Displacement Reactions

Some Standard Potentials

Weakest Oxidant	Oxidants		Reductants	E^0 (volts)	Strongest Reductant
|	$K^+ + e$	⇌	$K(s)$	-2.94	/\
|	$Ba^{2+} + 2e$	⇌	$Ba(s)$	-2.91	|
|	$Ca^{2+} + 2e$	⇌	$Ca(s)$	-2.87	|
|	$Na^+ + e$	⇌	$Na(s)$	-2.71	|
|	$Mg^{2+} + 2e$	⇌	$Mg(s)$	-2.36	|
|	$Al^{3+} + 3e$	⇌	$Al(s)$	-1.68	|
|	$Mn^{2+} + 2e$	⇌	$Mn(s)$	-1.18	|
|	$H_2O + e$	⇌	$\frac{1}{2}H_2(g) + OH^-$	-0.83	|
|	$Zn^{2+} + 2e$	⇌	$Zn(s)$	-0.76	|
|	$Fe^{2+} + 2e$	⇌	$Fe(s)$	-0.44	|
|	$Ni^{2+} + 2e$	⇌	$Ni(s)$	-0.24	|
|	$Sn^{2+} + 2e$	⇌	$Sn(s)$	-0.14	|
|	$Pb^{2+} + 2e$	⇌	$Pb(s)$	-0.13	|
|	**$H^+ + e$**	⇌	**$\frac{1}{2}H_2(g)$**	**0.00**	|
|	$SO_4^{2-} + 4H^+ + 2e$	⇌	$SO_2(aq) + 2H_2O$	0.16	|
|	$Cu^{2+} + 2e$	⇌	$Cu(s)$	0.34	|
|	$\frac{1}{2}O_2(g) + H_2O + 2e$	⇌	$2OH^-$	0.40	|
|	$Cu^+ + e$	⇌	$Cu(s)$	0.52	|
|	$\frac{1}{2}I_2(s) + e$	⇌	I^-	0.54	|
|	$\frac{1}{2}I_2(aq) + e$	⇌	I^-	0.62	|
|	$Fe^{3+} + e$	⇌	Fe^{2+}	0.77	|
|	$Ag^+ + e$	⇌	$Ag(s)$	0.80	|
|	$\frac{1}{2}Br_2(l) + e$	⇌	Br^-	1.08	|
|	$\frac{1}{2}Br_2(aq) + e$	⇌	Br^-	1.10	|
|	$\frac{1}{2}O_2(g) + 2H^+ + 2e$	⇌	H_2O	1.23	|
|	$\frac{1}{2}Cl_2(g) + e$	⇌	Cl^-	1.36	|
\\/	$\frac{1}{2}Cl_2(aq) + e$	⇌	Cl^-	1.40	|
	$MnO_4^- + 8H^+ + 5e$	⇌	$Mn^{2+} + 4H_2O$	1.51	
Strongest Oxidant	$\frac{1}{2}F_2(g) + e$	⇌	F^-	2.89	Weakest Reductant

Key Concepts

A more active metal will displace a least active metal from solution
A more active metal is a stronger reductant
(A reductant is a species that causes another species to be reduced, and is itself oxidised by losing electrons)
An *activity series* lists metals in order of decreasing strength as reductants
An example of an activity series of metals based on the Standard Potentials given would be:

K>Ba>Ca>Na>Mg>Al>Mn>Zn>Fe>Ni>Sn>Pb>Cu>Ag
In this series the most active metal is potassium (K) and the least active metal is silver (Ag)

Examples

a. If a strip of clean magnesium metal were placed in a copper sulphate solution, the magnesium metal would displace the copper from solution since magnesium is a more active metal than copper.

The magnesium metal strip would disintegrate while solid copper would be deposited

As Cu^{2+} ions are removed from solution, the solution would become a paler blue in colour.

The magnesium metal would donate electrons.

$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e \quad E^0 = +2.36V$
Magnesium metal is being oxidised (losing electrons) therefore it is the reductant.

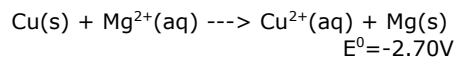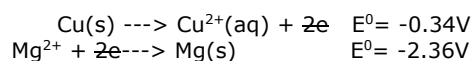The copper ions in solution would accept electrons forming solid copper.

$Cu^{2+}(aq) + 2e \longrightarrow Cu(s) \quad E^0 = +0.34V$
Copper ions are being reduced (gaining electrons) therefore the copper ions are the oxidant.

The overall equation would be:
$Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$
$E^0 = +2.36 \text{ V} + +0.34 \text{ V} = +2.70V$

The positive value for E^0 indicates a spontaneous reaction will occur.

b. If a strip of clean copper metal were placed in a magnesium chloride solution, no displacement reaction would occur since the copper is a less active metal than the

1

magnesium.

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e \quad E^0 = -0.34V$$
$$Mg^{2+} + 2e \longrightarrow Mg(s) \qquad E^0 = -2.36V$$

$$Cu(s) + Mg^{2+}(aq) \longrightarrow Cu^{2+}(aq) + Mg(s)$$
$$E^0 = -2.70V$$

The negative E^0 value indicates that the forward reaction will not proceed spontaneously.

Galvanic (Voltaic) Electrochemical Cells

Key Concepts

- Galvanic (Voltaic) Cell: converts chemical energy of oxidants and reductants into electrical energy
- Electrodes: are conductors used to permit the flow of electrons in an electrochemical cell.
    - One electrode is the anode, the other is the cathode.
- Anode: Oxidation occurs at anode
    - Anode is negative
    - Anode disintegrates
- Cathode: Reduction occurs at cathode
    - Cathode is positive
    - Solid deposits on cathode
- Salt bridge: allows for migration of ions to complete the electrical circuit
- Electron Flow: from anode to cathode
    - Electrons flow from negative to positive
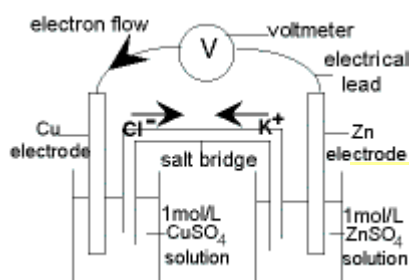- Spontaneous Reaction: E^0 for the galvanic cell is positive

Animated Tutorial

Example : The Daniell Cell

A galvanic (voltaic) cell is a device that uses REDOX reactions to produce electricity.

In 1836 Professor John Daniell adopted a two-cell approach to produce electricity.
The Daniell Cell is divided into 2 half-cells connected by a wire and a salt bridge to complete the electrical circuit.

Anode:	Zn	--->	$Zn^{2+} + 2e$	$E^o = +0.76V$
Cathode:	$Cu^{2+} + 2e$	--->	Cu	$E^o = +0.35V$
Cell:	$Zn + Cu^{2+}$	--->	$Zn^{2+} + Cu$	$E^o_{cell} = +1.11V$
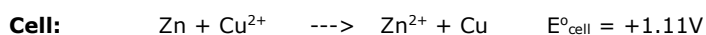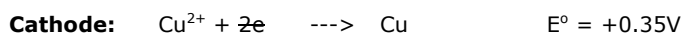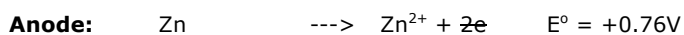
At the negative anode, zinc is oxidised to zinc ions.
The zinc anode disintegrates in time.
At the positive cathode, copper ions are reduced to copper atoms.
Copper is deposited on the copper cathode in time.
Electrons flow from the zinc anode to the copper cathode.

