SECOND-YEAR HL CHEMISTRY

Final Review: Electrochemistry

1. N03
2. M03
3. N02/4

A half-cell (A) is set up by placing a platinum electrode in a solution containing both Fe^{2+} and Fe^{3+} ions at a concentration of 1 mol dm^{-3}. This half-cell is then connected by means of a salt bridge to another half-cell (B) containing an iron electrode in a 1 mol dm^{-3} solution of Fe^{2+} ions.

(a) State the function of the salt bridge. *[1]*

(b) The two electrodes are connected externally.

(i) Use Table 15 of the Data Booklet to determine the cell standard electrode potential. *[2]*

(ii) Give the redox reactions that occur in each half-cell. *[2]*

A: .....

B: .....

(iii) State the direction of the electron flow in the external circuit. *[1]*

4. M02/ 6. (a) The electrolysis of sodium chloride can be carried out with the sodium chloride in the molten state or as a concentrated solution in water.

(i) Sketch a diagram showing how molten sodium chloride could be electrolysed in a beaker. Show clearly the polarity of the electrodes, the direction of the electron flow in the connecting wires and the products at each electrode. Write equations to show the formation of the product at each electrode. [5]

(ii) When aqueous sodium chloride solution is electrolysed, the product at one of the electrodes is different. Name this product and explain why it is formed instead of the product in (i). [3]

(iii) In the electrolysis of molten sodium chloride, a current of 5 A is passed for 1 hour. Calculate the mass of the product formed at the negative electrode. [3]

(b) Ethanedioate ions, $C_2O_4^{2-}$, can be oxidised to carbon dioxide by acidified dichromate(VI) ions.

(i) Deduce the half-equation for the oxidation of ethanedioate ions. Deduce the oxidation number of carbon in ethanedioate ions and in carbon dioxide, and use your values to explain why carbon is oxidised in this reaction. [4]

(ii) Using information from the Data Booklet, write the equation for the reaction between ethanedioate ions and dichromate(VI) ions in acid solution. [2]

(c) An electrochemical cell is constructed from two half-cells connected by a high-resistance voltmeter. One half-cell contains nickel in a solution of nickel nitrate, and the other contains silver in silver nitrate solution.

(i) State the conditions which must apply to the solutions for the measurements made to be described as standard. [2]

(ii) Outline how the two half-cells must be connected before any voltage readings can be made. [2]

(iii) Assuming that standard conditions apply, calculate the cell potential using information from the Data Booklet. Write the shorthand notation for the cell, including state symbols, and give the equation for the reaction occurring in the cell. [4]

5. N01/4. A blue aqueous solution of copper(II) sulfate,CuSO4, is electrolysed using copper electrodes.

(a) Write balanced half-equations, with state symbols, for the reactions occurring at the:

anode (positive electrode): ..... [1]

cathode (negative electrode): ..... [1]

(b) State whether or not the colour of the solution will change as the electrolysis proceeds. Explain your answer. [1]

(c) Write a balanced chemical equation for the products formed if the copper anode is replaced by a graphite anode. [2]

(d) State whether or not the colour intensityand pH ofthe solution will change asthe electrolysis proceeds in (c) above. [2]

(e) Calculate the mass of copper produced when a current of 0.180 amperes is passed through a 1.0 moldm−3 copper sulfate solution for 20 minutes 10 seconds. [3]

6. M01/7. (a) Redox equations may be balanced using changes in oxidation number. For the following redox equation calculate the oxidation number of manganese and carbon. Use these values to balance the equation.

$$MnO_4^-(aq) + C_2O_4^{2-}(aq) + H^+(aq) \rightarrow Mn^{2+}(aq) + CO_2(g) + H_2O(l)$$

[2]

(b) (i) Draw a cell diagram for the cell formed by connecting the following standard half-cells:

$$Ni(s)\,/\,Ni^{2+}(aq) \,\|\, Cd^{2+}(aq)\,/\,Cd(s)$$

(ii) Describe the key features of the standard hydrogen electrode. [3]

(c) Given:

$$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s) \qquad E^{\ominus} = -0.2\ V$$

$$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s) \qquad E^{\ominus} = -0.4\ V$$

(i) Write an equation for the reaction in each half-cell, identify the species which is oxidised and the oxidising agent. [4]

(ii) On the diagram of this cell drawn in (b) (i), label the anode (A), and show, with an arrow, the direction of electron flow in the external circuit. [2]

(iii) For the overall cell, calculate its voltage and state the sign of .G. [2]

(d) An aqueous solution of silver nitrate is electrolysed. Predict the product formed at each electrode. [2]

(e) A membrane cell is used to electrolyse aqueous sodium chloride. Hydrogen and sodium hydroxide are produced according to the following equation:

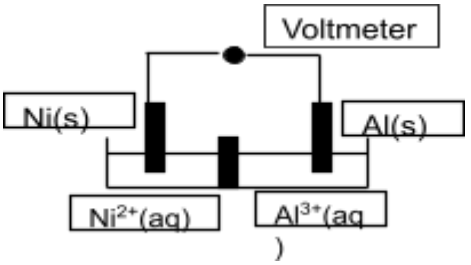
$2H2O(l) + 2e^- \rightarrow H2(g) + 2OH^-(aq)$

A current of 20 A is passed through the solution for 5 hours. Calculate the number of moles of OH^- produced, and the mass of sodium hydroxide formed. [4]

7. M or N 00/4. (a) [1] Define a reducing agent in terms of electrons.

(b) [2] In the following representation of a cell, label each electrode with a + or a - sign, as appropriate, and draw an arrow on the connecting wire to indicate the direction of electron flow. (Refer to Table 15 of the Chemistry Data booklet.)

(c) (i) [2] Write the balanced equation for the spontaneous reaction



in the cell.

(ii) [2] Calculate the standard cell potential.

(d) [5] Chromium is deposited from an acidic solution containing the dichromate(VI) ion, according to the equation

$Cr_2O_7^{2-}(aq) + 14\ H^+(aq) + 12\ e^- \rightarrow 2\ Cr(s) + 7\ H_2O(l)$

How many moles of Cr(s) can be deposited by a current of 8.00 A flowing for 2.00 hours.

8. 98/7. (a) [5] Define oxidation in terms of electrons.

Analyse what happens to chromium in the following two changes, giving your reasons:

(i) $Cr_2O_7^{2-} + 14\ H^+ + 6\ e^- \rightarrow 2\ Cr^{3+} + 7\ H_2O$;

(ii) $2\ CrO_4^{2-} + 2\ H^+ \rightarrow Cr_2O_7^{2-} + H_2O$

(b) [10] Explain how standard electrode potentials can be used to predict the spontaneity of a reaction. (I) using standard electrode potentials select a halogen that can be released from a halogen acid by dichromate(VI) solution and one that can not.

(i) Give your reasons and write a balanced equation for the release of the halogen.

(ii) It has been suggested that dichromate(VI) solutions are unstable with respect to decomposition to oxygen and Cr^{3+}. Discuss this in the light of relevant standard electrode potentials and comment on the stability of dichromate solutions.

(c) [5] A cell is constructed by joining a half cell consisting of a copper electrode in a solution of Cu^{2+} ions, via a salt bridge, to a half cell consisting of an inert electrode (Pt) in a solution containing Fe^{2+} and Fe^{3+} ions. There is a wire connecting the Pt to the Cu electrode. What is the cell reaction? Describe what happens in the cell paying particular attention to the reactions at electrodes, the wire and the ions in solution.

(d) [5] Metals are often protected by a layer of chromium which is deposited by electrolysis. The details are complex but the main source of chromium is Cr(VI). Calculate how many coulombs of electricity are required to deposit 1 g of chromium.

9. 95/8. This question deals with electrochemistry.

(a) (i) [2 *marks*] Define what is meant by a standard electrode potential.

(ii) [2 *marks*] Write a balanced equation for the reaction that is used as the reference point for the redox potential series and give the value of the potential for this reaction.

(iii) [4 *marks*] Describe briefly an experimental setup that could be used to measure a standard electrode potential for a metal that has just been isolated for the first time. Identify any other materials needed. (A sketch may be helpful.)

(b) For the hypothetical reaction: $2A(s) + B^{2+}(aq) \rightarrow 2A^+(aq) + B(s)$

(i) [2 *marks*]Describe how standard electrode potentials could be used to predict whether this reaction or the reverse reaction is spontaneous under standard conditions. Explain each prediction.

10. 92/3. You are given the following standard reduction potentials:

I $Ni^{2+}(aq) + 2e \square Ni(s)$ $E^o = -0.25$ volt

II $I_2(aq) + 2e \square 2I^-(aq)$ $E^o = +0.54$ volt

III $Mn^{2+}(aq) + 2e \square Mn(s)$ $E^o = -1.19$ volt

(a) [1 mark] Write a balanced equation for the overall reaction which will occur spontaneously when electrodes I and II are combined.

(b) [2 marks] When electrodes I and III are combined, what is the direction of movement of electrons? Explain your response.

(c) [1 mark] Calculate E^{o}_{cell} for an electrochemical cell formed using electrodes II & III.

(d) [3 marks] Draw and label a diagram of the electrochemical cell from part (c). Indicate the anode and cathode, and the direction of electron and ion flows (in the cell itself and in the salt bridge).

(f) [1 mark] Identify the strongest reducing agent from the species whose reduction potentials are given, and explain your answer.

(g) [1 mark] How many electrons does each molecule of the oxidizing agent gain?