Chemical Calculations Tutorial 8- Redox Titrations

1. The relative atomic mass of iron = 55.9

   (i) $MnO_4^-{}_{(aq)} + 8H^+{}_{(aq)} + 5e \rightarrow Mn^{2+}{}_{(aq)} + 4H_2O_{(l)}$

   (ii) $Fe^{3+}{}_{(aq)} + e \rightarrow Fe^{2+}{}_{(aq)}$

   (a) Construct the fully balanced redox ionic equation for the manganate(VII) ion oxidising the iron(II) ion

   (b) 24.3 cm^3 of 0.02 mol dm^{-3} $KMnO_4$ reacted with 20.0 cm^3 of an iron(II) solution acidified with dilute sulfuric acid.

   (i) Calculate the molarity of the iron(II) ion.

   (ii) How do you recognise the end–point in the titration?

   (c) Calculate the percentage of iron in a sample of steel wire if 1.51 g of the wire was dissolved in excess of dilute sulfuric acid and the solution made up to 250 cm^3 in a standard graduated flask. A 25.0 cm^3 aliquot of this solution was pipetted into a conical flask and needed 25.45 cm^3 of 0.02 mol dm^{-3} $KMnO_4$ for complete oxidation.

   (d) Suggest reasons why the presence of dil. sulfuric acid is essential for an accurate titration and why dil. hydrochloric and nitric acids are unsuitable to be used in this reaction.

   (e) 8.25g of an iron(II) salt was dissolved in 250 cm^3 of pure water. 25.0 cm^3 aliquots were pipetted from this stock solution and titrated with 0.02 mol dm^{-3} potassium manganate(VII) solution.

   The titration values obtained were 23.95 cm^3, 23.80 cm^3 and 23.85 cm^3.

   (i) What titration value should be used in the calculation and why?

   (ii) Calculate the moles of manganate(VII) used in the titration.

   (iii) Calculate the moles of iron(II) ion titrated

   (iv) Calculate the mass of iron(II) titrated

   (v) Calculate the total mass of iron in the original sample of the iron(II) salt.

   (vi) calculate the % iron in the salt. (94.2%)

2. Given the following two half–reactions

   (a) Given (i) $S_4O_6^{2-}{}_{(aq)} + 2e^- \rightarrow 2S_2O_3^{2-}{}_{(aq)}$

and (ii) $I_{2(aq)} + 2e^- \rightarrow 2I^-_{(aq)}$

construct the full ionic redox equation for the reaction of the thiosulfate ion $S_2O_3^{2-}$ and iodine I_2.

(b) What mass of iodine reacts with 23.5 cm^3 of 0.0120 mol dm^{-3} sodium thiosulfate solution?

(c) 25.0 cm^3 of a solution of iodine in potassium iodide solution required 26.5 mL of 0.0950 mol dm^{-3} sodium thiosulfate solution to titrate the iodine. What is the molarity of the iodine solution and the mass of iodine per dm^3?

3. 2.83 g of a sample of haematite iron ore [iron (III) oxide, Fe_2O_3] were dissolved in concentrated hydrochloric acid and the solution diluted to 250 cm^3. (Ans = 12.8 g cm^{-3})

   25.0 cm^3 of this solution was reduced with tin(II) chloride (which is oxidised to Sn^{4+} in the process) to form a solution of iron(II) ions. This solution of iron(II) ions required 26.4 cm^3 of a 0.0200 mol dm^{-3} potassium dichromate(VI) solution for complete oxidation back to iron(III) ions.

   (a) given the half reactions

   (i) $Sn^{4+}_{(aq)} + 2e \rightarrow Sn^{2+}_{(aq)}$

   and (ii) $Cr_2O_7^{2-}{}_{(aq)} + 14H^+_{(aq)} + 6e \rightarrow 2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$

   deduce the fully balanced redox equations for the reactions

   (i) the reduction of iron(III) ions by tin(II) ions

   (ii) the oxidation of iron(II) ions by the dichromate(VI) ion

   (b) Calculate the percentage of iron(III) oxide in the ore. (Ans = 89.4 %)

   (c) Suggest why potassium manganate(VII) isn't used for this titration?

