Arthur Vining Davis High Impact Fellows Projects   High Impact Curriculum Fellows

4-30-2014

AVD Tablet Labs: Virtual Chemistry Laboratories, Honors Chemistry, Grade:10

Caleb A. Arrington
Wofford College, arringtonca@wofford.edu

Tina Kanipe
Broome High School

Deep Sangani
Wofford College

Follow this and additional works at: http://digitalcommons.wofford.edu/avdproject

Part of the Chemistry Commons

Recommended Citation

Arrington, Caleb A.; Kanipe, Tina; and Sangani, Deep, "AVD Tablet Labs: Virtual Chemistry Laboratories, Honors Chemistry, Grade:10" (2014). *Arthur Vining Davis High Impact Fellows Projects.* Paper 20.
http://digitalcommons.wofford.edu/avdproject/20

High Impact Fellows Project Overview

Project Title, Course Name, Grade Level

AVD Tablet Labs: Virtual Chemistry Laboratories, Honors Chemistry, Grade:<u>10</u>

Team Members

Student(s):   *Deep Sangani*

High School Teacher: *Tina Kanipe*   School: Broome High School

Wofford Faculty: *Caleb Arrington*   Department: Chemistry

Brief Description of Project (<150 words)

With the increased prevalence of student computers in secondary schools, we created six virtual chemistry laboratories using the virtual lab environment created by Carnegie-Mellon University. Students, using their laptop or tablet computer, can manipulate typical laboratory instrumentation found in an introductory chemistry lab. While not designed to replace traditional wet laboratories, the virtual laboratory environment allows students to examine the nuances of experimental techniques in a low stakes, easily repeatable, environment. The six virtual laboratories developed for this project cover the introductory chemistry subjects of: 1.) density measurements and significant figures, 2.) reaction stoichiometry, 3.) measurements of grams and moles, 4.) limiting reactants, 5.) acid, bases and solution pH, and 6.) molar mass determination of an unknown acid.

List of Materials Required and Costs, if Applicable

The virtual laboratory is free from Carnegie-Mellon and runs over the web on any personal computer or tablet capable of running Java.

AVD Tablet Labs

Virtual Chemistry Laboratory

Overview

The students of District 3 have new tablet computers. We are eager to engage these students with their tablets in the study of laboratory chemistry. Using virtual laboratory software developed by the ChemCollective group at Carnegie Mellon University, our AVD High Impact Fellows team will develop simulated laboratory activities for the honors chemistry students in Mrs. Kanipe's class at Broome High School. Tablet lab experiments will have many possible solutions, allowing students to examine the merits of different laboratory techniques. A component of competition will be introduced in some of the activities that will allow students to be recognized for accurate laboratory work.

Virtual Lab Activities

Virtual Lab # 1: Density and Significant Digits

A laboratory activity created to measure the density of an unknown solution using different laboratory glassware. Introduces the idea of significant digits. *(South Carolina State Standard: C-1)*

Virtual Lab # 2: Gravimetric Analysis

This activity develops an experiment that uses a precipitation reaction to determine the amount of arsenic in a solution sample. Experimentation requires mastery of molar mass. *(South Carolina State Standard: C-4)*

Virtual Lab# 3: Stoichiometry

Experimentation that allows students to explore reaction stoichiometry by observing the mole ratios in which reactants combine and products form. Introduction to limiting reactant.*(South Carolina State Standard: C-4)*

Virtual Lab # 4: Acid Base Titration

Experimentation created to determine the concentration of acid in solution by the technique of titration. Lab work provides experience with molarity and solution dilution. *(South Carolina State Standard: C-6)*

Budget Request

$80.00 is requested for the purchase of eight $10.00 i-Tunes gift cards to be awarded to students for the most accurate work in a tablet lab.

Goals

• • •

- Engage students with their new tablet computers in the chemistry classroom.

- Learn laboratory techniques in volumetric solution preparation, titration, and precipitation.

- Develop laboratory questions covering the topics of: significant digits, precipitation, density, stoichiometry, acid-base neutralization, solution concentration and titration.

- Measure virtual lab success through testing between students with virtual lab experience and a control group.

Standards

• • •

The educational standards covered in these virtual laboratory activities are:

- The student will demonstrate an understanding of how scientific inquiry and technological design, including mathematical analysis, can be used appropriately to pose questions, seek answers, and develop solutions.

- The student will demonstrate an understanding of the types, the causes, and the effects of chemical reactions.

- The student will demonstrate an understanding of the nature and properties of various types of chemical reactions.

Project Timeline

September:
- Identify curriculum topics to be covered in the tablet labs.

October:
- Begin writing the tablet labs.

November:
- Continue writing new tablet labs.
- Create assessment questions for pre-lab and post-lab evaluation.
- First classroom visit by AVD Student Fellow to perform the first tablet lab. The visit will assess the appropriate level of the first virtual lab in the high school classroom.

December:
- Complete the writing phase of the tablet labs.
- Continue to create assessment material for pre-lab and post-lab lab evaluation.
- A second classroom visit, if time permits, to perform the second tablet lab.

January:
- Student Fellow works with two new classes of students. These students will conduct all four labs created by our team.
- Administer pre-lab assessment to the first class and post-lab assessment to the second class to determine the effectiveness of the virtual lab activities.

February:
- Continue working with students on the tablet labs, again perform pre-lab and post-lab assessment.

March:
- Broome students continue to use tablet labs.
- Evaluation of assessments and virtual labs.
- Make corrections to tablet labs and prepare labs for dissemination through the AVD program.
- Prepare symposium talk on the project.

April:
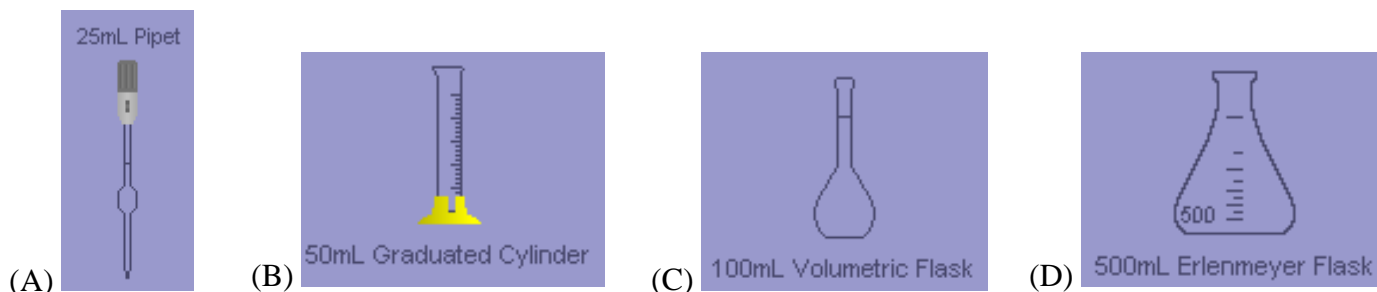- Complete tablet labs in the high school.
- Prepare the final AVD report.

Project Assessment Details

Mastery of the course content presented in each tablet lab will be assessed using pre-lab and post-lab assessment. Each lab activity will have a set of questions that will be administered to students using the classroom management software Edmodo. The assessment results will evaluate content mastery as well as examine pre-lab and post-lab performance. Mrs. Kanipe has two sections of Honors Chemistry that will perform the tablet labs. Multiple classes will allow both pre-lab and post-lab assessment to be administered.

Tablet Lab 01 Assessment

_____ 1. A procedure to make a tie dye solution calls for 20 mL of a potassium hydroxide solution. Which piece of glassware below should be used to measure the 20 mL of potassium hydroxide?

25mL Pipet

(A)   (B) 50mL Graduated Cylinder   (C) 100mL Volumetric Flask   (D) 500mL Erlenmeyer Flask

_____ 2. In filling a 100.00 mL volumetric flask, the flask to the right shows that the researcher has

(A) added exactly 100.0 mL of solution.
(B) added slightly too much solution.
(C) not added enough solution.

100mL Volumetric Flask
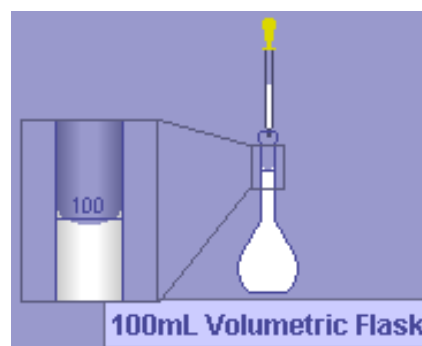
_____ 3. In a typical laboratory measurement of density, the volume and mass of the object must be determined. Which measurement has the greatest uncertainty when a graduated cylinder and an electronic balance are used?

(A) volume   (B) mass

_____ 4. To the nearest mL, what volume does the 50 mL graduated cylinder hold in the image below?

(A) 50 mL
(B) 44 mL
(C) 46 mL
(D) 34 mL
(E) 36 mL
(F) 45 mL

50mL Graduated Cylinder
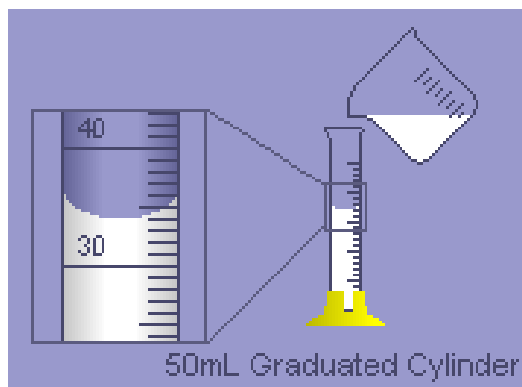
_____ 5. A solution volume of 16.2 mL is found to have a mass of 13.589 g. The density of this
solution is best reported as

(A) 1.19 g/mL   (B) 1.2 g/mL   (C) 0.839 g/mL   (D)0.83882 g/mL

# Tablet Lab: Virtual Lab Introduction and Solution Density Measurements

- To open the virtual lab on your tablet, go to "chemcollective.org" in a Webbrowser.
- From the Chemcollective homepage click on the link "Virtual Lab" on the right-hand side of the page.
- To launch the virtual lab, Java must open on your tablet. Windows will ask you if you want to run this application. Check "I accept the risk and want to run this app." Then click Run.
- The vitual lab will open. Scroll to put the virtual lab in the middle of your tablet.
- To begin exploring the virtual lab we will open the assignment entitled "Liquid Density Problem." To open this assignment, click on the "File" menu and then select "Load Homework". Click to open the folder "Molarity and Density". In that folder, select "Liquid Density Problem" and tap "OK"
- As the problem opens, the equipment available for the experiment is on the left-hand side of the virtual lab workbench. Tap and drag Compound A-1 into the center of the virtual lab workspace. The solution appears in an Erlenmeyer flask.

- All of the typical lab equipment can be found by clicking on the icons and . Glassware is found in the "beakers" icon. Click on the icon and select a 50 mL graduated cylinder. When you do so, a graduated cylinder will appear on your work space.

- An amount of compound A-1 can be poured into the graduated cylinder. Pouring is done by dragging the Compound A-1 to the top of the graduated cylinder and then releasing. When it is ready to pour, the glassware will look like the image below.
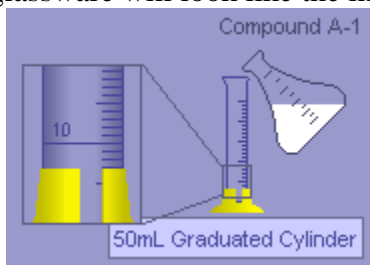
- Pouring is accomplished by clicking on the "pour" button at the bottom of the page. A sliding bar estimates how much is being added.
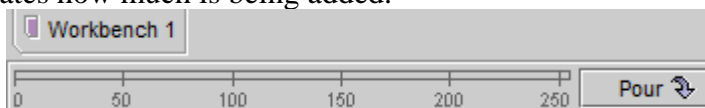
- Click and hold the pour button, but release it before the blue slide bar reaches approximatelyt 50 mL.
- Examine the amount of Compund A-1 in the 50 mL graduated cylinder. Remember, volumes are read from the bottom of the meniscus of the fluid. The next page shows a meniscus and volume estimate.
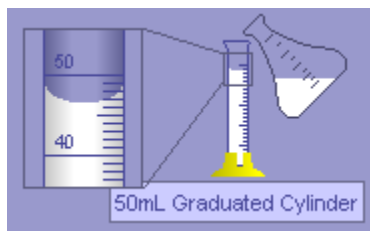
The volume in the above image should be reported as 46.8 mL.

- If your initial pour is between 30 and 50 mL you have good reflexes.
- If your volume is below 30 mL add a little more of Compound A-1 by tapping the pour button.
- Many of you might have added more than 50 mL. This is not a big deal.
- The virtual lab, fortunately, will not let you make a mess. If your volume is not over 50 mL, fill your graduated cylinder past 50 mL by holding down the pour button until the fill bar rises past 100 mL. The vitural lab cuts you off in the case of a massive over pour, and simply gives you a graduated cylinder with more than 50 mL. It would be nice if this were the case in the physical laboratory.
- When your 50 mL graduated cylinder is filled past 50 mL, move the Erlenmeyer flask of Compound A-1 to an empty spot on the table.
- Use the glassware icon to retrieve a 600 mL beaker. This will be a waste beaker during these experiments. It is important to label the glassware you are using so that the correct solution can be easily found. Label the 600 mL beaker by right clicking on the 600 mL beaker and selecting "Rename". In the dialog box, enter "Waste".
- Move the full 50 mL graduated cylinder over the waste beaker so a little of the A-1 solution can be poured out of the graduated cylinder. You may have to move the waste beaker to a lower portion of the workspace.
- Notice, when pouring from the graduated cylinder, that the pour bar has a much expanded scale. Try to pour 10 mL of Compound A-1 into your waste beaker from the 50 mL graduated cylinder.
- Move the graduated cylinder from the waste beaker. In the graduated cylinder Compound A-1 can be poured with some care. In order to read the volume in the graduated cylinder, right click on the graduated cylinder and click on "Detailed View".

Density Determination of Compound A-1:

- There are many ways to measure the density (d) of a liquid, but they all involve measuring the solution mass (m) and the solution volume (V) and solving for the m/V ratio.

$$d = \frac{m}{V}$$

- We have concentrated on volume measurements up to this point and will now turn our attention to mass measurements in the virtual lab. Click on the tools icon and select a scale.

- Use the glassware icon to bring a 25 mL graduated cylinder to the workbench.

