Analysis of Vitamin C

Minneapolis Community and Technical College
v.11.17

Objectives: To determine the amount of vitamin C in tablet and juice samples. Practice titration techniques and molar/mole/gram calculations using sequential reactions.

Prelab Questions: Read through this lab handout and answer the following questions before coming to lab. There will be a quiz at the beginning of lab over this handout and its contents.

1. Use chemical equation "a" below to determine the $I_2 : IO_3^-$ mole ratio
2. Use chemical equation "b" below to determine the $I_2 : C_6O_6H_8$ mole ratio
3. Use your results from 1 and 2 above to determine the $IO_3^- : C_6O_6H_8$ mole ratio
4. Why do manufacturers put more vitamin C in tablets than what is stated on the bottle label?
5. What color marks the endpoint of the titrations in today's experiments?
6. Why is it unnecessary to know the concentrations of I^- and H^+ in these experiments?
7. How many decimal places are required for all burette measurements?
8. Why are the vitamin C tablets first soaked in water?
9. Why is determining the endpoint of the juice titration difficult?
10. Why not use heat in today's titrations?

The Chemistry:

In today's experiment you will determining the amount of vitamin C in both tablets and in a juice sample. The procedure involves the careful addition potassium iodate (KIO_3) solution to the vitamin C containing sample via a titration. The KIO_3 is added to the reaction mixture *using a burett*e while stirring constantly with a magnetic stir plate. By knowing the concentration and volume of the KIO_3 required, the moles & grams of vitamin C can be determined.

As $\mathbf{KIO_3}$ is added to the reaction container, it reacts with loose $\mathbf{I^-}_{(aq)}$ ions to form $\mathbf{I_2}_{(aq)}$ (reaction "a" below). The $\mathbf{I_2}_{(aq)}$ product molecules are then chemically reduced by vitamin C, a strong reducing agent, to reform $\mathbf{I^-}_{(aq)}$ (reaction "b"). Thus, $\mathbf{I_2}_{(aq)}$ never builds up to appreciable levels because of the vitamin C that is present.

This process can't go on forever since eventually there will be no more vitamin C to react with $\mathbf{I_2}_{(aq)}$. At some point, when all of the vitamin C is completely gone (i.e. the limiting reactant), reaction "b" can no longer take place and the concentration of $\mathbf{I_2}_{(aq)}$ will increase. The end of the titration is signaled by the build up of $I_{2(aq)}$ and its reaction with excess I^- to form I_3^- (reaction "c") . This in turn reacts with starch to turn the solution **blue**. This point, when the solution turns *faintly* blue, is the endpoint of the titration.

Summarizing, we titrate with KIO_3. KI and HCl are present in excess and are responsible for the I^- and H^+ concentrations respectively. As long as there is vitamin C present no significant amount of I_2 can form. However, when all of the vitamin C is used up, I_2 concentrations rise and a blue color is observed. This process can be summarized by the following sequential reactions:

a. $\underset{\textit{(buret)}}{\mathbf{IO_3^-}} + \underset{\textit{KI}}{\mathbf{5I^-}} + \underset{\textit{HCl}}{\mathbf{6H^+}} \rightarrow \mathbf{3I_2} + \mathbf{3\,H_2O}$

b. $\underset{\textit{vitamin c}}{\mathbf{C_6O_6H_8}} + \mathbf{I_2} \rightarrow \mathbf{C_6O_6H_6} + \mathbf{2\,H^+} + \mathbf{2I^-}$

c. $\mathbf{I_2} + \mathbf{I^-} \rightarrow \mathbf{I_3^-}$

Experiment: Determining the milligrams of vitamin C in a tablet

You will be working with a partner today but turning in your own, individual lab report.

Because vitamin C degrades over time, manufacturers produce tablets with slightly more than the stated amount of vitamin C in them. This way, if the vitamins are purchased before the expiration date, they should have at least as much vitamin C as is stated on the bottle label.

After the expiration date there is no guarantee the tablet has the stated amount.

Do not be surprised if your tablet results are slightly greater than the bottle label claims.

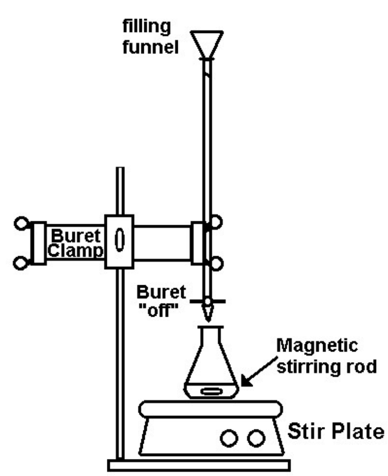
1. Obtain two 100 mg vitamin C tablets. Weigh each tablet and record the tablet masses in the data table.

   Be careful not to mix up the two tablets after you've weighed them.