4. An approximately 0.02 mol dm^{-3} potassium manganate(VII) solution was standardized against precisely 0.100 mol dm^{-3} iron(II) ammonium sulfate solution. 25.0 cm^3 of the solution of the iron(II) salt were oxidized by 24.15 cm^3 of the manganate(VII) solution. What is the molarity of the potassium manganate(VII) solution ? (Ans= 0.0207 M)

5. 10.0 g of iron(II) ammonium sulfate crystals were made up to 250 cm^3 of acidified aqueous solution. 25.0 cm^3 of this solution required 21.25 cm^3 of 0.0200 mol dm^{-3} potassium dichromate(VI) for oxidation.

   Calculate x in the formula $FeSO_4.(NH_4)_2SO_4.xH_2O$

   (Ans x = 6)

6. Given the half–reaction $C_2O_4^{2-}{}_{(aq)} \rightarrow 2CO_{2(g)} + 2e^-$

(a) write out the balanced redox equation for manganate(VII) ions oxidising the ethanedioate ion (or ethanedioic acid).

(b) 1.520 g of hydrated ethanedioic acid crystals, $H_2C_2O_4.2H_2O$, was made up to 250.0 cm^3 of aqueous solution and 25.00 cm^3 of this solution needed 24.55 cm^3 of a potassium manganate(VII) solution for oxidation.

Calculate the molarity of the manganate(VII) solution and its concentration in g dm^{-3}. (Ans = 3.11 g dm^{-3})

7. A standardization of potassium manganate(VII) solution yielded the following data: 0.150 g of the salt potassium tetraoxalatedihydrate, $KHC_2O_4.\ H_2C_2O_4.\ 2H_2O$ needed 23.20 cm^3 of the manganate (VII) solution. What is the molarity of the manganate(VII) solution? (Ans = 0.0203 M)

8. Given the half equations: (i) $O_{2(g)} + 2H^+{}_{(aq)} + 2e \rightarrow H_2O_{2(aq)}$

and (ii) $MnO_4^-{}_{(aq)} + 8H^+{}_{(aq)} + 5e \rightarrow Mn^{2+}{}_{(aq)} + 4H_2O_{(l)}$

(a) construct the fully balanced redox ionic equation for the oxidation of hydrogen peroxide by potassium manganate(VII)

(b) 50.0 cm^3 of solution of hydrogen peroxide were diluted to 1.00 dm^3 with water.

25.0 cm^3 of this solution, when acidified with dilute sulfuric acid, reacted with 20.25 cm^3 of 0.0200 mol dm^{-3} $KMnO_4$. Calculate the concentration of the original hydrogen peroxide solution in mol dm^{-3}? (Ans = 0.81 mol dm^{-3})

9. 13.2 g of iron(III) alum were dissolved in water and reduced to an iron(II) ion solution by zinc and dilute sulfuric acid. The mixture was filtered and the filtrate and washings made up to 500 cm^3 in a standard volumetric flask. If 20.0 cm^3 of this solution required 26.5 mL of 0.0100 mol dm^{-3} $KMnO_4$ for oxidation.

(a) Write the ionic equation for the reduction of iron(III) ions by zinc metal.

(b) Calculate the percentage by mass of iron in iron alum. (Ans = 14 %)

10. Calculate the concentration in mol dm^{-3} and g dm^{-3}, the salt sodium ethanedioate ($Na_2C_2O_4$) solution, 5.00 cm^3 of which were oxidized in acid solution by 24.50 cm^3 of a potassium manganate(VII) solution containing 0.05 mol dm^{-3}. (Ans = 0.613 mol dm^{-3} , 82.1 g dm^{-3})

