Analysis of commercial vitamin c tablet essay sample

Title: Determination of the vitamin C content (ascorbic acid) of a commercial vitamin C tablet and compare the result with the value specified by the manufacturer.

Theory :

In this experiment, we have to determine the vitamin C content of a commercial vitamin C tablet which is the mass of ascorbic acid in the tablet. As iodine is a weak oxidizing agent used mainly for the determination of strong reducing agents, ascorbic acid C6H8O6 (aq) can be rapidly and quantitatively oxidized by iodine in acidic condition as shown in the following equation:

I2(aq) + C6H8O6 (aq) → C6H6O6 (aq) + 2H+ (aq) + 2I- (aq)

This method involves the direct titration of the ascorbic acid with a standard iodine solution in an acidic medium. However, iodine is not very soluble in water (0. 001M), this makes the method less than ideal.

Instead of dissolving solid iodine in water, we can obtain excess quantity of iodine solution by adding an unmeasured excess of potassium iodide solution to a known volume of an acidified standard potassium iodate solution according to the following equation:

IO3-(aq) + 5 I- (aq) + 6 H+ (aq) 3I2 (aq) + 3H2O (l)

Then the experiment is immediately followed by a back titration of the excess liberated iodine with standard sodium thiosulphate solution as shown below:

$I_2 (aq) + 2S_2O_3^{2-} (aq) \rightarrow 2I^- (aq) + S_4O_6^{2-} (aq)$

With iodine solution, the thiosulphate ion is oxidized quantitatively to tetrathionate ion. The amount of iodine solution produced using this method can then be calculated.

The above method have to be repeated for two times. The first time without the ascorbic acid, but the second time with. By finding out the amount of iodine solution produced by this method and the actual amount of iodine solution reacted with the ascorbic acid, and with the help of the first chemical equation above, the mass of the ascorbic acid can then be calculated.

Experimental procedure :

A. Preparation of a standard solution of potassium iodate KIO_3

1. About 0. 65g of potassium iodate KIO_3 is weighed out accurately, then the exact reading is recorded in the result sheet.

2. The solids were dissolved in deionized water and were made up to 250 cm^3 in a volumetric flask.

B. Standardization of thiosulphate solution with iodate solution.

3. The burette is rinsed, and then it is filled with sodium thiosulphate solution.

4. 25. 00 cm^3 of the potassium iodate solution is pipetted into a 250 cm^3 conical flask by using a pipette and a pipette filler.

5. About 5 cm3 of 1M potassium iodide solution was added, then followed by adding about 8 cm3 of 0. 5M sulphuric acid into the conical flask.

6. The initial burette reading is recorded in table1. The resulting solution is immediately titrated with sodium thiosulphate solution.

7. After the reaction mixture had turned pale yellow, about 10 drops of freshly prepared starch solution were added to it as an indicator. The titration is continued until the dark blue colour of the reaction mixture changed to colourless. The final burette reading is recorded in table 1.

8. The titration is repeated for two times. Steps 3 to 6 are repeated.

9. The remaining standard solution of potassium iodate solution is transferred from the volumetric flask to a clean beaker.

C. Preparation of a standard solution of vitamin C tablet

10. One commercial vitamin C tablet is dissolved in about 150 cm3 of 0. 5M sulphuric acid in a beaker.

11. The resulting solution is transferred to a 250 cm3 volumetric flask. Deionized water was added to make up to 250 cm3.

D. Back titration of excess iodine with standard sodium thiosulphate solution

12. 25. 00 cm3 of the vitamin C solution is pipetted into a conical flask by using a pipette and a pipette filler.

13. About 5 cm3 of 1M potassium iodide solution was added, then followed by adding 25. 00 cm3 of the standard potassium iodate solution into the conical flask using a pipette and a pipette filler.

14. The initial burette reading is recorded in table 2. The resulting solution containing excess iodine is immediately back titrated with the standard sodium thiosulphate solution.

15. After the reaction mixture had turned pale brown, about 10 drops of freshly prepared starch solution were added as an indicator. The titration is continued until the reaction mixture turns pale orange. The final burette reading is recorded in table 2.

