

0.0089 mol/10mL/1000=<u>0.890M Acetic Acid</u>

Trial 2:

18.0mL-9.9mL=<u>8.1mL</u>

(8.1/1000)/(10mL/1000)=<u>0.81M</u>

Trial 3:

26.2mL-18.0mL=<u>8.2mL</u>

(8.2mL/1000)/(10mL/1000)=0.82M

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Trial 4:

34.4mL-26.2mL=8.2mL

(8.2/1000)/(10mL/1000)=0.82M

Trial 5:

42.65mL-34.4mL=8.25M

(8.25mL/1000)/(10mL/1000)=0.825M

Trial 6:

44.6mL-36.65mL=7.95mL

(7.95mL/1000)/(10mL/1000)=0.795M

Average Molarity:

(0.89M+0.81M+0.82M+0.82M+0.825M+0.795M)6=0.842M

% Deviation:

(0.825-0.810)/0.842*100=1.78%

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Conclusion:

The purpose of the lab was to determine the concentration of acetic acid in vinegar. I achieved this by combining NaOH with vinegar until the indicator turned pink. The average concentration across the trials was determined to be 0.842M. The percent deviation was 1.78%, taking the 3 closest trials. This shows that I could be more consistent between trials. It shows that I am not great, but not terrible at doing titrations. Some possible errors include the calculations for trial 1(in notebook), improperly read measurements, and contamination during trial 6. For those reasons, I would take trial 1 and 6 with a grain of salt, as they are most definitely inaccurate. If I were to do this lab again, I would be more patient. I would spend more tme taking measurements and mixing the reactants.

9. Molar Volume of a Gas

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9. Molar Volume of a Gas

Purpose:

> To practice calculating the molar volume of a gas by mixing magnesium metal and hydrochloric acid to form hydrogen gas, then using the mass of magnesium used and volume of hydrogen collected to calculate the volume of one mole of hydrogen gas.

Procedure:

1. Fill a 400mL beaker two-thirds full of room temperature water.
2. Cut a piece of magnesium ribbon 1-1.5cm long. Mass the magnesium ribbon.
3. Roll the magnesium ribbon into a loose coil. Tie it with a piece of string on one end,approximately 25cm in length.
4. This step requires the use of 6M hydrochloric acid, which is caustic and corrosive. Pour approximately 10mL of 6M HCl into a 50mL eudiometer.
5. While holding the eudiometer in a slightly yipped position, slowly pour distilled water into the eudiometer, being careful not to mix the HCl and water. Completely fill the eudiometer so that there is no air.
6. Lower the magnesium coil into the water in the eudiometer to a depth of about 5cm. Insert the rubber stopper into the open end of the eudiometer to hold the thread in place. The one-hole stopper should displace some water from the eudiometer.
7. Cover the hole of the stopper with your finger and invert the eudiometer into the 400mL beaker of water.

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Clamp the eudiometer into position on the ring stand. The acid flows down the tube(less dense) and reacts with the magnesium. The acid is more dilute.

8. When the magnesium has disappeared entirely and the reaction has stopped, cover the stopper with your finger and carefully transfer the eudiometer to a 1000mL graduated cylinder or other tall vessel filled with water. Adjust the level of the eudiometer so that the height of the liquid in the 2 vessels match. Read the volume of hydrogen as accurately as possible.

9. Record the temperature of the room and atmospheric pressure.

10. Find a table of water-vapor pressures and determine the vapor pressure of the water at the temperature of the room.

Data:

Temperature: 21.5 ^{o}C

Pressure: 29.00 inHg

Water-vapor pressure: 19.25 mmHg

Mass of Mg: 0.04g

Volume of hydrogen gas: 44.85 mL

Observations:

Nothing happens when adding magnesium to water. You can see the HCl move downwards when inverting the eudiometer. Took a minute or 2 before the HCl reached the magnesium and started reacting. Started bubbling more as time went on. Started to create an air pocket at the top(bottom) of the tube. There are bubbles coming out of the stopper.

