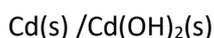


## Galvanic cells worksheet 3

- 1) The rechargeable, alkaline, nickel-cadmium cell is used to power small appliances such as portable computers. When the cell is being used, the electrode reactions are represented by the two half cells shown below.

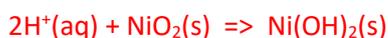


- a) Write the balanced half equations for the anode and the cathode.

At the cathode



Balanced for oxygen so now balance for hydrogen by adding  $\text{H}^+(\text{aq})$



Balance for charge by adding electrons to the most positive side



Add  $\text{OH}^-(\text{aq})$  to replace the  $\text{H}^+(\text{aq})$



At the anode



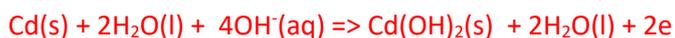
Balanced for oxygen by adding water



Balance for hydrogen by adding  $\text{H}^+(\text{aq})$



Add  $\text{OH}^-(\text{aq})$  to replace the  $\text{H}^+(\text{aq})$



Cancel two  $\text{H}_2\text{O}$  from both sides



- b) Write the overall equation



- c) What can you say about the pH of the electrolyte as the battery is discharging.

The pH does not change. The number of  $\text{OH}^-$  ions consumed at the cathode is equal to the number of  $\text{OH}^-$  ions produced at the anode.

- d) During recharging, what happens to the mass of the Cd electrode? Explain why.

Increases in mass due to the reversal of the half equation that takes place during discharge



- e) In the nickel cadmium cell, the reactive ingredients are held tightly in small pockets in an inert grid of nickel plated steel. The net result of this technique is that the reactive compounds stay where they belong and the cell lasts much longer. Explain why.

In order for a cell to be recharged the products must stay in contact with the electrodes. Over the course of the life of the battery electrodes are continuously being broken down

and reformed. This causes material to flake off and lose contact with the electrode as happens over the life of a lead acid accumulator.

- f)  $\text{NiO}_2$ , on its own, cannot be used as an electrode. As a consequence the electrode is composed of nickel plated steel impregnated with  $\text{NiO}_2$ . Explain why.

$\text{NiO}_2$  is an ionic substance and does not conduct electrons in the solid state.

- 2) Below is a diagram showing part of the electrochemical series.

$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+0.80
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.77
$\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{H}_2\text{O}_2(\text{aq})$	+0.68
$\text{I}_2(\text{s}) + 2\text{e}^- \rightleftharpoons 2\text{I}^-(\text{aq})$	+0.54
$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightleftharpoons 4\text{OH}^-(\text{aq})$	+0.40
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	+0.34
$\text{Sn}^{4+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Sn}^{2+}(\text{aq})$	+0.15

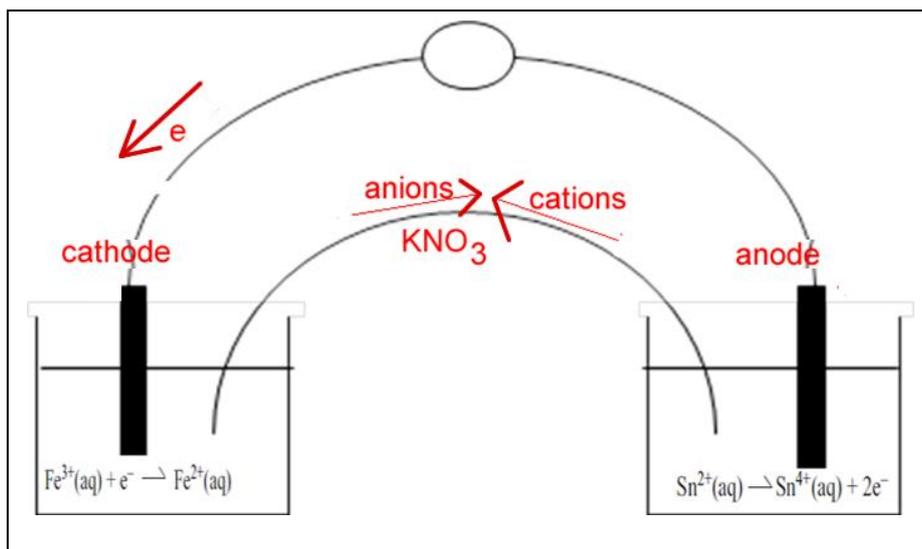
- a) The  $E^\circ$  values shown above are measured at standard conditions. What are the standard conditions that  $E^\circ$  values for each half cell are measured at?