16. The titration is repeated for two times. Steps 12 to 15 are repeated.

Data Collection and presentation :

Mass of potassium iodate KIO3 used : 0. 65g

Molarity of potassium iodide solution used : 1M

Molarity of sulphuric acid used: 0. 5M

Indicator used: about 10 drops of Starch solution

Table 1(vol of Na2S2O3 solution used in the first titration)

Trial 1 2

Final burette reading (cm3) 25. 95 44. 65 27. 50

Initial burette reading (cm3) 7. 10 25. 95 8. 85

Volume of Na2S2O3 solution added (cm3) 18. 85 18. 70 18. 65

Mean volume of Na2S2O3 solution added (cm3) (18. 70 + 18. 65) ÷ 2 = 18. 675 cm3

Table 2 (vol of Na2S2O3 solution used in the second titration)

Trial 1 2

Final burette reading (cm3) 12. 60 19. 65 26. 70

Initial burette reading (cm3) 5. 40 12. 65 19. 70

Volume of Na2S2O3 solution added (cm3) 7. 20 7. 00 7. 00

Mean volume of Na2S2O3 solution added (cm3) (7. 00 + 7. 00) ÷ 2 = 7. 00 cm3

Observation (colour change)

1st titration: brown à pale yellow à dark blue à colourless

2nd titration: brown à pale brown à dark blue à orange

(after adding iodide solution) (during titration) (with starch solution) (after the end point is reached)

Calculations :

1. IO3-(aq) + 5 I- (aq) + 6 H+ (aq) 3I2 (aq) + 3H2O (l)

2. $I_2 (aq) + 2S_2O_3^{2-} (aq) \rightarrow 2I^- (aq) + S_4O_6^{2-} (aq)$

3. $I_2(aq) + C_6H_8O_6 (aq) \rightarrow C_6H_6O_6 (aq) + 2H^+ (aq) + 2I^- (aq)$

From the results obtained, and the three equations of the reactions, we have:

For the first titration:

No. of moles of $IO_3^-(aq)$ in 25 cm^3 of the standard solution

= 0. 65g / (39. 1 + 127 + 16×3) g/ 10 = 0. 000304 mol

No. of moles of I_2 (aq) produced = 0. 000304 × 3 = 0. 000911 mol

From equation (2), no. of moles of $S_2O_3^{2-}$ (aq) = 2 no. of moles of I_2 (aq)

= 0. 000911 × 2 = 0. 00182 mol

∴ Molarity of $Na_2S_2O_3$ = 0. 00182 / 0. 0187 = 0. 0974 M

For the second titration:

No. of moles of $Na_2S_2O_3$ reacted = 0. 0974 × 0. 007 = 0. 000682 mol

No. of moles of I_2 (aq) = 1/2 No. of moles of $Na_2S_2O_3$ reacted = 0. 000341 mol

∴ no. of moles of $I_2(aq)$ reacted with ascorbic acid = 0. 000911 – 0. 000341 = 0. 000569 mol

no. of moles of $C_6H_8O_6$ (aq) = no. of moles of $I_2(aq)$ = 0. 000569 mol

mass of $C_6H_8O_6$ (aq) in 25 cm^3 of standard solution = 0. 000569 × (12 ×6 + 8 ×1 + 16 ×6) = 0. 1003 g

mass of vitamin C in 1 tablet = 0. 1003 × 10 = 1. 003 g = 1003 mg

Answers to the Questions to be discussed:

Q. 1 The mass of ascorbic acid per tablet is 1003mg as shown on the above.

Q. 2 (a). The starch solution is used as an indicator.

It gives a dark blue colour to the solution when excess oxidizing agent is present, in which the concentration ratio of iodine to iodide is high. And the dark blue colour disappears when there is excess reducing agent, that is when there is a higher concentration of iodide ions.

During the titrations of iodine with sodium thiosulphate, the amount of iodine is continuously decreasing, relatively, the amount of iodide ions is increasing, as a result, the dark blue colour is gradually diappeared. At the end-point, all the iodine has reacted with sodium thiosulphate, the colour changes sharply from dark blue to colourless. Thus, starch solution can serve as an indicator.

(b). The starch solution have to be added only when the reaction mixture becomes pale yellow since irreversible decomposition of starch will occur when starch is added to solutions with very high concentrations of iodine. When the colour of the reaction mixture becomes pale yellow, it indicates that the titration is nearly complete, the concentrations of iodine will not be

as high as that at the beginning of the titration. As a result, decomposition can be avoided and the colour change can be observed more clearly.

Q. 3 The exact volumes of potassim iodide solution and sulphuric acid used in the titration are not important. Although they are the reactants, both of them are not involved in the formation of the product I2 (aq) we want. Therefore, in the calculation of the no. of moles of I2 (aq), they are not to be included in the calculation. However, they are needed for the reaction to occur in the following ways.

For the sulphuric acid, it is used to acidify the mixture to let the reaction IO3- (aq) + 5 I- (aq) + 6 H+ (aq) 3I2 (aq) + 3H2O (l) to occur instantaneously.

For the potassium iodide solution, it is used as a solvent as iodine is dissolved in a small volume of moderately concentrated potassium iodide soluiton. Iodine is reasonably soluble in this medium as the following reaction: I2 (s) + I- I3-. As a result, their exact volumes are not important.

Q. 4 The titration has to be carried out immediately after the addition of sulphuric acid as the sulphuric acid will acidify the mixture and casue the reaction:

IO3-(aq) + 5 I- (aq) + 6 H+ (aq) 3I2 (aq) + 3H2O (l) to occur instantaneously.

