Lesson 11 –Revision of galvanic cells

1) Consider the galvanic cell shown on the right operating at standard conditions.

a) Identify the :

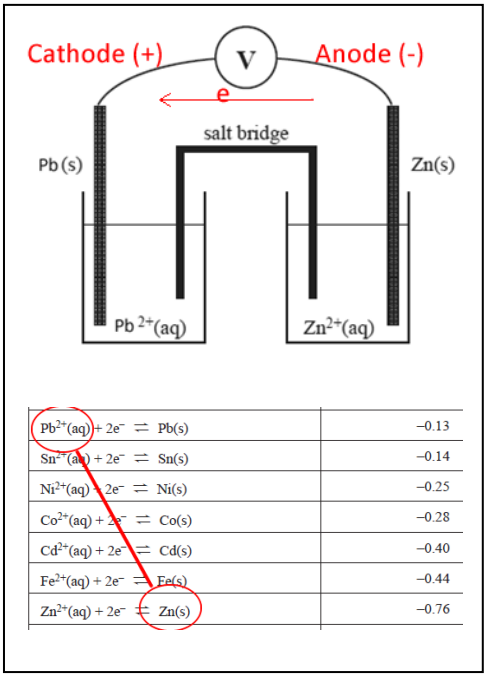
- oxidant = *$Pb^{2+}(aq)$*
- reductant = *Zn(s)*

Calculate the cell voltage.

-0.13 - - 0.76 = 0.66 V

b) On the diagram, label the :

- Anode *(oxidation occurs at the anode and is where the reductant is)* .
- Cathode *(reduction occurs at the cathode and is where the oxidant is)*



$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.25
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28
$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$	−0.40
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.44
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76

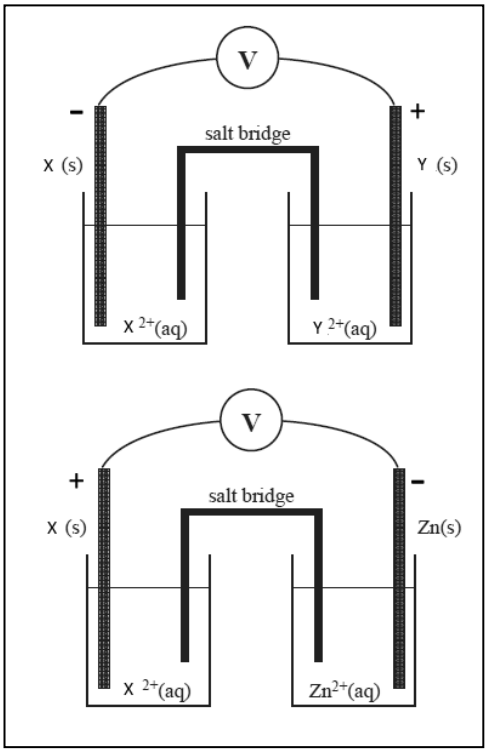
Consider the two galvanic cells shown on the right operating at standard conditions.

a) Which is the weakest oxidant out of the following species? Explain your answer with reference to the cells.

i) X(s), $X^{2+}(aq)$, Y(s), $Y^{2+}(aq)$, Zn(s), $Zn^{2+}(aq)$

b) Which is the weakest reductant out of the following species? Explain your answer with reference to the cells.

i) X(s), $X^{2+}(aq)$, Y(s), $Y^{2+}(aq)$, Zn(s), $Zn^{2+}(aq)$



The information on the right allows us to place the following half-cell reactions in the order they would appear in the electrochemical series(if placed in order of decreasing E^o values)

$Y^{2+}(aq) + 2e \rightarrow Y(s)$

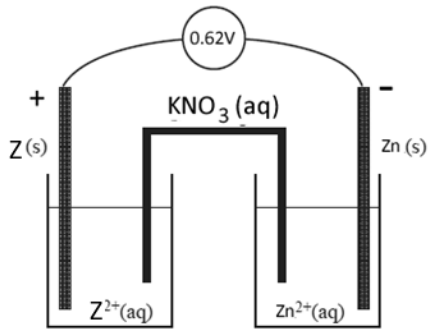
$X^{2+}(aq) + 2e \rightarrow X(s)$

$Zn^{2+}(aq) + 2e \rightarrow Zn(s)$

Hence, weakest oxidant is $Zn^{2+}(aq)$

While the weakest reductant is Y(s)

c) Consider the galvanic cell shown on the right.



$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$	+0.68
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(g)$	+0.14
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.23
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.44
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76

i. Identify the:

- anode *=Zn(s)*
- cathode *= Z(s)*
- direction of electron flow

from anode to cathode

- direction of cation flow.

Into the half-cell containing the cathode (Z)

ii. The salt bridge is initially composed of $KNO_3(aq)$, however, it is soon changed to $Fe(NO_3)_3(aq)$. Explain what changes, if any, occur in the galvanic cell and give the cell voltage.

$Fe^{3+}(aq)$ is now the strongest oxidant and will react directly with the Zn(s).
No longer will electrical energy be produced.
Heat energy is now produced as the following reaction occurs directly in the Zn/Zn^{2+} half-cell.
$2Fe^{3+}(aq) + Zn(s) \rightarrow 2Fe^{2+}(aq) + Zn^{2+}(aq)$.
No cell voltage

iii. Identify Y.

Using the E^o table can calculate the following
E^o oxidant – E^o reductant = cell voltage
Z^{2+} - Zn(s) = 0.62V
=> $Z^{2+} = 0.62 + E^o_{Zn} = 0.62 + (-0.76) = - 0.14$ V
=> $Z^{2+} + 2e \rightarrow Z(s)$ ---- $E^o = -0.014$V
=> Y = Sn