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Started slowing down as magnesium disappeared. Some HCl went out the stopper. Some bubbles stuck to the side of the tube and to the string. Magnesium eventually disappeared.

Calculations:

1. Number of moles of magnesium
   0.04/24.30 = 0.00165 moles of hydrogen reacted.
2. Partial pressure of hydrogen
   29.00*2.54/10 = 736.6 mmHg
   736.6-19.25 = 717.35
   717.35/760 = 0.944 atm of hydrogen
3. Volume at 1 atm
   $P_1V_1=P_2V_2$ $0.944*44.85=1.00*V_2$ $V_2=$42.3mL
4. Volume at 273.15 ^{O}K
   $V_1/T_1=V_2/T_2$ $42.3/(21.5+273.15)=V_2/273.15$
   $V_2=$39.2mL
5. Volume of 1 mol hydrogen at STP
   (29.2/1000)/0.00165=23.8 mol hydrogen

Conclusion:

The purpose of this lab was to calculate the molar volume of hydrogen. We achieved this by mixing a known amount of magnesium with HCl to form hydrogen gas in a eudiometer. We used he mass of the magnesium to find the moles of Mg and H_2, calculated the partial pressure of H_2, standardized it to STP, and converted that to density of H_2 at STP. It was calculated to be 23.8 mol.

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The formula for the Mg HCl reaction is $Mg+2HCl \rightarrow H_2+MgCl_2$. Since the moles of Mg equal the moles of H_2 in the equation, the moles are the same. We needed to adjust for water-vapor pressure because the water will evaporate to some degree during the experiment, mixing with the hydrogen. We measured and calculated a density of 23.8 L/mol. We should have gotten a value of 22.4 L/mol, making for a 6.25% error. We measured the level of the eudiometer in a bin of water so the pressure would be equal. Some potential errors stem from the mass being so low, making the balance less accurate to the decimal place, and bad reading of the eudiometer. If I were to do the lab again, I would use a balance that can measure to more decimal places, and have a better way of calculating the value of hydrogen produced.

10. Additivity of Heats: Hess's Law

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Partners: Jeff Somvorachith, Connor Engels

10. Additivity of Heats of Reaction: Hess's Law

Purpose:

To practice applying Hess's Law using a coffee cup calorimeter and confirming that the heat of one reaction should be equal to the sum of the heats for the other reactions.

Pre-Lab:

(1) $Na^{+}_{(aq)} + OH^{-}_{(aq)} \text{-->} NaOH_{(s)}$ $-\Delta H_1$

(2) $NaOH_{(s)} + H^{+}_{(aq)} + Cl^{-}_{(aq)} \text{-->} H_2O_{(l)} + Na^{+}_{(aq)} + Cl^{-}_{(aq)}$ ΔH_2

(3) $Na^{+}_{(aq)} + OH^{-}_{(aq)} + H^{+}_{(aq)} + Cl^{-}_{(aq)} \text{-->} H_2O_{(l)} + Na^{+}_{(aq)} + Cl^{-}_{(aq)}$ ΔH_3

Procedure:

1. Wear protective equipment.
2. Open Vernier Graphical Analysis. Connect the temperature probe.
3. Use a utility clamp and a slit stopper to suspend the temperature probe from a ring stand.
4. Place a Styrofoam cup into a 250mL beaker. Measure out 100.0mL of distilled water into the cup. Lower the temperature probe into the solution.
5. Weigh out around 2 grams of solid sodium hydroxide and record the mass to the nearest 0.01g.
6. Begin collecting data and obtain the initial temperature, t_1. After obtaining the initial temperature, add the NaOH to the cup. Stir continuously until the temperature has maximize and has begun to drop. Record the maximum temperature, t_2. After 3.3-4 minutes, data collection is completed.
7. Graph temperature vs. time(done automatically). Examine the data points along the curve to verify the values of t_1 and t_2.

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8. Rinse and dry the temperature probe, cup, and stirring rod. Dispose of the solution as directed.