- The electronic balance will record the mass of an object by placing the object over the balance. Obtain the mass of your 25 mL graduated cylinder. Record the mass and **units** on your data sheet (b).

25mL Graduated Cylinder

40.1212g TARE
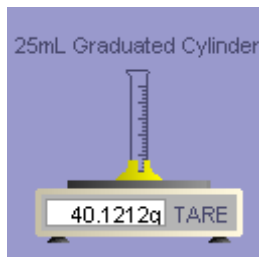
- Remove the 25 mL cylinder from the balance and set up the 50 mL graduated cylinder so you can pour Compound A-1 into the 25 mL graduated cylinder. It should look like the image below.

50mL Graduated Cylinder

25mL Graduated Cylinder
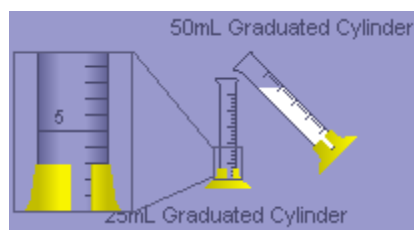
- Pour approximately 20 mL of Compound A-1 into the 25 mL graduated cylinder. After you have poured, record the actual volume and units transferred to the 25 mL graduated cylinder (c). If you filled past the 25 mL mark, pour a little into the waste beaker, until you can obtain a volume measurement of Compound A-1 in the 25 mL graduated cylinder.
- After recording the volume, move the 25 mL graduated cylinder to the balance and obtain the mass of the cylinder and the solution. Record this mass on your data sheet (d).
- From the mass measurements on your data sheet, determine the mass of solution in the 25 mL graduated cylinder. Record the solution mass on your data sheet (e).
- From the mass measurement and the volume measurement, determine the density of Compound A-1. Record all of the digits your calculator gives (f).
- Be sure the units of the measurement are included with your density.
- Your calculator, of course, gives too many digits. How many significant digits should be reported for the density measurement? (g)
- Which measurement, mass or volume is responsible for the number of significant digits that can be reported for density?(h)

Making a More Accurate Measurement of Density

- Several pieces of glassware can provide a more accurate measurement of volume. We will use a 100 mL volumetric flask to more accurately measure solution volume. Click the glassware icon and select "volumetrics" then select "100 mL volumetric flask". This flask has a mark on the neck that shows 100.0 mL. When a solution is added to this mark the volume is 100.0 mL. This provides four significant digits for the volume.
- Clean up your workspace by removing the 25 mL graduated cylinder. An object is

  removed by first clicking on the object and then clicking the "remove" icon, [icon]. Also remove the old Erlenmeyer of Compound A-1 and replace it with a new portion of Compound A-1. Empty the 50 mL graduated cylinder in the waste beaker.

- Weigh the empty 100 mL volumetric flask. Record the mass on your data page.(i)
- Remove the 100 mL volumetric flask from the electronic balance and place it on the workbench.
- Because of the added accuracy of the volumetric glassware, it is more difficult to fill to 100.0 mL mark. You will do an initial fill of the flask by adding two 50 mL portions from the 50 mL graduated cylinder. Fill the graduated cylinder to slightly below the 50 mL mark from the new Erlenmeyer flask of Compound A-1. If you overshoot the 50 mL mark, pour a little of the solution into the waste beaker. You should have about 48 mL of solution in the 50 mL graduated cylinder. Add the second 50 mL portion. The flask should have about 96-98 mL of solution.

50mL Graduated Cylinder

100

100mL Volumetric Flask
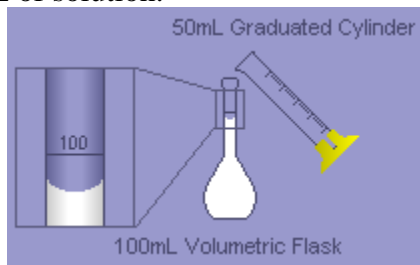
- If more than 100 mL have been added, the liquid line will be above the 100 mL mark, and a little will have to be poured into the waste beaker. There likely isn't enough liquid. In this case, no white liquid is observed in the flask neck. If this has happened, a little more liquid needs to be added.
    - The extra liquid should be added with a 10 mL graduated cylinder because it is easy to pour 1 or 2 mL. If you do not see liquid in the blowup of the flask neck, obtain a 10 mL graduated cylinder.
    - Fill the 10 mL graduated cylinder from the Erlenmeyer flask of Compound A-1.
    - Use the 10 mL graduated cylinder to fill the 100 mL volumetric flask until the volume level is similar to the image above.
- The filling of the volumetric flask to the 100 mL mark must be accomplished with a pipet. Use the glassware icon to select a "Disposable pipet" from the pipet list.

- Place the pipet over the Erlenmeyer flask of Compound A-1 to draw up liquid.

Disposable Pipet

Compound A-1

0.02   1   🔄 Withdraw   Pour ⤵
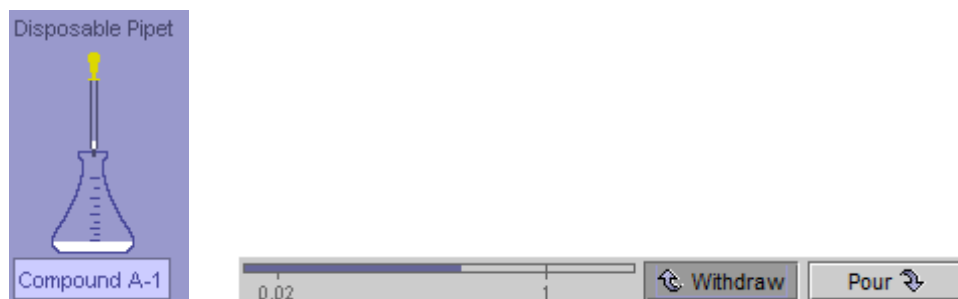
- A pipet must first draw up liquid. This is done by pressing "Withdraw" at the bottom of the page. Pull up at least one mL of solution into the pipet.
- The pipet can now be moved to the 100 mL volumetric flask.
- Add solution to the volumetric flask by clicking and holding the "Pour" button for the pipet. The pipet transfer process will have to be done three or four times to add the additional 3 to 4 mL. The last drops will have to be added by tapping the pour button. When you have filled the volumetric flask to the 100.0 mark the flask should look like the image below.

Disposable Pipet

100

100mL Volumetric Flask
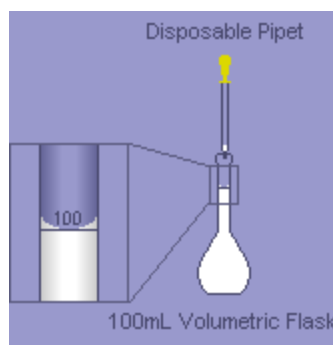
- If the bottom of the liquid meniscus is above the 100.0 mL mark, a little liquid can be removed by clicking the "Withdraw" button for the pipet.
- With the solution meniscus right on the 100.0 mL mark you have the volume measured to four significant digits. (That was a significant amount of work. Ha Ha!)
- Now measure the mass of the filled volumetric flask by placing it on the electronic balance. Record the mass of the filled flask on your data page (j).
- Subtract the mass of the filled flask from the mass of the empty flask to obtain the mass of the solution. Record the solution mass on your data page (k).
- Use the solution mass and the volume (100.0 mL) to determine the density of the solution. Record the solution density on your data page. Include all the digits on your calculator (l).
- How many digits should be used to report the density using this technique?

Comparing Density Measurements and Significant Digits

If there is time, the class will pool its density measurements to compare the precision of the two techniques for measuring solution density.

- Record your density on the white board. Record the density obtained using the 25 mL graduated cylinder in the table labeled "25 mL graduated cylinder". Record the second density measurement on the table labeled "100 mL volumetric flask". For both densities report at least six digits in the density.
- Examine the class data set on the board. Hopefully every measurement begins with a 1.2XX g/mL. If there are any densities with different digits in the ones and tenths place, these data should be removed from consideration. Some values may have been misread or misrecorded.
- For the "25 mL graduated cylinder" data, which digit was recorded most commonly in the hundredths place of the density? (n)
- How many of the density measurements have a different digit in the hundredths place than the most common digit? (o)
- Can a most common digit be found in the thousandths place of the density for the densities obtained by the method using the "25 mL graduated cylinder" or do the digits reported in the thousandths place appear completely random? (p)

Significant figure rules are designed to identify the smallest digit before the digits become completely random when a series of measurements are made. Application of the significant digit rules means that you know the last digit to report without having to make duplicate measurements showing where randomness begins. We do not report digits that would be random when a series of measurements are made.

- Focusing on the density data on the board obtained from the method using the 100 mL volumetric flask, which digit's place (tenths, hundredths, thousands, ten thousandths) do the reported digits appear to be random? (q)
- The digit's place one before the digit's place that is random is the last significant digit. What is the place of the last significant digit in the density obtained from the 100 mL volumetric flask method? (r)
- How does the number of significant digits in the density compare to the number of significant digits predicted from the significant digit rules? (s)

Virtual Lab Homework

In class you focused on Compound A-1. For homework determine the density of Compound A-2. You may use any method that you like. There are several other pieces of glassware in the virtual lab that can be used to obtain an accurate density. Explore the use of the 25.0 mL and the 10.0 mL pipet. These volume measuring instruments are very accurate and are easier to fill than the volumetric flask. In your lab notebook, record the steps that you perform in determining the density. For example, if you weigh a graduated cylinder, record:

- The mass of the 50 mL graduated cylinder was measured to be 73.0504 g.

Show the calculations for the density of Compound A-2.

Tablet Lab Data Sheet
Virtual Lab Introduction and Solution Density Measurement

(a) Solution volume _____ (include units)

(b) Mass of 25 mL graduated cylinder _____

(c) Volume of Compound A-1 _____

(d) Mass of cylinder + solution _____

(e) Solution mass _____

(f) Solution density _____

(g) How many digits should be reported? _____

(h) Which measurement is least well known? _____

(i) Mass of empty 100.0 mL volumetric flask _____

(j) Mass of filled 100.0 mL volumetric flask _____

(k) Mass of 100.0 mL of Compound A-1 _____

(l) Density of Compound A-1 _____

(m) How many digits should be reported? _____

(n) Most commonly recoded hundredths place digit _____

(o) Number of measurements not having the most common digit in the hundredths place _____.

(p) Is there a most common digit in the thousandths place? _____

(q) Which digit's place appears to be random in the second density? _____

(r) Which digit's place is the last significant digit in the second density? _____

(s) How do the number of significant digits determined from multiple samples compare to the number of significant digits determined from significant digit rules?

_____

Tablet Lab Assessment: Stoichiometry

_____ 1. An image of 100 mL of an aqueous solution of compound A is show below. Each letter A in the image represents 0.03 moles of compound A. What is the solution concentration of A in the beaker?
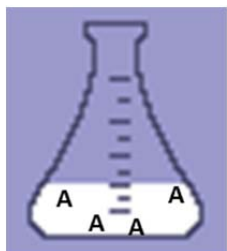

(A) [A] = 1.2 M   (B) [A] = 0.012 M   (C) [A] = 0.03 M   (D) [A] = 0.3 M

_____ 2. If 200 mL of water is added to 100 mL of 0.6 M NaOH solution, what will the concentration of NaOH be in 300 mL solution?

(A) [NaOH] = 0.3 M   (B) [NaOH] = 0.18 M   (C) [NaOH] = 0.2 M
(D) [NaOH] = 0.6 M

_____ 3. For the reaction A + 2B ⟶ 2C + D, what would be the best representation of the reaction mixture after 100 mL of Solution A was mixed with 100 mL of Solution B. Solutions A and B have the same initial concentration.

(A)   (B)   (C)   (D)
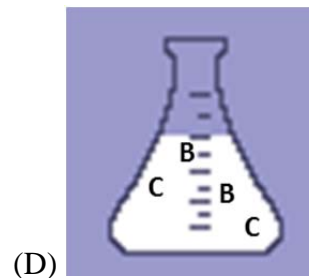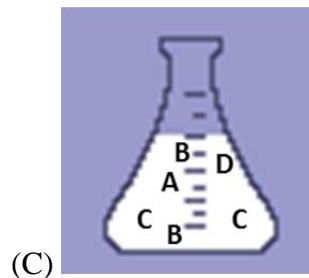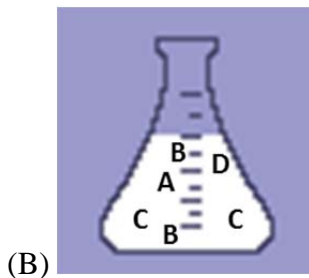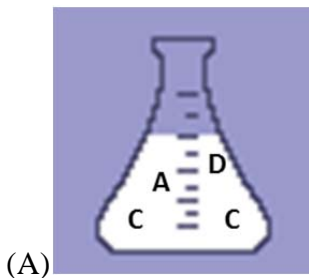
_____ 4. The reaction table below describes the reaction between compound B and compound D. The table shows the initial and final amounts of reactants and products. What is the balanced chemical equation for the reaction of B with D?

Compounds:	B(aq)	+ D(aq)		A(aq)	+ C(aq)
Initial (moles)	0.200	0.200		0.000	0.000
Final (moles)	0.000	0.100		0.300	0.100

(A) B + D $\longrightarrow$ A + C
(B) A + B $\longrightarrow$ C + D
(C) 2B + D $\longrightarrow$ 3A + C
(D) B + 2D $\longrightarrow$ A + 3 C

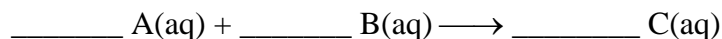Tablet Lab Work Sheet: Stoichiometry and Balanced Chemical Equations

1. Calculate the change in the number of moles for each substance in the reaction below. Enter the change in the appropriate column of the reaction table below.

	$N_2(g)$	+ $3H_2(g)$		$2NH_3(g)$
Initial (moles)	1.80	1.80		0
Final (moles)	1.2	0.00		1.2
Change in moles				
Whole number change in moles				

2. Complete the row for the whole number changes in moles above.

3. If each symbol (letter) in the beaker represents 0.2 moles, fill in the table below for final and initial moles. (Fill in just the first two rows at this point

	$A(aq)$	+ $B(aq)$		$C(aq)$
Initial (moles)				
Final (moles)				
Change in moles				
Whole number change in moles				

4. Calculate the change in moles by subtracting, $final_{moles} - initial_{moles}$. Complete the row for the change in moles above.

