REDOX TITRATIONS

Redox titrations involve the titration of an oxidant against a reductant. These reactions involve the transfer of electrons.

The procedure is identical to that described for acid-base titrations, however, in many redox titrations, an indicator is **NOT** required to detect the equivalence point of the reaction as a spontaneous colour change occurs around the equivalence point of the reaction.
i.e. many redox titrations are self-indicating.

Great care must be exercised when presented with a titration, so as to determine whether the titration is redox in nature. If the process is a redox titration, remember to use half equations to produce the overall equation. Do not attempt to combine anions and cations of reacting species.

Redox titrations may be identified by the following:

(a) There is a spontaneous colour change in the absence of an indicator.

(b) The words reduced or oxidised appear in the question.

(c) One or more of the chemical species has undergone a change in oxidation number during the **titration** process.

QUESTION 1

Given the half–cell equation $O_{2(g)} + 2H^+_{(aq)} + 2e^- \rightarrow H_2O_{2(aq)}$:

(a) Construct a fully balanced redox ionic equation for the oxidation of hydrogen peroxide by potassium manganate (VII) ($KMnO_4$) to form Mn^{2+}.

50.0 cm^3 of a solution of hydrogen peroxide was diluted to 1.00 L with water. 25.0 cm^3 of this solution, when acidified with dilute sulfuric acid, reacted with 20.25 cm^3 of 0.0200 M $KMnO_4$.

(b) What is the concentration of the original hydrogen peroxide solution in mol/L?

Solution

(a) Oxidation of hydrogen peroxide:

$$H_2O_{2(aq)} \rightarrow O_{2(g)} + 2H^+_{(aq)} + 2e^-$$

Reduction of $KMnO_4$:

$$MnO^-_{4(aq)} \rightarrow Mn^{2+}_{(aq)}$$
$$MnO^-_{4(aq)} \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$$
$$MnO^-_{4(aq)} + 8H^+_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$$
$$MnO^-_{4(aq)} + 8H^+_{(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$$

$$H_2O_{2(aq)} \rightarrow O_{2(g)} + 2H^+_{(aq)} + 2e^-$$ Multiply by 5

$$MnO^-_{4(aq)} + 8H^+_{(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$$ Multiply by 2

$$5H_2O_{2(aq)} \rightarrow 5O_{2(g)} + 10H^+_{(aq)} + 10e^-$$

$$2MnO^-_{4(aq)} + 16H^+_{(aq)} + 10e^- \rightarrow 2Mn^{2+}_{(aq)} + 8H_2O_{(l)}$$

$$2MnO^-_{4(aq)} + 6H^+_{(aq)} + 5H_2O_{2(aq)} \rightarrow 2Mn^{2+}_{(aq)} + 8H_2O_{(l)} + 5O_{2(g)}$$

(b) **Step 1:** Calculate the number of mole of known reactant.

$$n(KMnO_4) = cV = 0.0200 \times 0.02025 = 0.000405$$

Step 2: Use the mole ratio in the balanced equation to determine the number of mole of unknown reacting.

$$n(H_2O_2) = \frac{5}{2} \times n(KMnO_4) = 0.000405 \times 2.5 = 0.0010125$$

Step 3: Calculate the concentration of unknown used in the titration.

$$c = \frac{n}{V} = \frac{0.0010125}{0.025} = 0.0405\,M$$

Step 4: Take into account any dilutions performed.

Original molarity of H_2O_2 is $0.0405 \times \frac{1000}{50} = 0.810\,M$