9. Repeat steps 4-8, using 100.0mL of 0.5M hydrochloric acid instead of water. Use the same amount of NaOH.

10. Repeat steps 4-8, instead using 50.0mL of 1.0M HCl, and using 50.0mL of 1.0M NaOH.

Observations:

Combining the water and NaOH creates some bubbles. Combining 1.0M HCl and NaOH turns a litle pink.

Data:

	Reaction 1	**Reaction 2**	**Reaction 3**
Mass NaOH	2.00g	2.00g	
Mass Total	102g	102g	100g
Final Temperature	25.073 °C	31.593 °C	27.357 °C
Initial Temperature	20.675 °C	20.622 °C	20.504 °C
Change in Temperature	4.397 °C	10.972 °C	6.853 °C

Calculations:

Mass_{R1}= 100mL*1.00g/mol=100g +2.00g = 102g

Mass_{R2}= 100mL*1.00g/mol=100g +2.00g = 102g

Mass_{R3}= 50mL*1.00g/mol=50g *2=100g

q_{R1}= 4.18*102*4.397 =1875J /1000=1.875kJ

q_{R2}= 4.18*102*10.972 =4678J /1000=4.678kJ

q_{R3}=4.18*100*6.853 =2865J /1000=<u>2.865kJ</u>

ΔH=-q   q=-ΔH

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ΔH_{R1}= -1.875 kJ

ΔH_{R2}= -4.678 kJ

ΔH_{R3}= -2.865 kJ

mol_{R1}=2.00/(22.99+16.00+1.008)= 0.05 mol

mol_{R2}=2.00/(22.99+16.00+1.008)= 0.05 mol

mol_{R3}=1.0*50.1000= 0.05 mol

$\Delta H/mol_{R1}$=-1.875/0.05= -37.5 kJ/mol

$\Delta H/mol_{R2}$=-4.678/0.05= -93.6 kJ/mol

$\Delta H/mol_{R3}$=-2.865/0.05= -57.3 kJ/mol

$-\Delta H_{R1}$ ΔH_{R2} ΔH_{R3}

-37.5+-57.3= -94.8 kJ/mol

37.5+-93.6= -56.1 kJ/mol

|-94.8+93.6| / |-93.6| *100= 1.28% error

Conclusion:

The purpose of the lab was to practice applying Hess's Law and verifying that the heat of a reaction is equal to the sum of the other reactions. We achieved this by using a coffee cup calorimeter to combine NaOH with water and HCl and using a thermal probe to measure the initial and final temperature, which was used along with the mass or volume of NaOH, to find the heat released, which was used to find ΔH. We then used the mass of NaOH to find the ΔH/mol. We then added two reactions and compared them against the third to verify the data/calculations.

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We ended up with a 1.28% error, meaning some heat was probably lost, though it was close. This error could stem from inaccurate measuring of the NaOH or HCl, but it is more likely from losing heat to the Styrofoam cup or the atmosphere. If I were to do this lab again, I would use more accurate measuring tools, something that would absorb or lose less heat, like 2 coffee cups, and a more enclosed environment, such as a sealed container to minimize heat lost to the atmosphere.

11. Molecular Mass Determination Using Boiling and Freezing Point

Partners: Jeff Somvorachith

11. Molecular Mass Determination Using Boiling and Freezing Point

Purpose:

To practice calculating the molecular mass of solutes by comparing boiling and freezing point of two solutes with known concentrations.

Procedure:

A.

1. Measure 50g of distilled water into a 125ml flask. set up the flask on a hot plate and hang a thermometer in the flask using a clamp, not letting the thermometer touch the flask. Record the temperature to the nearest 0.5 ^oC after heated.
2. Prepare a solution of 10g potassium nitrate in 50g of distilled water. Determine the boiling point of the solution using the equation $m=\Delta T_{BP}/K_{BP}$ where K_{BP} =0.512 ^oC, m=molality, and ΔT=change in temperature. Dispose of the solution by rinsing it down a sink.
3. Repeat step 2, but replace the KNO_3 with $C6H12O_6$, Dextrose.

B.

