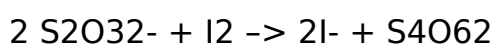
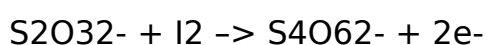
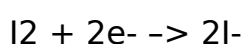


Titration to determine the concentration of an iodine solution



The volume of iodine solution needed to neutralize the sodium thiosulphate. You simply titrate the sodium thiosulphate against the solution of iodine using starch solution as an indicator. Then when you have calculated the concentration of the sodium thiosulphate can use it in turn as a standard solution, to find the concentration of the other substance. A redox reaction has occurred.



Safety

The iodine is strongly oxidising. At all times through the experiment you must wear eye protection and lab coats. When filling the pipette be careful it does not break. And when filling the burette make sure it is below eye level. If any spillages have been made clean them up and report it.

Apparatus

50 cm³ burette, 20 cm³ pipettes, dropping pipette, tile, 250 cm³ measuring beaker, and 6.12 grams of Na₂S₂O₃ · 5H₂O, 250 cm³ distilled water, clamp stand filter funnel, 250 cm³ conical flask, and starch solution.

Method

The solution of sodium thiosulphate was made by dissolving 6.12g of Na₂S₂O₃ · 5H₂O with 250 cm³ of distilled water. The sodium thiosulphate

has a concentration of 0.1 mol. dm³. Set up the apparatus. Wash equipment with solution three times. The burette with the sodium thiosulphate, and the conical flask and the pipette with the iodine solution. Measure 25 cm³ of the standard solution into a conical flask, using a pipette. Add the starch as the indicator. Pour the sodium thiosulphate into a 50 cm³ burette. Pipette 25 cm³ portion of the iodine solution into a conical flask. record the level. Drip the solution slowly into the conical flask, swirling the flask continuously. Rinse the wall of the reactor before the end point. Stop adding acid when a single drop turns the solution deep red to yellow. One drop accuracy. Record the new level of sodium thiosulphate in the burette. Calculate the volume of acid used. Carry out four titrations, one trial and three accurate. Give your answer to 2 decimal places.

Results

Titration

Initial reading

Final reading

volume used cm³

1

0.00

29.60

29.60

2

0.00

29.65

29.65

3

0.00

29.60

29.60

4

0.00

29.70

29.70

Calculations

Average volume of potassium manganate v11 used, from three accurate titrations:

$$29.65 + 29.60 + 29.70 = 29.65 \text{ cm}^3$$

Weighing board = 0.97

Altogether = 7.09

$\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$

$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$ reduction

$\text{S}_2\text{O}_3^{2-} + \text{I}_2 \rightarrow \text{S}_4\text{O}_6^{2-} + 2\text{e}^-$ oxidation

$2 \text{S}_2\text{O}_3^{2-} + \text{I}_2 \rightarrow 2\text{I}^- + \text{S}_4\text{O}_6^{2-}$

Mass 6.12

Moles * Mr $0.024 * 248$

Moles 0.024

Molarity * vol/dm³ ? * 250/1000

= 0.096

= 0.01

29.65 0.1 mol. dm³ $\text{S}_2\text{O}_3^{2-}$ = 25.00 cm³ I_2

$2 \text{S}_2\text{O}_3^{2-} + \text{I}_2 \rightarrow 2\text{I}^- + \text{S}_4\text{O}_6^{2-}$

1 mole $\text{S}_2\text{O}_3^{2-}$ = $2\text{I}^- + \text{S}_4\text{O}_6^{2-}$

$\text{S}_2\text{O}_3^{2-}$ =

$0.1 * 29.65/1000 = 2.9 * 10^{-3}$ moles

$2.9 * 10^{-3}$ moles = $1.48 * 10^{-3}$ moles

$$\text{conc} = 1.48 \times 10^{-3} \text{ moles}$$

$$25.00/1000 = 0.025 \text{ mol. dm}^{-3}$$

Discussion

The end point might be different due to human error, they might have added a couple of extra drops or they might not have added enough drops. Or they might have lost some of the solution. Another factor instrument error which can be built in, due to it will not give a precise point. The pipette could have been a different scale e. g. 10, 15, 25 cm³ etc. which might have had some sodium hydroxide left in its tip. The reading could not be exactly clear due human usage of the instrument time after time. Precision accuracy. The numbers could have been blurred so you can not read it to. The meniscus of the solutions could be out a little. We could have lost some of the solution when we were transferring it from the pipette to the conical flask. It could have remained on the side of the pipette and the conical flask. We could not collect all the solutions no matter how hard we tried. Due to equipment error.

$$\text{Instrument accuracy/ lowest reading} \times 100 = \text{ans}$$

$$0.05/28.40 \times 100 = 0.17\%$$

For the burette.

This shows that the equipment are not absolute accurate. Another factor could have been the air around us. Or standing near the window and the outside temperature could have its effect.

<https://assignbuster.com/titration-to-determine-the-concentration-of-an-iodine-solution/>

Other instruments could be used which have a better accuracy point. More sophisticated measuring equipment could also be another point to notice.

Conclusion

The molarity of the iodine solution was 0.0593 mol dm⁻³. The results were close together and the indicator changed quickly at the end point. If you were to repeat this experiment you could make it better by using a different indicator instead of starch, or using a different concentration of sodium thiosulphate.