**QUESTIONS**

**Question 1**: Given the following two half-reactions:

(i) **MnO4-(aq) + 8H+(aq) + 5e-==> Mn2+(aq) + 4H2O(l)**

and (ii) **Fe3+(aq) + e-==> Fe2+(aq)**

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

(b) **24.3** cm3 of **0.02** mol dm-3 KMnO4 reacted with **20.0** cm3 of an iron(II) solution.

(i) Calculate the molarity of the iron(II) ion. (ii) How do 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 sulphuric acid and the solution made up to **250** cm3 in a standard graduated flask.  **25.0** cm3 of this solution was pipetted into a conical flask and needed **25.45** cm3 of  **O.02** mol dm-3 KMnO4 for complete oxidation.

(d) Suggest reasons why the presence of dil. sulfuric acid is essential for an accurate titration

**Question 2**: Given the following two half-reactions

(a) Given (i) **S4O62-(aq) + 2e- ==> 2S2O32-(aq)**

and (ii) **I2(aq) + 2e- ==> 2I-(aq)**

construct the **full ionic redox equation** for the reaction of the thiosulphate ion S2O32-,and iodine.

(b) what mass of iodine reacts with **23.5** cm3 of **0.012** mol dm-3 sodium thiosulphate solution.

(c) **25**cm3 of a solution of iodine in potassium iodide solution required **26.5** cm3 of **0.095** mol dm-3 sodium thiosulphate solution to titrate the iodine.

What is the molarity of the iodine solution and the mass of iodine per dm3?

**Question 3**: **2.83** g of a sample of haematite iron ore [iron (III) oxide, **Fe203**] were dissolved in concentrated hydrochloric acid and the solution diluted to **250** cm3.

**25.00** cm3 of this solution was reduced with tin(II) chloride (which is oxidised to Sn4+ in the process) to form a solution of iron(II) ions. This solution required **26.4** cm3 of **0.02** mol dm-3 potassium dichromate(VI) for oxidation.

(a) given the half-cell reactions

(i) **Sn4+(aq) + 2e- ==> Sn2+(aq)**

and (ii) **Cr2O72-(aq) + 14H+(aq) + 6e- ==> 2Cr3+(aq) + 7H2O(l)**

deduce the fully balanced redox equations for the reactions (i) the reduction of iron(III) ions by tin(II) ions and (ii) the oxidation of iron(II) ions by the dichromate(VI) ion.

(b) Calculate the percentage of iron(III) oxide in the ore.

(c) Why isn't potassium manganate(VII) used for this titration? (as in Q1)

**ANSWERS**

**Question 1**: (a) **MnO4-(aq) + 8H+(aq) + 5Fe2+(aq)** ==> **Mn2+(aq) + 5Fe3+(aq) + 4H2O(l)**

(b) (i) mol MnO4- = 0.02 x 24.3 / 1000 = 0.000486, mol Fe2+ = 5 x 0.000486 (1 : 5 in equation) = 0.00243 in 20 cm3, so scaling up to 1 dm3, the **molarity of  Fe2+** = 0.00243 x 1000 / 20 = **0.122 mol dm-3**.

(ii) The end point is the **first faint permanent pink** due to a trace excess of KMnO4.

(c) mol MnO4- = 0.02 x 25.45 / 1000 = 0.000509, mol Fe = 5 x 0.000509 = 0.002545,

mass Fe = 0.002545 x 55.9 = 0.1423 g,

total Fe in wire = 0.1423 x 10 = 1.423 g (1/10th of the made up solution used in titration),

 so **% Fe** = 1.423 x 100 / 1.51 = **94.2 %**

(d) The choice of acid is fully discussed in

**Question 2**:(a) **2S2O32-(aq)  +  I2(aq)**  ==> **S4O62-(aq) + 2I-(aq)**

(b) mol S2O32- = 0.012 x 23.5 / 1000 = 0.000282, mole iodine as I2 = mol  S2O32- / 2 (1 : 2 in equation) = 0.000141,

**mass of iodine** = 0.000141 x 126.9 x 2 = **0.0358 g**

(c) mol S2O32- = 0.095 x 26.5 / 1000 = 0.002518,

mol of iodine = mole 'thio' / 2 = 0.002518 / 2 = 0.001259 in 25 cm3,

scaling up to 1 dm3 gives 0.001259 x 1000 /25 = **0.0504 mol dm-3 of molecular iodine I2**.

mass concentration of I2 = 0.0504 x 2 x 126.9 = **12.8 g dm-3 of iodine**

**Question 3**: (a) (i) **Sn2+(aq) + 2Fe3+(aq)** ==> **Sn4+(aq) + 2Fe2+(aq)**

(ii) **Cr2O72-(aq) + 14H+(aq) + 6Fe2+(aq)** ==> **2Cr3+(aq) + 6Fe3+(aq) + 7H2O(l)**

(b) for a 25cm3 aliquot titrated, mol Cr2O72- = 0.02 x 26.4 / 1000 = 0.000528,

mol Fe2+ titrated = 6 x Cr2O72- = 0.003168 (1 : 6 in equation),

mol Fe2O3 = mol Fe / 2 = 0.003168 / 2 =  0.001584,

Mr(Fe2O3) = 159.8, mass Fe2O3 = 0.001584 x 159.8 = 0.2531 g.

Total mass of Fe2O3 = 0.2531 x 10 (1/10th titrated) = 2.531 g. **% Fe2O3** = 2.531 x 100 / 2.83 **= 89.4%**

(c) Potassium manganate(VII) isn't used for this titration because it is strong enough to oxidise chloride ions (from the hydrochloric acid) to form chlorine, giving a completely false titration.