

Oxidation and Reduction



Topics covered in this chapter:

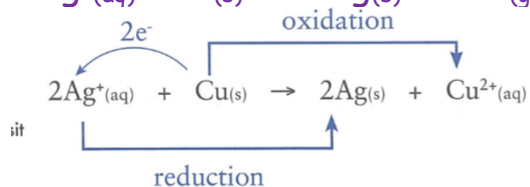
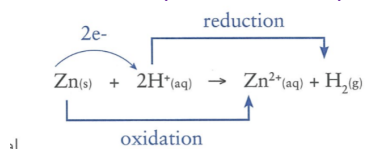
- 4.1 Oxidation and reduction
- 4.2 Identifying redox reactions
- 3.3 Balancing redox equation - using half-equations
- 3.4 Competition for electrons
- 3.5 Electrochemical cells
- 3.6 Electrolysis
- 3.7 Corrosion of metals

4.1 Oxidation and Reduction

Redox

- Redox reactions involve the transfer of electrons from one species to another.
 - Oxidation
 - Is
 - Loss of electrons
 - Reduction
 - Is
 - Gain of electrons

For example:



Oxidising and Reducing Agents

- Oxidation and reduction occur at the same time in a redox reaction.
- A redox reaction involves the transfer of electrons from the substance being oxidised to that being reduced.
- Oxidising agents or oxidants cause a substance to be oxidised. In order to do this, they have a tendency to accept electrons and are themselves reduced.
 - Common oxidising agents/oxidant are found top left hand side of the *Standard Reduction Potentials Table*

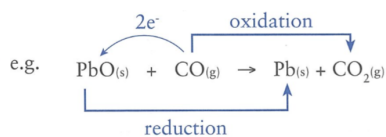
Standard Reduction Potentials at 25 °C

Half-reaction	E°(volts)
$\text{F}_2(\text{g}) + 2 \text{e}^- \rightleftharpoons 2 \text{F}^-(\text{aq})$	+ 2.89
$\text{H}_2\text{O}_2(\text{aq}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons 2 \text{H}_2\text{O}(\ell)$	+ 1.76
$\text{PbO}_2(\text{s}) + \text{SO}_4^{2-}(\text{aq}) + 4 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{PbSO}_4(\text{s}) + 2 \text{H}_2\text{O}(\ell)$	+ 1.69
$2 \text{HClO}(\text{aq}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Cl}_2(\text{g}) + 2 \text{H}_2\text{O}(\ell)$	+ 1.63
$\text{MnO}_4^-(\text{aq}) + 8 \text{H}^+(\text{aq}) + 5 \text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4 \text{H}_2\text{O}(\ell)$	+ 1.51
$\text{Au}^{3+}(\text{aq}) + 3 \text{e}^- \rightleftharpoons \text{Au}(\text{s})$	+ 1.50
$\text{HClO}(\text{aq}) + \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Cl}^-(\text{aq}) + \text{H}_2\text{O}(\ell)$	+ 1.49
$\text{PbO}_2(\text{s}) + 4 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Pb}^{2+}(\text{aq}) + 2 \text{H}_2\text{O}(\ell)$	+ 1.46
$\text{Cl}_2(\text{g}) + 2 \text{e}^- \rightleftharpoons 2 \text{Cl}^-(\text{aq})$	+ 1.36
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14 \text{H}^+(\text{aq}) + 6 \text{e}^- \rightleftharpoons 2 \text{Cr}^{3+}(\text{aq}) + 7 \text{H}_2\text{O}(\ell)$	+ 1.36
$\text{O}_2(\text{g}) + 4 \text{H}^+(\text{aq}) + 4 \text{e}^- \rightleftharpoons 2 \text{H}_2\text{O}(\ell)$	+ 1.23
$\text{Br}_2(\ell) + 2 \text{e}^- \rightleftharpoons 2 \text{Br}^-(\text{aq})$	+ 1.08
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+ 0.80
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+ 0.77
$\text{O}_2(\text{g}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{H}_2\text{O}_2(\text{aq})$	+ 0.70
$\text{I}_2(\text{s}) + 2 \text{e}^- \rightleftharpoons 2 \text{I}^-(\text{aq})$	+ 0.54
$\text{O}_2(\text{g}) + 2 \text{H}_2\text{O}(\ell) + 4 \text{e}^- \rightleftharpoons 4 \text{OH}^-(\text{aq})$	+ 0.40
$\text{Cu}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	+ 0.34
$\text{S}(\text{s}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{H}_2\text{S}(\text{aq})$	+ 0.17
$2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{H}_2(\text{g})$	0 exactly
$\text{Pb}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Pb}(\text{s})$	- 0.13
$\text{Sn}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Sn}(\text{s})$	- 0.14
$\text{Ni}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Ni}(\text{s})$	- 0.24
$\text{Co}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Co}(\text{s})$	- 0.28
$\text{PbSO}_4(\text{s}) + 2 \text{e}^- \rightleftharpoons \text{Pb}(\text{s}) + \text{SO}_4^{2-}(\text{aq})$	- 0.36
$\text{Cd}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Cd}(\text{s})$	- 0.40
$2 \text{CO}_2(\text{g}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{H}_2\text{C}_2\text{O}_4(\text{aq})$	- 0.43
$\text{Fe}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Fe}(\text{s})$	- 0.44
$\text{Cr}^{3+}(\text{aq}) + 3 \text{e}^- \rightleftharpoons \text{Cr}(\text{s})$	- 0.74
$\text{Zn}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Zn}(\text{s})$	- 0.76
$2 \text{H}_2\text{O}(\ell) + 2 \text{e}^- \rightleftharpoons \text{H}_2(\text{g}) + 2 \text{OH}^-(\text{aq})$	- 0.83
$\text{Mn}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Mn}(\text{s})$	- 1.18
$\text{Al}^{3+}(\text{aq}) + 3 \text{e}^- \rightleftharpoons \text{Al}(\text{s})$	- 1.68
$\text{Mg}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Mg}(\text{s})$	- 2.36
$\text{Na}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Na}(\text{s})$	- 2.71
$\text{Ca}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Ca}(\text{s})$	- 2.87
$\text{Sr}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Sr}(\text{s})$	- 2.90
$\text{Ba}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Ba}(\text{s})$	- 2.91
$\text{K}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{K}(\text{s})$	- 2.94

- Reducing agents or reductants cause a substance to be reduced. They have a tendency to donate electrons and are themselves oxidised.

- Common reducing agents/reductants are found on the bottom right hand side of the *Standard Reduction Potentials Table*

For example



PbO:

Accepts electrons

Is reduced to Pb

Is the oxidant

CO:

Donates electrons

Is oxidised to CO_2

Is the reductant

Oxidation numbers

- Oxidation numbers are a convenient way of determining if a substance has been oxidised or reduced.

RULES FOR ASSIGNING OXIDATION NUMBERS	
Species	Oxidation Number
1. Atoms in their elemental/pure state	0
2. Monatomic ions (Group 1 metals in combined state) (Group 2 metals in combined state)	Charge on the ion +1 +2
3. Oxygen in combined state Exception 1: peroxides e.g. H_2O_2 Exception 2: F_2O	-2 -1 +2
4. Hydrogen in combined state Exception: metal hydrides e.g. NaH	+1 -1
5. Polyatomic species - the sum of the oxidation numbers	Charge on the ion

Worked Example:

Determine the oxidation number of Mn in Mn_2O_3 and MnO_4^-



4.1 Oxidation and reduction Summary Questions

Question 1

The burning of magnesium metal in chlorine gas results in the formation of $\text{MgCl}_{2(s)}$.

