Galvanic cells - revision

1. Consider the following overall redox equations

   a. $Pb(s) + 2Fe^{3+}(aq) \rightarrow Pb^{2+}(aq) + 2Fe^{2+}(aq)$

   Oxidant ______*Fe^{3+}*__

   Reductant ___*Pb(s)*_____

   Oxidation reaction _____*$Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$*__________

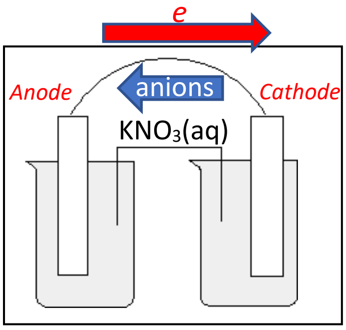
   Reduction reaction _____*$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$* ____

   Material anode is made from _____*Pb*________

   Material cathode is made from *C or an inert metal*

   Write the reactions taking place in each half cell and label the following on the image on the right.

   - Anode
   - Cathode
   - Direction of electron flow
   - Direction of negative ion flow.



$Pb(s) \rightarrow Pb^{2+}(aq) + 2e$ *$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$*

   b. $Fe(s) + 2Fe^{3+}(aq) \rightarrow 3Fe^{2+}(aq)$

   Oxidant ___*$Fe^{3+}(aq)$*_____

   Reductant ___*Fe(s)*_______

   Oxidation reaction __ *$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$*___

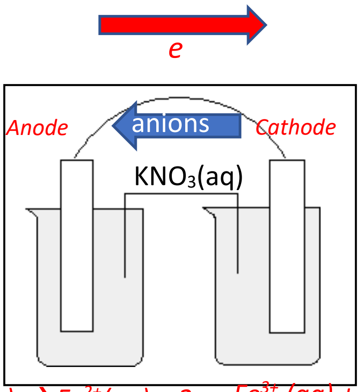
   Reduction reaction __ *$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$* ___

   Material anode is made from _______*Fe(s)*_______

   Material cathode is made from *C or an inert metal*

   Write the reactions taking place in each half cell and label the following on the image on the right.

   - Anode
   - Cathode
   - Direction of electron flow
   - Direction of negative ion flow.



$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ *$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$*

   c. $6Fe^{2+}(aq) + Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow 3Cr^{3+}(aq) + 7H_2O(l) + 6Fe^{3+}(aq)$

   The balanced overall equation, shown above, takes place in an acidic environment in a galvanic cell with the design shown on the right.

   Oxidant *$Cr_2O_7^{2-}(aq)$*

   Reductant ____ *$Fe^{2+}(aq)$* __

   Oxidation reaction ___ *$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$*

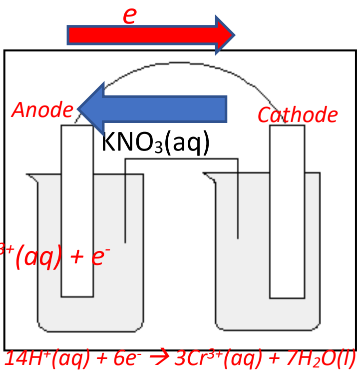
   Reduction reaction *$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 3Cr^{3+}(aq) + 7H_2O(l)$*

   Material anode is made from *C or an inert metal such as Pt*

   Material cathode is made from *C or an inert metal such as Pt*

   Write the reactions taking place in each half cell an label the following on the image on the right.

   - Anode
   - Cathode
   - Direction of electron flow



$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$

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

2. Consider the galvanic cell shown on the right.

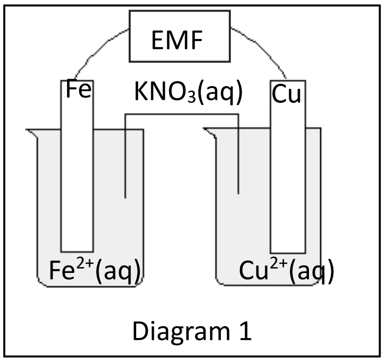


Diagram 1

   a. Calculate the theoretical voltage if all electrolytes are at 1 M and 25 °C

   _______EMF = +0.34 – -0.44 = 0.78V___________________

   ____

   ____

   b. When the two half cells are connected no observable reaction takes place. Offer an explanation for this observation.