5. Complete the row of "Whole number change in moles".

6. The values for the change in moles in the final row are the stoichiometric coefficients in the reaction. Complete the balanced chemical equation below.

_____ A(aq) + _____ B(aq) ⟶ _____ C(aq)

7. For a 400 mL solution of 0.80 M B, determine the number of moles of B.

_____ mol B

8. What is the new concentration of Compound A in the beaker labeled Solution A? _____

9. Have the number of moles of A changed in the beaker? _____

10. What is the molarity of a 500 mL solution that contains 0.3 moles of Compound A?

_____ M A

11. How many moles of B are in Solution B? _____

12. Did the concentration of Compounds A and B change when they were mixed? _____

13. Which solution is the best picture to describe the mixture of Solutions A and B? _____

14. Did the number of moles of Compound A and B change in the mixture "A+B"? _____

15. How many moles of A and how many moles of C are present in their respective flasks?

16. What is the best picture of the resulting solution when Solution A and C are mixed? _____

17. Calculate the <u>final</u> number of moles of each compound present in the mixture of A + C.
    (This is the most important calculation of the lab.)

_____ moles A   _____ moles B

_____ moles C   _____ moles D

18. From the reaction, which of the compounds can you conclude are reactants?

_____

19. From the reaction, which of the compounds are products?

_____

20. Complete the reaction table below.

	A(aq)	+	C(aq) ⟶	B(aq)	+	D(aq)
Initial (moles)						
Final (moles)						
Change in moles						
Whole number change in moles						

21. Finally, use the final row of the table to give the complete balanced chemical equation for the reaction.

_____

22. Which of the reactants is completely consumed? _____

23. What volumes of Solution A and Solution C would you need to mix together to have no reactant remaining in the final mixture? This would be the initial conditions such that **both** A and C are the limiting reactant.

_____ mL Solution A   _____ mL Solution B

24. Complete a reaction table that depicts the reaction you have run where A and C are both limiting

	A(aq)	+	2C(aq) ⟶	B(aq)	+	D(aq)
Initial (moles)				0.00		0.00
Final (moles)	0.00		0.00			
Change in moles						

25. How many moles of Compound B and D are present in the 600 mL flask?

_____moles B   _____ moles D

26. Record the two non-reacting tried in the virtual lab. _____

On your own lab work.

27. After you have found the reactants that mix to form products, complete the reaction table below. (It is possible that only one product is formed. In that case leave the last column blank.)

Reaction:			$\longrightarrow$		
Initial (moles)	0.10	0.10		0.00	0.00
Final (moles)					
Change in moles					
Whole number change in moles					

28. Use the table to determine the balanced chemical reaction. Your reaction is different from everyone else's, so avoid the urge to compare your answer with another student in the class.

_____

Tablet Lab: Stoichiometry and Balanced Chemical Equations

Laboratory Goals: Determine the balanced chemical equation for an unknown reaction using the stoichiometry of reactants and products.

Stoichiometry: There could not be a more nerdy word in all spoken language, however this nerdy word is incredibly useful in chemistry. Stoichiometry is the use of mole ratios to relate the amounts of reactants and products that change in a chemical reaction. We refer to the numbers in front of the compounds in a chemical reaction, such as $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, as the stoichiometric coefficients. These coefficients show the **change** in the number of **moles** of reactants and products during the chemical reaction. In math class we think of change as final – initial. It is the same in chemistry. A stoichiometric coefficient is $\Delta moles = moles_{final} - moles_{initial}$.

In the example of the reaction above we can keep track of the number of moles in a table. We will begin by reacting 1.8 moles of N_2 with 1.8 moles of H_2.

|	$N_2(g)$	+ $3H_2(g)$		$2NH_3(g)$
Initial (moles)	1.8	1.8		0
Final (moles)	1.2	0.0		1.2
Change in moles				
Whole number change in moles				

1. Calculate the change in the number of moles for each substance in the above reaction.

Your answer will have decimal numbers. Typically chemists like to express reaction coefficients with whole numbers. Convert the "Change in moles" to "Whole number change in moles " by dividing each value in the "Change in moles" by the smallest change, in this case (0.6).

2. Complete the row for the "Whole number change in moles".

Hopefully your calculations for the "Whole number change in moles" will match the stoichiometric coefficients of the reaction.

Now let's repeat this analysis using a reaction from the virtual lab. Below is the picture of a reaction. The reactants are A and B and the product is C.
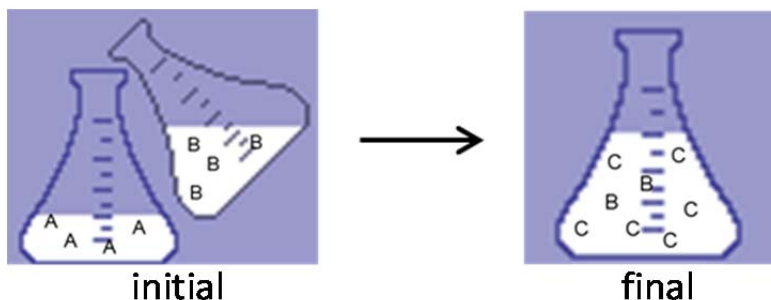
initial   final

3. If each symbol in the beaker represents 0.2 moles, fill in the table below for final and initial moles. (Fill in just the first two rows at this point.)

	A(aq)	+ B(aq)	⟶	C(aq)
Initial (moles)				
Final (moles)				
Change in moles				
Whole number change in moles				

The row for the "Change in moles" is the easiest row to complete.

4. Calculate the change in moles by subtracting, $\text{final}_{\text{moles}} - \text{initial}_{\text{moles}}$. Complete the row for the change in moles above.

You will notice that reactants will have a negative value for change in moles. This will always be the case because reactants are being consumed in a chemical reaction. We do not include the negative sign in the stoichiometric coefficients.

The final row, "Whole number change in moles" can be found by dividing each value for the "Change in moles" by the smallest value in the "Change in moles" row. This means that one of the values will be one. (Often after this division there might be a value such as 1.5. In that case each value in the "Whole number change in moles" row should be multiplied by 2.)

5. Complete the row of "Whole number change in moles".

6. The values for the change in moles in the final row are the stoichiometric coefficients in the reaction. Complete the balanced chemical equation below.

_____ A(aq) + _____ B(aq) ⟶ _____ C(aq)

We will next use this technique of calculating the change in moles to determine the stoichiometric coefficients in an unknown chemical reaction.

- Open your Tablet Lab workspace by going to chemcollective.org from your web browser
- Open the virtual lab by clicking on "Virtual Lab" under the Quick Links.
- A Java window will appear. Click "I accept the risk" and then click "Run".
- Under File in the virtual lab, select "Load Homework".
- Scroll down and choose the tab titled "Stoichiometry and Limiting Reagent".
- Select the assignment "Oracle Problem".

The "Oracle Problem" provides four different solutions, labeled Compound A, Compound B, Compound C, and Compound D. The design of the problem is that some combination of A, B, C, and D produce a chemical reaction. The reaction **might** be:

$$2B + D \longrightarrow A + 3C$$
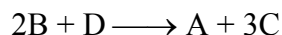
By, experimenting, mixing different reagents and observing the products, the balanced chemical reaction can be determined.

- Begin exploring the Oracle Problem by pulling Solution A onto the workbench.

When a solution is highlighted, the Virtual Lab supplies information about the solution in the "Solution Info…" window on the right. The "Solution Info…" window for Solution A is shown here. At the top, the solution volume is given. The concentration of Compound A is given at the bottom of the window, in units of molarity. The molarity of A is the number of moles of A divided by the liters of solution.

$$molarity\ A = \frac{moles\ A}{L\ sol}$$

In this case the molarity of A is 1.00 M. The molarity of the H^+ and OH^- ions are also given. Their concentrations are very small in water, $1x10^{-7}$ M. We will wait until we discuss acids and bases to examine these concentrations in more detail. The H^+ and OH^- concentration will remain unchanged in today's lab. The solution pH is also given. None of the compounds today are acids or bases so the solution remains pH = 7.
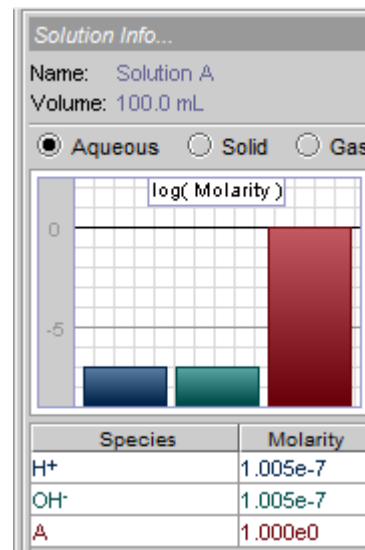
The molarity is useful, but if we are going to do stoichiometry it is really the number of moles of A that we need. The moles of a compound are found by multiplying molarity by volume, as is done below for Solution A.

$$1.00 \left(\frac{mol\ A}{L\ sol}\right)(0.100\ L\ sol) = 0.100\ mol\ A$$

7. For a 400 mL solution of 0.80 M B, determine the number of moles of B in the space below.

- Pull onto the workbench the distilled water.

- Add to the workbench a 250 mL Erlenmeyer Flask, using the glassware icon.
- Place the distilled water carboy over the 250 mL Flask, so that water can be added.
- Add 100 mL of water to the 250 mL flask by typing "100" in the "Transfer amount (mL):" bar at the bottom of the page. Then click the "Pour" button.

This method of pouring is of course not realistic. You could never get 100.00 mL from a large carboy.

- Right click on the 250 mL Erlenmeyer Flask and select "Rename". Change the name of the flask to "Water". There will be many flasks present and having them properly labeled will be necessary.

We are going to examine what happens when water is added to Solution A.

- Place the Water flask over the Solution A flask.

- Again type 100 as the transfer amount and click pour. Remove the water flask and then click on Solution A so that its properties are displayed to the right.

8. What is the new concentration of Compound A in the beaker labeled Solution A?
   _____

9. Have the number of moles of A changed in the beaker? _____

The number of moles of A naturally do not change just because water is added. The moles become more spread out and the concentration of Solution A decreases.
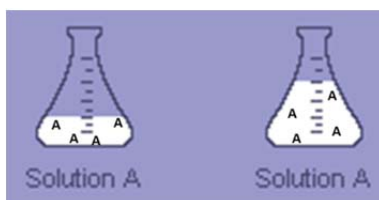
The concentration changes because of the change in solution volume, **not** because the number of moles changes. In this case

$$\frac{0.100 \; mole \; A}{0.200 \; L \; sol} = 0.500 \; M$$

10. What is the molarity of a 500 mL solution that contains 0.3 moles of Compound A?

- Place your solutions over to the side and pull from the stockroom on the left a new flask of Solution A and a flask of Solution B. We are interested in observing if these compound react.

11. How many moles of B are in Solution B? _____

- Typing in 100 mL each time you want to pour is tiresome. Change to the pour mode by selecting "Transfer Bar" from the "Tools" menu. Change the Transfer Bar setting from "Precise Transfer" to "Realistic Transfer"
- Place Solution B over Solution A, so that the two solutions can be mixed.
- Pour Solution B into Solution A, by holding down the Pour button until the scrolling volume moves past 100 mL.
- Pull Solution A away from Solution B.
- Relabel Solution A to "A + B".
- Examine the solution A + B in the Solution Info window.

12. Did the concentration of Compounds A and B change when they were mixed?
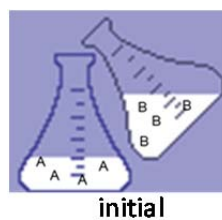
initial

13. For the reaction represented by the figure above _____ , which is the best picture to describe the final solution?

(A)

(B)
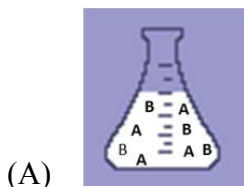
14. Did the number of moles of Compound A and B change in the mixture "A+B"?

Because the number of moles of A and B are not changed, we see that A and B do not react.

- Remove the flask labeled Solution B. This is done by tapping on the flask and clicking the remove icon.
- Let's try reaction A + C. Do this in the same way that you reacted A + B. Bring onto the workbench a new flask of A and C.

15. How many moles of A and how many moles of C are present in their respective flasks?
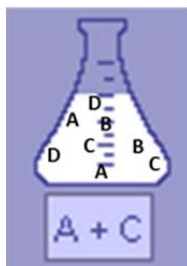
- Pour flask A into flask C and relabel the mixture "A + C".
- Examine the concentrations present in solution "A + C" using the Solution Info window.
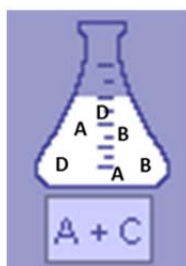

Solution C

Solution A
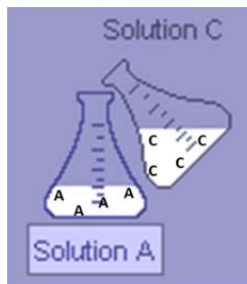
16. Given the picture of pouring C into A [image], what is the best picture of the mixture "A + C"? Talk to your neighbor and make sure you get this answer correct.

(A)   (B)   (C)

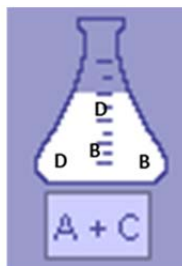17. Review Question 1 and calculate the <u>final</u> number of moles of each compound present in the mixture of A + C. (This is the most important calculation of the lab.)

_____ moles A   _____ moles B

_____ moles C   _____ moles D

18. From the reaction, which of the compounds can you conclude are reactants?

19. From the reaction, which of the compounds are products?

20. Complete the reaction table below.

	A(aq)	+ C(aq)	→	B(aq)	+ D(aq)
Initial (moles)					
Final (moles)					
Change in moles					
Whole number change in moles					

21. Finally, use the final row of the table to give the complete balanced chemical equation for the reaction.

22. Which of the reactants is completely consumed?

The reactant that runs out first in a chemical reaction is termed the "limiting reactant".

23. What volumes of Solution A and Solution C would you need to mix together to have no reactant remaining in the final mixture? This would be the initial conditions such that **both** A and C are the limiting reactant.

- Pull onto the workbench the amount of Solutions A and C that you will need to mix so that both reactants are limiting reactants.
- Bring onto the workbench a 600 mL beaker.
- Pour the flasks of Solution A and Solution C that you selected into the 600 mL flask.

