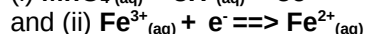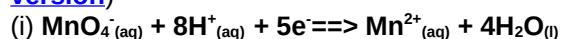**REDOX-ionic EQUATION CHECKS**

1. **Use of the correct 'species' (e.g. usually two given/chosen as/from half-cell equation data)**
2. **The 'species' direction change - which is oxidised or reduced? In other words get the half-cell equations the right way round! (If not indicated, might have to decide from E$^{\varnothing}$ data supplied, the more +ve half-cell is the reduction (oxidising agent). More in Equilibria Part 7 (being written)**
3. **The correct ratio of half-cell equations - the 'balance' must be based on oxidation number analysis or number of electrons transferred. The total increase in oxidation states = the total decrease in oxidation states or total electrons gained = total electrons lost by the species involved.**
4. **Add up the ion charges, the totals should be the same on both sides of the equation (I find this a handy extra check especially with stray H_2O's or H^+'s!).**
5. **'traditional' atom count - placed last because its not completely reliable with redox equations!**
6. **In some Q's the full equation may be given or the two half-cell equations to be put together the right way round and in the right ratio (see Redox Chemistry Part 2 (being written))**

Question 1: Given the following two half-reactions: (**Q1 can be done as an experimental 'word-fill' version**)
(i) $MnO_4^-{}_{(aq)} + 8H^+{}_{(aq)} + 5e^- ==> Mn^{2+}{}_{(aq)} + 4H_2O_{(l)}$
and (ii) $Fe^{3+}{}_{(aq)} + e^- ==> Fe^{2+}{}_{(aq)}$
(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.
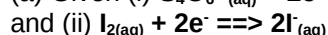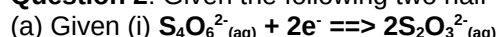(i) Calculate the molarity of the iron(II) ion. (ii) How do recognise the end-point in the titration?
(c) Calculate the percentage of iron in a sample of steel wire if **1.51** g of the wire was dissolved in excess of dilute sulphuric acid and the solution made up to **250** cm^3 in a standard graduated flask. **25.0** cm^3 of this solution was pipetted into a conical flask and needed **25.45** cm^3 of **O.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.

Question 2: Given the following two half-reactions
(a) Given (i) $S_4O_6^{2-}{}_{(aq)} + 2e^- ==> 2S_2O_3^{2-}{}_{(aq)}$
and (ii) $I_{2(aq)} + 2e^- ==> 2I^-{}_{(aq)}$
construct the **full ionic redox equation** for the reaction of the thiosulphate ion $S_2O_3^{2-}$, and iodine.
(b) what mass of iodine reacts with **23.5** cm^3 of **0.012** mol dm^{-3} sodium thiosulphate solution.
(c) **25**cm^3 of a solution of iodine in potassium iodide solution required **26.5** cm^3 of **0.095** mol dm^{-3} sodium thiosulphate solution to titrate the iodine.
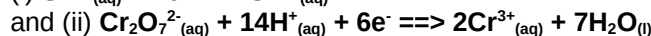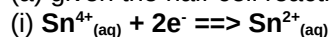What is the molarity of the iodine solution and the mass of iodine per dm^3?

Question 3: **2.83** g of a sample of haematite iron ore [iron (III) oxide, **Fe$_2$0$_3$**] were dissolved in concentrated hydrochloric acid and the solution diluted to **250** cm^3.
25.00 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 required **26.4** cm^3 of **0.02** mol dm^{-3} potassium dichromate(VI) for oxidation.
(a) given the half-cell reactions
(i) $Sn^{4+}{}_{(aq)} + 2e^- ==> Sn^{2+}{}_{(aq)}$
and (ii) $Cr_2O_7^{2-}{}_{(aq)} + 14H^+{}_{(aq)} + 6e^- ==> 2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)}$
deduce the fully balanced redox equations for the reactions (i) the reduction of iron(III) ions by tin(II) ions 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)

Question 4: An approximately **0.02** mol dm^{-3} potassium manganate(VII) solution was standardized against precisely **0.1** mol dm^{-3} iron(II) ammonium sulphate solution. **25.00** 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 ?

