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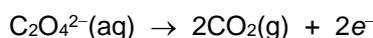
Mole Calculations Assignment Seven

Question One:

The ionic half-equation for the reduction of MnO_4^- (aq) to Mn^{2+} (aq) is given below:



The ionic half-equation for the oxidation of $\text{C}_2\text{O}_4^{2-}$ (aq) to CO_2 (g) is given below:



- a) i) Multiply the two ionic half-equations by simple whole numbers to ensure that the moles of electrons *lost* from $\text{C}_2\text{O}_4^{2-}$ (aq) equals the moles of electrons *gained* by MnO_4^- (aq).
ii) Combine and simplify the two ionic half-equations to give the overall ionic equation for the reaction between MnO_4^- (aq) and $\text{C}_2\text{O}_4^{2-}$ (aq).
- b) 25.00 cm³ of a 0.200 mol dm⁻³ aqueous solution of $\text{C}_2\text{O}_4^{2-}$ (aq) was found to react exactly with 16.00 cm³ of aqueous MnO_4^- (aq). Calculate the concentration of the aqueous MnO_4^- (aq) solution.

Question Two:

The ionic half-equation for the reduction of MnO_4^- (aq) to Mn^{2+} (aq) is given below:



The ionic half-equation for the oxidation of Fe^{2+} (aq) to Fe^{3+} (aq) is given below:



- a) i) Multiply the two ionic half-equations by simple whole numbers to ensure that the moles of electrons *lost* from Fe^{2+} (aq) equals the moles of electrons *gained* by MnO_4^- (aq).
ii) Combine and simplify the two ionic half-equations to give the overall ionic equation for the reaction between MnO_4^- (aq) and Fe^{2+} (aq).
- b) The reaction between MnO_4^- (aq) and Fe^{2+} (aq) can be used in the quantitative analysis of iron(II) salts. Ammonium iron(II) sulphate crystals have the following formula:



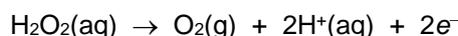
In an experiment to determine the value of n , 8.450 g of ammonium iron(II) sulphate crystals were dissolved in 250.0 cm³ of distilled water and sulphuric acid. 25.0 cm³ of this solution was further acidified and titrated against a 0.0150 mol dm⁻³ solution of potassium manganate(VII). A volume of 22.50 cm³ was required. Using this information, calculate the value of n .

Question Three:

The ionic half-equation for the reduction of $\text{MnO}_4^-(\text{aq})$ to $\text{Mn}^{2+}(\text{aq})$ is given below:



The ionic half-equation for the oxidation of $\text{H}_2\text{O}_2(\text{aq})$ to $\text{O}_2(\text{g})$ is given below:



- a) i) Multiply the two ionic half-equations by simple whole numbers to ensure that the moles of electrons *lost* from $\text{H}_2\text{O}_2(\text{aq})$ equals the moles of electrons *gained* by $\text{MnO}_4^-(\text{aq})$.
- ii) Combine and simplify the two ionic half-equations to give the overall ionic equation for the reaction between $\text{MnO}_4^-(\text{aq})$ and $\text{H}_2\text{O}_2(\text{aq})$.
- b) A solution of hydrogen peroxide was diluted 20.0 times. A 25.0 cm^3 portion of the diluted solution was acidified and titrated against $0.0150 \text{ mol dm}^{-3}$ potassium manganate(VII) solution. 45.70 cm^3 of the oxidant was required. Calculate the concentration of the hydrogen peroxide solution:
- i) In mol dm^{-3} .
- ii) In "volume concentration" *i.e.* the number of volumes of oxygen gas obtained from one volume of the solution.