- (a) Write the equation for this reaction.
- (b) Which species loses electrons?
- (c) Which species is oxidised?
- (d) Which species is the oxidant?

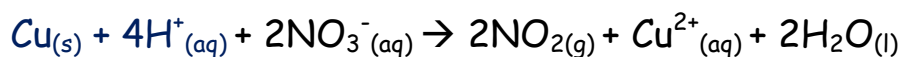
Question 2

Determine the oxidation number of each element in the following:



4.2 Identifying Redox Reactions

- By assigning oxidation numbers to all the atoms involved in a reaction, we can determine if oxidation or reduction has occurred.
- Oxidation - an increase in oxidation number
- Reduction - a decrease in oxidation number

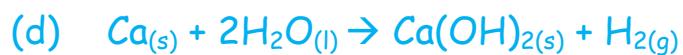
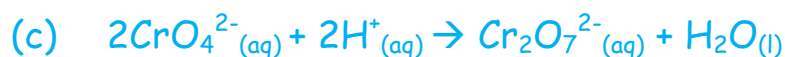
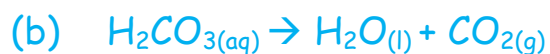
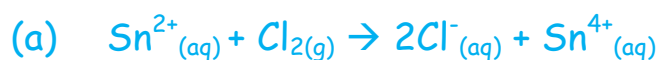


- Cu has been oxidised therefore Cu is the reducing agent/reductant.
- N has been reduced therefore NO_3^{-} is the oxidising agent/oxidant.
- All other species have retained the same oxidation number; hence they have neither been reduced nor oxidised.
- This is a redox reaction as both oxidation and reduction have taken place.

4.2 Identifying Redox Reaction Summary Questions

Question 1

Assign oxidation numbers to the following to determine if they are redox reactions?



3.3 Balancing Redox Equations – Using Half Equations

Half-Equations

- It is often more convenient to consider a redox reaction by writing two half equations.
- One equation represents oxidation while the other represents reduction.

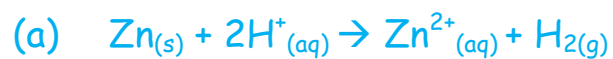
For Example:

- Oxidation half equation: $\text{Zn}_{(s)} \rightarrow \text{Zn}^{2+}_{(aq)} + 2e^{-}$
- Reduction half equation: $\text{Cu}^{2+}_{(aq)} + 2e^{-} \rightarrow \text{Cu}_{(s)}$
- Redox equation:
$$\text{Zn}_{(s)} + \text{Cu}^{2+}_{(aq)} \rightarrow \text{Zn}^{2+}_{(aq)} + \text{Cu}_{(s)}$$

3.3A Summary Questions

Question 1

Write half equations for:



Oxidation:

Reduction:



Oxidation:

Reduction:

Balancing Half-Equations

1. Use oxidation numbers to identify the element being oxidised or reduced.
2. Write a skeleton equation (reactants \rightarrow product)
3. Balance the number of atoms of the elements being oxidised or reduced.
4. Balance **oxygen** atoms by adding H_2O .
5. Balance **hydrogen** atoms by adding H^+ .
6. Balance the **charge** by adding electrons to the side that is most positive.

Worked Example:

Write a half equation for the oxidation of:

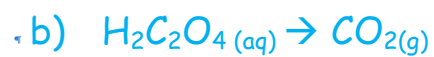
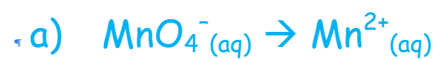
- (a) $\text{NO}_3^-_{(\text{aq})}$ to $\text{NO}_{(\text{g})}$ in acidic solution.

(b) $\text{PbO}_{2(s)}$ to $\text{Pb}^{2+}_{(aq)}$ in acidic conditions.

3.3B Summary Questions

Question 1

Complete the following half-equations for acidic conditions.



Balancing Redox Equations

- To obtain an overall reaction it is best to consider and balance each half equation first.
- These can then be added so that the number of electrons for each half equations is equal.

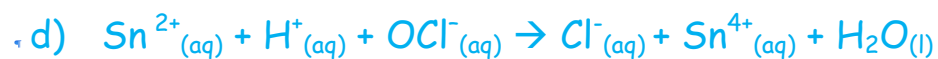
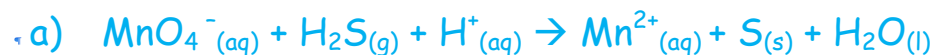
Worked Example

Balance the following redox reaction:



3.3C Summary Questions**Question 1**

Balance the following using half-equations.



3.4 Competition for electrons

- Redox reactions involve a competition for electrons.
- The species being reduced has a greater ability to gain electrons than the substance being oxidised.
- It is possible to set up an experiment which can be used to compare chemical species potential to be reduced or oxidised.

Standard reduction potentials (E° values)

- Are measured using the hydrogen half-cell as a reference cell
- Apply at STP (1.0 mol L⁻¹, 101.3 kPa gas pressure and 25 °C temperature)
- Are a measure of the relative tendency of a species to be reduced.
- Can be used to predict:
 - Electrochemical cell voltages
 - Whether particular redox reactions could occur.

Predicting the direction of redox reactions

The standard reduction potentials can be used to predict the likelihood of a particular redox reaction occurring as written.

The E° values for the two half equations are added.

- An overall + E° value means the reaction may proceed as written.
- An overall - E° value means the reaction will not occur as written.
- In general, substances listed high on the list of the Standard Reduction Potential table (on the left of the reaction arrow) will oxidise those listed low (on the right of the reaction arrow).