11. Calculate x in the formula $FeSO_4.xH_2O$ from the following data: 12.18 g of iron(II) sulfate crystals were made up to 500 cm^3 acidified with sulfuric acid. 25.0 cm^3 of this solution required 43.85 cm^3 of 0.0100 mol dm^{-3} $KMnO_4$ for complete oxidation. (Ans x = 7)

12. Given the half–reaction equation:

(i) $NO_3^-{}_{(aq)} + 2H^+{}_{(aq)} + 2e^- \rightarrow NO_2^-{}_{(aq)} + H_2O_{(l)}$

(ii) $MnO_4^-{}_{(aq)} + 8H^+{}_{(aq)} + 5e^- \rightarrow Mn^{2+}{}_{(aq)} + 4H_2O_{(l)}$

(a) give the ionic equation for potassium manganate(VII) oxidising nitrate(III) to nitrate(V)

(b) 24.2 cm^3 of sodium nitrate(III) [sodium nitrite] solution, added from a burette, were needed to discharge the colour of 25.0 cm^3 of an acidified 0.0250 mol dm^{-3} $KMnO_4$ solution. What was the concentration of the nitrate(III) solution in grams of anhydrous salt per dm^3 ? (Ans = 4.46 g dm^{-3})

13. 2.68 g of the salt iron(II) ethanedioate, FeC_2O_4, were made up to 500 cm^3 of acidified aqueous solution. 25.0 cm^3 of this solution reacted completely with 28.0 cm^3 of 0.0200 mol dm^{-3} potassium manganate(VII) solution. Calculate the mole ratio of $KMnO_4$ to FeC_2O_4 taking part in this reaction using the titration. Write the balanced reaction and justify your answer. (Ans = 3:5)

14. Given the half reactions:

(i) $I_{2(aq)} + 2e^- \rightarrow 2I^-{}_{(aq)}$

(ii) $IO_3^-{}_{(aq)} + 6H^+{}_{(aq)} + 5e^- \rightarrow {}^1/_2I_{2(aq)} + 3H_2O_{(l)}$

(a) Deduce the redox equation for iodate(V) ions oxidising iodide ions to iodine.

(b) What volume of 0.0120 mol dm^{-3} iodate(V) solution reacts with 20.0 cm^3 of 0.100 mol dm^{-3} iodide ion solution?

(c) 25.0 cm^3 of the potassium iodate(V) solution were added to about 15 cm^3 of a 15% solution of potassium iodide (ensures excess iodide ion). On acidification, the liberated iodine needed 24.1 cm^3 of 0.0500 mol dm^{-3} sodium thiosulfate solution to titrate it.

(i) Calculate the concentration of potassium iodate(V) in g dm^{-3} (Ans = 1.72 g dm^{-3})

$2S_2O_3^{2-}{}_{(aq)} + I_{2(aq)} \implies S_4O_6^{2-}{}_{(aq)} + 2I^-{}_{(aq)}$

(ii) What indicator is used for this titration and what is the colour change at the end–point?

15. Calculate the molarities of iron(II) and iron(III) ions in a mixed solution from the following data.

$$MnO_4^-{}_{(aq)} + 8H^+{}_{(aq)} + 5Fe^{2+}{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)} + 5Fe^{3+}{}_{(aq)} + 4H_2O_{(l)}$$

(i) 25.0 cm^3 of the original mixture was acidified with dilute sulfuric acid and required 22.5 cm^3 of 0.0200 mol dm^{-3} $KMnO_4$ for complete oxidation.

(ii) a further 25.0 cm^3 of the original iron(II)/iron(III) mixture was reduced with zinc and acid and it then required 37.6 cm^3 of the $KMnO_4$ for complete oxidation. (Ans: Fe^{+2} = 0.09 mol dm^{-3} ; Total 0.06 mol dm^{-3})

16. A piece of rusted iron was analysed to find out how much of the iron had been oxidised to rust [hydrated iron(III) oxide]. A small sample of the rusted iron was dissolved in excess dilute sulfuric acid to give 250 cm^3 of solution. The solution contains Fe^{2+} ions from the unrusted iron dissolving in the acid, and, Fe^{3+} ions from the rusted iron.