Question Four:

The ionic half-equation for the reduction of $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$ to $\text{Cr}^{3+}(\text{aq})$ is given below:



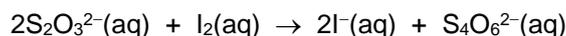
The ionic half-equation for the oxidation of $\text{Fe}^{2+}(\text{aq})$ to $\text{Fe}^{3+}(\text{aq})$ is given below:



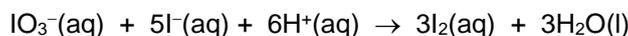
- a) i) Multiply the two ionic half-equations by simple whole numbers to ensure that the moles of electrons *lost* from $\text{Fe}^{2+}(\text{aq})$ equals the moles of electrons *gained* by $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$.
- ii) Combine and simplify the two ionic half-equations to give the overall ionic equation for the reaction between $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$ and $\text{Fe}^{2+}(\text{aq})$.
- b) A piece of iron wire of mass 2.225 g was put into a small conical flask containing dilute sulphuric acid. The flask was fitted with a bung carrying a Bunsen valve, to allow the hydrogen generated to escape, but prevent air from entering. The mixture was warmed to speed up the reaction. When all of the iron had reacted, the solution was cooled to room temperature and made up to 250.0 cm^3 with distilled water. With all of these precaution, the iron is converted to $\text{Fe}^{2+}(\text{aq})$ only, and no $\text{Fe}^{3+}(\text{aq})$ ions are formed. 25.0 cm^3 of the solution were acidified and titrated against a $0.0185 \text{ mol dm}^{-3}$ solution of potassium dichromate(VI). The volume required was 31.0 cm^3 . Calculate the percentage iron in the iron wire.

Question Five:

Thiosulphate ions (formula: $\text{S}_2\text{O}_3^{2-}(\text{aq})$) reduce iodine to iodide ions, forming tetrathionate ions (formula: $\text{S}_4\text{O}_6^{2-}(\text{aq})$) in the process:



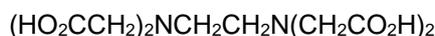
Salts of iodate(V) ions (formula: $\text{IO}_3^-(\text{aq})$) can be used to make primary standard solutions. Iodate(V) ions react with iodine in the presence of acid to form iodine:



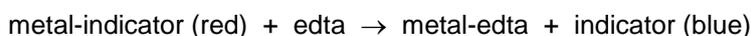
1.015 g of potassium iodate(V) are dissolved and made up to 250.0 cm^3 . To a 25.0 cm^3 portion are added an excess of potassium iodide and dilute sulphuric acid. The solution is titrated with a solution of sodium thiosulphate, with starch solution being used as the indicator. 29.80 cm^3 of thiosulphate solution are required. Calculate the concentration of the thiosulphate solution.

Question Six:

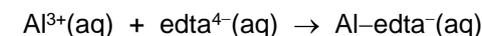
The formula of the compound bis[bis(carboxymethyl)amino]ethane is given below:



This complex molecule is often referred to as edta (an abbreviation of its old name). Edta forms stable complexes with a variety of different metal ions. Consequently, edta can be used to determine the concentration of metal ions by titration. The end-point of the titration is shown by an indicator which forms a coloured complex with the metal ion being titrated. If Eriochrome Black T is used as the indicator, the metal indicator colour of red is seen at the beginning of the titration. As the titrant is added, the metal ions are removed from the indicator and complex with the edta. At the end point, the colour of the free indicator (blue) is observed:



Find the value of n in the formula $\text{Al}_2(\text{SO}_4)_3 \cdot n\text{H}_2\text{O}$ from the following analysis. 2.000 g of aluminium sulphate hydrate was weighed out and made up to 250.0 cm^3 . A 25.0 cm^3 portion was allowed to complex with edta by being boiled with 50.0 cm^3 of edta solution of concentration $1.00 \times 10^{-2} \text{ mol dm}^{-3}$. The excess of edta was determined by adding Eriochrome Black T and titrating against a solution of $1.115 \times 10^{-2} \text{ mol dm}^{-3}$ zinc sulphate. 17.90 cm^3 of zinc sulphate solution were required to turn the indicator from blue to red. The reactions taking place are:



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http://www.chemist.sg/mole/assignments/mole_seven_ans.pdf