In the overall REDOX reaction zinc is donating electrons to copper ions.
This REDOX reaction occurs spontaneously, E^o is positive.
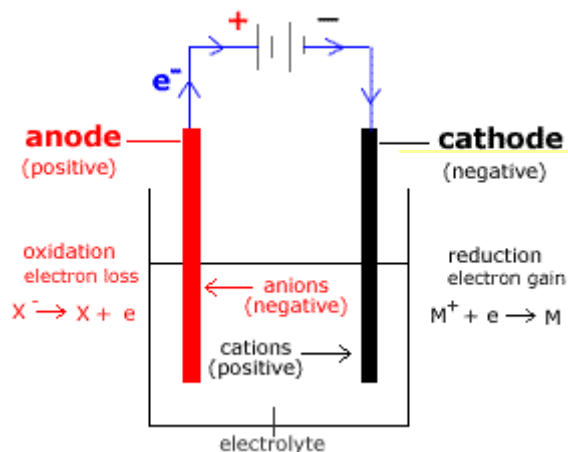This REDOX reaction produces 1.11V of electricity.

Electrolysis - Electrolytic Cells

Key Concepts

- Electrolysis: the process in which an electric current is used to bring about a chemical reaction which does not occur spontaneously.
- Electrolytic Cell: converts electrical energy into chemical energy.
- Electrodes : conductors used to permit the flow of electrons in an electrochemical cell.
  One electrode is the anode, the other is the cathode.
- Anode: Oxidation occurs at the anode.
  Anions (negatively charged ions) migrate to the anode.
  Anode is positive.
  Anode disintegrates.
- Cathode: Reduction occurs at the cathode.
  Cations (positively charged ions) migrate to the cathode.
  Cathode is negative.
  Solid deposits at the cathode.
- Electron flow: from anode to cathode.
  Electrons flow from positive to negative.
- If more than one reaction is possible, the reaction with the lowest E^o will occur.
- Non-spontaneous reaction: E^o for the electrolytic cell is negative.
- Applied emf must be greater than the emf for the cell, ie greater than E^o.
- Mass of substance produced electrolytically is proportional to the quantity of electricity flowing.

Electrolytic Cell

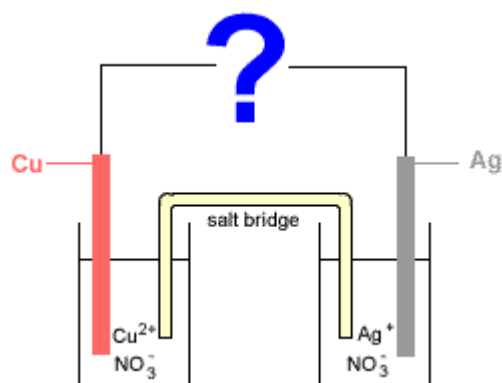A supply of electricity is required to drive the electrolytic cell reactions.

Electrolytic Cell Reactions			
anode:	X^-	--->	$X + e$ $\quad E^o_1$
cathode:	$M^+ + e$	--->	M $\quad E^o_2$
cell:	$X^- + M^+$	--->	$X + M \quad E^o_{cell} = -ve$

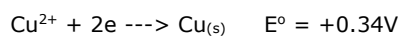$$E^o_{cell} = E^o_1 + E^o_1 = negative$$

EMF (volts) required to drive this non-spontaneous reaction is greater than E^o_{cell}

Example

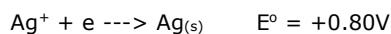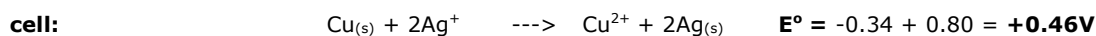Consider the following electrochemical cell:

The half-cell equations are:

$Cu^{2+} + 2e \text{ ---> } Cu_{(s)}$ $E^o = +0.34V$

and

$Ag^+ + e \text{ ---> } Ag_{(s)}$ $E^o = +0.80V$

Whether copper is oxidised or reduced depends on what the **?** in the diagram is.

? = voltmeter, galvanometer, or metal wire

If we replace the **?** in the diagram with a voltmeter or galvanometer to complete the circuit the reaction would proceed spontaneously:
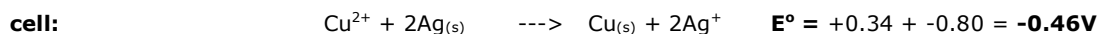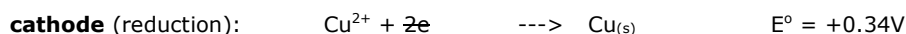
anode (oxidation):	$Cu_{(s)}$	--->	$Cu^{2+} + \cancel{2e}$	$E^o = -0.34V$
cathode (reduction):	$2Ag^+ + \cancel{2e}$	--->	$2Ag_{(s)}$	$E^o = +0.80V$
cell:	$Cu_{(s)} + 2Ag^+$	--->	$Cu^{2+} + 2Ag_{(s)}$	**$E^o = -0.34 + 0.80 = $ +0.46V**

Electrons would spontaneously flow from the negative copper anode to the positive silver cathode generating electricity which would be measured by the voltmeter or galvanometer.
The copper anode would disintegrate and silver would be deposited on the silver cathode.
This type of electrochemical cell is known as a galvanic or voltaic cell.

? = battery or other electricity supply with emf > 0.46V

If we replace the **?** in the diagram with a battery supplying more than 0.46V then the battery has greater electron pushing power than the spontaneous reaction in the electrochemical cell, forcing the reactions to go in the opposite direction to those above.

cathode (reduction):	$Cu^{2+} + \cancel{2e}$	--->	$Cu_{(s)}$	$E^o = +0.34V$
anode (oxidation):	$2Ag_{(s)}$	--->	$2Ag^+ + \cancel{2e}$	$E^o = -0.80V$
cell:	$Cu^{2+} + 2Ag_{(s)}$	--->	$Cu_{(s)} + 2Ag^+$	**$E^o = +0.34 + -0.80 = $ -0.46V**

Electrons are pulled out of the positive silver anode and flow to the negative copper cathode.
This reaction is not generating electricity it is using electricity!
The silver anode disintengrates while copper deposits on the copper cathode.
This type of electrochemical cell is known as an electrolytic cell.

Uses of Electrolysis

Recharging rechargable batteries
Plating one metal, eg, silver, gold or chromium, onto another metal (electroplating)
Production of sodium and chlorine from molten (fused) sodium chloride (Downs Cell)
Production of chlorine gas and sodium hydroxide from concentrated aqueous sodium chloride (Nelson or Diaphragm Cell)
Extraction of aluminium (Hall-Heroult Cells) and copper from their ores (electrowinning)
Refining of copper (electrorefining)

Calculating Cell EMF (voltage)

