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Redox Titration

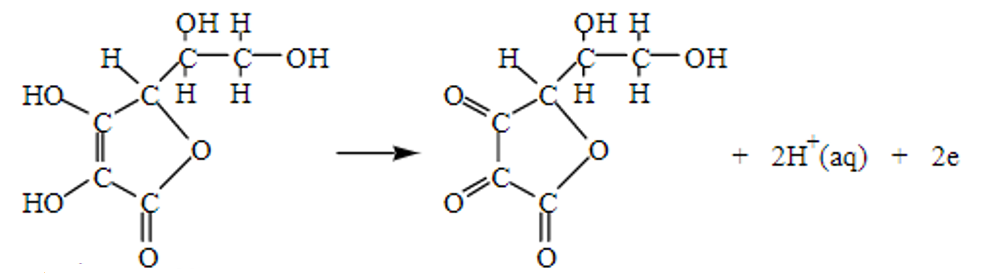
UNIT 3 PPA 3

**Introduction**

Vitamin C (ascorbic acid) is an important component of our diet. In its absence the protein collagen cannot form fibres properly and this results in a disease called Scurvey, causing skin lesions and blood vessel fragility.

Although vitamin C occurs naturally in many fruits and vegetables, many people take vitamin C tablets to supplement their intake

Vitamin C can undergo a redox reaction with iodine in which the vitamin C is oxidised.



The iodine molecules are reduced ( I2(aq) + 2e-  🡪 2I-(aq)  ) but this is not shown in the diagram).

The aim of this experiment is to determine the mass of vitamin C in a tablet by carrying out a redox titration using a solution of iodine of accurately known concentration and starch solution as an indicator.

You will need

|  |  |
| --- | --- |
| small beaker | 50 cm3 burette |
| pipette filler | 250 cm3 standard flask |
| 25 cm3 pipette | filter funnel |
| 50 cm3 burette | wash bottle |
| conical flask | pipette filler |
| standard solution of iodine | starch solution |
| vitamin C tablet | deionised water |

**Health & Safety**

The iodine solution irritates the eyes.

Wear eye protection and wash your hands if any iodine solution spills on them.

**Method**

1. Add a vitamin C tablet to the beaker.
2. Add some deionised water (approximately 50 cm3) to the beaker and stir the mixture until the tablet has dissolved.
3. Carefully add the resulting solution to the 250 cm3 standard flask. Rinse out the beaker several times with water and add the washings to the flask.
4. Add water to the standard flask to bring the volume of the solution up to the graduation mark on the neck.
5. Stopper the flask and invert it several times to make sure the solution is thoroughly mixed.
6. After rinsing the pipette with a little of the vitamin C solution: pipette 25 cm3 of it into the conical flask.
7. Add a few drops of starch solution to the vitamin C solution in the conical flask.
8. After rinsing the burette with a little iodine solution: fill the burette with the iodine solution.
9. Note the initial burette reading. Since the solution has a dark colour: it is difficult to see the bottom of the meniscus. Take the burette reading from the top of the meniscus.
10. Add the iodine solution slowly from the burette whilst gently swirling the solution in the conical flask. Initially you will see a blue/black colour as the iodine reacts with the starch but this will rapidly disappear as the iodine reacts with the vitamin C.
11. Near the end-point of the titration the colour disappears more slowly. At this point add the iodine solution drop by drop until the solution just turns a blue 'black colour and remains so.
12. This is the end-point of the titration i.e. all the vitamin C has reacted. Note the final burette reading.
13. Wash out the conical flask
14. Repeat the titrations until concordant results are obtained.

**Calculation**

1. Knowing the average volume and concentration of the iodine solution used in the redox titration, the number of moles of iodine can be calculated.
2. With the result from step (a) and the balanced equation for the redox reaction, we can work out the number of moles of vitamin C in 25 cm3 of the vitamin C solution. This can be scaled up to find the number of moles of vitamin C in 250 cm3 of the vitamin C solution.
3. Your final answer in step (b) will, of course: be equal to the number of moles of vitamin C in the tablet. Using this result and the mass of one mole of vitamin C (176 g) we can finally work out the mass of vitamin C in the tablet.

**Example calculation**

Suppose the average titre volume was 22.1 cm3 and the iodine solution had a concentration of 0.0250 mol l-1.

From the average titre volume in litres (V) and the concentration of the iodine solution (C), we can calculate the number of moles of iodine (n) used in the titration:

N(iodine) = C x V = 0.0250 x 0.0221

= 5.525 x 10-4 mol

We can now use the balanced redox equation to calculate the number of moles of vitamin C in a 25 cm3 sample of the vitamin C solution:

C6H8O6 + I2 🡪 C6H6O6 + 2H+ + 2I-

1 mol 1 mol

5.525 x 10-4 mol 5.525 x 10-4 mol

But there were 250 cm3 of vitamin C in total and so to determine the number of moles of vitamin C in the tablet we have to scale up our last answer:

N(vitamin C) per tablet = 10 x 5.525 x 10-4

= 5.525 x 10-3 mol

Vitamin C: C6H8O6

Mass of 1 mole = 6(12) + 8(1) + 6(16) = 176g

We can now calculate the mass of vitamin C per tablet:

Mass of vitamin C per tablet = 176 x 5.525 x 10-3

= 0.972g

**Technician Guide**

You will need

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| --- | --- |
| 1 small beaker | 1 50 cm3 burette |
| 1 pipette filler | 1 250 cm3 standard flask |
| 1 25 cm3 pipette | 1 filter funnel |
| 1 50 cm3 burette | 1 wash bottle |
| 1 conical flask | 1 pipette filler |
| 1 white tile |  |
| ~ 75 cm3 standard solution of iodine 0.025 mol l-1 \* | 1% starch solution (a few cm3 as an indicator) |
| vitamin C tablet 1g \*\* | ~ 250 cm3 deionised water |

\* 6.35g Iodine and 20g Potassium Iodide per litre

The iodine solution will need to be standardised. This can be done against a standard solution of sodium thiosulphate.

Alternatively, the iodine solution could be prepared from a commercial volumetric standard solution of iodine.

\*\* Both lemon and orange-flavoured effvescent vitamin C tablets are suitable. Despite the bright orange colour Of the latter the end-point of the titration is distinct.

**Health & Safety**

Check the risk assessment for this activity.

Solid iodine is harmful is swallowed and in contact with the skin and eyes and gives off irritating fumes. Work in a well-ventilated area.

Wear eye protection and wash your hands if any iodine solution spills on them.

**Method**

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