Half-reaction	E° (volts)
$F_2(g) + 2 e^- \rightleftharpoons 2 F^-(aq)$	+ 2.89
$H_2O_2(aq) + 2 H^+(aq) + 2 e^- \rightleftharpoons 2 H_2O(l)$	+ 1.76
$PbO_2(s) + SO_4^{2-}(aq) + 4 H^+(aq) + 2 e^- \rightleftharpoons PbSO_4(s) + 2 H_2O(l)$	+ 1.69
$2 HClO(aq) + 2 H^+(aq) + 2 e^- \rightleftharpoons Cl_2(g) + 2 H_2O(l)$	+ 1.63
$MnO_4^-(aq) + 8 H^+(aq) + 5 e^- \rightleftharpoons Mn^{2+}(aq) + 4 H_2O(l)$	+ 1.51
$Au^{3+}(aq) + 3 e^- \rightleftharpoons Au(s)$	+ 1.50
$HClO(aq) + H^+(aq) + 2 e^- \rightleftharpoons Cl^-(aq) + H_2O(l)$	+ 1.49
$PbO_2(s) + 4 H^+(aq) + 2 e^- \rightleftharpoons Pb^{2+}(aq) + 2 H_2O(l)$	+ 1.46
$Cl_2(g) + 2 e^- \rightleftharpoons 2 Cl^-(aq)$	+ 1.36
$Cr_2O_7^{2-}(aq) + 14 H^+(aq) + 6 e^- \rightleftharpoons 2 Cr^{3+}(aq) + 7 H_2O(l)$	+ 1.36
$O_2(g) + 4 H^+(aq) + 4 e^- \rightleftharpoons 2 H_2O(l)$	+ 1.23
$Br_2(l) + 2 e^- \rightleftharpoons 2 Br^-(aq)$	+ 1.08
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+ 0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+ 0.77
$O_2(g) + 2 H^+(aq) + 2 e^- \rightleftharpoons H_2O_2(aq)$	+ 0.70
$I_2(s) + 2 e^- \rightleftharpoons 2 I^-(aq)$	+ 0.54
$O_2(g) + 2 H_2O(l) + 4 e^- \rightleftharpoons 4 OH^-(aq)$	+ 0.40
$Cu^{2+}(aq) + 2 e^- \rightleftharpoons Cu(s)$	+ 0.34
$S(s) + 2 H^+(aq) + 2 e^- \rightleftharpoons H_2S(aq)$	+ 0.17
$2 H^+(aq) + 2 e^- \rightleftharpoons H_2(g)$	0 exactly
$Pb^{2+}(aq) + 2 e^- \rightleftharpoons Pb(s)$	- 0.13
$Sn^{2+}(aq) + 2 e^- \rightleftharpoons Sn(s)$	- 0.14
$Ni^{2+}(aq) + 2 e^- \rightleftharpoons Ni(s)$	- 0.24
$Co^{2+}(aq) + 2 e^- \rightleftharpoons Co(s)$	- 0.28
$PbSO_4(s) + 2 e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	- 0.36
$Cd^{2+}(aq) + 2 e^- \rightleftharpoons Cd(s)$	- 0.40
$2 CO_2(g) + 2 H^+(aq) + 2 e^- \rightleftharpoons H_2C_2O_4(aq)$	- 0.43
$Fe^{2+}(aq) + 2 e^- \rightleftharpoons Fe(s)$	- 0.44
$Cr^{3+}(aq) + 3 e^- \rightleftharpoons Cr(s)$	- 0.74
$Zn^{2+}(aq) + 2 e^- \rightleftharpoons Zn(s)$	- 0.76
$2 H_2O(l) + 2 e^- \rightleftharpoons H_2(g) + 2 OH^-(aq)$	- 0.83
$Mn^{2+}(aq) + 2 e^- \rightleftharpoons Mn(s)$	- 1.18
$Al^{3+}(aq) + 3 e^- \rightleftharpoons Al(s)$	- 1.68
$Mg^{2+}(aq) + 2 e^- \rightleftharpoons Mg(s)$	- 2.36
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	- 2.71
$Ca^{2+}(aq) + 2 e^- \rightleftharpoons Ca(s)$	- 2.87
$Sr^{2+}(aq) + 2 e^- \rightleftharpoons Sr(s)$	- 2.90
$Ba^{2+}(aq) + 2 e^- \rightleftharpoons Ba(s)$	- 2.91
$K^+(aq) + e^- \rightleftharpoons K(s)$	- 2.94

Worked Example

1. Could zinc react with 1.0 mol L^{-1} HCl acid?
2. Could copper react with 1.0 mol L^{-1} HCl
3. Magnesium is reaction with 1.0 mol L^{-1} HCl. Predict the reaction using E° values.

Metals and Metal ions – Displacement reaction

Metal displacement reactions will occur if a more reactive metal is placed in a solution of a less reactive metal.

Half-reaction	E°(volts)
$\text{F}_2(\text{g}) + 2 \text{e}^- \rightleftharpoons 2 \text{F}^-(\text{aq})$	+ 2.89
$\text{H}_2\text{O}_2(\text{aq}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons 2 \text{H}_2\text{O}(\text{l})$	+ 1.76
$\text{PbO}_2(\text{s}) + \text{SO}_4^{2-}(\text{aq}) + 4 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{PbSO}_4(\text{s}) + 2 \text{H}_2\text{O}(\text{l})$	+ 1.69
$2 \text{HClO}(\text{aq}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Cl}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{l})$	+ 1.63
$\text{MnO}_4^-(\text{aq}) + 8 \text{H}^+(\text{aq}) + 5 \text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4 \text{H}_2\text{O}(\text{l})$	+ 1.51
$\text{Au}^{3+}(\text{aq}) + 3 \text{e}^- \rightleftharpoons \text{Au}(\text{s})$	+ 1.50
$\text{HClO}(\text{aq}) + \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Cl}^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+ 1.49
$\text{PbO}_2(\text{s}) + 4 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Pb}^{2+}(\text{aq}) + 2 \text{H}_2\text{O}(\text{l})$	+ 1.46
$\text{Cl}_2(\text{g}) + 2 \text{e}^- \rightleftharpoons 2 \text{Cl}^-(\text{aq})$	+ 1.36
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14 \text{H}^+(\text{aq}) + 6 \text{e}^- \rightleftharpoons 2 \text{Cr}^{3+}(\text{aq}) + 7 \text{H}_2\text{O}(\text{l})$	+ 1.36
$\text{O}_2(\text{g}) + 4 \text{H}^+(\text{aq}) + 4 \text{e}^- \rightleftharpoons 2 \text{H}_2\text{O}(\text{l})$	+ 1.23
$\text{Br}_2(\text{l}) + 2 \text{e}^- \rightleftharpoons 2 \text{Br}^-(\text{aq})$	+ 1.08
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+ 0.80
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+ 0.77
$\text{O}_2(\text{g}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{H}_2\text{O}_2(\text{aq})$	+ 0.70
$\text{I}_2(\text{s}) + 2 \text{e}^- \rightleftharpoons 2 \text{I}^-(\text{aq})$	+ 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
$\text{S}(\text{s}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{H}_2\text{S}(\text{aq})$	+ 0.17
$2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{H}_2(\text{g})$	0 exactly
$\text{Pb}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Pb}(\text{s})$	- 0.13
$\text{Sn}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Sn}(\text{s})$	- 0.14
$\text{Ni}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Ni}(\text{s})$	- 0.24
$\text{Co}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Co}(\text{s})$	- 0.28
$\text{PbSO}_4(\text{s}) + 2 \text{e}^- \rightleftharpoons \text{Pb}(\text{s}) + \text{SO}_4^{2-}(\text{aq})$	- 0.36
$\text{Cd}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Cd}(\text{s})$	- 0.40
$2 \text{CO}_2(\text{g}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{H}_2\text{C}_2\text{O}_4(\text{aq})$	- 0.43
$\text{Fe}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Fe}(\text{s})$	- 0.44
$\text{Cr}^{3+}(\text{aq}) + 3 \text{e}^- \rightleftharpoons \text{Cr}(\text{s})$	- 0.74
$\text{Zn}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Zn}(\text{s})$	- 0.76
$2 \text{H}_2\text{O}(\text{l}) + 2 \text{e}^- \rightleftharpoons \text{H}_2(\text{g}) + 2 \text{OH}^-(\text{aq})$	- 0.83
$\text{Mn}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Mn}(\text{s})$	- 1.18
$\text{Al}^{3+}(\text{aq}) + 3 \text{e}^- \rightleftharpoons \text{Al}(\text{s})$	- 1.68
$\text{Mg}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Mg}(\text{s})$	- 2.36
$\text{Na}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Na}(\text{s})$	- 2.71
$\text{Ca}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Ca}(\text{s})$	- 2.87
$\text{Sr}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Sr}(\text{s})$	- 2.90
$\text{Ba}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Ba}(\text{s})$	- 2.91
$\text{K}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{K}(\text{s})$	- 2.94

For example, $\text{Mg}_{(\text{s})}$ will react with $\text{CuSO}_{4(\text{aq})}$ but $\text{Cu}_{(\text{s})}$ will not react with $\text{MgSO}_{4(\text{aq})}$

4. A strip of zinc metal is added to an aluminium nitrate solution. Predict the reaction, if any.

Halogens and halide ions – Displacement reaction

- All the halogen are good oxidising agents with fluorine being the strongest.
- The trend in oxidising ability can be found in relation to their position on the periodic table or standard reduction potential table.

