

Titration part 2

Iodine/thiosulfate

The iodine/thiosulfate titration is very common and gives a distinctive **blue-black to colourless** colour change at the end point.

This is a **redox** titration. Iodine is reduced and the thiosulfate ($\text{S}_2\text{O}_3^{2-}$) is oxidised. The equation is always:



- ✓ learn this equation. It is going to be there in every single iodine/thiosulfate question. Sometimes they give you it, sometimes they don't.

Starch is used as the indicator. It forms a blue/black colour with iodine. When the iodine has all reacted (the end point), the solution turns colourless as there is no iodine left.

- ✓ starch is only added towards the end of the titration as it can form an insoluble complex with high concentrations of iodine. So they wait until the I_2 concentration has gone down a bit.

To do this titration, iodine must be generated, which can be done in two ways:

1. Carry out a reaction where **iodine is the product** (usually by adding KI as a reactant). We can then titrate the iodine produced against the thiosulfate.
 2. Use **iodine as a reactant in excess**, and titrate the unreacted iodine.
- ✓ we looked at a similar method to point 2 above in the [AS titrations](#) by adding HCl in excess.

Type 1

In this example we want to work out the concentration of Cl_2 . The standard method to generate iodine is to oxidise iodide. Potassium iodide is added, which is oxidised by the $\text{Cl}_2 \rightarrow \text{I}_2$:



As was mentioned above in point 1, they deliberately try to generate I_2 , so it can then be titrated against the thiosulfate. This means there has to be a second equation involving iodine and thiosulfate:



Calculate the concentration of 10 cm³ of Cl₂. 9.05 cm³ of 0.01 mol dm⁻³ thiosulfate was added during the titration:

Step 1: moles of thiosulfate

$$n = 0.01 \times 9.05/1000 = \mathbf{9.05 \times 10^{-5} \text{ moles}} \text{ of thiosulfate (S}_2\text{O}_3^{2-}\text{)}.$$

Step 2: moles of I₂

✓ these first two steps almost always happen with these titrations.

Using the equation:



$$\text{Moles of I}_2 = 9.05 \times 10^{-5}/2 = \mathbf{4.525 \times 10^{-5} \text{ moles}}$$

Step 3: moles of Cl₂

Now that we have the I₂ moles, we can get the Cl₂ moles using the ratio of I₂:Cl₂:



$$1:1 \text{ ratio, therefore moles of Cl}_2 = \mathbf{4.525 \times 10^{-5} \text{ moles}}$$

✓ I₂ is always the “link” between the two equations

Step 4: answer the question

In this question, they are looking for a concentration, so you just use $n = c \times v$ again. The volume in the question was 10 cm³ of Cl₂:

$$c = n/v \rightarrow c = 4.525 \times 10^{-5}/0.01 \text{ (10/1000)}$$

$$= \mathbf{4.525 \times 10^{-3} \text{ mol dm}^{-3}}$$

✓ students often get confused which volume to use at this stage. It is always the **first** volume in the question. The original amount before anything is done to it.

Type 2

Ideally we would react the “unknown” with KI as above to generate the I₂. But in the example below we cannot oxidise I⁻ with Sn²⁺.

Instead we add I₂ as a **reactant**, which will oxidise the Sn²⁺. As I₂ is a reactant, we need to add an excess to ensure that there is some left to titrate with the thiosulfate.

Calculate the percentage of Sn (tin) existing as SnO₂, in a rock sample. The rock, 10.25 g, was dissolved in sulphuric acid. A reducing agent was added to convert Sn⁴⁺ ions to Sn²⁺ ions.

50cm³ of 0.25 moldm⁻³ I₂ solution was then added to the Sn²⁺ ions:



The excess I₂ was titrated against 0.1 moldm⁻³ thiosulfate. 11.6 cm³ of thiosulfate was added during the titration.

- ✓ a big problem with these questions is confusion and too much information. There are numbers everywhere! What the hell is going on! Try to keep it simple. They are all just simple redox reactions.

think of all the “extra” steps as preparation to get the solution into a form that can be titrated

For example, dissolving in sulphuric acid then reducing the ions etc. Just ignore it. It doesn't matter, it's just done to get the thing into solution ready for the titration. Just focus on the numbers, equations and **steps** that you know.

Step 1: moles of thiosulfate

$$\text{thiosulfate moles} = 11.6/1000 \times 0.1 = \mathbf{1.16 \times 10^{-3} \text{ moles}}$$

Step 2: moles of **unreacted** iodine

They gave the thiosulfate equation in the question:



Again divide by 2 → I₂ moles = **5.8 x 10⁻⁴ moles** (this is the unreacted I₂)

- ✓ remember this is the number of moles of I₂ left over (the excess) from the initial reaction of I₂ with Sn²⁺.

Step 3: moles of I₂ **reacted** with Sn²⁺

We need to know the original moles of I₂ using 50cm³ of 0.25 moldm⁻³ I₂.

$$\text{Original moles} = 50/1000 \times 0.25 = \mathbf{0.0125 \text{ moles}}$$

$$\text{Reacted moles} = 0.0125 - 5.8 \times 10^{-4} = \mathbf{0.0119 \text{ moles}}$$

Step 4: answer the question.

They want to know the mass of tin, so we do:

$$0.0119 \times 119 = \mathbf{1.42 \text{ g}}$$

Finally they want it as a percentage of the original 10.25 g rock sample:

$$1.42/10.25 \times 100 = \mathbf{13.8\%}$$

Exam advice

- ✓ You must be able to tell quickly which “type” of titration they are doing. Is it “normal” or is something in excess etc.
- ✓ Write out equations and put them one on top of the other (if there are multiple equations). It is much easier to see what is going on when you do this. Sometimes the equations aren’t given or are all over the place. Try to give it some sort of organisation.
- ✓ No matter how difficult it seems, always try to do the usual steps:
 - Calculate moles of the species of known concentration
 - Use the ratios to calculate the moles of the unknown
 - Check if you need to multiply x10 for example ($25 \text{ cm}^3 \rightarrow 250 \text{ cm}^3$) if they took a portion out to do the titration
 - Finish the question...they often get you to calculate percentage purity, calculate x in $x\text{H}_2\text{O}$, concentration or the M_r of a compound. All of which involve converting moles into something else.