1. Dissolve 5g of KNO_3 in 50g of water in a 125ml flask. Place the flask in a beaker with rock salt, with the beaker wrapped in a paper towel. Wait until crystals form in the solution and becomes slushy. Record the temperature. Rinse the solution down a drain.
2. Dissolve 10g $C_6H_{12}O_6$ in 50g of distilled water. Place the flask with the solution in a beaker filled with ice an rock salt and wait until the solution crystalizes. Measure the temperature and rinse down a drain.

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Data:

Solution	**BP(ºC)**	**ΔBP(ºC)**	**FP(ºC)**	**ΔFP(ºC)**
H_2O	98.5	XXXXX	0	XXXXX
$C_6H_{12}O_6$	102.0	3.50	-2.00	2.00
KNO_3	100.5	2.00	-3.50	3.50

Analysis:

Solution	**avg Molality(m)**	**avg experimental molar mass(g/mol)**	**actual molar mass(g/mol)**	**avg percent error(%)**
H_2O	xxxxx	xxxxx	xxxxx	xxxxx
$C_6H_{12}O_6$	3.94	107	180.156	40.6
KNO_3	1.45	104	101.11	2.97

KNO_3:

Boiling:

m=2.00/(2*0.512)

m=1.94m *(50/1000)=0.097 mol

10/0.097= <u>103 g/mol</u>

Freezing:

m=3.50/(2*1.85) =0.950m *(50/1000) =0.0475 mol

5/0.0475=<u>105 g/mol</u>

$C_6H_{12}O_6$:

Boiling:

m=3.50/0.515 =6.80m *(50/1000)=0.34 mol

10/0.34=<u>29.4 g/mol</u>

Freezing:

m=2.00/1.85 = 1.08m 8(50/1000)=0.054 mol

10/0.054=<u>185 g/mol</u>

% error:

KNO_3:

Boiling:

|103-101|/101 *100%=<u>1.98% error</u>

Freezing:

|105-101|/101 *100%=<u>3.98% error</u>

$C_6H_{12}O_6$:

Boiling:

|29.4-180|/180 *100%=<u>83.7% error</u>

Freezing:

|185-180|/180 *100%=<u>2.78% error</u>

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Values:

Solution	**BP(^{o}C)**	**ΔBP(^{o}C)**	**FP(^{o}C)**	**ΔFP(^{o}C)**	**m$_F$(m)**	**m$_B$(m)**
H_20	98.5	xxxxx	0	xxxxx	xxxxx	xxxxx
$C_6H_{12}O_6$	102	3.50	-2.00	2.00	1.08	6.80
KNO_3	100.5	2.00	-3.50	3.50	0.950	1.94

Solution	**exp g/mol (F)**	**exp g/mol (B)**	**Actual g/mol**	**% error (F)**	**% error (B)**
H_20	xxxxx	xxxxx	xxxxx	xxxxx	xxxxx
$C_6H_{12}O_6$	185	29.4	180.156	2.78	83.7
KNO_3	105	103	101.11	3.96	1.98

Conclusion:

> The purpose of this lab was to practice calculating molecular mass of solutes by comparing the boiling and freezing point of two solutions with a known concentration, and using that to find the molar mass. We achieved this by finding the boiling and freezing points, finding the molality, and then calculating the molar mass. We only needed to measure the boiling point of water and not the freezing point because it is easier to accurately measure the boiling point, as well as more consistent across measurements. The major source of error in our experiment was inaccurate measurements of the temperature and solution. Other sources of error include delayed boiling and crystallization and heat gain/loss to the environment. Ice melts with $CaCl_2$ when the outside environment is below the freezing point of water because adding the salt lowers the freezing as it dissociates and

bond with the water. $CaCl_2$ is more effective than NaCl because it produces more ions when dissolved per mole, leading to a larger decrease in the freezing point. If I were to do this lab again, I would measure out the solution more accurately and insulate the Ice better. The boiling trial for the potassium nitrate was the best of the two trials, giving a 1.98% error. The freezing point of the dextrose gave the best result of the two trials with a 2.78% error. These most likely resulted from inaccurate temperature and solution measurements. The molar mass of the boiling trial for dextrose was the only major outlier in % error.