Half-reaction	E°(volts)
$F_2(g) + 2 e^- \rightleftharpoons 2 F^-(aq)$	+ 2.89
$H_2O_2(aq) + 2 H^+(aq) + 2 e^- \rightleftharpoons 2 H_2O(l)$	+ 1.76
$PbO_2(s) + SO_4^{2-}(aq) + 4 H^+(aq) + 2 e^- \rightleftharpoons PbSO_4(s) + 2 H_2O(l)$	+ 1.69
$2 HClO(aq) + 2 H^+(aq) + 2 e^- \rightleftharpoons Cl_2(g) + 2 H_2O(l)$	+ 1.63
$MnO_4^-(aq) + 8 H^+(aq) + 5 e^- \rightleftharpoons Mn^{2+}(aq) + 4 H_2O(l)$	+ 1.51
$Au^{3+}(aq) + 3 e^- \rightleftharpoons Au(s)$	+ 1.50
$HClO(aq) + H^+(aq) + 2 e^- \rightleftharpoons Cl^-(aq) + H_2O(l)$	+ 1.49
$PbO_2(s) + 4 H^+(aq) + 2 e^- \rightleftharpoons Pb^{2+}(aq) + 2 H_2O(l)$	+ 1.46
$Cl_2(g) + 2 e^- \rightleftharpoons 2 Cl^-(aq)$	+ 1.36
$Cr_2O_7^{2-}(aq) + 14 H^+(aq) + 6 e^- \rightleftharpoons 2 Cr^{3+}(aq) + 7 H_2O(l)$	+ 1.36
$O_2(g) + 4 H^+(aq) + 4 e^- \rightleftharpoons 2 H_2O(l)$	+ 1.23
$Br_2(l) + 2 e^- \rightleftharpoons 2 Br^-(aq)$	+ 1.08
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+ 0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+ 0.77
$O_2(g) + 2 H^+(aq) + 2 e^- \rightleftharpoons H_2O_2(aq)$	+ 0.70
$I_2(s) + 2 e^- \rightleftharpoons 2 I^-(aq)$	+ 0.54
$O_2(g) + 2 H_2O(l) + 4 e^- \rightleftharpoons 4 OH^-(aq)$	+ 0.40
$Cu^{2+}(aq) + 2 e^- \rightleftharpoons Cu(s)$	+ 0.34
$S(s) + 2 H^+(aq) + 2 e^- \rightleftharpoons H_2S(aq)$	+ 0.17
$2 H^+(aq) + 2 e^- \rightleftharpoons H_2(g)$	0 exactly
$Pb^{2+}(aq) + 2 e^- \rightleftharpoons Pb(s)$	- 0.13
$Sn^{2+}(aq) + 2 e^- \rightleftharpoons Sn(s)$	- 0.14
$Ni^{2+}(aq) + 2 e^- \rightleftharpoons Ni(s)$	- 0.24
$Co^{2+}(aq) + 2 e^- \rightleftharpoons Co(s)$	- 0.28
$PbSO_4(s) + 2 e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	- 0.36
$Cd^{2+}(aq) + 2 e^- \rightleftharpoons Cd(s)$	- 0.40
$2 CO_2(g) + 2 H^+(aq) + 2 e^- \rightleftharpoons H_2C_2O_4(aq)$	- 0.43
$Fe^{2+}(aq) + 2 e^- \rightleftharpoons Fe(s)$	- 0.44
$Cr^{3+}(aq) + 3 e^- \rightleftharpoons Cr(s)$	- 0.74
$Zn^{2+}(aq) + 2 e^- \rightleftharpoons Zn(s)$	- 0.76
$2 H_2O(l) + 2 e^- \rightleftharpoons H_2(g) + 2 OH^-(aq)$	- 0.83
$Mn^{2+}(aq) + 2 e^- \rightleftharpoons Mn(s)$	- 1.18
$Al^{3+}(aq) + 3 e^- \rightleftharpoons Al(s)$	- 1.68
$Mg^{2+}(aq) + 2 e^- \rightleftharpoons Mg(s)$	- 2.36
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	- 2.71
$Ca^{2+}(aq) + 2 e^- \rightleftharpoons Ca(s)$	- 2.87
$Sr^{2+}(aq) + 2 e^- \rightleftharpoons Sr(s)$	- 2.90
$Ba^{2+}(aq) + 2 e^- \rightleftharpoons Ba(s)$	- 2.91
$K^+(aq) + e^- \rightleftharpoons K(s)$	- 2.94

Fluorine	9	F
Chlorine	17	Cl
Bromine	35	Br
Iodine	53	I
Astatine	85	At

This means, that $F_{2(g)}$ will displace all other halide ions (e.g. Cl^-)

5. Predict the reaction between $Cl_{2(g)}$ and a solution of $KBr_{(aq)}$.

Limitations of the use of E° tables

The standard reduction potential table is useful in predicting reaction tendency and the emf of electrochemical cells. **However they do have limitations:**

- The values of E° depends on concentration.
- It applies only to aqueous solutions.
- The emf of a cell can depend on temperature, pressure and acidity.
- The E° values give no indication of likely reaction rate.

3.4 Summary Questions

Question 1

Use standard reduction potentials to determine which of the following reactions are likely to proceed. Assume STP.

- a) Iron nails are placed in zinc sulfate solution.
- b) Copper is placed in hydrochloric acid solution.
- c) Bromine is added to sodium chloride solution.

- .d) Chlorine gas bubbled into hydrogen sulfide solution.
- .e) Potassium iodide is added to acidified potassium permanganate solution.
- .f) Calcium metal is added to water.
- .g) Acidified potassium permanganate solution is added to a solution of iron (II) sulfate.