2. Place each tablet in its own 250 mL Erlenmeyer flask and add approximately 50 mL of distilled water to each flask. Letting them soak for a few minutes makes it easier to break them up later.
3. Rinse (twice) a buret with 2-3 mL portions of KIO_3 solution.
4. Fill the buret and eliminate air bubbles that may have lodged in the valve assembly and tip.
5. Use a glass stirring rod to break up the first tablet and stir the mixture for several minutes.
6. Add approximately 1g KI to the first flask and continue stirring until completely dissolved.
7. Add approximately 5mL of 1 M HCl. Stir well.
8. Add approximately 2-3 mL of 0.5% starch solution. Stir well.
9. Place a magnetic stir bar in the Erlenmeyer flask and position the flask on the stir plate.
10. Start the magnetic stirrer (~400 rpm) No Splashing.

    Do not heat since it will speed up the decomposition of vitamin C.

11. Record the initial burette reading on your data sheet. (Remember 2 decimals)
12. Slowly add the 0.0100 M KIO_3 until a persistent faint blue color is observed.
13. Record the final burette reading on your data sheet and determine how the volume of KIO_3 that was dispensed.



14. Repeat the above procedure a second time using the second 100mg vitamin C tablet.
15. From your tablet results, calculate the mass of vitamin C contained in each tablet in mg. Record your result both on your data sheet and on the whiteboard.

Experiment: Determining the milligrams of vitamin C in a juice sample

1. Measure out approximately ~50 mL of a fruit juice into a clean 250 mL Erlenmeyer flask.

2. Accurately record the amount of juice you dispensed with one decimal place, the type and name of the juice on your data sheet.

3. Record the name of your juice in your data table.

4. Locate the nutrition information on the carton and record the % daily vitamin C on your data table.

5. Add approximately 1g KI to the juice sample and continue stirring until dissolved.

6. Add approximately 5mL of 1 M HCl to the mixture and thoroughly stir.

7. Add approximately 2-3 mL of 0.5% starch solution to the mixture and thoroughly stir.

8. Titrate the mixture as you did above with the KIO_3 solution.

   Note: The expected color change will be disguised by the original color of the juice. Keep a small amount of fresh juice nearby as a reminder of the original juice color. Do the best job you can in determining the endpoint of the titration.

9. Repeat the titration a second time with another portion of the same juice.

Answers to Prelab Questions:

1. 3 I_2 : 1 IO_3^- mole ratio
2. 1 I_2 : 1 $C_6O_6H_8$ mole ratio
3. 1 IO_3^- : 3 $C_6O_6H_8$ mole ratio
4. Because vitamin C decomposes over time. Manufacturers anticipate this and produce tablets with slightly higher than originally claimed vitamin C content.
5. Faint blue
6. I^- and H^+ are in excess (i.e. not limiting reactants)
7. Two
8. To soften them up and make it easier to break them up with a stirring rod.
9. The color change is masked by the color of the juice.
10. It decomposes the vitamin C. Thus, the titration would detect less vitamin C than was present initially.

C1151 Data Sheet
Determination of Vitamin C
Instructor Initials_____

Name__________________________
Date of Exp. _________________
Lab Section _________________

Your <u>individual</u> experimental report will be due at the beginning of class next week.
Data Table: (All entries must be in written in ink before you leave the lab).

Juice information:

Juice ID __________________ **Serving size…** _________**mL** **Claimed % vitamin C** ___________**%**

	Tablet 1	Tablet 2	Tablet 3	Juice 1	Juice 2	Juice 3
Concentration of KIO_3 (M)						
Tablet mass (g)						
Juice volume (mL)						
Initial burette reading (mL)						
Final burette reading (mL)						
mL of KIO_3 *dispensed* (mL)						
moles of IO_3^-						
moles of $C_6O_6H_8$						
Mass of $C_6O_6H_8$ **(mg)**	Round here	Round here	Round here	Round here	Round here	Round here
Male % vitamin C (see below)						
Female % vitamin C(See below)						

A third tablet or juice titration is only necessary if something goes wrong with one of the first two trials.

<u>Questions</u>:

1. In the space below, show how you calculated the mass ascorbic acid (mg) for the two tablets ***beginning with the original buret measurements***. Report your results in the data table rounded to the correct number of significant figures.

2. In the space below, show how you calculated the mass of ascorbic acid (mg) for the two juice trials ***beginning with the original buret measurements***. Report your results in the data table rounded to the correct number of significant figures.

3. Use your two juice volumes and vitamin C masses to determine the mass of vitamin C in a single serving (240 mL) for each titration.

4. The % vitamin C is based on its "recommended daily allowance" (RDA), which is approximately 75 mg for the average adult female and 90 mg for the average adult male. Calculate the % vitamin C based on your results from part 3 for the average male and average female. Report your results in the data table rounded to the correct number of significant figures.

5. How does the % vitamin C value you calculated compare with the advertised values? Are the manufactures advertising their product accurately?

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