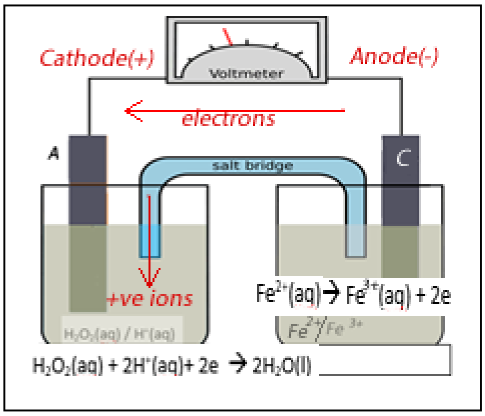
Quiz 6 – revision of redox and galvanic cells.




1. Consider the galvanic cell shown on the right.
   a. What is the theoretical EMF produced by the cell if it is run at standard conditions. *+1.77 - -0.77 = 2.54 V*
   b. Write the half equations to each half-cell in the space provided.
   c. On the diagram label the:
      i. Direction of electron flow.
      ii. Direction of positive ions flow from the salt bridge
      iii. Anode and give its polarity
      iv. Cathode and give its polarity.
   d. What will happen to the mass of each electrode as the galvanic cell is allowed to operate. Explain you answer.
      *No change in the mass of electrode "A" but a decrease in the mass of the anode (Fe)*
   e. What material should electrode "A" be made from? *Carbon or platinum*

2. Below are two redox equations. For each equation identify the :
   - atom being reduced (justify your answer using oxidation numbers)
   - atom being oxidised (justify your answer using oxidation numbers)
   - oxidant
   - reductant

a. $IO_3^-(aq) + 2H_2O_2(aq) + H^+(aq) \rightarrow HOI(aq) + 2O_2(g) + 2H_2O(l)$

*The iodine atom (I) in $IO_3^-$ has an oxidation state of +5 and is reduced to +1 in HOI. While the oxygen atom in $H_2O_2$ has an oxidation state of -1 and is oxidised to 0 in $O_2$.*

b. $C_2H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

*The carbon atom (C) in $C_2H_6$ has an oxidation state of +3 and is oxidised to +4 in $CO_2$. While the oxygen atom in $O_2$ has an oxidation state of 0 and is reduced to -2 in both $CO_2$ and $H_2O$.*

3. Give the oxidation number of the underlined atoms in each of the following substances.
   a. $\underline{Mn}O_4$ *=> Mn + 4 X -2 = 0 => Mn = +8*
   b. $H_2\underline{O}_2$ *=> 2O + 2 X +1 = 0 => O = -1*
   c. $\underline{Mg}(OH)_2$*=> Mg + (-2 + 1 ) X 2 = 0 => Mg = +2*
   d. $\underline{Cr}O_4$*=> Cr + 4 X -2 = 0 => Cr = +8*
   e. $\underline{C}_3H_8O$*=> 3C + 8 -2 = 0 => C = +2*

4. Consider the following unbalanced redox reactions. Write balanced half equations for each, states not included, and identify each one as either an oxidation or reduction.
   a. *$2H^+ + 2e + Cr_2O_3 \rightarrow 2CrO + H_2O$ ------ reduction (electrons accepted) Cr (+3) in $Cr_2O_3$ is reduced to an oxidation state of +2 in CrO.*

   b. *$5e + 6H^+ + MnO_4^- \rightarrow MnO + 3H_2O$ ------ reduction (electrons accepted) Mn (+7) in $MnO_4^-$ is reduced to an oxidation state of +2 in MnO*

   c. *$2e + 2H^+ + H_2O_2 \rightarrow 2H_2O$ ------ reduction (electrons accepted) O (-1) in $H_2O_2$ is reduced to an oxidation state of -2 in $H_2O$*

5. Using the E° series shown on the right predict if a spontaneous reaction will occur when:
   i. Tin(Sn) metal is placed in a 0.1 M $AgNO_3$ solution. Yes
   ii. Pure manganese metal is placed in a solution of 0.2M $Al(NO_3)_3$ No reaction
   iii. Nickel metal is placed in a solution of $Fe(NO_3)_3$. Yes
   iv. Lithium metal is placed in a 1.0 M solution of $Zn(NO_3)_2$. Yes
   v. Iron metal is placed in a 0.10 M HCl solution. Yes
   vi. Magnesium metal is placed in a 1.0M HCl solution. Yes

6. For each spontaneous reaction that occurs in question 5, above, give the
   i. oxidant and reductant taking part in the reaction
   ii. Write the balanced half equations, with states
   iii. Write the balanced overall equation, with states.

Tin(Sn) metal is placed in a 0.1 M $AgNO_3$ solution.
Sn(s) is the reductant $Ag^+(aq)$ is the oxidant.
$Ag^+(aq) + e \rightarrow Ag(s)$ – reduction
$Sn(s) \rightarrow Sn^{2+}(aq) + 2e$ – oxidation
$2Ag^+(aq) + Sn(s) \rightarrow Sn^{2+}(aq) + 2Ag(s)$

| Reaction | Standard electrode ($E^\circ$) in volts at : |
|---|---|
| $F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$ | +2.87 |
| $H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$ | +1.77 |
| $Au^+(aq) + e^- \rightleftharpoons Au(s)$ | +1.68 |
| $Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$ | +1.36 |
| $O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$ | +1.23 |
| $Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$ | +1.09 |
| $Ag^+(aq) + e^- \rightleftharpoons Ag(s)$ | +0.80 |
| $Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$ | +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 |
| $2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$ | −0.83 |
| $Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$ | −1.03 |
| $Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$ | −1.67 |
| $Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$ | −2.34 |
| $Na^+(aq) + e^- \rightleftharpoons Na(s)$ | −2.71 |
| $Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$ | −2.87 |
| $K^+(aq) + e^- \rightleftharpoons K(s)$ | −2.93 |
| $Li^+(aq) + e^- \rightleftharpoons Li(s)$ | −3.02 |

Nickel metal is placed in a solution of $Fe(NO_3)_3$. . Ni(s) = reductant, $Fe^{3+}$(aq). is the oxidant

$Ni(s) \rightarrow Ni^{2+}(aq) + 2e$ – oxidation
$Fe^{3+}(aq) + e \rightarrow Fe^{2+}(aq)$ – reduction
$Ni(s) + 2Fe^{3+}(aq) \rightarrow 2\,Fe^{2+}(aq) + Ni^{2+}(aq)$

Lithium metal is placed in a 1.0 M solution of $Zn(NO_3)_2$. Li(s) = reductant, $Zn^{2+}$(aq). is the oxidant

$Li(s) \rightarrow Li^{+}(aq) + e$ – oxidation
$Zn^{2+}(aq) + 2e \rightarrow Zn(s)$ – reduction
$2Li(s) + Zn^{2+}(aq) \rightarrow 2Li^{+}(aq) + Zn(s)$

Iron metal is placed in a 0.10 M HCl solution. Fe(s) = reductant, $H^{+}$(aq) is the oxidant
$Fe(s) \rightarrow Fe^{2+}(aq) + 2e$ – oxidation
$2H^{+}(aq) + 2e \rightarrow H_2(g)$ – reduction
$Fe(s) + 2H^{+}(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$

Magnesium metal is placed in a 1.0M HCl solution. Mg(s) = reductant, $H^{+}$(aq) is the oxidant
$Mg(s) \rightarrow Mg^{2+}(aq) + 2e$ – oxidation
$2H^{+}(aq) + 2e \rightarrow H_2(g)$ – reduction
$Fe(s) + 2H^{+}(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$