(a) 25.0 cm^3 of this solution required 16.9 cm^3 of 0.02 mol dm^{-3} $KMnO_4$ for complete oxidation of the Fe^{2+} ions.

Given: $MnO_4^-{}_{(aq)} + 8H^+{}_{(aq)} + 5Fe^{2+}{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)} + 5Fe^{3+}{}_{(aq)} + 4H_2O_{(l)}$

Calculate the moles of Fe^{2+} ions in the sample titrated. (Ans = 0.00169 mol)

17. 25.0 cm^3 of an iodine solution was titrated with 0.100 mol dm^{-3} sodium thiosulfate solution and the iodine reacted with 17.6 cm^3 of the thiosulfate solution.

(a) give the reaction equation.

(b) what indicator is used? and describe the end–point in the titration.

(c) calculate the concentration of the iodine solution in mol dm^{-3} and g dm^{-3}. (Ans = 8.94 g dm^{-3})

18. 1.01g of an impure sample of potassium dichromate(VI), $K_2Cr_2O_7$, was dissolved in dil. sulfuric acid and made up to 250 cm^3 in a calibrated volumetric flask. A 25.0 cm^3 aliquot of this solution pipetted into a conical flask and excess potassium iodide solution and starch indicator were added. The liberated iodine was titrated with 0.100 mol dm^{-3} sodium thiosulfate and the starch turned colourless after 20.0 cm^3 was added.

(a) Construct the full balanced equation for the reaction between the dichromate(VI) ion and the iodide ion.

(b) Write the balanced redox equation for the reaction between the thiosulfate ion and iodine.

(c) Calculate the moles of sodium thiosulfate used in the titration and hence the number of moles of iodine titrated.

(d) Calculate the moles of dichromate(VI) ion that reacted to give the iodine titrated in the titration.

(e) Calculate the formula mass of potassium dichromate(VI) and the mass of it in the 25.0 cm^3 aliquot titrated.

(f) Calculate the total mass of potassium dichromate(VI) in the original sample and hence its mass percentage purity. (Ans = 97 %)

19. This question involves titrating ethanedioic acid (*oxalic acid*), $H_2C_2O_4$ or $(COOH)_2$ (i) with standard sodium hydroxide solution and then with potassium manganate(VII) solution (*potassium permanganate*, $KMnO_4$).

The titration data is as follows:

10 cm^3 of a $H_2C_2O_4$ solution required 8.50 cm^3 of a 0.20 mol dm^{-3} solution of sodium hydroxide for complete neutralisation using phenolphthalein indicator (first permanent pink end-point).

10 cm^3 of the same $H_2C_2O_4$ solution required 8.20 cm^3 of a $KMnO_4$ solution for complete oxidation to carbon dioxide and water in the presence of dilute sulfuric acid to further acidify the ethanedioic acid solution (first permanent pink end-point).

(a) Write an equation for the neutralisation reaction of ethanedioic acid with sodium hydroxide.

(b) Calculate the moles of $H_2C_2O_4$ in the solution and the molarity of the ethanedioic acid solution.

(c) Write the full redox titration equation for the oxidation of ethanedioic acid by potassium manganate(VII).

(d) From the equation in (c) and the titration data, deduce the molarity of the potassium manganate(VII) solution. (Ans = 0.415 M)

20. Lawn sand containing the salt iron(II) sulfate is used to treat moss.

2.50 g of the lawn sand was mixed with dilute sulfuric acid to extract the iron(II) salt. In a volumetric titration, it required 24.50 cm^3 of a 0.0200 mol dm^{-3} solution of potassium manganate(VII) to fully oxidise the iron salt.

(a) Write out the equation for the redox reaction involved in the titration.

