Redox Titrations - Descriptions

1. Potassium manganate and iron (II). In this reaction, manganate (VII) ions are reduced to Mn^{2+} and iron (II) ions are oxidised to iron (III) ions. Write balanced half equations and a full ionic equation for the reaction.

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

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

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

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

Potassium manganate is dark purple so the top of the meniscus is read in the burette. The end point is marked by the first hint of a permanent pink colour in the solution.

2. Sodium thiosulfate and iodine. In this reaction, sodium thiosulfate ($S_2O_3^{2-}$) is oxidised to $S_4O_6^{2-}$ and iodine is reduced to iodide ions. Write balanced half equations and a full ionic equation for the reaction.

$$2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-$$

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

$$2S_2O_3^{2-} + I_2 + \cancel{2e^-} \rightarrow S_4O_6^{2-} + \cancel{2e^-} + 2I^-$$

$$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$$

In this reaction sodium thiosulfate is usually in the burette. The solution in the conical will contain iodine and so be brown in colour. As the titration proceeds the brown colour fades to pale yellow as fewer iodine molecules remain. At this point starch is added to indicate the remaining iodine. A blue-black colour is seen. The titration is then continued until the solution turns colourless.

Note – if starch is added at the start it will bind the iodine and cause spurious results.

It is common to see this reaction used in back titrations. See the examples in the video on redox titration calculations.