Some Standard Potentials

Weakest Oxidant	Oxidants	Reductants	E^0 (volts)	Strongest Reductant
|	$K^+ + e$	$K(s)$	-2.92	/\
|	$Ba^{2+} + 2e$	$Ba(s)$	-2.90	|
|	$Ca^{2+} + 2e$	$Ca(s)$	-2.87	|
|	$Na^+ + e$	$Na(s)$	-2.71	|
|	$Mg^{2+} + 2e$	$Mg(s)$	-2.36	|
|	$Al^{3+} + 3e$	$Al(s)$	-1.66	|
|	$Mn^{2+} + 2e$	$Mn(s)$	-1.18	|
|	$H_2O + e$	$\frac{1}{2}H_2(g) + OH^-$	-0.83	|
|	$Zn^{2+} + 2e$	$Zn(s)$	-0.76	|
|	$Fe^{2+} + 2e$	$Fe(s)$	-0.41	|
|	$Ni^{2+} + 2e$	$Ni(s)$	-0.23	|
|	$Sn^{2+} + 2e$	$Sn(s)$	-0.14	|
|	$Pb^{2+} + 2e$	$Pb(s)$	-0.13	|
|	**$H^+ + e$**	**$\frac{1}{2}H_2(g)$**	**0.00**	|
|	$SO_4^{2-} + 4H^+ + 2e$	$SO_2(aq) + 2H_2O$	0.17	|
|	$Cu^{2+} + 2e$	$Cu(s)$	0.35	|
|	$\frac{1}{2}O_2(g) + H_2O + 2e$	$2OH^-$	0.40	|
|	$Cu^+ + e$	$Cu(s)$	0.52	|
|	$\frac{1}{2}I_2(s) + e$	I^-	0.54	|
|	$\frac{1}{2}I_2(aq) + e$	I^-	0.62	|
|	$Fe^{3+} + e$	Fe^{2+}	0.77	|
|	$Ag^+ + e$	$Ag(s)$	0.80	|
|	$\frac{1}{2}Br_2(l) + e$	Br^-	1.07	|
|	$\frac{1}{2}Br_2(aq) + e$	Br^-	1.09	|
|	$\frac{1}{2}O_2(g) + 2H^+ + 2e$	H_2O	1.23	|
|	$\frac{1}{2}Cl_2(g) + e$	Cl^-	1.36	|
\/	$\frac{1}{2}Cl_2(aq) + e$	Cl^-	1.40	|
	$MnO_4^- + 8H^+ + 5e$	$Mn^{2+} + 4H_2O$	1.51	
Strongest Oxidant	$\frac{1}{2}F_2(g) + e$	F^-	2.87	Weakest Reductant

Key Concepts

- The overall electrochemical cell reaction can be written as 2 half-equations:
  1 equation for the reduction reaction and 1 equation for the oxidation reaction
- The number of electrons gained in the reduction half reaction must equal the number of electrons lost in the oxidation half reaction
- E^o is a type of energy per electron so it remains unchanged even if we double the numbers of reactants and products in the reaction
- If an equation is reversed (so that the reactants become the products), the sign of E^o is also reversed
- The cell's emf (electromotive force, often referred to as the cell voltage), is calculated by adding together the E^o values for each half reaction:
  $E^o_{cell} = E^o_{reduction} + E^o_{oxidation}$
- The reaction is spontaneous in the direction as written if
  $E^o_{cell} > 0$
  (E^o_{cell} positive)
- The reaction is spontaneous in the reverse direction to that written if
  $E^o_{cell} < 0$
  (E^o_{cell} negative)
- A galvanic cell (voltaic cell) produces electricity so the overall cell reaction must have a positive E^o_{cell} value
  ($E^o_{cell} > 0$)
- In practice, cell emf depends on temperature and concentration of reactants and products.
  If the concentration of reactants increases relative to products, the cell reaction becomes more spontaneous and the emf increases.
  As the cell operates, the reactants are used up as more product is formed causing the emf to decrease.
- An electrolytic cell requires an input of electricity so the overall cell reaction must have a negative E^o_{cell} value
  ($E^o_{cell} < 0$)
- If an electrolytic cell is operating in aqueous solution, then the water is also being reduced at the cathode or formed at the anode, so these equations must also be incorporated into the calculation of E^o_{cell}
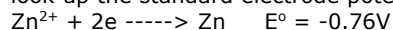
Calculating Cell EMF Examples

a. Calculate the emf (voltage) for the following reaction:
   $Zn(s) + Fe^{2+} -----> Zn^{2+} + Fe(s)$

   Write the 2 half reactions:
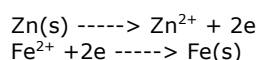
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Zn(s) -----> Zn^{2+} + 2e
Fe^{2+} + 2e -----> Fe(s)

look up the standard electrode potentials in the table above
Zn^{2+} + 2e -----> Zn E^o = -0.76V
This equation needs to be reversed, so the sign of E^o will also be reversed.
Zn(s) -----> Zn^{2+} + 2e E^o = +0.76V
Fe^{2+} + 2e -----> Fe(s) E^o = -0.41V

Add the two equations together:

Zn(s)	----->	Zn^{2+} + ~~2e~~	E^o = +0.76V
Fe^{2+} + ~~2e~~	----->	Fe(s)	E^o = -0.41V
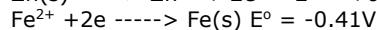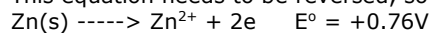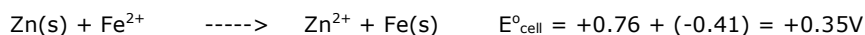
Zn(s) + Fe^{2+} -----> Zn^{2+} + Fe(s) E^o_{cell} = +0.76 + (-0.41) = +0.35V
E^o_{cell} > 0 (E^o_{cell} positive) so the reaction is spontaneous as written

b. Calculate the cell emf (voltage) for the following reaction:
Br_2(aq) + $2Fe^{2+}$ -----> $2Br^-$ + $2Fe^{3+}$
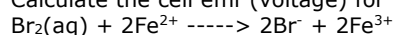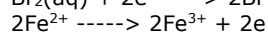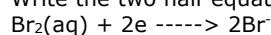
Write the two half equations:
Br_2(aq) + 2e -----> $2Br^-$
$2Fe^{2+}$ -----> $2Fe^{3+}$ + 2e

Look up the standard electrode potentials in the table above
½Br_2(aq) + e -----> Br^- E^o = +1.09V
This equation needs to be multiplied by 2, however, the value of E^o remains the same
Br_2(aq) + 2e -----> $2Br^-$ E^o = +1.09V
Fe^{3+} + e -----> Fe^{2+} E^o = +0.77V
This equation needs to be reversed, the sign of E^o will also be reversed
Fe^{2+} -----> Fe^{3+} + e E^o = +0.77V
This equation needs to be multiplied by 2, however, the value of E^o remains the same
$2Fe^{2+}$ -----> $2Fe^{3+}$ + 2e E^o = +0.77V

Add the two equations together:

Br_2(l) + ~~2e~~	----->	$2Br^-$	E^o = +1.09V
$2Fe^{2+}$	----->	$2Fe^{3+}$ + ~~2e~~	E^o = -0.77V
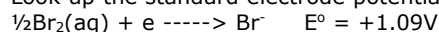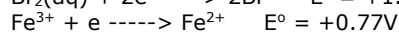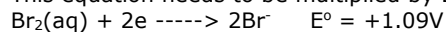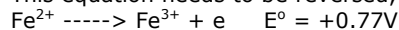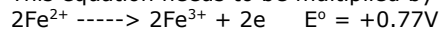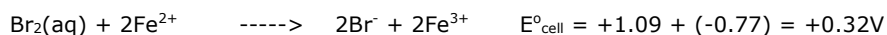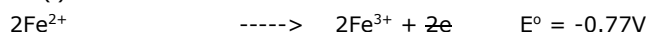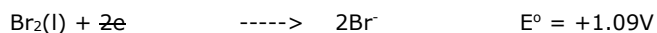
Br_2(aq) + $2Fe^{2+}$ -----> $2Br^-$ + $2Fe^{3+}$ E^o_{cell} = +1.09 + (-0.77) = +0.32V
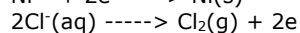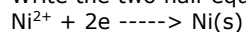E^o_{cell} is positive (E^o_{cell} > 0) so the reaction is spontaneous in the direction as written

c. Calculate the cell emf for the following reaction:
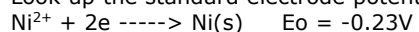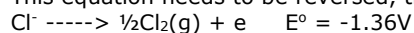Ni^{2+} + $2Cl^-$(aq) -----> Ni(s) + Cl_2(g)
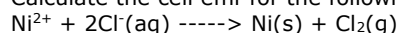
Write the two half equations:
Ni^{2+} + 2e -----> Ni(s)
$2Cl^-$(aq) -----> Cl_2(g) + 2e