24. Complete a reaction table that depicts the reaction you have run where A and C are both limiting

	A(aq)	+ 2C(aq)	→	B(aq)	+ D(aq)
Initial (moles)				0.00	0.00
Final (moles)	0.00	0.00			
Change in moles					

25. How many moles of Compound B and D are present in the 600 mL flask? Remember the Solution Info window reports the molarity of Compound B and D.

- We tried reacting Solution A and Solution C after only one failed reaction attempt (the reaction of A + B). Pull onto the workbench any two solutions that you expect to not produce a reaction.
- Pour these two solutions together and verify that no reaction occurs.

26. Record the two non-reacting tried in the virtual lab.

Believe it or not this is just the pre-lab assignment! It is now time for you to determine the balanced chemical reaction for your own reaction.

- To access this assignment tap "home" above the virtual lab on the line that says, "You are here." Alternatively, you can type, "chemcollective.org" into your browser.
- On the lower left hand side of the screen find the link named "Autograded Problems" under the "Resources By Type". Tap the link to access these problems.
- From the list of problem types select "Stoichiometry".
- From the Stoichiometry problems scroll down and select, "Determining Stoichiometric Coefficients".
- Allow Java to run on your tablet and let the virtual workbench open.

You are now free to mix solutions. Mix two solutions at a time and keep track of what you are mixing, by relabeling mixture flasks with names such as, "A + B". Look to identify the compounds that are reactants and products. The virtual lab takes numbers to extremes. In this lab you will see concentration such as 2.44e-24 M. This concentration is so small that we consider the concentration to be 0.000 M

(Unfortunately, this lab "glitches" occasionally and you may find that no solution combinations react. If this happens to you, tap the reload button on the browser and you will get a new set of solutions that should work. No one is unlucky enough to have two "glitched" labs in a row. If you experience two "glitched" labs, I would not buy a lottery ticket today. You have some really bad luck.)

27. After you have found the reactants that mix to form products, complete the reaction table below. (It is possible that only one product is formed. In that case leave the last column blank.)

Reaction: $\longrightarrow$

Initial (moles)	0.10	0.10		0.00	0.00
Final (moles)					
Change in moles					
Whole number change in moles					

28. Use the table to determine the balanced chemical reaction. Your reaction is different from everyone else's, so avoid the urge to compare your answer with another student in the class.

The virtual lab will check your work. Enter the balanced reaction at the bottom of the lab. An **example** of how the reaction is submitted is shown below. Once you have entered the equation in the light blue box, tap Check.

How do these substances react? [ ] A + [ ] B + 2   C + 1   D --> 1   A + 2   B + [ ]
C + [ ]   D

Check

You are about to submit: 2 C + 1 D --> 1 A + 2 B

If your reaction is incorrect, check the calculation for the number of moles. You should continue to experiment and resubmit a new reaction.

When you have successfully determined the chemical reaction for your lab, show your teacher the screen indicating that you have successfully completed the lab.

Hopefully, you are now a little more comfortable with **stoichiometry**: the <u>change</u> in the number of moles of a substance in a chemical reaction.

Tablet Lab 03 Grams and Moles

_____ 1. How many grams of potassium bromide, KBr, are present in on mole of KBr(s)?

(A) 53 g   (B) 106 g   (C) 119 g   (D) 110 g

_____ 2. How many grams of calcium nitrate are required to prepare a solution containing 0.65 moles of calcium nitrate?

(A) 164 g   (B) 107 g   (C) 64 g   (D) 252 g

_____ 3. A 0.20 mole sample of sodium phosphate, Na_3PO_4, will contain how many moles of sodium ions when dissolved in water?

(A) 0.20 moles   (B) 11 moles   (C) 102 moles   (D) 0.60 moles

_____ 4. How many moles of carbonate are present in a 26.5 g sample of sodium carbonate, Na_2CO_3?

(A) 0.25 moles   (B) 0.50 moles   (C) 2810 moles   (D) 4.0 moles

Tablet Lab: Grams and Moles of Soluble Salts

Objective: Weigh out masses of soluble salts to give solutions with known numbers of moles of ions.

In chemistry we spend significant time trying to determine how many molecules or ions are present in a sample. Chemists have developed the gram molar mass as a method of weighing out a certain number of molecules. This tablet lab will help you practice converting mass into moles for several different ionic compounds.

- Open the virtual lab by going to the address: www.chemcollective.org\chem\wofford
- Enter your first and last name in the spaces on the page and then click on the link to Limiting Reagent Activities.
- Click on the NaCl bottle and drag it on to the Workbench. Bring the large carboy of Distilled H_2O on the Workbench as well.

- In addition use the equipment icon, , to bring onto the Workbench a 1000 mL (1 1iter) volumetric flask.

1000mL Volumetric Flask

1. Determine the number of grams in one mole of NaCl. Record this value in your lab notebook.

- Place the bottle of NaCl over the 1000 mL volumetric flask.
- Enter the mass of one mole of NaCl in the Transfer amount box at the bottom of the Workbench.
- Click Withdraw. This button scoops out the amount you entered on to the spatula.
- Now click Pour to place this amount into the flask.
- Pull the NaCl bottle away from the flask. Click back on the flask to observe the contents.

2. In the Solution Info window, on the right, the volume of NaCl in the container is given. Record the volume in your lab notebook as Question # 2.

3. Determine how much water must be added to completely fill the flask. This is a 1000 mL flask, so perform the subtraction. 1000 mL – 27.00 mL = 973.00 mL. Show this calculation in your lab notebook.

- Pull the Distilled H_2O carboy over the 1000 mL flask.
- Type 973.00 into the mL transfer bar then click pour.
- Check the volume of the container. If the volume of the container is not very close to 1000 mL, ask your neighbor or teacher for help.

In a real lab volumes are not additive in this way. The volume of the solid is different than the volume of the dissolved ions. This is an aspect of laboratory work the virtual lab does not get correct.

- Pull the water away from the flask and then click back on the NaCl solution you have prepared in the volumetric flask.
- Under the edit menu click Rename. Give the solution the name "1.00 M NaCl(aq)". Then click okay.
- In the Solution Info window click on the aqueous button.

This window shows how many moles of ions are floating around in the solution. The units of ion concentration are molarity. Molarity is the moles present in 1 liter of solution. Since we prepared one liter of solution the molarity is the same as the number of moles.

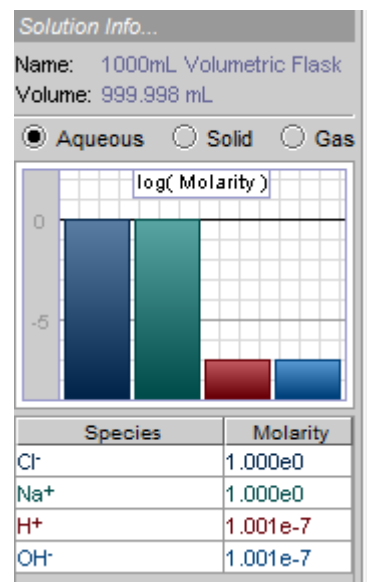4. Does the Solution Info window indicate that you have 1.0 moles of Cl^- and 1.0 moles of Na^+ in the solution?

The Solution Info window also shows very small concentration 1×10^{-7} M of H^+ ion and OH^- ion. These small concentrations are always present in water and will become important as you study acid base chemistry, but are not the focus of this lab.

Solution Info...

Name: 1000mL Volumetric Flask
Volume: 999.998 mL

◉ Aqueous   ○ Solid   ○ Gas

log(Molarity)

Species	Molarity
Cl-	1.000e0
Na+	1.000e0
H+	1.001e-7
OH-	1.001e-7

Solution Info Window

If you wanted to prepare a solution with 1.35 moles of NaCl, you would have to perform a slightly more complex calculation. To calculate how many grams of NaCl are needed for 1.35 moles the following can be done.

$$? \underline{\hspace{1cm}} \text{ g NaCl} = 1.35 \text{ moles NaCl} \times \frac{58.45 \ g \ NaCl}{1 mole \ NaCl} = 78.91 \ g \ NaCl$$

5. Using this example, perform a calculation to determine how many grams of $AgNO_3$, silver nitrate, are needed to prepare a 0.750 mole sample of $AgNO_3$. Perform this calculation in your lab notebook.

- Set the NaCl(aq) solution over to the side. Remove NaCl solid from the Workbench, by highlighting the bottle and then clicking the removal icon, ⌁.
- Retrieve a new 1000 mL volumetric flask and the $AgNO_3$(s) reagent bottle and bring them on to the workbench.
- Prepare a 1000 mL solution containing 0.750 moles of $AgNO_3$(aq) , employing the techniques used in preparing the previous solution.

At this point work on your own and remember that you need to subtract the volume of the solid from the 1000 mL container volume to determine the volume of water required.

6. Be sure to record the mass calculations necessary to prepare this solution in your lab notebook. Label this calculations as # 6.

- After preparing the solution check the Solution Info window and see that you have 0.75 moles of Ag^+ and 0.75 moles of NO_3^-.
- Use the Rename feature to name this solution "0.75 M AgNO3(aq)" and move it to the side.

You will prepare a third salt solution, this time using the salt sodium carbonate.

- Remove the solid $AgNO_3$ from the Workbench. Set the $AgNO_3$ solution to the side.
- Bring onto the Workbench sodium carbonate and a new 1000 mL volumetric flask.

7. In your note book calculate the grams of Na_2CO_3 required to obtain 0.25 moles of Na_2CO_3.

- Using your calculations prepare a 1000 mL solution containing 0.25 moles of sodium carbonate. Be sure to properly label the solution flask. (Examine the other solutions for help with the labeling.)

8. After preparing the sodium carbonate solution use the Solution Info widow to record the moles of CO_3^{2-} and the moles of Na^+ in the solution in your lab notebook.

9. Are the moles of these species the same as the moles of sodium carbonate you measured?

10. Explain any differences in the number of moles of ions in solution compared to the number of moles of salt.

Question 10 is very important, be sure you understand why the moles of Na^+ are not the same as the moles of CO_3^{2-}.

The last solution we will prepare together is a 0.0500 molar solution of lead nitrate, $Pb(NO_3)_2$. We will make this solution preparation a little more realistic by changing the transfer attributes of the virtual lab.

- Click the Tools menu. Under this menu select Transfer Bar then select Significant Figures Transfer.

Significant Figures Transfer will allow you to type the number of grams you wish to add, for example 50, but when you click pour you won't get exactly 50. You will have to use a few more lab tools to get the correct mass and the correct volume for the solution.

11. Make the mass calculation for the number of grams of $Pb(NO_3)_2$ required to obtain 0.050 moles of $Pb(NO_3)_2$.

- Bring onto the Workbench the bottle of $Pb(NO_3)_2$, a new carboy of water and a new 1000 mL volumetric flask. Remove the nearly empty water carboy.
- To weigh the necessary amount of $Pb(NO_3)_2$ will take more care in the Significant Figures Transfer mode. Bring onto the Workbench a weighing boat, found with the tools icon, .
- Pull the bottle of $Pb(NO_3)_2$ over the weighing boat.
- Type 10 g into the Transfer window.
- Tap Withdraw and the tap Pour.
- In the Solution Info window, select Solid. This shows the grams actually transferred. You should have less $Pb(NO_3)_2$ than you need.
- Withdraw another 10 grams onto the spatula. Adding all of the mass would be too much $Pb(NO_3)_2$.
- Change the amount in the Transfer window to 1 g.
- Tab the Pour button once to tap a little onto the weighing boat.
- Continue to tap the Pour button until you get close to your measurement.
- Assuming that you go over your measurement, a little solid will have to be removed from the weigh boat.
- Bring onto the Workbench a 250 mL beaker as a waste beaker for $Pb(NO_3)_2$. Label this beaker, "Solid Waste".
- Place the weigh boat over the waste beaker.
- Type 1 in the transfer bar window.
- Tap the Pour button. To remove a little of the solid from the weigh boat.

You may go back and forth a few times between bottle and waste beaker. In the end you should have the mass of $Pb(NO_3)_2$ required to at least ½ a gram of your calculated mass. Of course the closer your mass is to your calculated mass, the better your result.

- Once you have the necessary mass, place the weigh boat over the empty 1000 mL volumetric flask.
- Type a number in the transfer bar that is much larger than the actual number of grams of $Pb(NO_3)_2$ (ex. 50 g). Then tap the Pour button. This will transfer all of the solid.

Adding the volume of water required will also require care.

- Place the carboy over the 1000 mL volumetric flask. Type 900 mL into the Transfer window and click Pour.
- Bring onto the Workbench another empty 250 mL beaker.
- Fill this beaker with enough water to fill the volumetric flask, say 200 mL.
- Bring the beaker of water over the 1000 mL volumetric flask. Add 50 mL of water to the flask. (Be sure that this is well below the fill line.)

- Change the Transfer window to 10 mL.
- Add water to the flask in 10 mL increments until you have slightly more than 980 mL in the 1000 mL flask.
- Bring onto the Workbench a 10 mL graduated cylinder.
- From the beaker of water fill this graduated cylinder to 10 mL.
- Place the 10 mL graduated cylinder over the 1000 mL flask and enter 1.0 in the Transfer bar.
- Tap the Pour button to add 1 mL of water. Continue to add 1mL increments until the total volume in the 1000 mL flask is just above 999 mL. (If the graduate cylinder empties before the volume is reached, refill the graduated cylinder to 10 mL.)
- Once you are close to 999 mL, change the Transfer bar to a volume of 0.1 mL.
- Tap the Pour button until the volume is 1000.0 ± 0.05 mL.
- Rename this flask, "0.05 M Pb(NO3)2(aq)."

12. Check the Solution Info window and record the moles of $NO_3{}^-$ and Pb^{2+} in the solution. Record the exact concentration in your lab notebook.

13. Do you expect the moles of Pb^{2+} and $NO_3{}^-$ to be the same? What is the molar relationship between Pb^{2+} and $NO_3{}^-$ in this salt?

14. We should examine the results of the class to see whose solution is the closest to 0.500 moles Pb^{2+}. Record your Pb^{2+} solution concentration on the whiteboard if directed to do so.

You now have four salt solutions. Let's examine if any reaction takes place when these solutions are combined. Don't get too excited, no explosion will happen when these salts react. We will simply see if some of these ions have a strong enough ionic attraction to stick together and form a solid.

