

Writing Half Equations

Introduction

At GCSE and in Y12 A level chemistry, you have learned how to write simple half-equations (electron equations) showing oxidation and reduction of simple species.

For example, the **reduction** of aluminium ions and the **oxidation** of oxide ions in the electrolysis of alumina to form liquid aluminium and oxygen (before it reacts with the carbon anodes).



These are relatively trivial to write, once you understand them.

However, in Y13, most of you will need to write half-equations that are quite a bit more involved.

For example, you may need to write an equation for the **reduction** of manganate (VII) ions (MnO_4^-) to manganese (II) ions (Mn^{2+}).

Here are some rules that will help you solve these equations, simply and reliably.

You can use them with more difficult examples such as the **oxidation** of ethanol to ethanal.

Let's use the **reduction** of MnO_4^- to Mn^{2+} as the example.

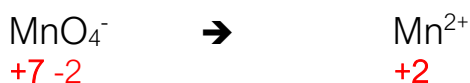
The first thing to do is to write down the species that are being oxidised or reduced and what these become, after the reaction. In this case



Now we need to fill in the rest!

Step 1

Work out the oxidation states (numbers) of the species involved. If you are unsure how to do this you can check out the guide to doing this in **TOOLS** on *CramNow*



Step 2

Work out the number of electrons that are involved and add them to the left or right hand side of the equation.



Step 3

Check to see if there is an imbalance of **oxygen** atoms. There are obviously 4 oxygens on the left but no oxygens on the right. We need to balance this by adding oxygen to the right-hand side. You may think that this could be solved by adding 2O_2 to the right. HOWEVER! If you added O_2 molecules, you would have changed the oxidation state of oxygen (from 0 to -2). You must remember that in a half-equation, there is only one element that undergoes a change in oxidation state, and it is manganese! So, oxygen must appear on the right with an oxidation state of -2. This is done by **balancing oxygens using water molecules**.



Step 4

Check to see if there is an imbalance of **hydrogen** atoms. There are obviously 8 **hydrogens** on the right but no **hydrogens** on the left. We need to balance these by adding **hydrogen** to the left-hand side. Again, you may think that this could be solved by adding 4H_2 to the left. BUT, If you added H_2 molecules, you would have changed the oxidation state of **hydrogen**.

You must remember that in a half-equation, there is only one element that undergoes a change in oxidation state, and it is manganese!

So, **hydrogen** must appear on the left with an oxidation state of +1. This is done by balancing **hydrogen** using hydrogen ions, H^+ .



I personally believe that this method is preferable to others that you may come across. I think that there are two reasons.

Firstly, it applies oxidation states to solve the problem. It would seem crazy to learn about oxidation numbers without applying them in this ultimately practical situation.

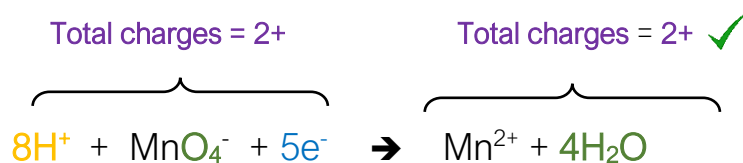
Secondly, this method leaves you with a rare opportunity to know whether you have solved it correctly! See step 5!

Step 5

The final thing that you should do it to **CHECK THE CHARGES**.

In any equation in chemistry (full equations, half-equations, ionic equations, steps in an organic mechanisms),

ALWAYS REMEMBER THAT THE TOTAL CHARGES ON BOTH SIDES MUST BE THE SAME!



If the charges balance, you have almost certainly solved it correctly. If you have made either a chemical or careless error during steps 1-4, this will tell you so that you can go back and check!

