CHEM 3.7 Practise Electrochemistry Questions

a) Using the following reduction potentials discuss which species would undergo spontaneous reaction with **bromine, Br2**, under standard conditions. Justify your answer and write balanced equations for the reactions occurring.

*E*°(Br2/Br–) = +1.07 V, *E*°(Sn4+/Sn2+) = +0.15 V,

*E*°(Cu2+/Cu) = +0.34 V, *E*°(H2O2/H2O) = +1.77 V

**To undergo spontaneous reaction the bromine would need to be reduced. This means it would need to react with a reductant in the half-cell with a more negative reduction potential ie. it could react with both Sn2+ and also with Cu but not with H2O.**

**For reaction with Cu the reactions occurring are:**

 **Cu→ Cu2+ + 2e- and Br2 + 2e- → 2Br- and overall Cu + Br2 → Cu2+ + 2Br-**

**The *Eocell* for this reaction = 1.07 V – 0.34 V = +0.73 V and the positive value shows the reaction is spontaneous under standard conditions.**

**Reaction with Sn2+ is:**

**Sn2+ → Sn4+ + 2e- and Br2 + 2e- → 2Br- and overall Sn2+ + Br2 → Cu2+ + Sn4+**

**The *Eocell* for this reaction = 1.07 V – 0.15 V = +0.92 V which again is positive indicating a spontaneous reaction will occur.**

***Eocell* for reaction between Br2 and water = 1.07 V – 1.77 V = -0.70 V. Because the value is negative reaction will not occur.**

b) Cobalt is a transition metal that exists in both the +2 and +3 oxidation states.

 A piece of cobalt metal is reacted with acidified **potassium dichromate** solution.

 Using the relevant reduction potentials, determine if the cobalt ion produced in this reaction is Co2+ or Co3+.

*E*°(Co3+/Co2+) = +1.82 V *E*°(Co2+/Co) = –0.28 V *E*°(Cr2O72-/Cr3+) = +1.36 V

**The first reaction will be between C2O72- (reduced) and Co (oxidised). The**

***Eocell*  = 1.36 V – (-0.28) V = 1.64 V**

**Since the *Eocell*  is positive the reaction will occur and Co2+ will be produced.**

**The second reaction will be between C2O72- (reduced) and Co2+ (oxidised).**

***Eocell*  = 1.36 V – 1.82 V = - 0.46 V**

**Since the *Eocell*  is negative the reaction not will occur and Co3+ will not be produced.**

c) The basis for all batteries is electrochemical cells. Below is a partly completed drawing of the cell: Fe⏐Fe2+⏐⏐Sn2+⏐Sn

*E*°(Sn2+/Sn) = –0.14 V and *E*°(Fe2+/Fe) = –0.44 V



(i)On the diagram above

complete the cell including the external circuit,

fully label each half-cell (including charge of each electrode),

identify where oxidation and reduction occur

give balanced equations for the half-reactions occurring at each electrode

indicate the direction of flow of charged particles in both the internal and external circuits

(ii) Discuss the features that you have drawn on the cell-diagram above, using the electrode potentials to support your answer.

**The half cell on the right is where reduction occurs, it is the cathode. This is because the *Eo*(Sn2+/Sn) half cell has a more positive reduction potential. It is the electrode that gains electrons hence the direction of electron flow in the external circuit is towards the Sn electrode. To balance the charge in this cell positive ions move into this half-cell from the salt bridge. The more negative *Eo* for Fe2+/Fe is where oxidation occurs and electrons are lost. To balance the loss of negative charge nitrate ions move through the salt bridge into this half-cell.**