Look up the standard electrode potentials in the table above:
Ni^{2+} + 2e -----> Ni(s)   Eo = -0.23V
½Cl_2(g) + e -----> Cl^- E^o = +1.36V
This equation needs to be reversed, the sign of E^o will also be reversed:
Cl^- -----> ½Cl_2(g) + e E^o = -1.36V
This equation needs to be multiplied by 2, however, the value of E^o will remain the same:
$2Cl^-$ -----> Cl_2(g) + 2e E^o = -1.36V

Add the two equations together:

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$Ni^{2+} + 2e$ -----> $Ni(s)$ $E^o = -0.23V$
$2Cl^-$ -----> $Cl_2(g) + 2e$ $E^o = -1.36V$

$Ni^{2+} + 2Cl^-(aq)$ -----> $Ni(s) + Cl_2(g)$ $E^o_{cell} = -0.23 + (-1.36) = -1.59V$
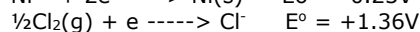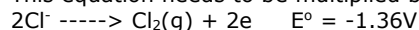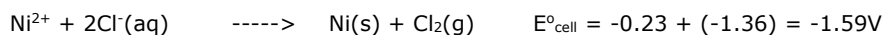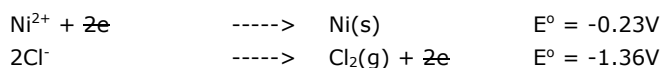
E^o_{cell} is negative, ($E^o_{cell} < 0$), so the reaction is NOT spontaneous in the direction as written.
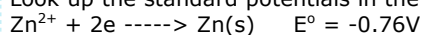The reaction would be spontaneous if reversed:
$Ni(s) + Cl_2(g)$ -----> $Ni^{2+} + 2Cl^-$ $E^o_{cell} = +1.59V$
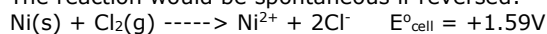
Calculating Galvanic Cell (Voltaic Cell) EMF (voltage)
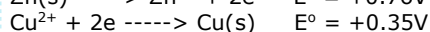
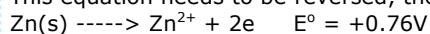The Daniell Cell is a typical Galvanic Cell: $Zn|Zn^{2+}||Cu^{2+}|Cu$
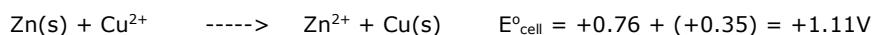
Write the two half equations:
$Zn(s)$ -----> $Zn^{2+} + 2e$
$Cu^{2+} + 2e$ -----> $Cu(s)$
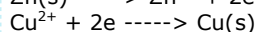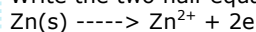
Look up the standard potentials in the table above:
$Zn^{2+} + 2e$ -----> $Zn(s)$ $E^o = -0.76V$
This equation needs to be reversed, the sign of E^o will also be reversed:
$Zn(s)$ -----> $Zn^{2+} + 2e$ $E^o = +0.76V$
$Cu^{2+} + 2e$ -----> $Cu(s)$ $E^o = +0.35V$

Add the two equations together:

$Zn(s)$ -----> $Zn^{2+} + 2e$ $E^o = +0.76V$
$Cu^{2+} + 2e$ -----> $Cu(s)$ $E^o = +0.35V$

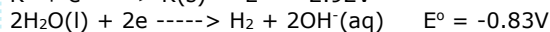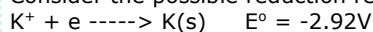$Zn(s) + Cu^{2+}$ -----> $Zn^{2+} + Cu(s)$ $E^o_{cell} = +0.76 + (+0.35) = +1.11V$
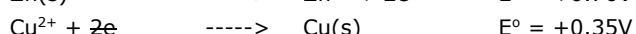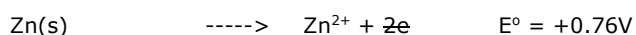
E^o_{cell} must be positive ($E^o > 0$) in a Galvanic Cell

Calculating EMF (voltage) for Electrolytic Cells containing Aqueous Solutions

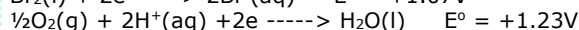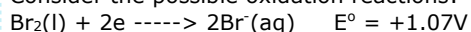A solution of potassium bromide (KBr) can be electrolysed.

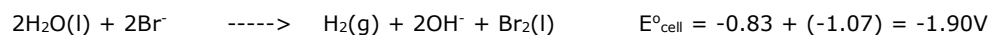Consider the possible reduction reactions:
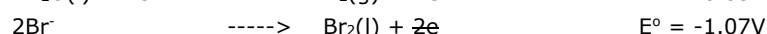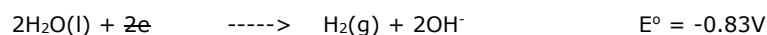$K^+ + e$ -----> $K(s)$ $E^o = -2.92V$
$2H_2O(l) + 2e$ -----> $H_2 + 2OH^-(aq)$ $E^o = -0.83V$
The strongest oxidant (weakest reductant) is water (H_2O). So $H_2(g)$ will be produced at the cathode.

Consider the possible oxidation reactions:
$Br_2(l) + 2e$ -----> $2Br^-(aq)$ $E^o = +1.07V$
$\frac{1}{2}O_2(g) + 2H^+(aq) + 2e$ -----> $H_2O(l)$ $E^o = +1.23V$
The strongest reductant (weakest oxidant) is Br^-. So, $Br_2(l)$ will be produced at the anode.

Add the equations together:

$2H_2O(l) + 2e$ -----> $H_2(g) + 2OH^-$ $E^o = -0.83V$
$2Br^-$ -----> $Br_2(l) + 2e$ $E^o = -1.07V$

$2H_2O(l) + 2Br^-$ -----> $H_2(g) + 2OH^- + Br_2(l)$ $E^o_{cell} = -0.83 + (-1.07) = -1.90V$

E^o_{cell} for an electrolytic cell must be negative ($E^o_{cell} < 0$).
A minimum of 1.90V must be applied to the cell in order for the electrolytic reaction to occur.

Balancing Half Equations

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Key Concepts

1. Write a skeletal equation for the oxidation or reduction equation based on the information provided.
2. Balance the half-reaction equation according to the following sequence:
    a. Balance all atoms other than H and O by inspection
    b. Balance O atoms by adding H_2O to the appropriate side
    c. Balance the H atoms. The way this is done depends on whether the solution is acidic or basic:
        ▪ Acidic Solution:
           add the appropriate number of H^+ to the side deficient in H
        ▪ Basic Solution:
           add one H_2O molecule to the side deficient in H AND one OH^- ion to the opposite side, for each H atom needed.
           You may need to cancel out H_2O molecules duplicated on each side at this point.
    d. Balance the charge by adding electrons (e) to the side deficient in negative charge.