- Remove all items from the Workbench except for the four salt solutions you have prepared.
- Bring onto the Workbench a 250 mL Erlenmeyer flask to be used as the reaction flask.
- Add 100 mL of the NaCl solution to the Erlenmeyer flask.
- Add 100 mL of the $AgNO_3$ solution to the Erlenmeyer flask.
- Remove the $AgNO_3$ solution from the reaction flask and then click on the Erlenmeyer reaction flask so that it is highlighted.
- In the Solution Info window, click on the Solid button. Does any solid appear in the Solution Info window?

Solid that is formed when solutions are mixed is called a precipitate.

15. In your lab notebook, identify the compound that forms a precipitate.

The chemical reaction that occurred can be expressed as a chemical equation. In this case the reaction is:

$$NaCl(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$$
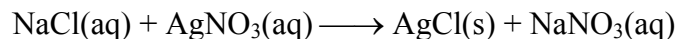
The phase designations aqueous (aq) and solid (s) are important in this reaction. Aqueous means dissolved in water, while solid indicates that the compound is not part of the solution.
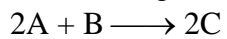
16. The final assignment in this lab is for you to select two more reactions to run by combining any two salt solutions as you did for $NaCl(aq)$ and $AgNO_3(aq)$. Run the reactions in a new 250 mL Erlenmeyer flask. Use the Solution Info window to determine if any solid has formed. If no solid forms, then no reaction has occurred. For each reaction you run, write the chemical reaction in molecular form as is found above.

If no reaction is observed, the following should be written in your lab notebook.
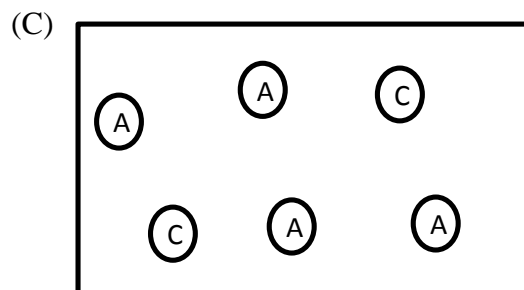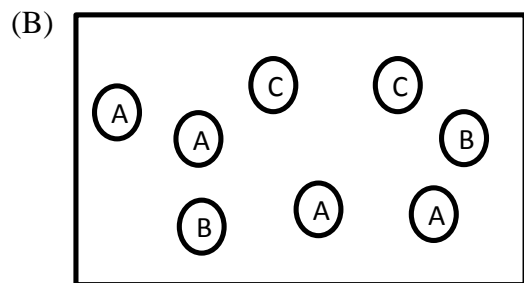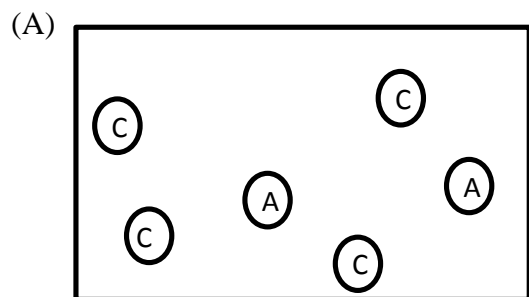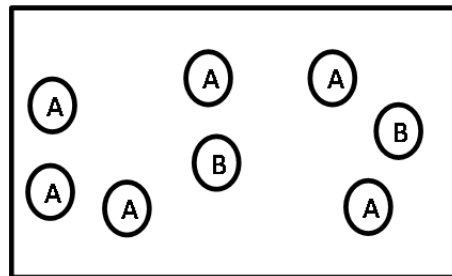
$$NaCl(aq) + NaI(aq) \longrightarrow N.R.$$
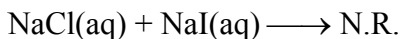
Tablet Lab Assessment: Limiting Reactant

_____ 1. The initial amounts of reactants for the reaction
below are shown to the right.

$$2A + B \longrightarrow 2C$$

What image below best shows the system after this
reaction has come to completion?

Initial system diagram:
Box containing: A, A, A, A, A, A, B, B

(A)

Box containing: C, C, C, C, A, A

(B)

Box containing: A, A, A, A, C, C, B, B

(C)

Box containing: A, A, A, A, C, C

_____ 2. Consider the reaction stated below:

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$
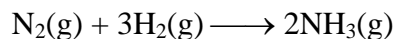
Suppose you have 4 moles of N_2 and 6 moles of H_2, which is the limiting reactant?

(A) $N_2(g)$ (B) $H_2(g)$ (C) $NH_3(g)$ (D) both $N_2(g)$ & $H_2(g)$

_____ 3. For the reaction

$$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$$
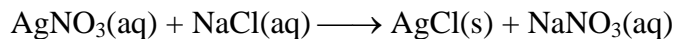
How many grams of NaCl are need to completely react with 4.15 g of $AgNO_3$?

(A) 4.15   (B) 5.14   (C) 1.43   (D) 3.49   (E) 2.07

_____ 4. For the reaction:

$$2AgNO_3(aq) + Na_2CO_3(aq) \longrightarrow Ag_2CO_3(s) + 2NaNO_3(aq)$$
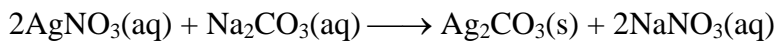
If initially 0.80 moles of $AgNO_3$ are reacted with 0.60 moles of Na_2CO_3 in a 500 mL aqueous solution how many moles of $Ag_2CO_3(s)$ will precipitate?

(A) 0.400   (B) 0.60   (C) 0.80   (D) 1.4   (E) 2.2

Tablet Lab: Examining the Role of the Limiting Reactant

Laboratory Goals: Determine the amount of product produced in a series of reactions. Understand which reactants are responsible for the amount of product produced.

One of the reactions you will study in this Tablet Lab is silver nitrate reacting with sodium chloride. This is a great precipitation reaction to study in the virtual lab, because currently a 125g bottle of silver nitrate costs $ 1,095.

1. Label a new page in your laboratory notebook for this tablet lab on Limiting Reactants.

2. Write the double displacement reaction for $AgNO_3$(aq) + NaCl(aq) in your laboratory notebook.

- Go to the website http://chemcollective.org/chem/wofford/
- In the blanks for your name enter your first and then last name
- Select "Limiting Reagents Activities"
- Allow Java to run on your tablet and open the Virtual Lab.

In the stockroom there are several water soluble salts NaCl(s), NaI(s), $AgNO_3$(s), $Pb(NO_3)_2$, and Na_2CO_3(s). We will use these reagents to investigate limiting reactanst in precipitation reactions.

3. What is characteristic of a precipitation reaction? Give one sentence in you lab notebook that describes a precipitation reaction.

- Pull onto the workspace the solid labeled "NaCl" and the flask labeled "100mL H_2O".
- Bring onto the virtual workspace a weigh boat that can be found by tapping the tools icon
- Place the reagent bottle of NaCl over the weigh boat. Enter 1.00 grams in the transfer bar below the workspace. Tap withdraw to place the amount on the spatula. Now tap pour and add this amount to the weigh boat.
- Pull the weigh boat over the "100mL H_2O" flask.
- Enter 1.00 g in the Transfer bar at the bottom of the Workbench. Then tap pour.

Wow, it would really make lab work easy if you could obtain exact masses this easily! This ads 1.00 g of NaCl to the solution.

Now we will react this sodium chloride solution with silver nitrate.

- Pull onto the workspace the solid $AgNO_3$.
- Place the solid $AgNO_3$ over a weigh boat.
- In the Transfer amount window type 1.00, then tap Withdraw and then Pour.
- Pull the weigh boat over the "100mL H_2O" flask.
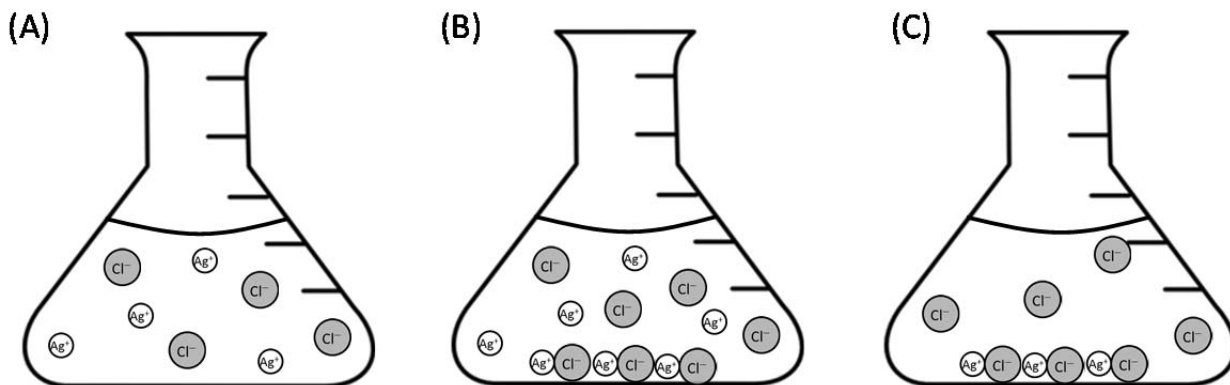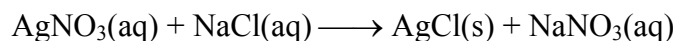- Enter 1.00 g in the Transfer bar at the bottom of the Workbench. Then tap pour.

The reaction you have just performed is: $AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$

- Remove the flask from the $AgNO_3(s)$ weigh boat.
- Click on the flask and select Solid in the Solution Info window.
- Determine the grams of solid formed. Be sure to indicate which compound is a solid.

4. Record the mass of $AgCl(s)$ in your lab notebook.

- Switch the Solution Info view to Aqueous
- Examine the ion concentrations in this solution.

5. Which ion is present in greatest concentration, Cl^- or Ag^+?

6. Which image below best depicts the species present in the reaction flask? In the images the spectator ions Na^+ and NO_3^- are not shown.

(A)   (B)   (C)

7. Based on you answers to Questions 5 and 6, did $AgCl(s)$ stop precipitating because chloride ions or silver ions are completely reacted away?

8. Determine the mass of $AgCl(s)$ that would be made using 1.00 g of NaCl and unlimited $AgNO_3$. Show your calculation in your lab notebook.

$$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$$

9. Repeat the mass calculation of $AgCl(s)$ that would be produced by the reaction of 1.00 g $AgNO_3(s)$ in the presence of unlimited NaCl.

10. Which calculation of $AgCl(s)$ mass matches the mass you recorded from the virtual lab?

Hopefully, by the end of the analysis we can conclude that it is the $AgNO_3$ that limits the amount of $AgCl(s)$ that will precipitate in the reaction you have performed.

11. Determine the mass of $AgNO_3$ that will completely react with 1.00 g of NaCl. Show this calculation in your notebook.

- Remove from the Workbench all items that are not $AgNO_3(s)$ and $NaCl(s)$.
- Bring onto the Workbench a new 100mL flask of H_2O.
- Weigh out 1.00 g of NaCl and add it to the flask of H_2O.
- Weigh out the amount of $AgNO_3$ required to exactly react with 1.00 g of NaCl.
- Add the $AgNO_3$ to the NaCl solution.
- Examine the mass of AgCl formed in the Solution Info window.

The Ag^+ and Cl^- concentrations should be very small, below $1x10^{-4}M$. These were both consumed in the reaction. Both the Na^+ and NO_3^- are spectator ions and will be present in significant concentrations.

12. Perform the reverse calculation, by determining the grams of NaCl required to react with exactly 1.00 g of $AgNO_3(aq)$. Perform this calculation in your laboratory notebook.

13. Determine the number of grams of $AgCl(s)$ that will form as precipitate when 1.00 g of $AgNO_3(aq)$ is completely reacted with $NaCl(s)$. Again in your laboratory notebook.

- Remove from the Workbench all items that are not $AgNO_3(s)$ and $NaCl(s)$.
- Bring onto the Workbench a new 100mL flask of H_2O.
- Weigh out 1.00 g of $AgNO_3$ and add it to the flask of H_2O.
- Weigh out the amount of NaCl required to exactly react with 1.00 g of $AgNO_3$.
- Add the NaCl to the $AgNO_3$ solution.
- Examine the mass of AgCl formed in the Solution Info window.

14. Record the mass of $AgCl(s)$ formed in the reaction flask in your lab notebook.

15. How does the number of grams of $AgCl(s)$ formed compare to your calculation in Problem 13?

The virtual lab has several advantages over real laboratory chemistry. To obtain the mass of $AgCl(s)$ in a solution you would have to filter the solid away from the liquid and then dry the solid and then finally weigh the solid. The Solution Info window is a wonderful virtual lab feature.

- Remove every item from the workbench except the solid $AgNO_3(s)$ and $NaCl(s)$.

We will examine the same reaction with larger quantities of reactants. Begin the reaction with 7.00 g of $AgNO_3$ reacting with 5.00 g of NaCl.

16. Use the molar masses of each compound to calculate the initial moles of $AgNO_3$, NaCl in you lab notebook.

Now that you know the number of moles of the reactants, we need to determine which reactant is limiting. Hopefully, by inspecting the number of moles we see that there are more moles of NaCl than $AgNO_3$. This means that $AgNO_3$ will be the limiting reactant. It is important to note that mass is not enough information to determine the limiting reactant.

17. To verify that AgNO3 is the limiting reactant, in your notebook perform the calculation, how many moles of $AgNO_3$ are required to react with 0.0855 mol NaCl. The calculation will take the form:
$$0.0855 \; mol \; NaCl \times \frac{1 \; mol \; AgNO_3}{1 \; mol \; NaCl} = 0.0855 \; mol \; AgNO_3$$

We interpret this calculation as, "We need 0.855 mol of $AgNO_3$ to react all of the NaCl, but since we only have 0.0412 moles of $AgNO_3$ there will be excess NaCl. The reactant that runs out will be $AgNO_3$.

There are two quantities we want to determine for the reaction. 1.) How many grams of NaCl remain in solution? 2.) How many grams of AgCl(s) are formed?

18. To determine the grams of NaCl remaining in solution, begin by calculating how many moles NaCl will react with 7.00 g (0.0412 mol) $AgNO_3$. In your lab notebook perform the stoichiometric calculation, how many of NaCl will react with 0.0412 moles $AgNO_3$?

19. Now in your lab notebook perform the subtraction of the moles of NaCl consumed (0.0412 mol) from the original number of moles of NaCl (0.0855 moles). This subtraction is the remaing moles of NaCl in solution.

