DETERMINATION OF VITAMIN C CONTENT IN TABLETS

INTRODUCTION

The active ingredient in tablets can be mixed with fillers, binders, flavours and colours. This makes them more attractive to consumers and allows the required dose to be given in a tablet that is reasonably sized and does not crumble in the container. In this experiment we will use a redox titration to determine the amount of ascorbic acid, or Vitamin C, in a commercial Vitamin C tablet.

Ascorbic acid has complex structure that can be represented $C_6H_4O_2(OH)_4$. The solid is quite stable, but in solution it is easily oxidised by atmospheric oxygen and the Vitamin C content drops slowly.

The half equations are:

$$C_6H_4O_2(OH)_4(aq) \rightarrow C_6H_4O_4(OH)_2(aq) + 2H^+(aq) + 2e^-$$
$$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$$

Since iodine reacts readily with ascorbic acid, the amount of ascorbic acid in a solution can be determined by titrating it against a standard iodine solution.

The half equation for the reaction of iodine is:

$$I_2(aq) + 2e^- \rightarrow 2I^-(aq)$$

However, iodine is quite volatile and it is best to standardise a solution immediately before use.

It can be titrated against sodium thiosulphate, which reacts with the half equation:

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

AIM

To calculate the mass of Vitamin C in commercial Vitamin C tablets.

MATERIALS

- 3 x 300 ml conical flasks
- 20.0 ml pipette
- 100 ml measuring cylinder
- 10 ml measuring cylinder
- Burette and stand
- Wash bottle
- White tile
- Stirring rod

METHOD AND SAMPLE RESULTS

1. Record the manufacturer's name, product name and the mass of ascorbic acid as stated on the label.

   For Example:

   A typical label could state that the product contains a mixture of sodium ascorbate, calcium ascorbate and ascorbic acid, and provide 'the equivalent of 1000 mg of Vitamin C'.

2. Using the half equations provided, derive an equation for the reaction of iodine with thiosulphate ions in aqueous solution

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

   Overall equation:

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

   Using the half equations provided, derive an equation for the oxidation of ascorbic acid by iodine

$$C_6H_4O_2(OH)_4(aq) \rightarrow C_6H_4O_4(OH)_2(aq) + 2H^+(aq) + 2e^-$$
$$I_2(aq) + 2e^- \rightarrow 2I^-(aq)$$

   Overall equation:

$$C_6H_4O_2(OH)_4(aq) + I_2(aq) \rightarrow C_6H_4O_4(OH)_2(aq) + 2H^+(aq) + 2I^-(aq)$$

3. Record the precise concentration of the standard sodium thiosulphate solution.

   For Example: Concentration of $Na_2S_2O_3$ = 0.382 M

4. Rinse and fill a burette with the iodine solution provided, and record the initial reading of the burette.

5. Use a pipette to transfer 20.00 ml of the sodium thiosulphate solution into a conical flask and add 2 drops of starch indicator solution. Titrate the contents of the flask with the iodine solution until the first permanent blue colour appears in the flask. Record the Final reading of the burette and calculate the volume of the titre.

6. Repeat steps 4 and 5 until three concordant titres have been obtained. Calculate the average volume of iodine from these results.

   For Example: Average titre = 19.16 ml

7. Calculate the concentration of the iodine solution provided.

i. $Na_2S_2O_3(aq) \rightarrow 2Na^+(aq) + S_4O_6{}^{2-}(aq)$

$$n(S_4O_6{}^{2-}) = n(Na_2S_2O_3)$$
$$\therefore C(S_4O_6{}^{2-}) = C(Na_2S_2O_3) = 0.382M$$

ii. $n(S_4O_6{}^{2-}) = C \times V$

$$= 0.382 \times 0.02000 = 0.00764mol$$

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

$$n(I_2) \quad = \frac{1}{2} \times n(S_4O_6{}^{2-})$$
$$= \frac{1}{2} \times n(S_4O_6{}^{2-}) = \frac{1}{2} \times 0.00764 = 0.00382mol$$

iv. $C(I_2) \quad = \frac{n}{V}$

$$= \frac{0.00382}{0.01916} = 0.199M$$

8. Dissolve one tablet in about 50 ml of distilled water in a conical flask. If necessary, crush the tablet with a glass stirring-rod and stir the solution to help the tablet dissolve. Add about 1 ml of starch indicator solution.