Two half-reaction equations (one oxidation and one reduction) can be added together (as for calculating cell EMF).
In this case, multiply each balanced half equation by an appropriate number in order to balance the electrons lost in the oxidation reaction with the electrons gained in the reduction reaction.
Subtract, or cancel out, anything that appears on both sides. The electrons should cancel out in this step.

Examples

a. Write a balanced half equation for the oxidation of Fe^{2+} to Fe^{3+}
    1. Write the skeletal equation: $Fe^{2+} \longrightarrow Fe^{3+}$
    2. Balance atoms:
        a. Balance all atoms other than H or O: only 1 Fe^{2+} and 1 Fe^{3+} are required:
            $Fe^{2+} \longrightarrow Fe^{3+}$
        b. Balance O atoms: none present
        c. Balance H atoms: none present
        d. Balance charge: left hand side charge = +2, right hand side charge = +3, 1 electron is required on the right hand side in order to balance the charge:
            $Fe^{2+} \longrightarrow Fe^{3+} + e$
b. Write a balanced half equation for the reduction of Al^{3+} to Al(s)
0. Write the skeletal equation: $Al^{3+} \longrightarrow Al(s)$
    1. Balance atoms:
        a. Balance all atoms other than H or O: only 1 Al^{3+} and 1 Al are required:
            $Al^{3+} \longrightarrow Al(s)$
        b. Balance O atoms: none present
        c. Balance H atoms: none present
        d. Balance charge: left hand side charge = +3, right hand side charge = 0, 3 electrons are required on the left hand side in order to balance the charge:
            $Al^{3+} + 3e \longrightarrow Al(s)$
c. Write a balanced half equation for the reduction of $Cr_2O_7^{2-}$ to Cr^{3+} in acidic solution
0. Write the skeletal equation: $Cr_2O_7^{2-} \longrightarrow Cr^{3+}$
    1. Balance atoms:
        a. Balance all atoms other than H or O:
            2 Cr atoms are on the left hand side, only 1 Cr is present on the right hand side hand of the equation.
            $Cr_2O_7^{2-} \longrightarrow 2Cr^{3+}$
        b. Balance O atoms by adding H_2O to the side deficient in O
            $Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + H_2O$
            7 O atoms are present on the left hand side, 1 on the right hand side of the equation:
            $Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$
        c. Balance H atoms as for an acidic solution by adding H^+ to the side deficient in H
            $Cr_2O_7^{2-} + H^+ \longrightarrow 2Cr^{3+} + 7H_2O$
            There is 1 H atom on the left hand side, and 14 H atoms on the right hand side of the equation:
            $Cr_2O_7^{2-} + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O$
        d. Balance charge:
            charge on left hand side = -2 + 14 = +12
            charge on right hand side = 2 x +3 = +6
            difference in charge = +12 - +6 = +6 So 6 electrons are required on the left hand side of the equation:
            $Cr_2O_7^{2-} + 14H^+ + 6e \longrightarrow 2Cr^{3+} + 7H_2O$
d. Write a balanced half equation for the reduction of CrO_4^{2-} to CrO_2^- in basic solution
Write the skeletal equation: $CrO_4^{2-} \longrightarrow CrO_2^-$

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1. Balance atoms:
    a. Balance all atoms other than H or O:
       1 Cr atom is on the left hand side, 1 Cr is present on the right hand side hand of the equation.
       CrO_4^{2-} -----> CrO_2^-
    b. Balance O atoms by adding H_2O to the side deficient in O
       CrO_4^{2-} -----> CrO_2^- + H_2O
       4 O atoms are present on the left hand side, 3 on the right hand side of the equation:
       CrO_4^{2-} -----> CrO_2^- + $2H_2O$
    c. Balance H atoms as for a basic solution by adding H_2O to the side deficient in H and OH^- to the opposite side
       CrO_4^{2-} + H_2O-----> CrO_2^- + $2H_2O$ + OH^-
       There are 2 H atoms on the left hand side, and 5 H atoms on the right hand side of the equation:
       CrO_4^{2-} + $4H_2O$-----> CrO_2^- + $2H_2O$ + $4OH^-$
       Cancel out H_2O molecules appearing on both sides of the equation:
       CrO_4^{2-} + $2H_2O$-----> CrO_2^- + $4OH^-$
    d. Balance charge:
       charge on left hand side = -2
       charge on right hand side = -1 + -4 = -5
       difference in charge = -2 - -5 = +3 So 3 electrons are required on the left hand side of the equation:
       CrO_4^{2-} + $2H_2O$ + 3e -----> CrO_2^- + $4OH^-$

Faraday Laws of Electrolysis

Key Concepts

○ First Law

The quantity of a substance produced by electrolysis is proprotional to the quantity of electricty used.

○ Second Law

For a given quantity of electricity the quantity of substance produced is proportional to its weight.

○ The quantity of electricity or charge contained in a current running for a specified time can be calculated:

$Q = I \times t$

Q = quantity of electricity or charge in coulombs (C)
I = current in amps (A)
t = time (seconds)

○ The Faraday constant, F, is the quantity of electricity carried by one mole of electrons.

F = Avogadro's Number x charge on electron in coulombs

$F = 6.022 \times 10^{23}$ mol^{-1} x 1.602192×10^{-19} C
F = 96,484 C mol^{-1}
This is usually rounded off to 96,500 C mol^{-1} for calculations in chemistry.

○ The quantity of electricity required to deposit an amount of metal can be calculated:

$$Q = n(e) \times F$$

Q = quantity of electricity in coulombs (C)
n(e) = moles of electrons
F = Faraday constant = 96, 500 C mol^{-1}

○ Electrical Energy, E, can be calculated:

$$E = Q \times V$$

E = electrical energy in joules (J)
Q = quantity of electricity in coulombs (C)
V = voltage (or EMF) in volts (V)

○ 1 kilowatt-hour, kWH, is a unit of electrical energy.

$$1 \text{ kWH} = 3.6 \times 10^6 \text{ J}$$

Examples

a. $Q = I \times t$

   Calculate the quantity of electricity, Q, obtained when a current of 25 amps runs for 1 minute.

   Q = ? C
   I = 25 A
   t = 1 minute = 60 seconds

   Q = 25 x 60 = 1,500 C

b. $I = Q \div t$

   Calculate the current needed to provide 30,000 coulombs of electricity in 5 minutes.

   Q = 30,000 C
   I = ? A
   t = 5 minutes = 5 x 60 = 300 seconds

   I = Q ÷ t = 30,000 ÷ 300 = 100 amps

c. $t = Q \div I$

   Calculate the time required to produce 12,000 C of electricity using a current of 10 amps.
   Q = 12,000 C
   I = 10 A
   t = ?

   t = Q ÷ I = 12,000 ÷ 10 = 1,200 seconds = 1,200 ÷ 60 = 20 minutes

d. $Q = n(e) \times F$

   Calculate the quantity of electricity obtained from 2 moles of electrons
   Q = n x F
   Q = ?
   n = 2 mol
   F = 96,500 C mol^{-1}

Q = 2 x 96,500 = 193,000 C

e. n(e) = Q ÷ F

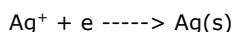Calculate the moles of electrons obtained from 250 C of electricity
n(e) = ? mol
Q = 250 C
F = 96,500 C mol^{-1}

n(e) = 250 ÷ 96,500 = 2.59 x 10^{-3} mol

f. Calculate the time required to deposit 56g of silver from a silver nitrate solution using a current of 4.5A.
   1. Calculate the moles of electrons required for the reaction:

     $Ag^+ + e ----> Ag(s)$
     moles of Ag(s) deposited, n(Ag) = moles of electrons required, n(e)
     moles of Ag = n(Ag) = mass ÷ MM
     mass Ag deposited = 56g
     MM = 107.9 g mol^{-1} (from Periodic Table)
     n (Ag) 56 ÷ 107.9 = 0.519 mol = n(e)

   2. Calculate the quantity of electricity required: Q = n(e) x F
     Q = ? C
     n(e) = 0.519 mol
     F = 96,500 C mol^{-1}
     Q = 0.519 x 96,500 = 50,083.5 C
   3. Calculate the time required: t = Q ÷ I
     Q = 50,083.5 C
     I = 4.5 A
     t = 50,083.5 ÷ 4.5 = 11,129.67 seconds
     t = 11,129.67 ÷ 60 = 185.5 minutes
     t = 185.5 ÷ 60 = 3.1 hours

g. What mass of copper could be deposited from a copper (II) sulphate solution using a current of 0.50 A over 10 seconds?
0. Calculate the quantity of electricity: Q = I x t
   I = 0.50 A
   t = 10 seconds
   Q = 0.50 x 10 = 5.0 C
1. Calculate the moles of electrons: n(e) = Q ÷ F
Q = 5.0 C
F = 96,500 C mol^{-1}
n(e) = 5.0 ÷ 96,500 = 5.18 x 10^{-5} mol
2. Calculate mass of copper: mass = n x MM
$Cu^{2+} + 2e ----> Cu(s)$
1 mole of copper is deposited from 2 moles electrons
n(Cu) = ½n(e) = ½ x 5.18 x 10^{-5} = 2.59 x 10^{-5} mol
MM = 63.55 g mol^{-1} (from Periodic Table)
mass (Cu) = (2.59 x 10^{-5}) x 63.55 = 1.65 x 10^{-3} g = 1.65 mg
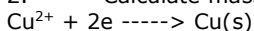
h. An EMF of 4.5 V produces 1 kg of sodium metal by the electrolysis of Na$^+$.
Calculate the minimum number of kilowatt-hours of electricity needed to produce the sodium metal.
0. Calculate the moles of electrons, n(e), required
   Write the equation for the electrolysis of Na$^+$:
   $Na^+ + e ----> Na(s)$
   moles of Na(s) = moles of electrons used n(e)
   n(e) = n(Na) = mass ÷ MM
   mass = 1kg = 1,000g
   MM = 22.99 g mol^{-1} (from Periodic Table)
   n(e) = 1,000 ÷ 22.99 = 43.497 mol
1. Calculate the quantity of electricity required: Q = n(e) x F
n(e) = 43.497 mol
F = 96,500 C mol^{-1}
Q = n(e) x F = 43.497 x 96,500 = 4.2 x 10^6 C
2. Calculate the electrical energy: E = Q x V
Q = 4.2 x 10^6
V = 4.5 V
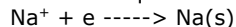
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E = 4.2 x 10^6 x 4.5 = 1.89 x 10^7 J
3.   Convert Electrical Energy to kilowatt-hours: kilowatt-hours = E ÷ 3.6 x 10^6
E = (1.89 x 10^7) ÷ (3.6 x 10^6) = 5.25 kWH

Corrosion

Key Concepts

- Corrosion: the oxidation of metals by certain substances such as water and oxygen in the environment.
- Rusting of iron is a common form of corrosion.
- Noble Metal: a metal resistant to corrosion and found uncombined in nature, eg, gold and silver. These are also known as Cathodic, or Protected, Metals.
- Ignoble Metal: a metal that corrodes. These are also known as Anodic, or Corroding, Metals.
- Passivating metal: a reactive metal that forms an inactive coating as a result of reacting with substances such as oxygen and water, eg, aluminium and chromium.
- In general*, corrosion is a spontaneous electrochemical process which requires:

a.   an anode where oxidation occurs
b.   a cathode where reduction occurs
c.   a metal path to allow for the flow of electrons
d.   an electrolyte

- Common methods of corrosion prevention are:

a.   Painting
b.   Plating with another metal
c.   Cathodic protection
d.   Alloying with another metal

*Biological activity can also corrode metals.

Example: Rusting of Iron
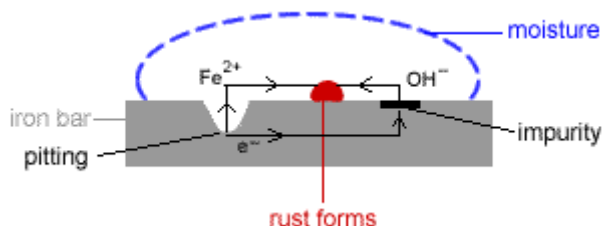
$Fe_{(s)}$ is oxidised to Fe^{2+} and a pit develops on the surface of the iron.

At the site of an impurity oxygen is reduced to hydroxide ions.

Electrons flow through the iron bar from the pit to the impurity.

Ions flow through the water electrolyte to the rusting site.

At the rusting site, iron (II) hydroxide is formed which readily undergoes oxidation to form hydrated iron (III) oxide (rust).

Anode (oxidation)	pitting develops	$Fe_{(s)}$	--->	$Fe^{2+}_{(aq)}$ + 2e
Cathode (reduction)	site of impurity	$O_{2(g)}$ + $2H_2O_{(l)}$ + 4e	--->	$4OH^-_{(aq)}$
Cell Reactions	site of rust	$Fe^{2+}_{(aq)}$ + $2OH^-_{(aq)}$	--->	$Fe(OH)_{2(s)}$
REDOX Reaction		$4Fe(OH)_{2(s)}$ + $O_{2(g)}$	--->	$2Fe_2O_3.H_2O_{(s)}$ + $2H_2O_{(l)}$

Factors Influencing the Rate of Rusting

Purity of the iron:
Pure iron rusts very, very slowly.
Rusting occurs very quickly if a less active metal impurity is present, eg, copper or tin.
Amount of Stress:
Bending, cutting and sharpening the iron place it under stress which distorts the arrangement of metal ions in the lattice.
Fe^{2+} ions can escape from the metal lattice more easily, so rusting occurs faster.
Electrolyte Composition:
Pure water is a poor electrical conductor so rusting occurs more slowly.
Salt water is a much better electrical conductor so rusting occurs more rapidly.
Oxygen Concentration:
A large concentration of oxygen at the surface of iron will speed up the corrosion of another area with lower oxygen concentration if there is a path for electrons and ions to follow between the two areas.

## Rust Prevention Methods

Painting
Prevents oxygen and water coming into contact with the iron so rust can't form.
If the paint is scratched or chipped, oxygen and water can come into contact with the iron so rust will then form.
Plating with Another Metal
Plating with a less active metal like tin is just like painting. The tin covering prevents the iron coming into contact with water and oxygen, but, if the tin plate is scratched, rust will readily form.

Plating with a more active metal like zinc in the galvanising process, firstly acts like painting in that water and oxygen can not get to the iron surface, but, if the surface is scratched this sets up a new galvanic electrochemical cell:

$$Zn_{(s)} + Fe^{2+}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + Fe_{(s)}$$
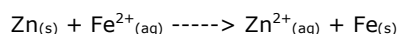
The iron is reformed in this process.
Zinc ions will form zinc oxide (or carbonate) which is less permeable to water and oxygen than iron oxides so this slows down the rate of corrosion.

Cathodic Protection:
A Sacrifical Anode is a block of active metal such as zinc or magnesium attached to the hull of a ship, or to tanks and pipelines buried in moist earth.
The more active zinc sacrifices itself by corroding in preference to the iron.
Sacrifical anodes must be replaced when they have corroded in order for the protection of iron to be maintained.