Question 5: **10.0** g of iron(II) ammonium sulphate crystals were made up to **250** cm^3 of acidified aqueous solution. **25** cm^3 of this solution required **21.25** cm^3 of **0.02** mol dm^{-3} potassium dichromate(VI) for oxidation.
Calculate **x** in the formula $FeSO_4.(NH_4)_2SO_4.xH_2O$

Question 6: Given the half-reaction $C_2O_4{}^{2-}{}_{(aq)}$ **- 2e$^-$ ==> 2CO$_{2(g)}$**
or **H$_2$C$_2$O$_{4(aq)}$ - 2e$^-$ ==> 2CO$_{2(g)}$ + 2H$^+{}_{(aq)}$**
(a) write out the balanced redox equation for manganate(VII) ions oxidising the ethanedioate ion (or ethane-dioic acid).
(b) **1.520** g of ethanedioic acid crystals, **H$_2$C$_2$O$_4$.2H$_2$O**, was made up to **250** cm^3 of aqueous solution and **25.0** 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}.

Question 7: A standardization of potassium manganate(VII) solution yielded the following data:
0.15 g of potassium tetroxalate (tetraoxalate?), **KHC$_2$O$_4$.H$_2$C$_2$O$_4$.2H$_2$O** needed **23.2** cm^3 of the manganate(VII) solution.
What is the molarity of the manganate(VII) solution? Use the equation and reasoning from Q6 to help you.

Question 8: Given the half-cell equation **O$_{2(g)}$ +2H$^+{}_{(aq)}$ + 2e$^-$ ==> H$_2$O$_{2(aq)}$**
(a) construct the fully balanced redox ionic equation for the oxidation of hydrogen peroxide by potassium manganate(VII)
(b) **50** cm^3 of solution of hydrogen peroxide were diluted to **1** dm^3 with water. **25.0** cm^3 of this solution, when acidified with dilute sulphuric acid, reacted with **20.25** cm^3 of **0.02** mol dm^{-3} KMnO$_4$.
What is the concentration of the original hydrogen peroxide solution in mol dm^{-3}?

Question 9: **13.2** g of iron(III) alum were dissolved in water and reduced to an iron(II) ion solution by zinc and dilute sulphuric acid. The mixture was filtered and the filtrate and washings made up to **500** cm^3 in a standard volumetric flask. **20.0** cm^3 of this solution required **26.5** cm^3 of **0.01** mol dm^{-3} KMnO$_4$ for oxidation.
(a) give the ionic equation for the reduction of iron(III) ions by zinc metal.
(b) Calculate the percentage by mass of iron in iron alum.

Question 10: Calculate the concentration in mol dm^{-3} and g dm^{-3}, of a sodium ethanedioate (**Na$_2$C$_2$O$_4$**) solution, **5.00** cm^3 of which were oxidized in acid solution by **24.5** cm^3 of a potassium manganate(VII) solution containing **0.05** mol dm^{-3}.

Question 11: Calculate **x** in the formula $FeSO_4.xH_2O$ from the following data:
12.18 g of iron(II) sulphate crystals were made up to **500** cm³ acidified with sulphuric acid.
25.0 cm³ of this solution required **43.85** cm³ of **0.01** mol dm⁻³ KMnO₄ for complete oxidation.

Question 12: Given the half-reaction $NO_3^-{}_{(aq)} + 2H^+{}_{(aq)} + 2e^- ==> NO_2^-{}_{(aq)} + H_2O_{(l)}$
(a) give the ionic equation for potassium manganate(VII) oxidising nitrate(III) to nitrate(V)
(b) **24.2** cm³ of sodium nitrate(III) [sodium nitrite] solution, added from a burette, were needed to discharge the colour of **25** cm³ of an acidified **0.025** mol dm⁻³ KMnO₄ solution.
What was the concentration of the nitrate(III) solution in grammes of anhydrous salt per dm³?