I2 (aq) is then rapidly produced which lacks stability due to the volatility of the solute.

Iodine can loss from an open vessel in a relatively short time. For this reason, titration has to be carried out immediately to prevent any side reactions and loss of iodine from happening.

Q. 5 Firstly, rinse the beaker with deionized water. Secondly, remove the stopper and pour a few cm3 of the standard solution from the volumetric flask to the beaker for rinsing. After rinsing, use a stirring rod to direct the flow of liquid from the volumetric flask to the beaker slowly. The KIO3 solution is then transferred to a beaker.

Q. 6 To investigate that ascorbic acid (vitamin C) deteriorates on heating and on exposure to the air, we can set up two experiments as the same as the one done on the above. However, there will be differences on the two commercial vitamin C tablets.

Both of the tablets will be first dissolved in the sulphuric acid in a beaker reaspectively. Then one of the resulting solution will be left on exposure to the air for 24 hours. And the other one will be heated in the water bath for 5-10 minutes. To do this, we have to set up a Bunsen burner with a tripod and wire gauze to hold the beaker on it. Afterwards, the two beakers of solution will be used to replace the vitamin C solution prepared by dissolving the vitamin C tablet directly in the sulphuric acid. The experiment will then be the same as the above.

And the result will show that the amount of iodine reacted with the ascorbic acid will be almost zero, the volume of thiosulphate solution recorded for the first tiration and the second titration would be almost the same for the two

experiments. This indicates that the ascorbic acid deteriorates, so that no ascorbic acid can be reacted with iodine solution.

Overall Comment to the Experiment:

1. Uncertainty / Error Analysis :

1. The Na2S2O3 in the burette turned milky while the last titration was undergoing. This causes errors to the experimental result. It is due to the following chemical reaction:

$S_2O_3^{2-}$ (aq) + H^+ (aq) = HSO_3^- (aq) + S (s)

Although sodium thiosulphate solution is resistant to air oxidation, they tend to decompose to give sulphur and hydrogen sulphate ion. The solution turned to milky due to the suspension of the solid sulphur. However, this process occurs at a very low rate. The rate of the decomposition reaction increases significantly when the solution becomes acidic and when there are bacteria.

In the experiment, some sulphuric acid may have been wrongly mixed with the sodium thiosulphate solution in the burette to give rise to this error or there may be some bacteria.

2. When the vitamin C tablet was added to the sulphuric acid, there was effervescence. The tablet floated on the surface and moved around with fizzy sound. Some of the resulting solution might have jumped out of the beaker due to the effervescence and this may cause an error that some of the mass of the ascorbic acid may have lost away.

3. The colour change at the end point for the second titration with the ascorbic acid present is very difficult to observe because the colour changed from brown to very pale brown, and then back to the original colour orange which is due to the present of the pigment of vitamin C. Excess $Na_2S_2O_3$ (aq) solution may have been added to the reaction mixture due to this reason and this causes an error.

4. Iodine solution lacks stability due to the volatility of the solute. Iodine solution can loss from an open vessel in a relatively short time. Consequently, the amount of iodine solution produced and the amount of iodine solution reacted can not be known accurately enough with some loss of iodine solution. This error may casue deviation to the answers.

5. The iodide ions present in the solution can undergo air oxidation which can cause changes in the molarity of an iodine solution:

$4I^- + O_2 (g) + 4H^+ \rightarrow 2I_2 + 2 H_2O$

This reaction causes the molarity of the iodine to increase.

2. Discussion and Conclusion (ways to minimize errors) :

Conclusion:

The value of the mass of ascorbic acid calculated from the experiment is slightly greater than the value specified by the manufacturer. It is probably due to the errors stated above.

To improve the experiments, errors should be minimized as shown below:

1. To minimize the first problem stated above, standard solution of sodium thiosulphate should be prepared under reasonably sterile conditions to prevent bacterial activity. Moreover, cleaning of the apparatus is critical to prevent any acids from mixing up with the thiosulphate soltuion.

2. To prevent jumping out of solution from the beaker during dissolving the tablet in the sulphuric acid, we can use a larger beaker or a conical flask.

3. To prevent air oxidation of iodide ion and the loss of iodine solution, the titrations have to be done very quickly and immediately after . This can minimize the uncertainty.

Modification of the experiment:

In the experiment, potassium iodate solution is used to standardize a thiosulphate solution as iodine solution is produced for titration with sodium thiosulphate.

Instead of using potassium iodate solution, we can use iodine solution directly for titration. Although the iodine has low solubility, it can still be prepared by dissolving iodine in a small volume of concentrated solution of potassium iodide. To ensure complete solution, care should be taken to avoid dilution of the concentrated solution until the last piece of solid iodine has disappeared. This can be done by filtering the solution through a sintered glass crucible before standardization.