Lesson 18 – Half equations

Redox reactions can be represented by two half equations, one showing the oxidation process and one showing the reduction process. Oxidation half equations represent the loss of electrons, so electrons will be found on the right hand side of the arrow. For reduction half equations the electrons are written to the left of the arrow, as reduction is the gain of electrons.

Example - Displacement

$$2KI + Cl_2 \rightarrow 2KCl + I_2$$

In the reaction above, iodide is being oxidised as it is increasing in oxidation state from -1 in KI to 0 in I_2. Chlorine is being reduced as it is decreasing in oxidation state from 0 in Cl_2 to -1 in KCl.

To write the half equation for the oxidation, we simply need to show the conversion from I^- to I_2. Electrons will be included on the right hand side of the half equation, as they are being removed. This also has the effect of balancing the charge: the total charge on the left hand side of the half equation must equal the total charge on the right hand side.

$$2I^- \rightarrow I_2 + 2e^-$$

We use the same method to write the half equation for the reduction, showing the conversion of Cl_2 to Cl^-. In this case the electrons will be on the left hand side, as they are being gained by the chlorine atoms. Once again, the charge must be balanced in the half equation.

$$Cl_2 + 2e^- \rightarrow 2Cl^-$$

Writing more complex half equations

Many redox processes take place in solution. Quite often, they may involve the exchange of oxygen and/or hydrogen atoms from water as well as electrons. In this case, the half equations are often more complex! Redox processes can take place in acidic conditions, in which case H^+ may be involved, or they make take place in alkaline conditions, in which case OH^- may be involved. For our A level course, you only need to be able to write half equations in acidic conditions, which can be achieved using the step by step method below.

1. Identify the two forms of the element being oxidised or reduced. You need to include the whole molecule or ion in which the element is found.
2. Make sure you have equal numbers of atoms of this element on the left and right of the equation.
3. If necessary, balance O by adding H_2O to the left or right hand side of the half equation.
4. If necessary, balance H by adding H^+ to the left or right hand side of the half equation.
5. Balance the charge by adding electrons to the least negative side of the half equation.

Once you have finished, you can check you are correct by looking at the number of electrons being gained or lost: this should correspond with the total change in oxidation state of the element being oxidised or reduced.

Example

Acidified potassium manganate (VII) can be used as an oxidising agent in organic chemistry. When acting as an oxidising agent, the manganate (VII) ions (MnO_4^-) are themselves reduced to Mn^{2+} ions. To write the half equation for the reduction of manganate (VII) ions, follow the steps given above:

1. MnO_4^- (oxidation state of Mn is +7) is converted to Mn^{2+} (oxidation state of Mn is +2)
2. Balance Mn: $MnO_4^- \rightarrow Mn^{2+}$
3. Balance O by adding H_2O: $MnO_4^- \rightarrow Mn^{2+} + 4H_2O$
4. Balance H by adding H^+: $MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$
5. Balance charge by adding electrons: Total charge on the left is currently +7 and on the right the total charge is +2, difference of 5: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

Five electrons have been added, this corresponds with the change in oxidation state of manganese from +7 to +2.

Combining half equations

Once you have written the half equations for both the oxidation and reduction processes of a redox reaction, they can be combined to write the overall ionic equation for the reaction. To do this, you must first check that the number of electrons in each half equation is the same, if not simply multiply the balancing numbers in the half equation up so that it is! After this, combine the equations by writing out the reactants from both half equations on the left, and the products from both half equations on the right. The electrons will cancel out, and depending on the half equations you are using you may be able to cancel out some of the other species involved, such as H_2O and H^+.

Example

To work out the overall ionic equation for the reaction in which acidified potassium manganate (VII) oxidises Fe^{2+} to Fe^{3+}: first write out the two half equations:

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

As it is currently written, there are 5 electrons needed by the manganate (VII) ion, but only one being generated by the reaction of Fe^{2+}. To “balance” the electrons, all the balancing numbers in the iron half equation need to be multiplied by 5.

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

$$5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$$

The half equations are now ready to be combined:

$$MnO_4^- + 8H^+ + 5e^- + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+} + 5e^-$$

The electrons can be cancelled out, leaving the overall ionic equation:

$$MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$$

References

AQA Chemistry – Lister and Renshaw

A level Chemistry – Exam Board AQA (CGP)