Question 2

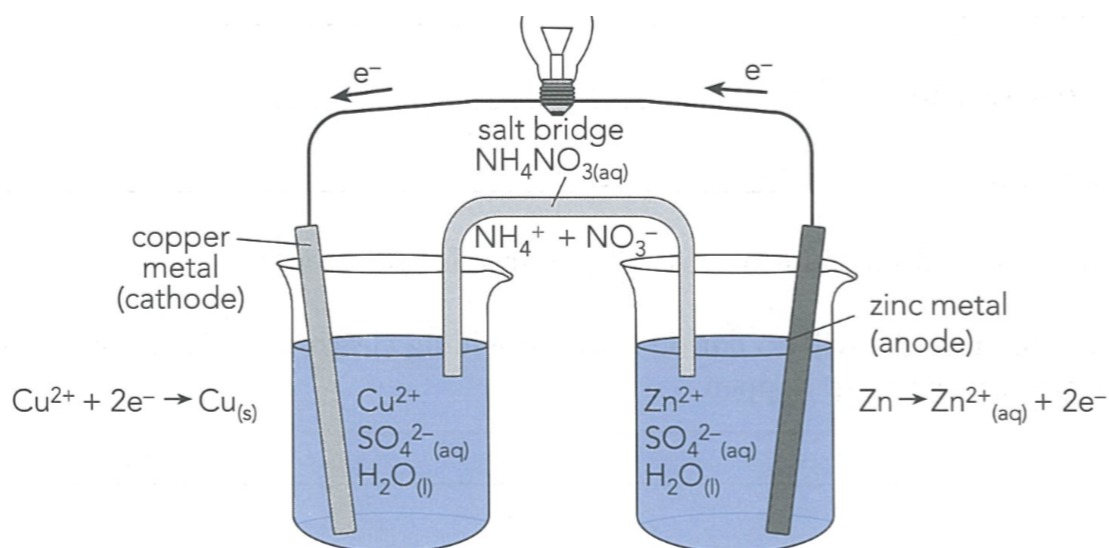
A storage plant designer needs to determine if a steel tank can be used to store copper (II) sulfate solution. Use E° values to determine if this is a good idea.

3.5 Electrochemical Cells

- Electrochemical cells are an excellent example of redox process put to commercial use.
- There are two types of electrochemical cells:
 - Galvanic cells (such as those used to power a torch)
 - Electrolytic cells - which are typically used for the electrolysis of solutions.
- Galvanic cells produce an electric current due to the competition for electrons between two different materials.
- The reaction is spontaneous.
- Electrolytic cells require an electrical current to produce a chemical reaction that would not otherwise occur.
- The process is the reverse of what occurs in a galvanic cell.
- In both types of cells a transfer of electrons is involved, that is, a redox reaction.

Galvanic cells

- In a galvanic cell, the reactants are separated and electrons are forced along an external path.
- A spontaneous redox reaction takes place; however, the chemical energy from the reaction produces an external electric current.
- In the laboratory, it is possible to construct a galvanic cell using two beakers and a salt bridge as shown below.



- Zinc is oxidized and electrons flow through the external circuit to the copper.

- Copper ions in the Cu/Cu^{2+} cell are reduced.
- The salt bridge (which can simply be filter paper soaked in ammonium nitrate solution) allows ions to migrate between the two half cells and so that each cell remains electrically neutral.
- The choice of salt used in the salt bridge is such as to ensure no precipitation.

Worked example / common questions on galvanic cells

1. Why is a salt bridge necessary?
2. Mark on the diagram the movement of ions in the salt bridge.
Show:
 - a. Which ions move away from the Zn/Zn^{2+} half-cell.
 - b. Which ions move away from the Cu/Cu^{2+} half-cell.
3. Describe the changes that would occur to:
 - a. The zinc electrode
 - b. The copper electrode

c. The colour of the solution in the Zn/Zn^{2+} cell

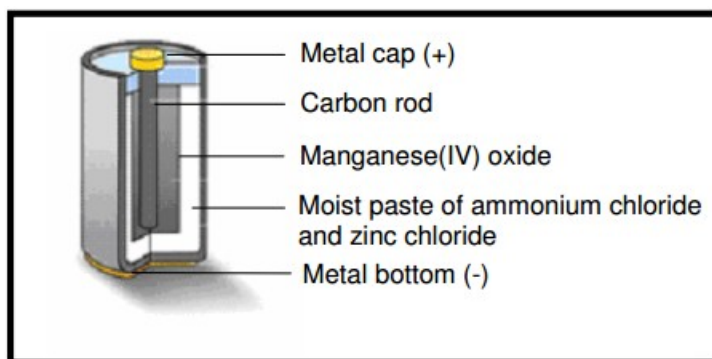
d. The colour of the solution in the Cu/Cu^{2+} cell

Determining cell voltage

The emf (voltage) produced by a cell is easily determined using the standard reduction potential table.

Commercial cells - the dry cell

- The dry cell is fairly inexpensive and quite portable.
- Its compact size allows for many uses such as torches, portable radios and cameras.
- The dry cell consist of:
 - An outer case of zinc (anode)
 - A carbon rod surrounded by a paste of MnO_2 (cathode)
 - An electrolyte paste of NH_4Cl and ZnCl_2 .



Anode reaction: $\text{Zn}_{(s)} \rightarrow \text{Zn}^{2+}_{(aq)} + 2e^{-}$

Cathode reaction: $2\text{MnO}_{2(s)} + 2\text{H}^{+}_{(g)} + 2e^{-} \rightarrow \text{Mn}_2\text{O}_{3(s)} + \text{H}_2\text{O}_{(l)}$

Overall reaction:

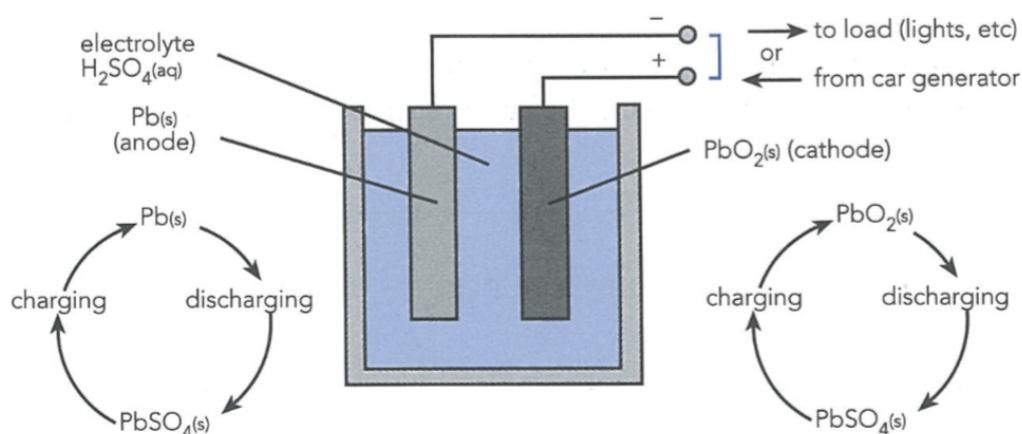
The lead-acid accumulator

- Lead-acid batteries are relatively inexpensive and can store large quantities of charge.
- Although they are a little bulky, their major advantage is that they can be recharged.
- They are widely used in motor vehicles and isolated farmhouses.

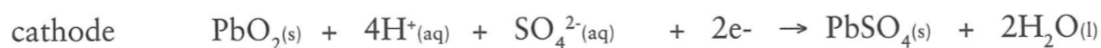
Important features of the lead-acid battery are:

- Electrodes
 - Anode - spongy lead - large surface area

- Cathode - lead (IV) oxide packed on a metal grid
- Electrolyte is concentrated sulfuric acid (4.5 mol L^{-1})
- A typical car battery consists of 6 cells placed in series (12V total)
- It can be recharged - hence why it is referred to as a secondary cell.
- The density of the electrolyte indicates the state of charge.