Step 1: oxidation states

Step 2: electrons

Step 3: oxygens

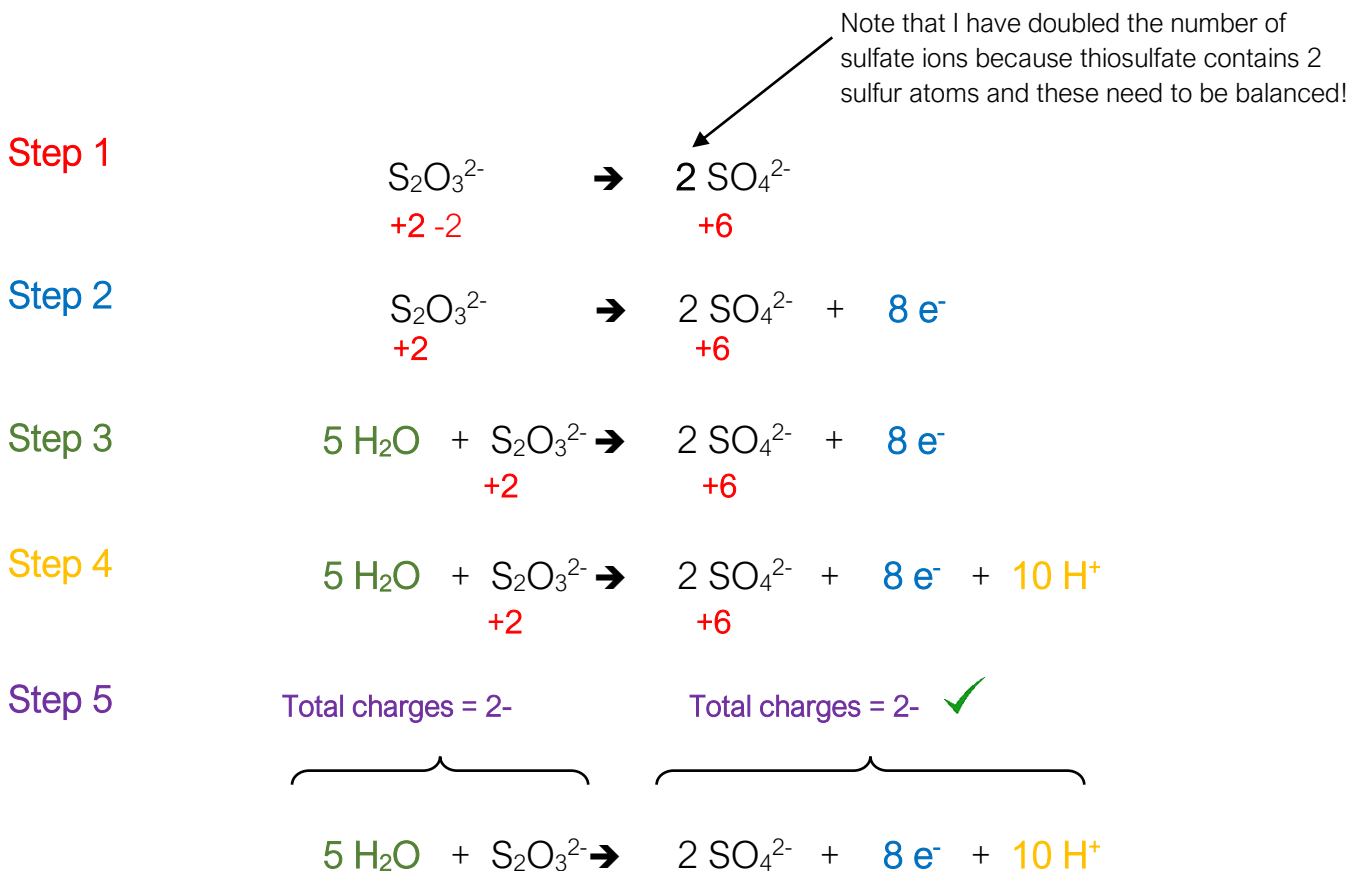
Step 4: hydrogens

Step 5: CHECK CHARGES! ✓

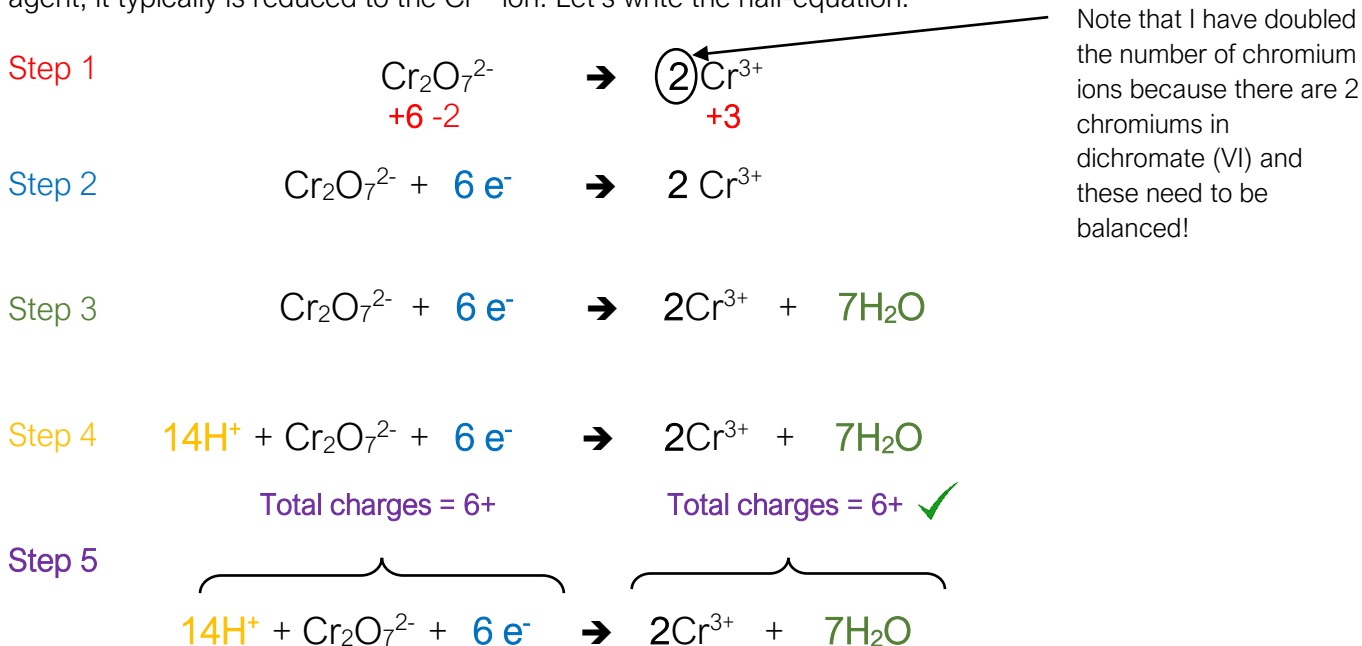
Let's work through another example

This second example involves the common reducing agent, the **thiosulfate ion**, $\text{S}_2\text{O}_3^{2-}$. This is typically found in the form of sodium thiosulfate solution, $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$.

