6 Electrochemistry

This chapter contains a small amount of material about reduction and oxidation (**redox**) and the use of oxidation numbers. For further material examined at A Level, see Chapter 28.

KEY TERM

Redox The word **redox** comes from two familiar words – **red**uction and **ox**idation – and refers to what happens in chemical reactions in which electrons are gained or lost.

When a metal is oxidised, it loses electrons, for example:

$$Mg \rightarrow Mg^{2+} + 2e^-$$

When magnesium is burned in air, the electrons released by magnesium atoms are picked up by oxygen atoms, reducing them to oxide ions:

$$\tfrac{1}{2}O_2 + 2e^- \rightarrow O^{2-}$$

Oxidation and reduction do not only occur with elements – for example, species such as iron(II) ions can be oxidised to iron(III) ions:

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

Reduction of one ion to form another ion can also occur, such as in a manganate(VII) titration:

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

Did you notice the three sets of Roman numerals used in the last two examples? These numbers show the magnitude of the oxidation number of the element concerned. So in manganate(VII), the manganese has an oxidation number of +7.

Oxidation numbers

To work out the oxidation number of an element in a compound or ion there are some simple rules to follow:

- Atoms of uncombined elements have an oxidation number of zero.
- Simple ions have an oxidation number equal to the charge on the ion.
- In complex ions, the sum of the oxidation numbers in all the elements present equals the overall charge on the ion.

You may be given equations and asked for the change in oxidation number (oxidation state) of one of the elements present. In these cases it is often easier to think of oxidation numbers on a line, as shown in Figure 6.1.

Figure 6.1

Using this idea, if you mark the reduced and oxidised species at the correct points on the line, you can see the change in oxidation number, and the number of electrons lost or gained.

You can use changes in oxidation numbers to help balance chemical equations or, if you have the equation, to deduce the change in oxidation number of an element.

WORKED EXAMPLE

Consider the oxidation of hydrogen sulfide to form sulfur dioxide:

$$2H_2S(g) + 3O_2(g) \rightarrow 2SO_2(g) + 2H_2O(l)$$

What is the change in oxidation number of sulfur?

Answer

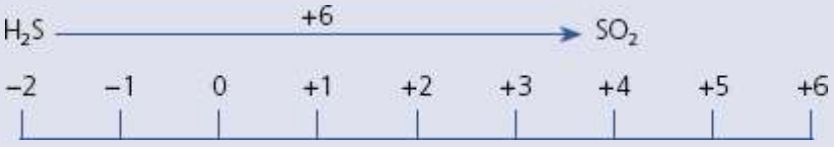

Figure 6.2

As you can see from Figure 6.2, you count the distance between the two oxidation numbers. In this case the change is from −2 to +4; or a change of +6.

NOW TEST YOURSELF

1 What is the oxidation number of iron in each of the following substances?

a Fe_2O_3

b Fe

c Na_2FeO_4

2 What are the oxidation numbers of manganese in this equation?

$$2KMnO_4 + MnO_2 + 4KOH \rightarrow 3K_2MnO_4 + 2H_2O$$

REVISION ACTIVITY

a What are the oxidation numbers of chlorine in the following reaction?

$$3ClO^-(aq) \rightarrow 2Cl^-(aq) + ClO_3^-(aq)$$

b Write half-equations to show how chlorine is both reduced and oxidised in the reaction above.

END OF CHAPTER CHECK

By now you should be able to:

- understand and use electron transfer and changes in oxidation number (oxidation state) to explain redox processes