9. Re-fill the burette with the standard iodine solution, and record the initial reading of the burette. Titrate the ascorbic acid solution with the iodine until the first permanent blue colour appears through the solution. Record the final reading of the burette and calculate the volume of iodine solution used during the titration of ascorbic acid solution.

   For Example: Titre = 28.16 ml

10. Calculate the mass of ascorbic acid in the tablet

i. $n(I_2) \quad = C \times V$

$$= 0.199 \times 0.02816 = 0.00561mol$$

ii. $C_6H_4O_2(OH)_4(aq) + I_2(aq) \rightarrow C_6H_4O_4(OH)_2(aq) + 2H^+(aq) + 2I^-(aq)$

$$C_6H_4O_2(OH)_4 \qquad = n(I_2)$$
$$= 0.00561mol$$

iii. $mass(C_6H_4O_2(OH)_4) = n \times M$

$$= 0.00561 \times 176 = 0.988g$$

11. Repeat the procedure for a further two Vitamin C tablets and record the mass of ascorbic acid in each of the tablets.

    For Example:

Tablet 1	Tablet 1	Tablet 3
0.988 g	0.985 g	0.991 g

12. Calculate the average mass of ascorbic acid in one tablet and compare your result with the manufacturer's claims, as stated on the label.

$$average\, mass\ of\ ascorbic\, acid = \frac{0.988 + 0.985 + 0.991}{3} = 0.988g = 988mg$$

The average mass of ascorbic acid in one tablet was found to be 988 mg, which is less than the stated value of 1000 mg per tablet.

DISCUSSION QUESTIONS

QUESTION 1
If the average mass of one of these Vitamin C tablets is 1.12 g, calculate the average percentage mass of ascorbic acid in the tablets.

Solution

QUESTION 2
Explain why iodine is generally not favoured as a primary standard even when it is available in a very pure state.

Solution

QUESTION 3
How would the results have been affected if the iodine were not standardised immediately before each of the titrations?

Solution

QUESTION 4
Use the information provided to write an equation for the oxidation of ascorbic acid in air

Solution

QUESTION 5
It has been suggested that this method could also be used to determine the rate of oxidation of ascorbic acid in vitamin enriched fruit juices. Briefly explain what tests would be necessary and identify one reason why this may **NOT** give satisfactory results.

Solution

CONCLUSION

The average mass of ascorbic acid found in these tablets was found to be 988 mg, which is slightly less than the manufacturer's claim of 1000 mg.

ANSWERS TO DISCUSSION QUESTIONS

QUESTION 1

$$\%(ascorbic\,acid) = \frac{mass\,of\;ascorbic\,acid\,(g)}{mass\,of\;tablet\,(g)}\times 100$$

$$= \frac{0.988}{1.12}\times 100 \quad = 88.2\,\%(w/w)$$

QUESTION 2

Iodine is volatile and the concentration therefore tends to decrease with time.

QUESTION 3

Iodine is volatile and so the concentration decreases over time. Consequently:

- A larger volume of iodine would have been required during the titration.
- The calculated amount, in mole, of iodine would be greater.
- Due to the use of mole ratios, the calculated amount of ascorbic acid would be greater.
- The calculated mass of ascorbic acid would be greater than the true figure.

QUESTION 4

$$2C_6H_4O_2(OH)_4(aq)\rightarrow 2C_6H_4O_4(OH)_2(aq)+4H^+(aq)+4e^-$$
$$O_2(g)+4H^+(aq)+4e^-\rightarrow 2H_2O(l)$$

Overall equation:

$$2C_6H_4O_2(OH)_4(aq)+O_2(g)\rightarrow 2C_6H_4O_4(OH)_2(aq)+2H_2O(l)$$

QUESTION 5

A sample of the fruit juice would need to be tested at several time intervals to show whether the amount of ascorbic acid changed with time. For example, a sample might be tested at the same time each day for seven days, with the juice stored in a refrigerator between titrations.

However, the iodine might react with other components of the juice to give misleading results. And the concentration of the other components might also change over time to further affect results.