   There is a very slow rate of reaction

   ____

   ____

   c. **Explain** the possible outcomes when the following changes to the galvanic cell in diagram 1 are made.

      i. The copper electrode is replaced with a zinc electrode.

   There will be a redox reaction taking place in the half cell with the zinc electrode. Heat will be released as the following overall reaction takes place. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

   ____

      ii. The iron electrode is replaced with a carbon (graphite) electrode.

   ____

   No reaction will occur as the reductant, Fe(s) is removed.

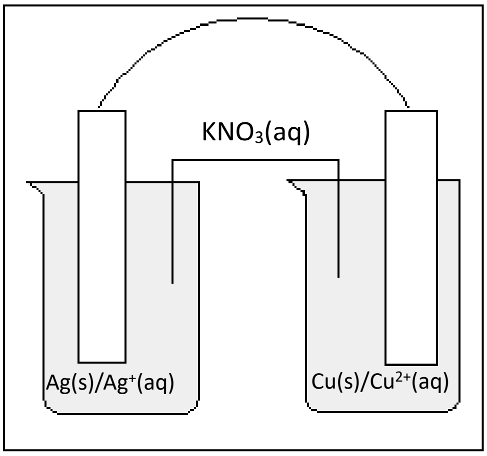
      iii. The KNO_3 present in the salt bridge is replaced with a solution of $Mg(NO_3)_2$.

   No observable difference to the overall cell reaction and to the cell EMF. The $Mg^{2+}(aq)$ *is very weak oxidant and will not take part in any side reactions.*

   ____

3. On the next page is a range of labelled galvanic cells. Complete the questions for each cell.

Solutions are shown on the video .



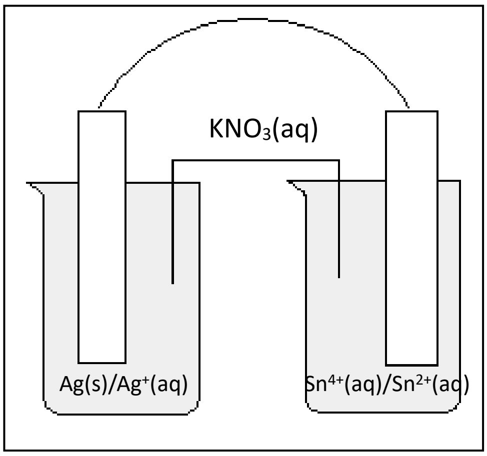
Show the following on the diagram.

Anode (-)
Cathode (+)
Electron flow.
Anion flow.
Oxidation half equation____________________
Reduction half equation____________________

Overall equation __________________________

What material is the anode made from? __________

What material is the cathode made from? _________



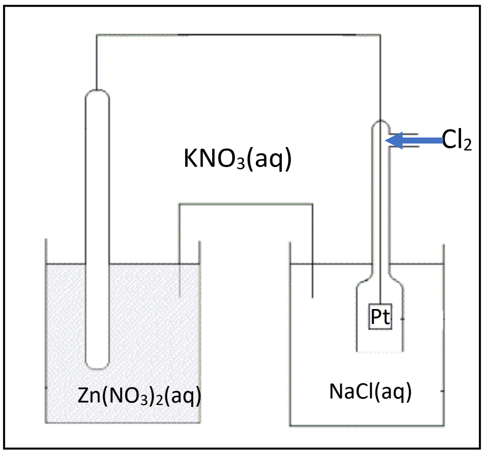
Show the following on the diagram.

Anode (-)
Cathode (+)
Electron flow.
Anion flow.
Oxidation half equation____________________
Reduction half equation____________________

Overall equation __________________________

What material is the anode made from? __________

What material is the cathode made from? _________



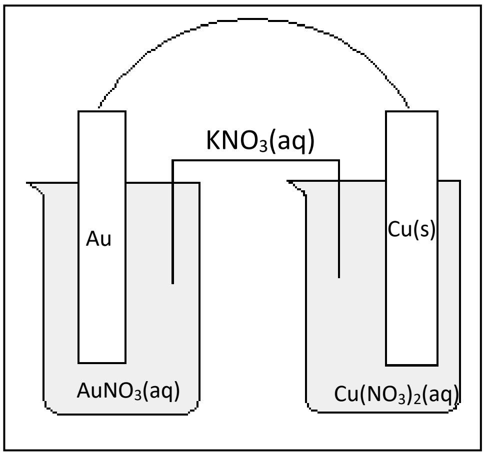
Show the following on the diagram.