Question 13: **2.68** g of iron(II) ethanedioate, FeC_2O_4, were made up to **500** cm³ of acidified aqueous solution. **25.0** cm³ of this solution reacted completely with **28.0** cm³ of **0.02** mol dm⁻³ potassium manganate(VII) solution.
Calculate the mole ratio of KMnO₄ to FeC₂O₄ taking part in this reaction. Give the full redox ionic equation for the reaction.

Question 14: Given the half-cell reaction $IO_3^-{}_{(aq)} + 6H^+{}_{(aq)} + 5e^- ==> {}^1/_2I_{2(aq)} + 3H_2O_{(l)}$ (see also Q2)
(a) Deduce the redox equation for iodate(V) ions oxidising iodide ions.
(b) What volume of **0.012** mol dm⁻³ iodate(V) solution reacts with **20.0** cm³ of **0.100** mol dm⁻³ iodide solution?
(c) **25.0** cm³ of the potassium iodate solution were added to about 15 cm³ of a 15% solution of potassium iodide (ensures excess iodide ion). On acidification, the liberated iodine needed **24.1** cm³ of **0.05** mol dm⁻³ sodium thiosulphate solution to titrate it.
(i) Calculate the concentration of potassium iodate(V) in g dm⁻³
(ii) What indicator is used for this titration and what is the colour change at the end-point?

Question 15: Calculate the molarities of iron(II) and iron(III) ions in a mixed solution from the following data.
(i) **25.0** cm³ of the original mixture was acidified with dilute sulphuric acid and required **22.5 cm³** of **0.02** mol dm⁻³ KMnO₄ for complete oxidation.
(ii) a further **25.0** cm3 of the original iron(II)/iron(III) mixture was reduced with zinc and acid and it then required **37.6** cm³ of the KMnO₄ for complete oxidation.

Question 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 iron was dissolved in excess dilute sulphuric acid to give 250 cm³ 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** cm³ of this solution required **16.9** cm³ of **0.020** mol dm⁻³ KMnO₄ for complete oxidation of the Fe^{2+} ions.
Calculate the moles of Fe^{2+} ions in the sample titrated.
(b) To a second **25.0** cm³ of the rusted iron solution an oxidising agent was added to convert all the Fe^{2+} ions present to Fe^{3+} ions. The Fe^{3+} ions were titrated with a solution of $EDTA^{4-}{}_{(aq)}$ ions and **17.6** cm³ of **0.10** mol dm⁻³ EDTA were required.
Assuming 1 mole of EDTA reacts with 1 mole of Fe^{3+} ions, calculate the moles of Fe^{3+} ions in the sample.
(c) From your calculations in (a) and (b) calculate the ratio of rusted iron to unrusted iron and hence the percentage of iron that had rusted.

Question 17: 25.0 cm^3 of an iodine solution was titrated with 0.1 mol dm^{-3} sodium thiosulphate solution and the iodine reacted with 17.6 cm^3 of the thiosulphate 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}.

Question 18: **1.01g** of an impure sample of potassium dichromate(VI), **$K_2Cr_2O_7$**, was dissolved in dil. sulphuric acid and made up to **250 cm^3** in a calibrated volumetric flask. A **25 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 **0.1 mol dm^{-3} sodium thiosulphate** and the starch turned colourless after **20.0 cm^3** was added.
(a) Using the half-equations from Q3(a)(ii) and Q2(a)(ii), construct the full balanced equation for the reaction between the dichromate(VI) ion and the iodide ion.
(b) Using the half-equations from Q2(a) construct the balanced redox equation for the reaction between the thiosulphate ion and iodine.
(c) Calculate the moles of sodium thiosulphate 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 cm^3 aliquot titrated.
(f) Calculate the total mass of potassium dichromate(VI) in the original sample and hence its % purity.