Thiosulfate can be oxidised to several species. One common ion that it is oxidised to is the sulfate (VI) ion, SO_4^{2-}

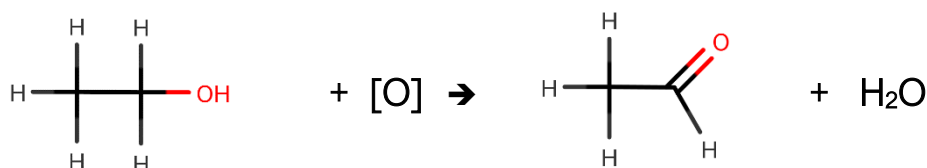


A common oxidising agent is **potassium dichromate** (VI), $\text{K}_2\text{Cr}_2\text{O}_7$. When the $\text{Cr}_2\text{O}_7^{2-}$ ion reacts as an oxidising agent, it typically is reduced to the Cr^{3+} ion. Let's write the half-equation.

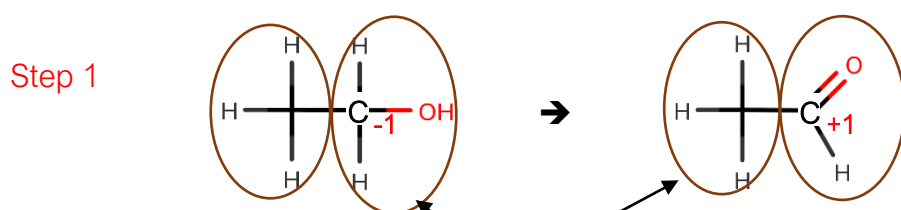


We can apply this to **organic molecules** too.