(b) Calculate the % by mass of Fe^{2+} ions in the lawn sand. Atomic mass of iron = 55.8

(Ans = 5.47 %)

21. Analysing a medicinal iron tablet containing an iron(II) salt. (For people showing signs of iron deficiency e.g. low red blood cell count, suffering from anaemia). A 1.02 g iron tablet was dissolved in dilute sulfuric acid and titrated with a standard solution of potassium dichromate(VI). With a suitable redox indicator, it took 23.80 cm^3 of a 0.02 mol dm^{-3} solution of the dichromate reagent to completely oxidise all the Fe^{2+} ions in the tablet.

    (a) Give the full redox equation of the titration.

    (b) Calculate the percentage by mass of Fe^{2+} ions in the tablet. Atomic mass of iron = 55.8

(Ans = 15.6 %)

22. Consider the following redox reaction in acidic solution:

$$MnO_4^- + H_2O_2 \rightarrow Mn^{2+} + O_2 \text{ (acidic)}$$

    a) Write a balanced equation for the above reaction.

    b) The above reaction was used for a redox titration. At the equivalence point 5.684×10^{-4} mol $KMnO_4$ was required to titrate 5.00 mL of H_2O_2 solution. Calculate the concentration of H_2O_2.

23. The data below were obtained in a redox titration of a 25.00 mL sample containing Sn^{2+} ions using 0.125 mol dm^{-3} $KMnO_4$ according to the following reaction:

$$2\ MnO4^- + 16\ H^+ + 5\ Sn^{2+} \rightarrow 2\ Mn^{2+} + 8\ H_2O + 5\ Sn^{4+}$$

    If the volume required for the titration is 5.2 mL then calculate the concentration of Sn^{2+} in the original sample.

24. An impure sample of CaC_2O_4 weighing 0.803 g is titrated with 15.70 mL of 0.101 mol dm^{-3} $KMnO_4$. What is the percent by mass of the CaC_2O_4 in the original sample?

25. In a redox titration 12.50 mL of 0.0800 mol dm^{-3} $K_2Cr_2O_7$ (aq) was used in acidic solution to oxidize Sn^{2+} (aq) ions to Sn^{4+} (aq) ions. The volume of $K_2Cr_2O_7$ (aq) used was just sufficient to oxidize all the Sn^{2+} (aq) in 10.0 mL of the solution. Calculate the concentration of the Sn^{2+} (aq) ions in the solution according to the following unbalanced equation.

$$Cr2O7^{2-} (aq) + Sn^{2+} (aq) \rightarrow Sn^{4+} (aq) + Cr^{3+} (aq)$$

(Ans: 0.300 mol/L)

26. The copper (II) ions in a solution can be converted to copper metal by trickling the solution over scrap iron. The reaction produced iron (II) ions from scrap iron. If the process produces 25.00 L of solution containing 0.00200 mol dm^{-3} of $Fe^{2+}{}_{(aq)}$ ions, what mass of copper is produced?

$$Cu^{2+}{}_{(aq)} + Fe_{(s)} \rightarrow Fe^{2+}{}_{(aq)} + Cu_{(s)}$$

(Ans: 3.18g)

27. What volume of 0.0500 mol dm^{-3} $KMnO_{4\,(aq)}$ is needed to oxidize all the $Br^{-}{}_{(aq)}$ ions in 25.0 mL of an acidic 0.200 mol dm^{-3} $NaBr_{(aq)}$ solution according to

$$MnO_4^{-}(aq) + Br^{-}(aq) \rightarrow Br_2\,(aq) + Mn^{2+}\,(aq)$$

(Ans: v = 20.0mL)

28. Aqueous solutions of hydrogen peroxide sold in pharmacies are usually approximately 3% H_2O_2 by mass. However, in solution, hydrogen peroxide decomposes into water and oxygen.

What is the percent by mass of a solution of hydrogen peroxide, H_2O_2, prepared from 1.423 g of H_2O_2 which is titrated with 40.22 mL of 0.01143 mol dm^{-3} $KMnO_{4}(aq)$. The reaction occurs in an acidified solution.

(Ans: 2.747%)