An Impressed Current System is used to protect immersed hulls of ships using noble metal anodes and the ship's own electrical system.
A reference cell such as silver/silver chloride is mounted to detect voltage differences between itself and the hull.
This voltage difference is measured and a reactor rectifier is used to control the current from the ship's electrical system to the external anodes in order to prevent corrosion.

Alloying with Another Metal
Alloying iron with chromium produces stainless steel.
The chromium readily forms an impermeable (impervious) oxide layer that adheres strongly to the iron surface preventing the formation of rust.

Electrochemical Cells & Batteries
Key Concepts

Battery: cells in a series

Cells:

- Primary cells: cannot be recharged

    examples: Leclanché, Alkaline cell, Lithium battery, Button cell

- Secondary cells: can be recharged by applying voltage greater than that of the spontaneous cell reaction which reverses the half reactions

    examples: Lead-acid battery, Nickel-cadmium cell, Vanadium redox battery

- Fuel cells: those that function continuously due to continued supply of reactants

    examples: hydrogen-oxygen fuel cell

Examples

PRIMARY CELLS

Leclanché or dry cell : $Zn|(ZnCl_2),NH_4Cl|MnO_2,C$

- Anode (-): zinc case
  $Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$
- Cathode (+): carbon rod in contact with carbon & MnO_2
  $NH_4^+(aq) + MnO_2(s) + e^- \longrightarrow NH_3(aq) + MnO(OH)(s)$
- Electrolyte: Paste of NH_4Cl, $ZnCl_2$ (acid electrolyte)
- Voltage: 1.5 volts
- Uses: low drain appliances such as radios, torches, liquid crystal calculators, toys

Alkaline cell : $Zn,ZnO|KOH(paste)|MnO_2,C$

- Anode (-): brass rod in contact with powdered zinc
  $Zn(s) + 2OH^-(aq) \longrightarrow ZnO(s) + H_2O(l) + 2e^-$
- Cathode (+): steel case in contact with carbon & MnO_2
  $MnO_2(s) + H_2O(l) + e^- \longrightarrow MnO(OH)(s) + OH^-(aq)$
- Electrolyte: paste of 7M KOH (alkaline electrolyte)
- Voltage: 1 volt
- Uses: High current or high total capacity appliances such as electronic photographic flash units, tape recorders, radios, shavers
  Can deliver higher current than Leclanché cell without severe voltage drop.

Mercury button cell

- Anode (-): steel cap in contact with powdered zinc
  $Zn(s) + 2OH^-(aq) \longrightarrow ZnO(s) + H_2O(l) + 2e^-$
- Cathode (+): steel case in contact with powdered HgO
  $HgO(s) + H_2O(l) + 2e^- \longrightarrow Hg(l) + 2OH^-(aq)$
- Electrolyte: paste of KOH (alkaline electrolyte)
- Voltage: 1.3 volts
- Uses: watches, pacemakers, hearing aids, microphones

Silver button cell: $Zn,ZnO|KOH (paste)|Ag_2O,Ag$

- Anode (-): steel cap in contact with powdered zinc
  $Zn(s) + 2OH^-(aq) \longrightarrow ZnO(s) + H_2O(l) + 2e^-$
- Cathode (+): steel case in contact with powdered Ag_2O
  $Ag_2O(s) + H_2O(l) + 2e^- \longrightarrow 2Ag(s) + 2OH^-(aq)$

- Electrolyte: paste of KOH (alkaline electrolyte)
- Voltage: 1.3 volts
- Uses: watches, pacemakers, hearing aids
  As cell discharges, concentrations do NOT change as reactants & products are solids.
  [OH^-] in electrolyte does NOT change.
  So, the voltage of this cell remains constant throughout its lifetime.

Lithium cell : $Li|Li^+(nonaqueous)|KOH(paste)|MnO_2,Mn(OH)_3,C$

- Anode (-): lithium
  $Li ----->Li_+ + e^-$
- Cathode (+): carbon in contact with manganese(III),manganese(IV) electrode
  $MnO_2+2H_2O+e^- ----->Mn(OH)_3+OH^-$
- Electrolyte: paste of KOH (alkaline electrolyte)
- Voltage:
- Uses: auto-wind, auto-focus cameras
  Can deliver high current.
  Have long shelf-life.

Lithium iodide solid state cell : $Li|LiI(s)|I_2(in\ polymer)$

- Anode (-): lithium
  $Li(s) -----> Li^+ + e^-$
- Cathode (+): iodine
  $\frac{1}{2}I_2(s) + e^- -----> I^-$
- Electrolyte: lithium iodide
- Voltage: 2 volts
- Uses: cardiac pacemaker, watches, calculators
  Lasts about 10 years

SECONDARY CELLS

Lead-acid battery (or accumulator) : $Pb,PbSO_4|H_2SO_4(aq)|PbO_2,Pb$

- Anode (-): lead plates
  $Pb(s)+HSO_4^-(aq) ----->PbSO_4(s)+H^+(aq)+2e^-$
- Cathode (+): lead plates covered with PbO_2
  $PbO_2(s)+3H^+(aq)+HSO_4^-(aq)+2e^- ----->PbSO_4(s)+2H_2O(l)$
- Electrolyte: 4M H_2SO_4
- Voltage: 2.1 volts per cell (6 cells in series makes up a 12volt car battery)
- Uses: motor vehicle batteries
  As cell delivers current, lead sulfate deposits on electrodes & [H_2SO_4] falls.
  Six 2V cells connected in series make up a 12V battery.
  To obtain high current output from the battery, the surface area of the electrodes in contact with the electrolyte must be large & the plates close together.

Nickel-cadmium cell (NiCad) : $Cd,Cd(OH)_2|KOH|NiO_2,Ni(OH)_2,Ni$

- Anode (-): grid covered with cadmium
  $Cd(s)+2OH^-(aq) ----->Cd(OH)_2(s)+2e^-$
- Cathode (+): grid covered with nickel(III)hydroxide
  $NiO(OH)(s)+H_2O(l)+e^- ----->Ni(OH)_2(s)+OH^-(aq)$
- Electrolyte: KOH
- Voltage: 1.25volts
- Uses: video cameras, phones, cordless drills, laptop computers

FUEL CELLS

promising future
more efficient than coal or gas-fired power stations
non-polluting

porous electrodes lined with catalysts, eg, Ni or Pt, making it expensive
systems required to continuously supply reactants & remove products
CH_4, CH_3OH, C_3H_8 can be used instead of $H_2(g)$

Alkaline fuel cell

Anode (-): porous carbon anode with catalyst in contact with $H_2(g)$
$H_2(g)+2OH^-(aq)----->2H_2O(l)+2e^-$
Cathode (+): porous carbon cathode with catalyst in contact with $O_2(g)$
$O_2(g)+2H_2O(l)+4e^----->4OH^-(aq)$
Electrolyte: hot KOH(aq)
Voltage: ~1volt

Acid fuel cell

Anode (-): porous carbon anode with catalyst in contact with $H_2(g)$
$H_2(g)----->2H^+(aq)+2e^-$
Cathode (+): porous carbon cathode with catalyst in contact with $O_2(g)$
$O_2(g)+4H^+(aq)+4e^----->2H_2O(l)$
Electrolyte: hot $H_3PO_4(aq)$
Voltage: ~1volt

Sodium Hydroxide Production by Electrolysis

Key Concepts

- Sodium hydroxide, NaOH, is also known as lye or caustic soda.
- Sodium hydroxide is a commonly used base.
- Electrolysis of concentrated sodium chloride solutions (brine) produces chlorine gas, hydrogen gas and aqueous sodium hydroxide.