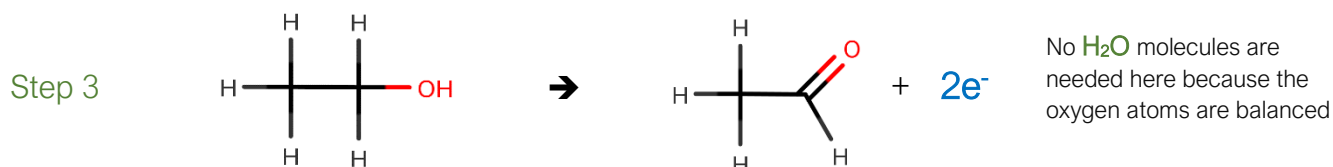
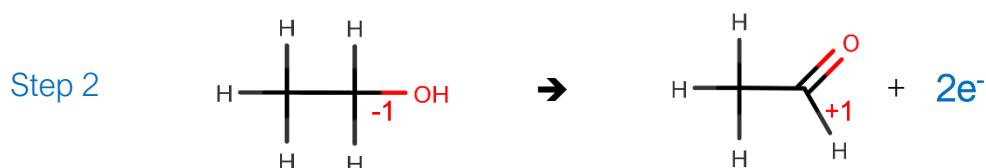
Consider the oxidation of ethanol to ethanoic acid. In Y12 Chemistry, we tend to represent this oxidation of ethanol to ethanoic acid by adding the oxidising agent to the equation as $[O]$. See the equation below.



Now let's consider writing a half-equation for the oxidation process



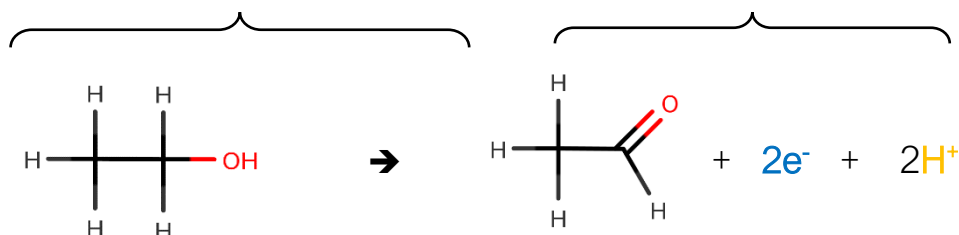
The sum of all the oxidation states on all of the atoms in each of the bubbles = 0
This allows us to calculate the oxidation state of the C that gets oxidised.



Total charges = 0

Total charges = 0 ✓

Step 5

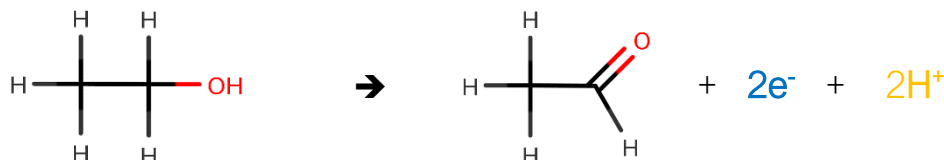


Writing half-equations is only a means to an end. The point of writing half-equations is so that we can put them together to form full equations. Once we have a full (or overall) equation, we can find the **ratio** of oxidising agent to reducing agent, *i.e.* the **stoichiometry** of the reaction. This is essential if we intend to combine the chemicals in a quantitative way, for example, a REDOX titration.

Let consider the last two examples of half-equations on the previous page.

One was for the **oxidation of ethanol** and the other was for the **reduction of dichromate (VI)**. This is the most common oxidising agent for the oxidation of ethanol in the A-Level chemistry laboratory.

Equation 1: the OXIDATION PROCESS:

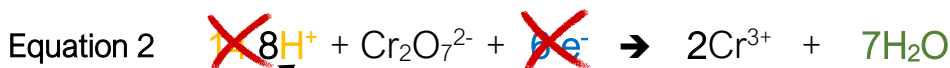
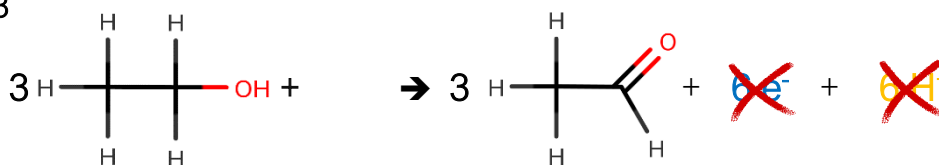