20. Convert the number of moles of excess NaCl to grams of NaCl. Show your calculation in your lab notebook.

Now for the important calculation of how many grams of precipitate are formed.

21. Use a stoichiometric calculation to determine the number of grams of AgCl(s) formed in the reaction. Show the full calculation in your lab notebook.

Having performed all of the theoretical calculations of the reaction, use the virtual lab to run the reaction.

- Bring onto the Workbench the materials needed to react 7.00 g of $AgNO_3$ with 5.0 g of NaCl in 100 mL of H_2O.

- Perform the above reaction in 100 mL of water.
- Examine the products of the reaction in the Solution Info window, tap on the Solid button.

22. Record the mass of AgCl(s) formed during the reaction in your lab notebook.

- Return to the Solution Info window and select the Aqueous button.

The Aqueous button shows how many ions are left floating in the solution. The units for the ion concentration are molarity.

$$\text{Molarity} = \frac{moles}{liter}$$

Molarity is how many moles there would be in one liter of solution. In our case we only have 100 mL or $1/10^{th}$ of a liter. To obtain the moles in solution you must divide the molarity given in the Solution Info window by 10.

23. Record the moles of NO_3^- and Cl^- from the Solution Info window by dividing the molarity by 10.

24. How does the moles of Cl^- from the solution window compare to the moles of NaCl that you calculated for the reaction

A new precipitation reaction that we will study is lead(II) nitrate reacting with sodium iodide.

$$Pb(NO_3)_2(aq) + 2NaI(aq) \longrightarrow PbI_2(s) + 2NaNO_3(aq)$$
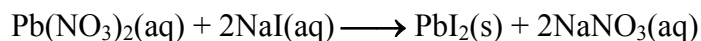
This is an interesting reaction because the stoichiometry is not 1:1.

We will study the specific reaction of 15.0 g of $Pb(NO_3)_2$ reacting with 15.0 g of NaI. First you will perform the mass and mole calculations and then perform the reaction in the virtual lab.

25. As you did in the previous reaction, begin by determining the moles of each reactant. In you lab notebook calculate the moles of both 15.0 g of $Pb(NO_3)_2$ and 15.0 g of NaI.

In this case you cannot simply examine the moles of reactants and decide the limiting reactant based on the reactant with the smallest number of moles. Instead you must perform the calculation that determines how many moles of NaI are required to react with 0.0452 moles of $Pb(NO_3)_2$.
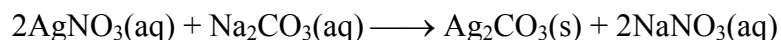
26. In your lab notebook perform the calculation asking how many moles of NaI are required to react with 0.0452 moles of $Pb(NO_3)_2$. You can examine Question # 17 for guidance in this calculation.

27. Based on your calculation is there enough NaI initially present? For this reaction what is the limiting reactant?

28. Using the moles of the limiting reactant determine how many grams of $PbI_2(s)$ are formed in the reaction. Show your stoichiometric calculations in your lab notebook.

- Completely clear the Workbench.
- Bring onto the Workbench the materials needed to react 15.0 g of $Pb(NO_3)_2$ with 15.0 g of NaI in 100 mL of H_2O.
- Following the same techniques of previous reactions, perform the above reaction in the virtual lab.
- Examine the products of the reaction in the Solution Info window, tap on the Solid button.

29. Record the grams of $PbI_2(s)$ produced in this reaction.

Hopefully, you are beginning to get comfortable calculating limiting reactant problems in the virtual lab environment. The final precipitation reaction of this experiment is the reaction of silver nitrate with sodium carbonate:

$$2AgNO_3(aq) + Na_2CO_3(aq) \longrightarrow Ag_2CO_3(s) + 2NaNO_3(aq)$$

30. Without instruction, complete in your lab notebook the calculations for the reaction of 9.00 grams of $AgNO_3$ with 4.00 g of Na_2CO_3. Be sure that you calculate the grams of solid formed and the moles of the excess reagent.

Tablet Lab: Examining the Role of the Limiting Reactant

Laboratory Goals: Determine the amount of product produced in a series of reactions. Understand which reactants are responsible for the amount of product produced.

One of the reactions you will study in this Tablet Lab is silver nitrate reacting with sodium chloride. This is a great precipitation reaction to study in the virtual lab, because currently a 125g bottle of silver nitrate costs $ 1,095.

1. Label a new page in your laboratory notebook for this tablet lab on Limiting Reactants.

2. Write the double displacement reaction for $AgNO_3$(aq) + NaCl(aq) in your laboratory notebook.

- Go to the website http://chemcollective.org/chem/wofford/
- In the blanks for your name enter your first and then last name
- Select "Limiting Reagents Activities"
- Allow Java to run on your tablet and open the Virtual Lab.

In the stockroom there are several water soluble salts NaCl(s), NaI(s), $AgNO_3$(s), $Pb(NO_3)_2$, and Na_2CO_3(s). We will use these reagents to investigate limiting reactanst in precipitation reactions.

3. What is characteristic of a precipitation reaction? Give one sentence in you lab notebook that describes a precipitation reaction.

- Pull onto the workspace the solid labeled "NaCl" and the flask labeled "100mL H_2O".
- Bring onto the virtual workspace a weigh boat that can be found by tapping the tools icon

- Place the reagent bottle of NaCl over the weigh boat. Enter 1.00 grams in the transfer bar below the workspace. Tap withdraw to place the amount on the spatula. Now tap pour and add this amount to the weigh boat.
- Pull the weigh boat over the "100mL H_2O" flask.
- Enter 1.00 g in the Transfer bar at the bottom of the Workbench. Then tap pour.

Wow, it would really make lab work easy if you could obtain exact masses this easily! This ads 1.00 g of NaCl to the solution.

Now we will react this sodium chloride solution with silver nitrate.

- Pull onto the workspace the solid $AgNO_3$.
- Place the solid $AgNO_3$ over a weigh boat.
- In the Transfer amount window type 1.00, then tap Withdraw and then Pour.
- Pull the weigh boat over the "100mL H_2O" flask.
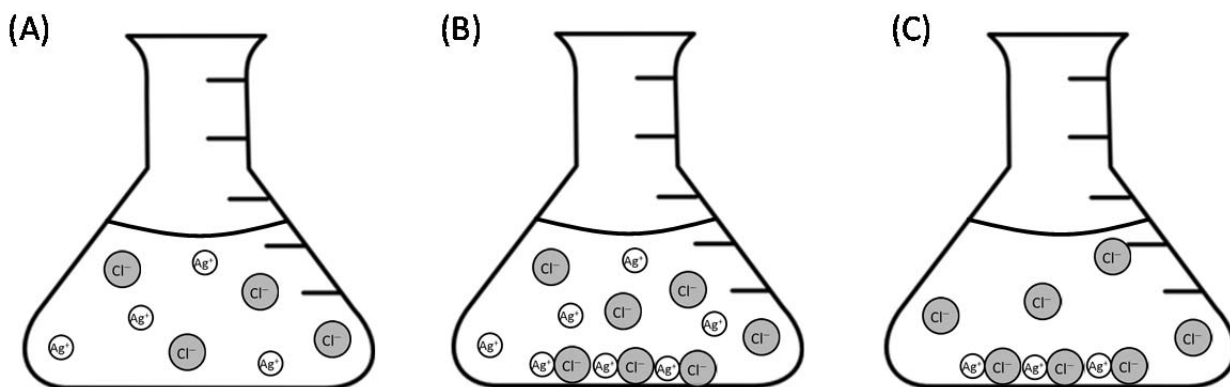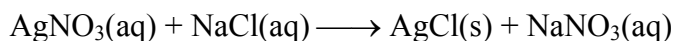- Enter 1.00 g in the Transfer bar at the bottom of the Workbench. Then tap pour.

The reaction you have just performed is: $AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$

- Remove the flask from the $AgNO_3(s)$ weigh boat.
- Click on the flask and select Solid in the Solution Info window.
- Determine the grams of solid formed. Be sure to indicate which compound is a solid.

4. Record the mass of $AgCl(s)$ in your lab notebook.

- Switch the Solution Info view to Aqueous
- Examine the ion concentrations in this solution.

5. Which ion is present in greatest concentration, Cl^- or Ag^+?

6. Which image below best depicts the species present in the reaction flask? In the images the spectator ions Na^+ and NO_3^- are not shown.

7. Based on you answers to Questions 5 and 6, did $AgCl(s)$ stop precipitating because chloride ions or silver ions are completely reacted away?

8. Determine the mass of $AgCl(s)$ that would be made using 1.00 g of NaCl and unlimited $AgNO_3$. Show your calculation in your lab notebook.

$$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$$

9. Repeat the mass calculation of $AgCl(s)$ that would be produced by the reaction of 1.00 g $AgNO_3(s)$ in the presence of unlimited NaCl.

10. Which calculation of $AgCl(s)$ mass matches the mass you recorded from the virtual lab?

Hopefully, by the end of the analysis we can conclude that it is the $AgNO_3$ that limits the amount of $AgCl(s)$ that will precipitate in the reaction you have perfomed.

11. Determine the mass of $AgNO_3$ that will completely react with 1.00 g of NaCl. Show this calculation in your notebook.

- Remove from the Workbench all items that are not $AgNO_3(s)$ and $NaCl(s)$.
- Bring onto the Workbench a new 100mL flask of H_2O.
- Weigh out 1.00 g of NaCl and add it to the flask of H_2O.
- Weigh out the amount of $AgNO_3$ required to exactly react with 1.00 g of NaCl.
- Add the $AgNO_3$ to the NaCl solution.
- Examine the mass of AgCl formed in the Solution Info window.

The Ag^+ and Cl^- concentrations should be very small, below 1×10^{-4}M. These were both consumed in the reaction. Both the Na^+ and NO_3^- are spectator ions and will be present in significant concentrations.

12. Perform the reverse calculation, by determining the grams of NaCl required to react with exactly 1.00 g of $AgNO_3(aq)$. Perform this calculation in your laboratory notebook.

13. Determine the number of grams of $AgCl(s)$ that will form as precipitate when 1.00 g of $AgNO_3(aq)$ is completely reacted with $NaCl(s)$. Again in your laboratory notebook.

- Remove from the Workbench all items that are not $AgNO_3(s)$ and $NaCl(s)$.
- Bring onto the Workbench a new 100mL flask of H_2O.
- Weigh out 1.00 g of $AgNO_3$ and add it to the flask of H_2O.
- Weigh out the amount of NaCl required to exactly react with 1.00 g of $AgNO_3$.
- Add the NaCl to the $AgNO_3$ solution.
- Examine the mass of AgCl formed in the Solution Info window.

14. Record the mass of $AgCl(s)$ formed in the reaction flask in your lab notebook.

15. How does the number of grams of $AgCl(s)$ formed compare to your calculation in Problem 13?

The virtual lab has several advantages over real laboratory chemistry. To obtain the mass of $AgCl(s)$ in a solution you would have to filter the solid away from the liquid and then dry the solid and then finally weigh the solid. The Solution Info window is a wonderful virtual lab feature.

- Remove every item from the workbench except the solid $AgNO_3(s)$ and $NaCl(s)$.

You will be introduced to a tabular method for examining chemical reactions. This is a tabular method used by organic chemists. The reaction table below is helpful for keeping track of all species before and after a reaction.

We will use the table below to examine the reaction of 7.00 g of $AgNO_3$ reacting with 5.00 g of NaCl.

	$AgNO_3(aq)$	+ $NaCl(aq)$	$\longrightarrow$	$AgCl(s)$	+ $NaNO_3(aq)$
Initial amount (g)	7.00	5.00		0.00	0.00
Initial amount (mol)					
Loss / Gain (X)	−X			+X	
Final amount (X)					
Final amount (mol) X = _____					
Final amount (g)					

The chart looks intimidating at first, but with practice you might find you like the organization of the reaction that the chart provides.

16. Use the molar masses of each compound to calculate the initial moles of $AgNO_3$, NaCl, $AgCl(s)$, and $NaNO_3$. Copy this table into your lab notebook. Place the mole values in the second row of the reaction table. (The initial moles of products will be zero.)

The row labeled "Loss / Gain" expresses the loss of reactant and gain of product, taking into account the stoichiometric coefficients. In this reaction the stoichiometry is all 1 mole:1 mole. Reactants are decreasing and their change is shown as –X, while the products are increasing in amount and their change is shown as +X. X has units of moles in a reaction table.

17. In the row for Loss / Gain, insert X with the appropriate sign.

The row labeled "Final amount (X)" is maybe the easiest row to determine in the table. Final amount is simply, Initial amount (mol) + Loss / Gain (X). For the case of $AgNO_3$ "Final amount (X)" would be:

$$\text{Final amount } AgNO_3 = 0.0412 \text{ mol} + (-X)$$

18. Complete the row for "Final amount (X)".

Any time you see X in an expression you know some math is going to follow. Yes, we need to solve for X. Fortunately, the algebra is simple. We know that one (or both) of the reactants is going to run out, reaching 0.00 moles. We need to solve for X assuming the $AgNO_3$ is going to run out and then separately solve for X assuming that NaCl is going to run out.

To calculate X, assuming $AgNO_3$ is limiting, the following algebra is used.

$$0.0412 \text{ mol} - X = 0.000 \text{ mol}$$

$$X = 0.0412 \text{ mol.}$$

19. Repeat the calculation for X assuming that NaCl will equal 0.00 mol.

Now that X has been determined for the two reactant species, you need to examine the reasonableness of the value of X. X cannot be a value that results in a negative number of moles of any reactant.

20. Calculate the final moles of $AgNO_3$ and NaCl if X = 0.0412 mol.

21. Calculate the final moles of $AgNO_3$ and NaCl if X = 0.0855 mol.

22. Which value of X results provides a reasonable result for both $AgNO_3$ and NaCl? Plug this value of X into the reactant table in the "Final amount (mol)" location.

23. Complete the row "Final amount (mol)" by using the correct value of X.

24. The final row in the table is to report the final amount in units of grams rather than units of moles. This fortunately only requires you to multiply the amount of moles by the molar mass. Complete the final row by multiplying the final moles by the molar mass of each species.

Wow, this was a ridiculous amount of work. Yes, creating the table takes some time, but it provides a complete picture of the reaction.

- Bring onto the Workbench the materials needed to react 7.00 g of $AgNO_3$ with 5.0 g of NaCl in 100 mL of H_2O.
- Perform the above reaction.
- Examine the products of the reaction in the Solution Info window, tap on the Solid button.