The electrode reaction for **Discharge** (when the battery is being used) is as follows:



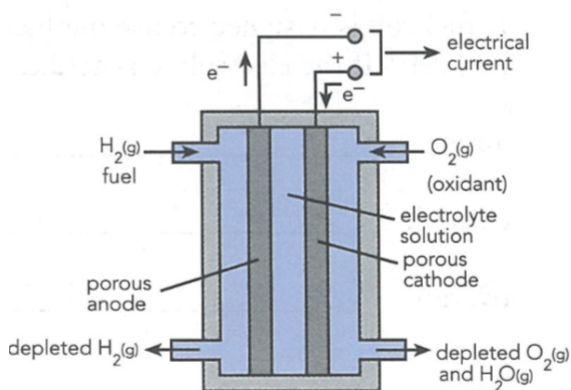
Overall _____
(complete this)

The fuel cell

- A fuel cell is different from both primary and secondary cells as it does not store the reactants or products.
- A fuel cell converts the energy of a chemical reaction directly and continuously into electrical energy.

Important features of a fuel cell are:

- Electrodes consist of porous platinum or graphite
- Electrolytes can be either basic or acidic
- Reactants are usually gaseous (e.g. $\text{H}_2(\text{g})$, $\text{O}_2(\text{g})$)
- They supply electricity as reactants are fed into them.



The major advantages of fuel cells are their high efficiency of energy conversion and their ability to generate electricity continuously for a very long time.

- However fuel cells are not widely used because of their high cost and slow rate of reaction.



overall

(complete this)

3.5 Summary Questions

Question 1

(a). Draw a fully label diagram to represent a Mg/Mg^{2+} , Cu/Cu^{2+} galvanic cell. Label the anode, cathode, ions, e^- movement, ion movement...etc

(b)

Anode reaction:

Cathode reaction:

Redox reaction:

(c). What is the expected emf of this cell (1.0 mol L^{-1} solution)?

(d). Eventually this cell will go flat (i.e. stop providing electrons). Detail two conditions that may cause this to happen?

Question 2

(a). The $H^+_{(aq)}$ ions required in the cathode reaction of the dry cell are provided by the $NH_4^+_{(aq)}$ in the electrolyte paste. Write an equation to show how this occurs.

(b). Why is the Leclanché cell referred to as a primary cell?

(c). List the major advantages and disadvantages of this type of cell.

Question 3

Alkaline dry cells are also commercially available and are better performers than the acid form. An important example is the silver oxide/zinc cell commonly used in watches and calculators.

The zinc is oxidised to $\text{Zn}^{2+}_{(\text{aq})}$ while the silver oxide is reduced to silver. Write out the half equation and the overall redox reaction.

Question 4

Write the overall equation for the charging of a lead-acid accumulator.

Question 5

When an accumulator becomes flat, both electrodes become $\text{PbSO}_{4(\text{s})}$.

- (i) Explain why no emf would be possible under these conditions.

- (ii) What other change occurs which would more easily help us to decide if the battery is flat?

Question 6:

(a). Use E° values to determine the theoretical emf of the the fuel cell (assume 1.0 mol L^{-1} solution).

(b). Write the half-equation for this cell in basic solution:

(c) What is the overall reaction in both cases?

(d) How does this overall reaction compare with the combustion of hydrogen gas?

Question 7

A fuel cell is designed to use methane (CH_4) and oxygen in a similar manner to the H_2/O_2 fuel cell. If the electrolyte is acidic, determine the electrode reactions.

3.6 Electrolysis

- Electrolysis refers to the chemical changes that take place when an electrical current is passed through a molten substance or solution.
- Electrolysis occurs in an electrolytic cell and is the opposite process to that which occurs in a galvanic cell.
- The reactions that occur during electrolysis are *forced reactions*, that is, they are not spontaneous reactions.
- By applying a suitable voltage, electrical energy is converted to chemical energy.

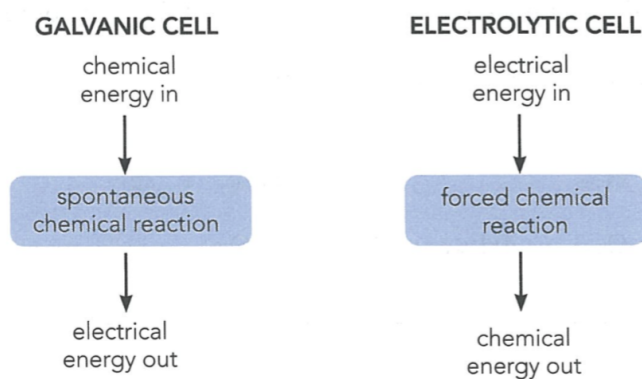


Figure 3.7 Comparing galvanic and electrolytic cells. Galvanic cells produce electrical energy due to spontaneous reactions. Electrolytic cells have an applied external voltage which causes forced reactions to occur (electrolysis). In both cases the reactions are redox reactions.

Electrolytic cells

- Systems that bring about electrolysis reactions are called electrolytic cells.
- An applied voltage causes reactions at the two electrodes.

ANODE: Where oxidation occurs

CATHODE: Where reduction occurs

This is the same for all electrochemical cells

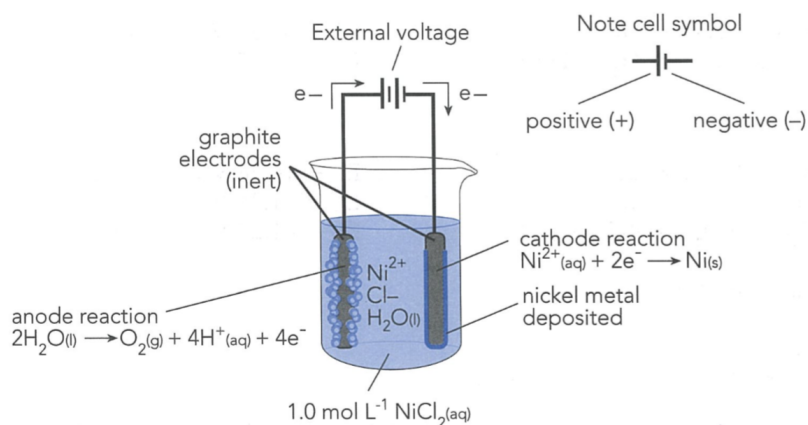


Figure 3.8 A typical electrolytic cell. An external applied voltage supplies electrons to the cathode where a reduction reaction takes place. At the anode, an oxidation reaction occurs which results in the same number of electrons being produced as those consumed at the cathode. Note that only ions flow through the solution. As they do, they complete the electrical circuit.

Predicting electrode reactions

- Often more than one reaction is possible at each electrode of an electrolytic cell.
- For each electrode we need to consider all substances present and whether the electrode is inert or reactive.
- We can use E° values to predict reactions and voltage required.

- The concentration of solutions can also have an effect on E values and products formed but for now we will assume all 1.0 M and inert electrodes (we will look at the effects of electrodes and concentration later).

For the electrolytic cell shown above:

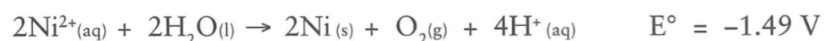
Possible cathode reaction. Note that only reduction reactions can occur here:



Possible anode reactions. Note that only oxidation reactions can occur here:



Hence overall reaction:

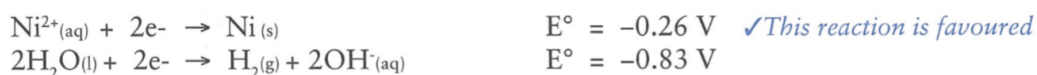


- The least negative half reaction is always favoured
- A minimum voltage of + 1.49 V must be applied to cause a reaction.
- Graphite electrodes can be considered inert and do not react.