Anode (-)
Cathode (+)
Electron flow.
Anion flow.
Oxidation half equation____________________
Reduction half equation____________________

Overall equation __________________________

What material is the anode made from? __________

What material is the cathode made from? _________



Show the following on the diagram.

Anode (-)
Cathode (+)
Electron flow.
Anion flow.
Oxidation half equation____________________
Reduction half equation____________________

Overall equation __________________________

What material is the anode made from? __________

What material is the cathode made from? _________

4. The following overall redox reaction occurs during the discharge of a primary cell using an acidic electrolyte.

$$MnO_4^{2-}(aq) + 4H^+(aq) + Zn(s) \rightarrow MnO_2(s) + Zn^{2+}(aq) + 2H_2O(l)$$

a. Identify the:

i. Oxidant ________ *$MnO_4^{2-}(aq)$*_____________

ii. reductant ___________ *$Zn(s)$*__________________________

b. Write a balanced ionic equation, states included, for the oxidation reaction.

_________________________ *$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$* _____________________________________

c. Write a balanced ionic equation, states included, for the reduction reaction.

_____________ *$MnO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow MnO_2(s) + 2H_2O(l)$*

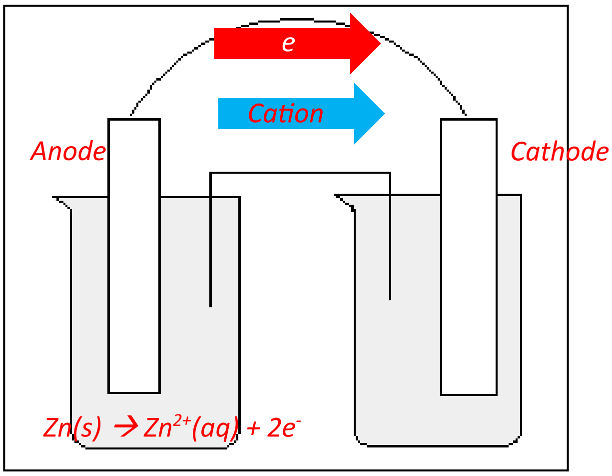
d. In the diagram below write the reactions taking place in each half cell and correctly label the following:

i. Anode

ii. Cathode

iii. Direction of electron flow.

iv. Direction of **cation**, flow



$MnO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow MnO_2(s) + 2H_2O(l)$

e. Give one possible material which could be used to make the

i. anode

_________*Zn*___________

ii. Cathode

________*C or Pt*__

f. Is it possible to use MnO_2 as an electrode? Justify your answer.

________*No, ionic compounds do not conduct electricity.*_____________

____

____

g. Suggest the possible changes in the mass of each electrode by circling one of the the three alternatives given. Offer an explanation to support your choice.

i. anode . ***Increase,*** ***stay the same ,*** [✓ ***decrease***]

_________________ *$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$* _________________________________

____

ii. cathode. [✓ ***Increase,***] ***stay the same ,*** ***decrease***

___________________*Deposition of $MnO_2(s)$*___________________________

____

h. A research scientist suggested that an alkaline electrolyte be used for the same battery to improve the shelf life of the battery. Write the balanced equation, states included, for the overall reaction taking place in the new **alkaline** version.

$$MnO_4^{2-}(aq) + 2H_2O(l) + Zn(s) \rightarrow MnO_2(s) + Zn^{2+}(aq) + 4OH^{-}(aq)$$

i. Given that the cell EMF is 1.36 V, in an acidic electrolyte, calculate the standard electrode potential (E^O), in volts, of the half cell MnO_4^- (aq)/MnO_2(s). Place the reaction and its E^0 in the appropriate position in the table below. Two possible locations are highlighted for you.

Reaction	E^o
$MnO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow MnO_2(s) + 2H_2O(l)$	+0.60
$2H^+(aq) + 2e^- \rightarrow H_2^{(g)}$	0.00
$Zn^{2+}(aq) + 2e \rightarrow Zn(s)$	-0.76

EMF = $E^o_{oxidant} - E^o_{reductant}$

=> 1.36 – 0.76 = $E^o_{oxidant}$ = + 0.60 V