25 °C and 1.00 M solution

- b) Using the diagram below as a template construct a galvanic cell that will deliver 0.62V under standard conditions.

Clearly label or give the following

- i. The direction of electron flow
- ii. The anode and cathode
- iii. The polarity of each electrode
- iv. The material each electrode is made up of. Carbon
- v. The substance forming the salt bridge ( $\text{KNO}_3$ )
- vi. Direction of anion and cation movement
- vii. The oxidant  $\text{Fe}^{3+}(\text{aq})$
- viii. The reductant  $\text{Sn}^{2+}(\text{aq})$
- ix. The oxidation half equation
- x. The reduction half equation
- xi. Overall equation  $2\text{Fe}^{3+}(\text{aq}) + \text{Sn}^{2+}(\text{aq}) \Rightarrow \text{Sn}^{4+}(\text{aq}) + 2\text{Fe}^{2+}(\text{aq})$



- 3) A galvanic cell consists of one half cell that is made up of an inert graphite electrode in a solution containing 1.0 M  $\text{Sn}^{2+}(\text{aq})$  and 1.0 M  $\text{Sn}^{4+}(\text{aq})$  at 25°C. Which one of the following could be used as the second half cell so that the polarity of the electrode in this second half cell is positive?
- a lead electrode in a solution of 1.0 M  $\text{Pb}^{2+}(\text{aq})$
  - a silver electrode in a solution of 1.0 M  $\text{Ag}^+(\text{aq})$
  - A cobalt electrode in a solution of 1.0 M  $\text{Co}^{2+}(\text{aq})$
  - an inert graphite electrode in a solution of 1.0 M  $\text{Br}^-(\text{aq})$

In order for the second cell to have a positive electrode it must form the cathode. It must therefore host the reduction reaction and hence have an oxidant that is above  $\text{Sn}^{4+}$  on the electrochemical series, in other words, an oxidant stronger than  $\text{Sn}^{4+}$ .

$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}(\text{l})$	i. $\text{Pb}^{2+}$ is a weaker oxidant than $\text{Sn}^{4+}$
$\text{Br}_2(\text{l}) + 2\text{e}^- \rightleftharpoons 2\text{Br}^-(\text{aq})$	ii. $\text{Ag}^+$ is a stronger oxidant than $\text{Sn}^{4+}$ and hence will form a half-cell with a positive electrode.
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	iii. $\text{Co}^{2+}$ is a weaker oxidant than $\text{Sn}^{4+}$ and hence will form a half cell with a negative electrode.
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	iv. This is not an acceptable half-cell. A half-cell must contain an appropriate electrode an oxidant and its conjugate reductant, as shown on the electrochemical series.
$\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{H}_2\text{O}_2(\text{aq})$	
$\text{I}_2(\text{s}) + 2\text{e}^- \rightleftharpoons 2\text{I}^-(\text{aq})$	
$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightleftharpoons 4\text{OH}^-(\text{aq})$	
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	
$\text{Sn}^{4+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Sn}^{2+}(\text{aq})$	
$\text{S}(\text{s}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{H}_2\text{S}(\text{g})$	
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{H}_2(\text{g})$	
$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Pb}(\text{s})$	
$\text{Sn}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Sn}(\text{s})$	
$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Ni}(\text{s})$	
$\text{Co}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Co}(\text{s})$	