Equation 2: the REDUCTION PROCESS



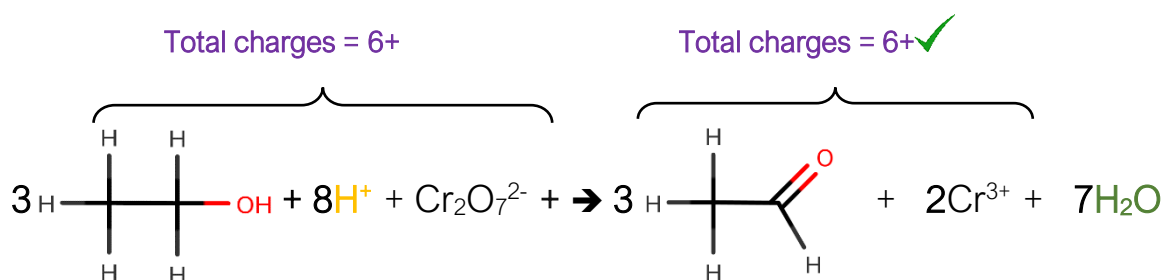
We need to **add these two equations** in a way that allows us to cancel the electrons. This requires us to multiply the equation 1 by 3.

Equation 1 x3



reduced in number for 14 to 8 because 6H⁺ have been cancelled in both half-equation

Now we **add the equations together** having **cancelled** the **electrons**, and **6H⁺** on each side

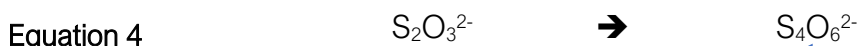
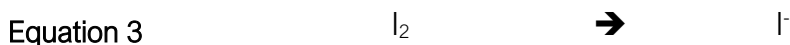


We can see that it requires **1** dichromate (VI) ion to fully oxidise **3** ethanol molecules (to ethanal).

Note also the requirement for this reaction to be done in the **presence of hydrogen ions**. This is typically done by carrying the reaction out in an excess of sulfuric acid

Now, you have a go at writing half-equations and full equations for the following examples.

Just stick to the steps.



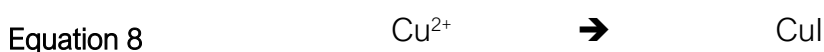
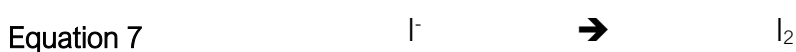
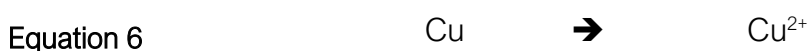
Do not be concerned about the non-integer value for the O.N. of S in $\text{S}_4\text{O}_6^{2-}$. The fact is that the non-integer value for the O.N. of S is simply because there are 2 different O.N. for the 4 sulfurs and 2.5 is simply the **average!** It still works.



These 2 overall equations can be used in the analysis of bleach to determine the amount of available chlorine. Typical household bleach contains the chlorate (I) ion, ClO^- .

In the first step, an **excess** of iodide ions is added to a sample of accurately diluted **bleach** (known volume!). The **iodine** that is produced can then be titrated against a solution of **sodium thiosulfate** (of known concentration) to determine the moles of iodine and therefore, the moles of ClO^- present in the original sample. To indicate when all the iodine has been reduced in the Stage 2, a little **starch** is used to identify the remaining unreacted iodine. When the dark blue/black colour is finally removed, the end point has been reached.

Here are some more equations to solve.



Equation 8 may seem a little tricky! It is the **copper** that changes oxidation state. Therefore, this needs **iodide** adding to the left-hand side!



In stage 2, **iodine** is made, just like in the overall reaction for stages 3 and 4!

The concentration of iodine can then be determined by the same method, *i.e.* by a **thiosulfate titration!**

This reaction of **copper with nitric acid** to produce Cu^{2+} ions (Overall reaction in stage 1) and the subsequent oxidation of iodide to make iodine is the chemistry involved in the analysis of **copper content in brass**.

The ANSWERS are on in the CramNow Inorganic Folder. There are also some quizzes on this topic on **CramNow** Module 5.