25. Record the mass of AgCl(s) formed in your lab notebook.

- Return to the Solution Info window and select the Aqueous button.

The Aqueous button shows how many ions are left floating in the solution. The units for the ion concentration are molarity.

$$\text{Molarity} = \frac{moles}{liter}$$

Molarity is how many moles there would be in one liter of solution. In our case we only have 100 mL or $1/10^{th}$ of a liter. To obtain the moles in solution you must divide the molarity given in the Solution Info window by 10.

26. Record the moles of NO_3^- and Cl^- from the Solution Info window by dividing the molarity by 10.

27. How does the moles of NO_3^- from the solution window compare to the moles of $NaNO_3$ that you calculated in the reaction table?

A new precipitation reaction that we will study is lead(II) nitrate reacting with sodium iodide.

$$Pb(NO_3)_2(aq) + 2NaI(aq) \longrightarrow PbI_2(s) + 2NaNO_3(aq)$$
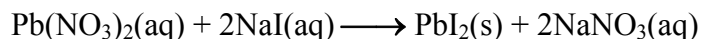
This is an interesting reaction because the stoichiometry is not 1:1.

We will study the specific reaction of 15.0 g of $Pb(NO_3)_2$ reacting with 15.0 g of NaI. Again use the reaction table below to help organize the reaction.

	$Pb(NO_3)_2$(aq)	**+ 2 NaI(aq)**	$\longrightarrow$	**PbI_2(s)**	**+ 2 NaNO_3(aq)**
Initial amount (g)	15.00	15.00		0.00	0.00
Initial amount (mol)					
Loss / Gain (X)					
Final amount (X)					
Final amount (mol) X = _____					
Final amount (g)					

28. Begin by creating this reaction table in your lab notebook and using the molar mass of each compound to determine the initial amount present in moles. Enter the initial moles in row two of your table.

29. To express the change in number of moles in terms of X, we must examine the stoichiometric coefficients of the reaction. If the change in $Pb(NO_3)_2$ is expressed as –X, then the change in NaI will be expressed as −2X. Complete the row of the table for the loss / gain expressed as X.

30. Recall thatthe Final amount (X) is simply the sum of Initial Amount + Loss / Gain (X). Complete this row of the table.

31. Final amount (mol) is the most difficult row to solve. We know that the reaction will proceed until one of the reactants reaches 0.000 moles. Solve for X assuming that $Pb(NO_3)_2$ will be 0.00 moles this would be:

$$0.0453 - X = 0.00$$

32. Next solve for X assuming that NaI will be 0.000 moles.

33. One of the solutions for X is unreasonable. If the larger value of X was used in Row 5 some of the cells would have negative values of moles. Use the smaller of the two X values from Problems 31 and 32 to solve for the Final amount in moles.

34. Determine the final grams at the end of the reaction by using the molar mass to convert moles to grams. Complete the final row of the table.

- Completely clear the Workbench.
- Bring onto the Workbench the materials needed to react 7.00 g of $AgNO_3$ with 5.0 g of NaCl in 100 mL of H_2O.
- Perform the above reaction.
- Examine the products of the reaction in the Solution Info window, tap on the Solid button.

35. Record the grams of $PbI_2(s)$ produced in this reaction.

The final precipitation reaction of this experiment is the reaction of silver nitrate with sodium carbonate:

$$2AgNO_3(aq) + Na_2CO_3(aq) \longrightarrow Ag_2CO_3(s) + 2NaNO_3(aq)$$
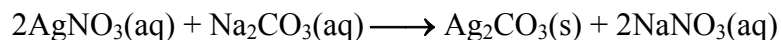
36. Without instruction, complete in your lab notebook the table for the reaction of 9.00 grams of $AgNO_3$ with 6.00 g of Na_2CO_3.

	2 $AgNO_3$(aq)	+Na_2CO_3(aq)	$\longrightarrow$	Ag_2CO_3(s)	+ 2 $NaNO_3$(aq)
Initial amount (g)	9.00	6.00		0.00	0.00
Initial amount (mol)					
Loss / Gain (X)					
Final amount (X)					
Final amount (mol) X = _____					
Final amount (g)					

Tablet Lab: Acids, Bases and Solution pH

Laboratory Objectives: Examine different acidic and basic solutions, recording and calculating the solution pH. Develop and understanding of the relationship between concentrations of hydrogen ions and hydroxide ions in aqueous solutions. Calculate the hydrogen ion concentration from the solution pH.

Acids and bases are two of the oldest chemical classifications of chemical compounds. A simple definition of an acid is a compound that donates a hydrogen cation, H^+, (we call this a proton). A base can be identified as a compound that generates hydroxide ions, OH^-, in aqueous solution.

We will examine some acid and base solutions in the Virtual Lab.

- Open the virtual lab on your tablet by going to "chemcollective.org" in a web browser.
- From the Chemcollective homepage (chemcollective.org) tap on the link "Virtual Lab" on the right-hand side of the page.
- To launch the virtual lab, Java must open on your tablet. Windows will ask you if you want to run this application. Check "I accept the risk and want to run this app." Then click Run.
- The virtual lab will open. Scroll to put the virtual lab in the middle of your tablet.
- In the "Stockroom Explorer" on the left, open the cabinet for Strong-acids. Pull on to your Workbench a 0.1 M solution of HCl.
- Examine the Solution Info window for this solution.

1. Record in your lab notebook the concentration of each species in solution. (A short hand for molar concentration is [ ]. $[NO_3^-]$ means the concentration of nitrate ions in units of molarity, $\frac{mol}{L}$.)

2. Which two species have the same concentration?

The reaction that occurs when hydrogen chloride gas is added to water is:

$$HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$

Hopefully, you did not observe any HCl(aq) in solution. No HCl molecules are observed because HCl is an acid that completely breaks apart into its ions in water. Acids that dissociate 100% to ions are termed **strong acids**. You might expect that **weak acids** will behave differently.

- Go to the Weak-acids cabinet and withdraw 1.0 M HF.
- Examine the species present in the 1.0 M HF solution in the Solution Info window.

3. Record in your lab notebook the concentration of each species in the 1.0 M HF solution.

4. What is the most abundant specie in the HF(aq) solution?

The acid reaction is very similar for the hydrogen fluoride gas.

$$HF(g) \xrightarrow{H_2O} H^+(aq) + F^-(aq)$$

The big difference is that while 100% of the HCl dissociated in water less than 3 % of the HF molecules dissociated into ions when HF is added to water. The HF(aq) concentration is 0.975 M, compared to its initial concentration of 1.00 M. Acids that only dissociate (or ionize) a little in water are termed **weak acids**.

There are far more weak acids than strong acids. The list of strong acids is short. It is likely that you will be asked to memorize a handful of common strong acids.

- Remove your solution of HF from the Workbench, using the remove icon, . We will work with your HCl solution first.
- Click on your HCl solution so that its information is available in the Solution Info window.

5. Record the solution pH for the 0.1 M HCl solution.

If you have discussed this in class, you know the pH is simply another way to express the concentration of hydrogen ions in solution. In your HCl solution the hydrogen ion concentration is $[H^+] = 0.100$ M. Solution pH is calculated by the following:

$$pH = -\log([H^+])$$

On your calculator be sure that you can calculate a solution pH, by completing the operation below.
$$pH = -\log(0.100) = 1.0$$

The log you are using on your calculator is the log base 10. This means that a 10 times change in $[H^+]$ will result in a pH change of 1. Let's examine this change.

- Using the glassware icon , pull onto your Workbench an empty 250 mL Erlenmeyer flask and a 600 mL beaker.
- Additionally, bring onto the Workbench the large container of distilled water found in the Stockroom Explorer.
- Place the distilled water over the 600 mL beaker.
- Type "500" into the Transfer bar at the bottom of the Workbench. Then tap "Pour".
- Tap on your water beaker and select "Rename" from the Edit menu. Label the beaker, "Water".
- Place the 0.1 M HCl flask over the empty Erlenmeyer flask.

- Pour 10 mL of the HCl solution into the empty flask, using the Transfer bar.

6. Has the hydrogen ion concentration, [H$^+$], changed in either container?

- From your water beaker, add 90 mL of water to the 250 mL flask that contains 10 mL of 0.1 M HCl.

7. Use the relationship, $M_1V_1 = M_2V_2$, to determine the concentration of HCl in your new mixture. (Show your calculation in your lab notebook.)

- Check your dilution calculation, by examining the Solution Info window for this diluted acid solution.
- Rename your diluted acid flask to, "0.01 M HCl"

By creating a 100 mL solution using 10 mL of acid, you have performed a 10:1 dilution (10-fold dilution).

8. Record the pH of the 0.01 M HCl solution.

- Pull onto the Workbench a new 250 mL Erlenmeyer Flask.
- Perform a 10:1 dilution of your 0.01 M HCl in this new beaker.

9. What is the concentration of this new acid solution?

10. Record the pH of this new acid solution?

- Continue to prepare 10:1 acid dilutions using the most recently prepared solution. Create sequentially more dilute HCl solutions until you have acid solutions labeled from 0.1 M to 1E-6 M, as shown below.

0.1M HCl   0.01 M HCl   0.001 M HCl   1E-4 M HCl   1E-5 M HCl   1E-6 M HCl

11. Create a table in your lab notebook, like the one below, which allows you to record pH, [H$^+$], [OH$^-$] for the acid solutions you have prepared. Add a final column that will be the product (multiplication) of the H$^+$ and OH$^-$ ion concentrations. Complete the table.

HCl Solution	pH	[H$^+$]	[OH$^-$]	[H$^+$]·[OH$^-$]
0.10 M				
0.01 M				
0.001 M				
1 x10^{-4} M				
1 x10^{-5} M				
1 x10^{-6} M				
1 x10^{-7} M				

To complete the final column of the table, multiply the H^+ ion concentration by the OH^- ion concentration this will give a small number.

12. Examine your data table. Write in your notebook the correct statement below regarding pH and solution acidity.

(A) As the solution pH increases the solution acidity increases.
(B) As the solution pH increases the solution acidity decreases.

13. Examine your data table. Write in your notebook the correct statement below regarding pH and solution acidity.

(A) As the H^+ concentration in solution increases the OH^- concentration decreases.
(B) As the H^+ concentration in solution increases the OH^- concentration increases.

14. For each acid solution what is the product of the H^+ ion and OH^- ion concentrations.

Hopefully, from your examination of the data table, you have notice the inverse relationship between $[H^+]$ and $[OH^-]$ in solution. When either of the species is in large concentration the other species is present in a small concentration. If you know either $[H^+]$ or $[OH^-]$ in solution you can determine the other concentration. Mathematically, this inverse relationship can be expressed as:

$$[H^+] = \frac{1 \times 10^{-14}}{[OH^-]} \quad or \quad [OH^-] = \frac{1 \times 10^{-14}}{[H^+]} \quad or \quad [H^+][OH^-] = 1 \times 10^{-14}$$

15. For an acid solution with a hydrogen ion concentration of $[H^+] = 4 \times 10^{-5}$M, calculate the hydroxide ion, OH^-, concentration and the solution pH

How can the $[H^+] = 4 \times 10^{-5}$M solution be prepared? Probably by dilution since that is how all of the previous solutions have been prepared.

16. Record in your notebook the solutions below that could be used to prepare a $[H^+] = 4 \times 10^{-5}$M solution by dilution.

(A) 1×10^{-4}M HCl   (B) 1×10^{-5} M HCl   (C) 1×10^{-6} M HCl

In order to perform a dilution you must have a sample that is more concentrated than the solution you are trying to prepare.

- Pull onto your Workbench a new 250 mL flask. Refill the water beaker if it is running low.

17. Use the dilution relationship, $M_1V_1 = M_2V_2$, to calculate how much 1 x 10^{-4} M HCl you require to prepare 100 mL of 4 x 10^{-5} M HCl solution.

- Using your dilution calculation, prepare a 100 mL solution with $[H^+] = 4 \times 10^{-5}$M. Be sure to properly label this new acid solution.

- Examine the Solution Info window for your new acid solution. Do the pH and OH^- concentration match your calculated values?

It is true that the OH^- concentration is a little bit off, but only in the third decimal place. If you want all of the decimals to match you will need to use the more accurate relationship, $[H^+][OH^-] = 1.01 \times 10^{-14}$.

There is one unfilled row in your data table. You have not prepared a 1E-7 M HCl solution.

18. Before you prepare this solution make a prediction about the pH of a 1E-7 M HCl solution in your lab notebook.

- Bring onto the Workbench a new 250 mL Erlenmeyer flask and prepare a 1E-7 M HCl solution, by performing a 10:1 dilution of the 1E-6 M HCl solution.
- Relabel the new acid solution "1E-7 M HCl".
- Examine the solution pH and species concentrations in the Solution Info window.

19. Does the pH match your prediction?

20. What on your Workbench has a pH of 7.00?

Hopefully, you identified that the water on the Workbench has a pH = 7.00. In this solution $[H^+]$ = $[OH^-]$ = 1×10^{-7}. Anytime the $[H^+]$ is the same as the $[OH^-]$ the solution is considered neutral and the solution pH = 7.00. If you add any acid to pH = 7.00 water, even if it is only a little bit, the pH will drop below 7.00 and the solution will be acidic.

Let's examine the pH of a basic solution.

- The Workbench has gotten crowded. Remove all of the HCl flasks using the remove

  icon . Leave on the bench your water beaker and carboy.
- Bring onto the bench a flask of 0.1M NaOH.
- Examine the solution properties in the Solution Info window.

21. Record the solution concentrations and pH in your lab notebook for this base solution.

Recall that one of the definitions for a base is a substance that produces hydroxide, OH^- ion in solution. Sodium hydroxide is a soluble salt that fully dissociates by the reaction.

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

Sodium hydroxide is termed a **strong base** because it dissociate completely. The pH of 0.1 M NaOH solution is pH = 13.0. This pH is quite a bit larger than the neutral pH = 7 solution. Solutions with a pH greater than 7.0 are considered basic.

- Obtain a clean 250 mL Erlenmeyer flask and perform a 10:1 dilution of the 0.1 M NaOH solution. This dilution will prepare 100 mL 0f 0.01 M NaOH.