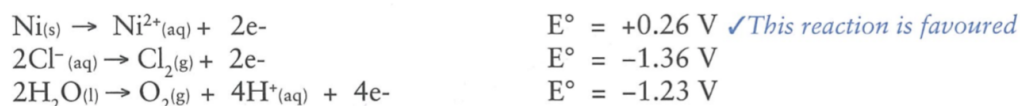
Using reactive electrodes

- Where reactive electrodes are used, they need to be considered in predicting reactions.
- For instance, if we used nickel metal electrodes, we would have:

Possible cathode reactions:



Possible anode reactions:



Hence overall reaction:



- By using nickel electrodes, one of the preferred reactions is different from the previous example.
- At the anode, nickel metal reacts rather than the water.
- Effectively, nickel metal is oxidised at the anode and the resulting ions are reduced back to the metal at the cathode.
- This type of reaction is very important in electroplating and electrefining of metals like copper.

Aqueous solutions – concentration effect

- The concentration of an aqueous solution can affect the preferred reaction that occur in an electrolytic cell.
- E° values apply only to 1.0 M solutions.
- At other concentrations the E values her competing change and other competing reactions may be more favoured.
- In all aqueous solutions, water is always a possible reactant; at either electrode.
- In fact, if a low concentration solution is used, then electrolysis of water will occur.
- The product would be oxygen gas at the anode and hydrogen gas at the cathode.

Worked Example

- Let us consider the electrolysis of NaCl solution at different concentration. Assuming inert electrodes.
- The possible electrode reactions and E° values (1.0 M solutions) are listed and for neutral solutions (i.e. non standard) E value for the water reaction is also given.

Possible cathode reactions:



Possible anode reactions:



Aqueous NaCl (1.0 M solution).

- At the cathode, water is reduced rather than the sodium ions as indicated by the E° values.
- Hydrogen gas is the only product at both high and low concentrations of solutions.
- At the anode we can see that for 1.0 M solutions the E° values for the possible reactions are very similar.
- Oxygen gas is produced in preference to chlorine gas although some chlorine gas may also be produced.
- At low concentration only oxygen gas is produced.

Aqueous NaCl (concentrated solution).

- At high concentration only chlorine gas is produced at the anode.
- Hydrogen gas is still produced at the cathode.
- The electrolysis of a concentrated aqueous solution of sodium chloride is an important industrial process.

Electrolysis of molten salts

- The electrolysis of aqueous solutions often involves the reaction of water itself.
- Solid ionic salts do not conduct electricity.
- However if they are heated sufficiently, they become molten and can be electrolysed.
- In this situation, the ions of the ionic salt will react at the electrodes.
- For instance, if sodium chloride is heated to above 800 °C it melts and can be electrolysed as shown below.

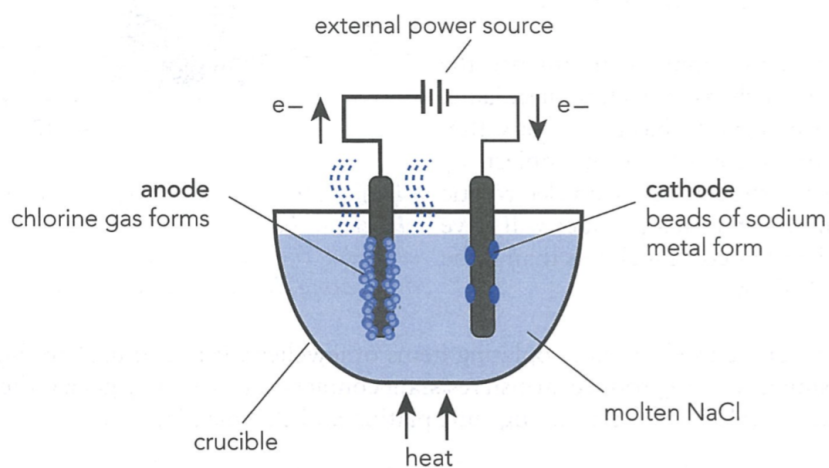


Figure 3.9 Electrolysis of molten sodium chloride. Only sodium and chloride ions are available to react at the inert electrodes. Sodium liquid is produced at the cathode and chlorine gas at the anode.

Electroplating

- Metals may be electroplated by adding a very thin layer of another metal by the process of electrolysis.
- This important industrial process is used in order to improve the appearance of more base metals such as iron and also improve their resistance to corrosion.
- Silver plating, for example, can enhance the looks of items such as bracelets, necklaces, cutlery and other objects.
- A very thin layer of silver is coated on the object by attaching them to the cathode of an electrolytic cell.
- The finished product will have the look of silver but cost much less than if the item was solid silver.

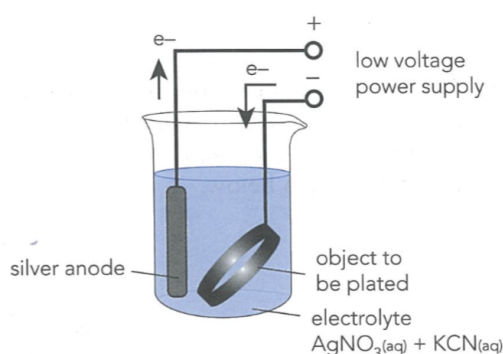


Figure 3.10 Electroplating with silver. The metal object to be plated is always attached to the cathode. The anode is made of the metal used for the coating. It does not need to be pure.

- Gold similarly can be used for electroplating items of jewellery. It is also used in the electronics industry in a similar way to produce tarnish resistant contacts and switching gear.
- Other important examples of electroplating are tin plating, zinc plating and chrome plated steel.

Electrolytic refining of copper

- Copper is produced from its ores by a series of processes.
- The final smelt contains about 98.5 % of copper.
- However the purity required by electrical wiring, one of the main use of copper, is very high (~ 99.9 %).
- To achieve this high purity, crude copper or 'blister' copper is refined.
- This electrolytic process is very similar to electroplating.
- The copper and other reactive metals within the anode are oxidised and become part of the electrolyte solution.
- The high cost of electrefining is partly offset by the recovery of valuable impurities such as silver and gold.
- The use of very low voltage (approximately 0.3 V) ensures that only copper is deposited at the cathode and that the less active silver and gold are not oxidised at the cathode but are left behind as part of the anode mud.

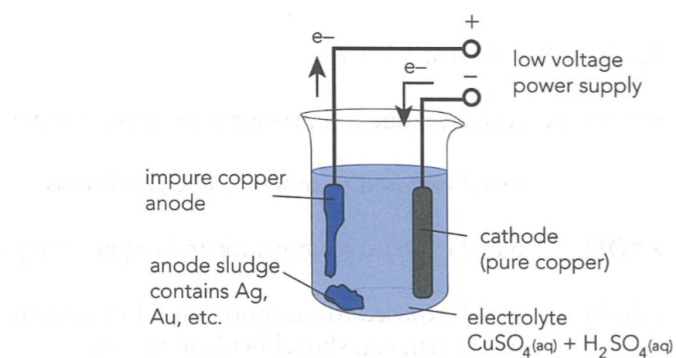


Figure 3.11 Electrolytic refining of copper. Industrially, copper anodes consist of large copper slabs produced from crude copper ($\approx 98.5\%$) or recycled high grade copper scrap. Impurities mostly include gold, silver, lead and nickel. These impurities are recovered from the anode mud or from solution.

