Electrode potential – key ideas

- Reducing agent = e^- donor (e^- provider)
- Oxidising agent = e^- acceptor

- Two oxidising agents cannot react with each other, an oxidising agent (e- acceptor) can only react with a reducing agent (e- donor)

- A half-cell comprises an element in two oxidation states, e.g. Cu(s) and Cu^{2+} (aq), Fe^{2+}(aq) and Fe^{3+} (aq), …

- A reaction can only happen if the E^o of the cell is positive. The half-cell with the most positive E^o undergoes reduction and the one with the least positive E^o oxidation

Look at the following half equations, also called redox systems:

Redox system				$E^\ominus$ / V
1	$Mg^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Mg(s)$	–2.37
2	$Cu^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Cu(s)$	+0.34
3	$Al^{3+}(aq) + 3e^-$	$\rightleftharpoons$	$Al(s)$	–1.66
4	$Fe^{3+}(aq) + e^-$	$\rightleftharpoons$	$Fe^{2+}(aq)$	+0.77
5	$I_2(aq) + 2e^-$	$\rightleftharpoons$	$2I^-(aq)$	+0.54
6	$Cl_2(g) + 2e^-$	$\rightleftharpoons$	$2Cl^-(aq)$	+1.36
7	$ClO^-(aq) + 2H^+(aq) + e^-$	$\rightleftharpoons$	$\frac{1}{2}Cl_2(g) + H_2O(l)$	+1.63

The **oxidising agents** (e- acceptors) are all on the **left-hand side** of the equations. i.e.: Mg^{2+}, Cu^{2+}, Al^{3+}, Fe^{3+}, I_2, Cl_2, ClO^-. The **reducing agents** are on the **right-hand side** of the equations: Mg, Cu, Al, Fe^{2+}, I^-, Cl^- and Cl_2. Remember that **an oxidising agent can only react with a reducing agent**.

Notice than Cl_2 can behave as an oxidising agent (system 6) and as a reducing agent (system 7), if as an oxidising agent then it becomes Cl^- (system 6), if as a reducing agent then it becomes ClO^- (system 7)

Q) An electrochemical cell is made of systems 1 and 2. Predict the reaction that would take place.

E^o (Cu^{2+}/Cu) > E^o (Mg^{2+}/Mg) therefore the Cu half cell undergoes reduction and the Mg one oxidation, the reaction that would happen is $Cu^{2+} + Mg \rightarrow Mg^{2+} + Cu$

1. If you mix system 1 and 3, would a reaction take place? If so, write it down

2. If you mix Mg^{2+} (not the Mg-half cell which would contain Mg as well) with Al^{3+}, would a reaction take place? Justify your answer

3. If you mix Mg^{2+} with Al, would a reaction take place? Justify your answer

4. If you mix Mg with Al^{3+}, would a reaction take place? Justify your answer

Standard electrode potentials for eight redox systems are shown in **Table 6.1**.

You will need to use this information throughout this question.

redox system	half-equation	$E^{\ominus}$/V
1	$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
2	$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
3	$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
4	$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
5	$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
6	$CO_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons HCOOH(aq)$	–0.22
7	$HCOOH(aq) + 2H^+(aq) + 2e^- \rightleftharpoons HCHO(aq) + H_2O(l)$	+0.06
8	$Cr^{3+}(aq) + 3e^- \rightleftharpoons Cr(s)$	–0.74

Table 6.1

See Table 6.1 and notice that HCOOH can behave as a reducing or oxidising agent

1. A cell is made based on systems 1 and 6. Write the equation for the reaction that would take place

2. A cell is made based on systems 1 and 7. Write the equation for the reaction that would take place

3. Write the overall balanced equation when Cu^{2+} is reduced by HCOOH

4. Write the overall balanced equation when Cu is oxidised by HCOOH

5. Which species can reduce $Cr_2O_7^{2-}$ but cannot reduce Fe^{3+}? Justify your answer