12. Equilibrium and LeChatlier's Principle

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5/8/25

Partner: Jeff Somvorachith

Lab 12. Equilibrium and Le Chatlier's Principle

Purpose:

To investigate Le Chatlier's Principle and how it relates to equilibrium/how equilibrium shifts.

Procedure:

1. Thoroughly rinse and dry a 50mL beaker with a paper towel and then use the markings on the side to measure out 20mL of ethanol into the beaker.
2. Rinse and dry 5 test tubes of about the same size.
3. Examine the solid Cobalt(II) chloride, noting both it's color and the formula for the compound as shown on the label.
4. Place a small amount of the solid in the beaker of ethanol. Stir to dissolve the solid until solution is a purple color.
5. Fill the 5 test tubes with 2-3mL of the alcoholic cobalt solution. Get roughly equal amounts. Leave some solution in the beaker for steps 8/9. One test tube will be your control.
6. In one of the test tubes, add drops, or until there is a change, of distilled water, one at a time while recording observations after each drop. Gently swirl to mix the contents. Duplicate the process with the other 3 test tubes until all 4 are the same color.
7. A.
   Take one of the test tubes from step 6 to a fume hood. Use a dropper to add 5 drops(or until there is a change) of 12M hydrochloric acid, one drop at a time, to the solution in the test tube. Gently swirl the test tube, recording observations.

   B.
   In a second test tube, add 2 small clumps of solid calcium chloride. Gently stir to dissolve, taking observations.

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C.
To a third test tube, add 25 drops of acetone until a permanent color change occurs. Record observations.

D.
To a fourth test tube, add 5 drops of silver nitrate, $AgNO_3$, one at a time. Gently swirl to mix. Record observations.

8. To the solution remaining in the beaker, add distilled water to get a purple color, about half way between the blue and pink. Place the beaker on a hot plate until a color change occurs. Record observations.

9. Chill the beaker in an ice bath to see if the color change is reversible. Record observations.

Observations:

step 4: Ethanol and cobalt chloride solution is a dark blue color.

step 6: Changed from deep blue to to deep purple, then to pink.

step 7a: Solution turned light blue, then settle to a dark blue at the bottom.

step 7b: Calcium chloride turned a light blue and the solution around it turned a dark blue.

step 7c: Solution turned a dark purple. then turned dark blue, then dark purple.

step 8: Turned a dark blue when heated

step 9: Returned to a pink when chilled.

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Conclusion:

Net-ionic equation for the equilibrium:

(pink complex) $Co(H_2O)_6^{2+}(aq) + 4Cl^-(aq) \leftrightarrow CoCl_4^{2}$-(aq) + $6H_2$)(l) (blue complex)

1. The pink cobalt complex was favored when adding water because the increase in water shifted the equilibrium to the left, favoring the reverse reaction.
2. The blue cobalt complex, $CoCl_2$, was favored when the HCl or Calcium chloride was added because it dissociates in the solution, causing an increase in Cl^- ions, which is a common ion in both of the reagent, shifting the equilibrium to the right, favoring the forward.
3. When acetone was added, it effectively removed some of the water from the system, shifting the equilibrium to the right to replace the water, favoring the forward reaction and the blue cobalt complex.
4. When silver nitrate was added, it decreased the Cl^- ions because when the silver nitrate dissociate, it releases Ag^+ ions, which react with the free Cl^- ions forming AgCl. This causes the equilibrium to shift to the left, favoring the reverse reaction and the pink cobalt complex to form.
5.
6. Heating the solution favored the blue cobalt complex and cooling it favored the pink cobalt complex.
   $Co(H_2O)_6^{2+}(aq) + 4Cl^-(aq) + heat \leftrightarrow CoCl_4^{2}$-(aq) + $6H_2$)(l)   ΔH=+50 KJ/mol
   Since it turned blue when heating and pink when cooling, that means that heat is a reactant and the solution is endothermic. Adding heat causes the equilibrium to shift right, favoring the forward reaction and the blue cobalt complex. Removing heat shifts the equilibrium to the left, favoring the reverse reaction and the pink cobalt complex.