3.6 Summary Questions

Question 1

(a). If an electrolytic cell was set up by placing a copper electrode in a solutions of H_2O and Cl^- . Consider the possible electrode reactions to determine the overall result.

Possible cathode reaction:

Possible anode reactions:

Overall reactions:

(b). The reactions within this electrolytic cell may change after some time. Explain why.

Question 2

Consider the electrolysis of sodium chloride solution at different concentrations. Assuming inert electrodes give the electrode reaction and list the products formed for the following concentrations.

(a). Aqueous NaCl (1.0 mol L^{-1})

Anode reaction:

Cathode reaction:

Products are:

(b). Aqueous NaCl (concentrated solutions)

Anode reaction:

Cathode reaction:

Products are:

Question 3

A 1.0 mol L^{-1} solution of $\text{HCl}_{(\text{aq})}$ is to be electrolysed using inert electrodes.

(a), Draw and fully label a diagram showing:

- Chemical species
- Electron flow
- The anode and cathode

(b). Determine the possible reactions at each electrode and give the overall reaction. Include E° values.

Anode reaction:

Cathode reaction:

Overall:

Question 4

Consider the electrolysis of molten sodium chloride. Assuming inert electrodes, give the electrode reactions for the following examples.

(a). Molten NaCl

Anode reaction:

Cathode reaction:

Products are:

(a). Molten PbI_2

Anode reaction:

Cathode reaction:

Products are:

Question 5

For the electroplating of silver using the electrolyte $\text{AgNO}_{3(\text{aq})}$ and $\text{KCN}_{(\text{aq})}$ give the relevant reactions.

Anode reaction:

Cathode reaction:

Overall:

Question 6

To get a really smooth finish when electroplating with silver, the silver ions concentration in solution must be very low. This is why KCN is added to the electrolyte solution. An equilibrium is established as follows:



Explain how this equilibrium reaction helps.

Question 7

(a). For the electrorefining of copper using a electrolyte solution of $\text{CuSO}_{4(\text{aq})}$ and H_2SO_4 . Give the relevant reactions.

Anode reaction:

Cathode reaction:

Overall:

(b). The gold and silver impurities not oxidized at the anode simply fall to the bottom and become part of the anode mud.

(i). Explain why this is the case. Refer to E° values in your answer.

(ii). What use might the anode mud be?

(c). Oxidised metal impurities such as Ni^{2+} and Pb^{2+} are not deposited at the cathode.

(i). Explain why this is the case. Refer to E° values in your answer.

(ii). Suggest how the nickel and copper ions may be removed or recovered?

3.7 Corrosion Of Metals

- The corrosion of metals is an electrochemical process which occurs when metals are oxidised by substances in their environment.
- The metal oxidises at the anode
 - $\text{Metal} \rightarrow (\text{Metal ion})^{x+} + xe^{-}$
- At the cathode, reduction occurs
- - Some common reduction processes are:
 - $2\text{H}^{+} + 2e^{-} \rightarrow \text{H}_2$
this occurs when metal is exposed to moisture in acidic conditions.
 - $\text{O}_2 + 2\text{H}_2\text{O} + 4e^{-} \rightarrow 4\text{OH}^{-}$
metal exposed to moisture and oxygen (very common)
 - $2\text{H}_2\text{O} + 2e^{-} \rightarrow \text{H}_2 + 2\text{OH}^{-}$
metal exposed to moisture but at a low concentration of oxygen (e.g. Blade of a shovel in soil)



metal in contact with another metal with a greater reduction potential.

Preventing corrosion is often about limiting the contact of metal with these four processes.

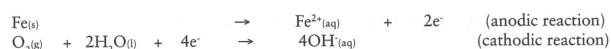
- When some metals oxidise, they form a tough, protective oxide coating that greatly limits the ability of the oxidising agent to come in contact with the metal.
- Aluminum is a common example.
- Even though Al oxidises rapidly, the oxide coating protects the underlying Al from further corrosion.
- Iron, however forms a oxide coating that is easily penetrated by oxygen and water and subsequently the oxide layer does not protect the Fe from further corrosion.

Rust – a special example of Corrosion

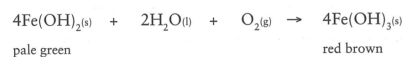
- The most common and important example of corrosion is the oxidation of iron.
- This is caused by the action of oxygen and water vapour in the air.

The formation of rust involves a series of reactions.

- The initial oxidation of the iron.



- The further oxidation of the $\text{Fe(OH)}_2(\text{s})$ formed.



78

- The partial dehydration of the $\text{Fe(OH)}_3(\text{s})$ to rust. One possible reaction is shown.

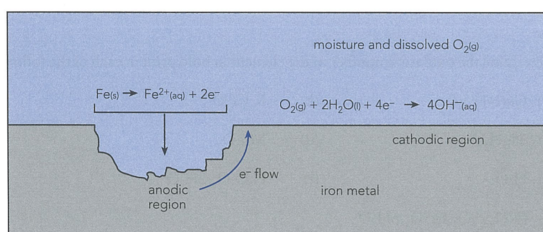


Figure 3.12 Corrosion of iron by the action of oxygen and water.

- This is an electrochemical process.
- Areas of stress in the iron become anodic while areas of high oxygen concentration become cathodic.
- Electrons flow through the metal from the anode to the cathode.
- The moisture acts as the electrolyte.

Corrosion Prevention

- Prevention of corrosion is of great economic importance.
- To prevent corrosion, the reactions described earlier must be inhibited.

This can be done by:

- Excluding air and/or water from the metal surface by
 - Protecting the iron surface with paint, grease or plastic
 - Plating the iron surface with metals such as chromium or metallic tin
- Using a sacrificial anode (more reactive metal) by
 - Galvanising iron with zinc - zinc corrodes in preference
 - Attaching magnesium and aluminum (to ships)
- Using cathodic protection so that jetties and pipelines are rendered negative by a low voltage dc current. The anodes may be made of scrap iron (e.g. Old engine blocks) or titanium coated inert electrode.

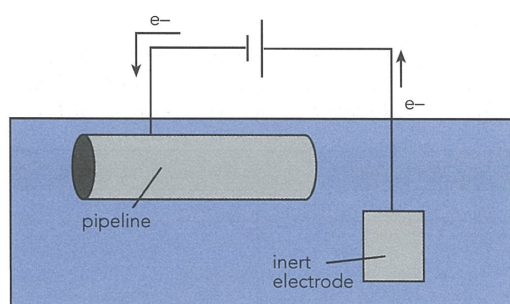


Figure 3.13 Cathodic protection. The pipeline is made negative to prevent its oxidation. Instead, H_2O is reduced at the pipeline, and oxidised at the inert electrode.

3.7 Summary Questions

Question 1

List the following metals in order of increasing tendency to oxidized.

Fe, Zn, Au, Mg, Na, Ag, Al, Pb, Cu

Question 2

The corrosion of iron to form rust is a complicated process and it is often best to break it down into several simpler stages. Explain the process of rusting.

(Question 29 on review)

Question 3

Tin is often used to coat iron cans which are subsequently called tin cans.

(a) Why is tin used as a coating for iron?

- (b) Use half equations and E° values to explain why this coating will only work while it is coherent and no iron is exposed.
- (c) Use half equations and E° values to explain why a zinc coating is much more effective in preventing the corrosion of iron.

(*Question 31)