You now have two basic solutions; one pH = 12.00 and the other pH = 13.00. Calculating the pH of a basic solution is done in same way as an acidic solution.

$$pH = -\log([H^+])$$

However, in a 0.01 M NaOH solution the OH⁻ concentration is equal to 0.01 M and the H⁺ ion concentration must be calculated using the relationship below, before you can plug into the pH equation.

$$[H^+] = \frac{1 \times 10^{-14}}{[OH^-]}$$

There is a relationship similar to pH known as pOH. pOH is minus the log of the hydroxide ion concentration.

$$pOH = -\log([OH^-])$$

22. Calculate the pOH for the two base solutions on the Workbench. Show your work in your lab notebook.

Because [H⁺] is mathematically related [OH⁻], pH is mathematically related to pOH. The relationship is

$$pH + pOH = 14 \qquad or \qquad pH = 14 - pOH$$

23. From your pOH values for the two base solutions, calculate the pH of each NaOH solution in your lab notebook.

24. Calculate what the solution pH will be when 5 mL of 0.01 M NaOH is added to 145 mL of water. In your lab note book report the [OH⁻], pOH, pH and [H⁺].

- In the virtual lab obtain a new 250 mL Erlenmeyer flask and prepare the solution describe in the problem above.

Problems involving acids and bases are often expressed in the following form: How much 1M HCl is required to make 1L of a pH = 4.5 solution?

In this case we are given pH and must find the hydrogen ion concentration, [H⁺]. The log function has an inverse. The log inverse is expressed in the following relationship.

$$[H^+] = 10^{-(pH)}$$

Any pH = 4.5 solution will have [H⁺] calculated by the operation

$$[H^+] = 10^{-(4.5)} = 3.13 \times 10^{-5} M$$

25. Calculate the hydrogen ion concentration, [H⁺], for a solution that is pH = 3.72.

- In the virtual lab remove all basic solutions.

- Use any of the acids in the Strong-acids cabinet to prepare 200 mL of an acidic solution that is pH =3.72.

26. Show the calculations required to prepare this pH = 3.72 solution in your lab notebook.

The final assignment for this lab is to prepare a basic solution of a given pH. A similar relationship between [OH−] and pOH is given below.

$$[OH^-] = 10^{-(pOH)}$$

27. Calculate the hydroxide ion concentration, [OH⁻], for a solution that is pH = 9.81.

- Use any concentration of sodium hydroxide solution in the Strong-bases cabinet to prepare 500 mL of a basic solution of pH = 9.81.

28. Record in your lab notebook, the calculations you require to prepare the above basic solution.

Leave your Workbench a complete mess. You have completed this tablet lab.

Tablet Lab Assessment: Acids, Bases and Solution pH

_____ 1. As the $[H^+]$ ion concentration is increased in a solution which pair of changes will be observed?

(A) pH will increase and $[OH^-]$ will increase
(B) pH will decrease and $[OH^-]$ will increase
(C) pH will increase and $[OH^-]$ will decrease
(D) pH will decrease and $[OH^-]$ will decrease

_____ 2. For a solution with $[H^+] = 1.8 \times 10^{-4}$ M the $[OH^-]$ will equal

(A) 4.1×10^{-4} M   (B) 1.8×10^{-4} M   (C) 5.6×10^{-11} M   (D) 1.0×10^{-7} M
(E) $[OH^-]$ cannot be determined without the solution volume

_____ 3. What is the pH of a 3.1×10^{-5} M HCl(aq) solution?

(A) 3.1×10^{-5} (B) 4.51   (C) 10.4   (D) 3.10   (E) 9.81

_____ 4. A solution with the label 1.0×10^{-7} M HCl should be considered

(A) acidic   (B) basic   (C) neutral

_____ 5. How many mL of 0.20 M NaOH are needed to prepare 500 mL of 6.8×10^{-3} M NaOH solution?

(A) 5.8 mL   (B) 147 mL   (C) 367 mL   (D) 0. 272 mL (E) 17 mL

Tablet Lab Assessment: Determining the Molar Mass of a Monoprotic Acid

_____ 1. In a titration of an acid with 1.00 M NaOH(aq) how many moles of NaOH are supplied
by the buret if the initial reading is $V_1 = 2.35$ mL and the final buret reading is $V_2 =$
28.00 mL?

   (A) 2.35 moles   (B) 0.028 moles   (C) 0.02565 moles   (D) 2.35×10^{-4} moles

_____ 2. During a laboratory titration of a weak acid and strong base the end-point of the
titration is determined by

   (A) reaching the correct buret volume
   (B) the solution pH becomes 7.00
   (C) stoichiometric coefficients of the reaction
   (D) the titration indicator changes color

_____ 3. In the determination of the molar mass by titration which measurement limits the
significant digits that can be reported for the molar mass?

   (A) moles of the acid
   (B) mass of the acid
   (C) volume of water used
   (D) weigh boat mass

_____ 4. In the laboratory a 5.138 g sample of an unknown acid was titrated to its end-point by
the addition of 18.3 mL of 1.00 M NaOH. From these laboratory measurements the
molar mass of the unknown acid is best reported as?

   (A) 281 g/mol   (B) 356 g/mol   (C) 94.02 g/mol   (D) 183 g/mol

# Tablet Lab: Determining the Molar Mass of a Monoprotic Acid

Laboratory Goals: Use the technique of acid base titration to determine the molar mass of an unknown monoprotic acid.

One of the essential properties of a compound is its molar mass. Determining the g/mol of a molecule is one of the first measurements made in characterizing an unknown a compound.

1. Title and date a new page in your laboratory notebook for this tablet lab on Molar Mass Determination.

- Go to the website http://chemcollective.org/chem/wofford/
- In the blanks for your name enter your first and then last name.
- Select "Unknown Solid Acid Experiment (Version 1)".
- Allow Java to run on your tablet and open the Virtual Lab.

In the virtual lab stockroom there are two acids. The first acid, KHP, (potassium hydrogen phthalate, $KC_8H_4O_4H$, molar mass 204.2 g/mol) is an acid with a known molecular weight that we will use to learn the titration process. The second acid is an unknown compound that you will experiment with to measure the molecular weight. The stockroom also contains the base NaOH at 1.00 M concentration and the acid base indicator phenolphthalein.

Begin experimentation by titrating a known amount of the KHP aicd. The acid base neutralization reaction that you will be performing is:

$$KC_8H_4O_4H(aq) + NaOH(aq) \longrightarrow NaKC_8H_4O_4(aq) + H_2O(l)$$
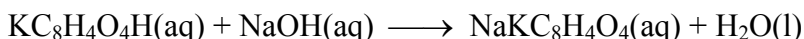
- Bring onto the Workbench by clicking and dragging: the bottle of KHP, the flask of 1.00 M NaOH, the phenolphthalein indicator and the flask of 100 mL H_2O.

- From the glassware icon, [icon], select a 50 mL buret and a 250 mL beaker (for waste).

- From the equipment icon, [icon], bring onto the workbench a scale and a weigh boat.

In a titration, a buret is used to carefully add a known amount of reactant (NaOH) to an unknown amount of a second reactant (the acid).

- Prepare the buret by filling it with the 1.00 M NaOH solution. To accomplish this, drag the NaOH solution over the buret and release.
- In the Transfer amount (mL) line at the bottom of the Workbench enter 50 mL and press Pour.

A problem is that you will not get exactly 50 mL of NaOH. You may get 45 mL or you may get 55 mL or something in between. If your buret is fille above the 50 mL mark, you must remove some of the excess liquid.

- If your buret is above the 50 mL mark, drag the buret over the 250 mL waste beaker. Use the transfer window to remove some of the liquid, usually 5.00 mL will be enough.

Your buret is now ready to use. In the lab you would have had to clean the buret before you added any NaOH. Virtual glassware is always clean; isn't that nice? Now we must prepare the KHP acid solution.

- Place the weigh boat onto the electronic balance and tap tare.
- Move the bottle of KHP over the balance.
- In the Transfer amount(g) window enter 10 and tap Withdraw. (This takes a 10 grams scoop from the bottle.)
- You want between 5 and 6 grams of sample. Change the amount Transfer amount(g) window to 1 g.
- Tap the Pour button. As you tap the Pour button, watch the mass on the balance.
- Tap the Pour button until the mass of KHP in the weigh boat is between 5 and 6 grams.

2. In your lab notebook record the exact mass of KHP.

Next, dissolve the solid KHP in 100 mL of water.

- Drag the weigh boat of KHP over your 100 mL Erlenmeyer flask of water.
- In the Transfer amount (g) window enter a number much greater than 5.0 g (ex. 20).
- Then tap the Pour button, this will transfer all of the KHP to the flask.

The final step of preparation is to add the indicator to the acid solution.

- From the glassware icon on the Workbench find a disposable pipet.
- Pull the pipet over the bottle of Phenolphthalein.
- Use the Transfer amount (mL) window to withdraw 1.0 mL of the indicator.
- Bring the pipet over the KHP solution.
- Dispense the indicator into the KHP solution using the Transfer amount window and the Pour button.

The phenolphthalein indicator is a compound that is clear when the solution is acidic and turns pink in a basic solution. In lab work you would only use 3-4 drops of the indicator. In the virtual lab, the color change is not as vibrant as in the lab and more indicator is needed.

It is now time to titrate the acid sample with the sodium hydroxide.

- Bring the buret and the acid solution into the middle of the Workbench.
- Drag the buret over the flask containing the acid.
- Make the initial buret reading, V_1. In the example to the right the buret reading would be recorded as $V_1 = 0.50$ mL.
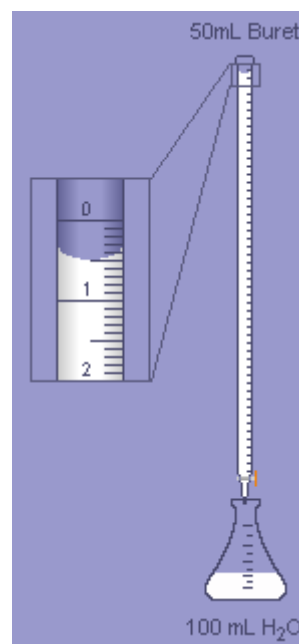
3. Record your initial buret reading in your lab notebook.

Notice that the pH is reading in the ball park of pH = 3. This is showing the pH of the acid solution. It will be important to keep an eye on the pH meter as you titrate.

- Begin to add base using the buret by typing 1.00 mL into the Transfer amount (mL) window.
- Tap the Pour button to deliver close to 1.0 mL.
- Continue to tap the Pour button to add additional amounts of NaOH. 1.0 mL is a lot of base to add. Add 1.0 mL NaOH increments until the solution pH changes by 0.3 pH units when 1 mL is added. This will be around pH = 6.0 for KHP.
- When a pH change of 0.3 units is observed, change the Transfer amount window to read 0.10 mL.
- Tap the Pour button until the pH is equal to 7.0. This again will take several taps.

Even though you are at pH = 7.0 the solution still has more moles of acid than moles of base. It is pH 7.0 because NaOH is a strong base and KHP is a weak acid.

- Change the Transfer amount window to 0.05 mL.
- Continue to add NaOH to the acid solution by tapping the Pour button, but this time keep an eye on the color of the acid solution. As soon as the solution turns a faint pink, stop adding NaOH from the buret.
- Record the final buret reading.

4. In your lab notebook record the final volume reading, V_2, of the buret.

At this point you have collected the experimental data needed to determine the molar mass of the acid, KHP. The lab goal is to determine the molar mass of the monoprotic acid. The grams of acid are determined from weighing the acid and the moles are determined from the titration with NaOH. The figure below shows how the acid mass and number of moles combine to give the molar mass in units g/mol.
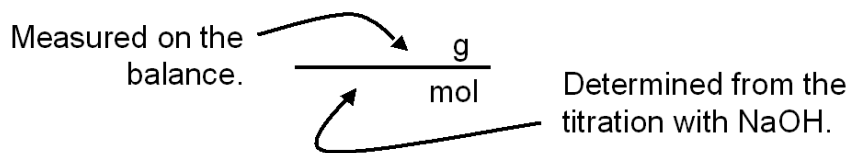
At the titration end point (when the solution turns color), the moles of OH$^-$ (base) are equivalent to the moles of H$^+$ (acid) in the sample. The moles of OH$^-$ added to the solution from a buret are calculated from the concentration of the base (M_{OH}) and the volume of base (V_{OH}) added, as:

$$M_{OH^-}\left(\frac{mol}{L}\right) V_{OH^-}(L) = n_{OH^-}(mol) \qquad \text{Eqn. 1}$$

Because KHP is a monoprotic acid, the moles of acid are equal to the moles of OH⁻ at the endpoint of the titration.

$$n_{OH^-}(mol) = n_{acid}(mol) \qquad \text{Eqn. 2}$$

5. To begin calculating the moles of acid titrated, first determine the volume of NaOH added by the buret. This is a subtraction, $\Delta V = V_2 - V_1$. Record ΔV in your notebook in both mL and L.

The virtual lab has made the mole calculation easier by using 1.00 M NaOH. For a change in volume of 23.2 mL NaOH the mole calculation would be:

$$(0.0232\ L)\left(1.00\frac{mol\ OH^-}{L}\right) = 0.0232\ mol\ OH^-$$

6. Convert the volume of NaOH into a number of moles of NaOH delivered by the buret. Record the number of moles in your lab notebook.

For acids with a single proton (monoporotic) the moles of OH⁻ are equal to the moles of acid, KHP.

7. Use your determination of the mass of KHP and the moles of KHP to determine the molar mass of KHP. The figure on the previous page will help you with the calculation. Record the molar mass of KHP in your lab notebook. To how many significant digits do you feel comfortable reporting the molar mass?

8. After obtaining the molar mass of KHP, if directed by your instructor, go to the classroom whiteboard and record your value of the molar mass of KHP.

Clean up the virtual Workbench and prepare to determine the molar mass of the unknown acid. You will use your own procedure for this determination, but it is likely to be similar to the procedure you used for the determination of molar mass of the acid KHP. KEEP AN EYE ON THE pH during the titration. When the pH changes by 0.3 units after the addition of 1 mL of NaOH, change the volume of NaOH from 1.0 mL to 0.10 mL or 0.05 mL

In your individual work be sure to record the important measurements and perform all calculations in your lab notebook.

When you have determined the molar mass of your unknown acid, enter the value in the window at the bottom of the page. After entering the molar mass (use 4 significant digits) click the "Check" button. Hopefully you will have performed your work carefully and the molar mass will be correct. If you have entered the incorrect molar mass you may try titrating again with a new sample of the unknown acid. If you want to try a different acid entirely you can click the refresh button on your web browser. This will provide you with a different unknown acid sample.