$$2NaCl_{(aq)} + 2H_2O_{(l)} -----> H_{2(g)} + Cl_{2(g)} + 2NaOH_{(aq)}$$

- $Cl_{2(g)}$ is produced at the anode (positive electrode).
- $H_{2(g)}$ and $NaOH_{(aq)}$ are produced at the cathode (negative electrode).
- Three types of electrolytic cell are used to produce sodium hydroxide from brine:

Castner-Kellner Cell (Mercury Process)
Nelson Diaphragm Cell
Membrane Cell

Properties and Uses of the Electrolysis Products

sodium hydroxide NaOH	hydrogen H_2	chlorine Cl_2
Properties		
- strong base - water soluble - dissolving NaOH in water is an exothermic reaction - deliquescent : absorbs moisture from the air - absorbs carbon dioxide from the air	- flammable gas	- poisonous yellow gas

Uses

domestic oven and drain cleaner	margarine	PVC
pulp and paper production	nylon	dry-cleaning solvent
manufacture of rayon and other synthetic fibres	HCl (gas and acid)	HCl (gas and acid)
soap and detergent production		bleaches
extraction of alumina from bauxite (Bayer Process)		weed-killers
removal of sulfur dioxide and hydrogen sulfide in petroleum refining		paints and dyes
		anti-bacterial agent in water

Electrolytic Cells for the Production of Sodium Hydroxide

Castner-Kellner Cell (Mercury Process)
- o Anode (positive electrode): titanium
  Anode reaction (oxidation):
  $2Cl^-_{(aq)} \longrightarrow Cl_{2(g)} + 2e$
- o Cathode (negative electrode): mercury flowing along bottom of cell
  Cathode reaction (reduction):
  $Na^+_{(aq)} + e \longrightarrow Na_{(s)}$
- o $Na_{(s)}$ dissolves in the liquid mercury to form an amalgam which is removed to the decomposer.
- o In the decomposer the amalgam reacts with water to form sodium hydroxide, hydrogen gas and mercury.
  $2Na/Hg + 2H_2O_{(l)} \longrightarrow 2Na^+ + 2OH^- + H_{2(g)} + 2Hg_{(l)}$
  Sodium hydroxide solution and hydrogen gas are collected.
  Mercury is recycled through the electrolytic cell.
- o Mercury is toxic so care must be taken to prevent mercury losses.
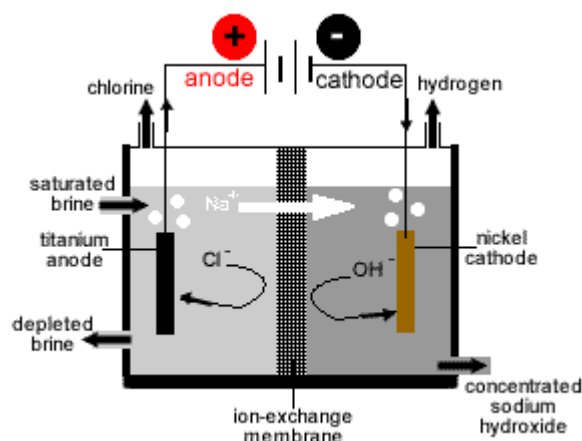

Nelson Diaphragm Cell

- o Porous diaphragm of asbestos or metal oxide with polymer separates anode and cathode compartments.
- o Diaphragm prevents hydroxide ions entering anode compartment and prevents chloride ions entering cathode compartment.
- o Saturated brine enters anode compartment where chlorine gas is produced.
- o Anode (positive electrode): carbon (graphite) or titanium coated with Ru-Ti oxide.
- o Anode reaction (oxidation):
  $2Cl^-_{(aq)} \longrightarrow Cl_{2(g)} + 2e$
- o Cathode (negative electrode): steel mesh
- o Cathode reaction (reduction):

$2H_2O_{(l)} + 2e \longrightarrow H_{2(g)} + 2OH^-_{(aq)}$
- o Na^+ migrates across diaphragm to cathode compartment combining with OH^- to form NaOH.
- o Overall cell reaction (showing Na^+ spectator ions):
  $2H_2O_{(l)} + 2Cl^-_{(aq)} + 2Na^+_{(aq)} \longrightarrow 2Na^+_{(aq)} + 2OH^-_{(aq)} + H_{2(g)} + Cl_{2(g)}$
- o Product contains sodium chloride and sodium hydroxide. $NaOH_{(s)}$ can be crystallised out.
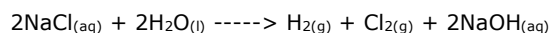
Membrane Cell

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- o Favoured method for producing sodium hydroxide as there is no hazardous waste, very pure sodium hydroxide is produced and it has the lowest energy requirements.
- o Ion-exchange membrane selectively allows Na^+ and water to flow to the cathode compartment but prevents products from moving between compartments.
- o Saturated brine enters the anode compartment where chlorine gas is formed.
- o Anode (positive electrode): titanium
- o Anode reaction (oxidation): $2Cl^-_{(aq)} \longrightarrow Cl_{2(g)} + 2e$
- o Cathode (negative electrode): nickel
- o Cathode reaction (reduction): $2H_2O_{(l)} + 2e \longrightarrow H_{2(g)} + 2OH^-_{(aq)}$
- o Na^+ migrates across the membrane to cathode compartment combining with OH^- to form NaOH.
- o Overall cell reaction (showing Na^+ spectator ions): $2H_2O_{(l)} + 2Cl^-_{(aq)} + 2Na^+_{(aq)} \longrightarrow 2Na^+_{(aq)} + 2OH^-_{(aq)} + H_{2(g)} + Cl_{2(g)}$
- o Product is concentrated sodium hydroxide.

Sodium Hydroxide Production by Electrolysis

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Properties and Uses of the Electrolysis Products

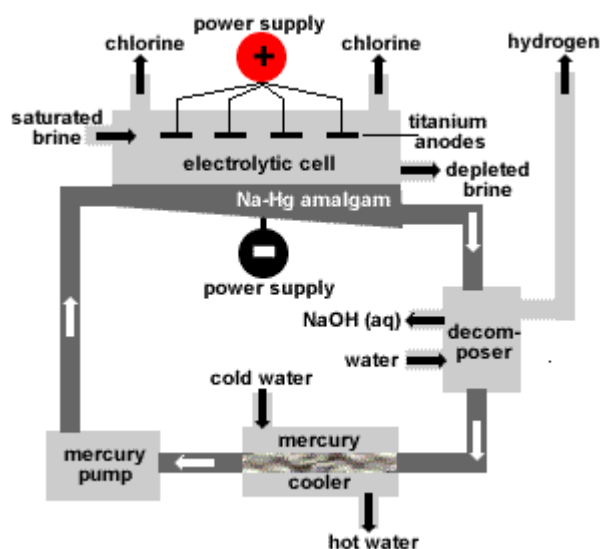| sodium hydroxide NaOH	hydrogen H_2	chlorine Cl_2
Properties - strong base - water soluble - dissolving NaOH in water is an exothermic reaction - deliquescent : absorbs moisture from the air - absorbs carbon dioxide from the air	- flammable gas	- poisonous yellow gas

Uses

- domestic oven and drain cleaner
- pulp and paper production
- manufacture of rayon and other synthetic fibres
- soap and detergent production
- extraction of alumina from bauxite (Bayer Process)
- removal of sulfur dioxide and hydrogen sulfide in petroleum refining

- margarine
- nylon
- HCl (gas and acid)

- PVC
- dry-cleaning solvent
- HCl (gas and acid)
- bleaches
- weed-killers
- paints and dyes
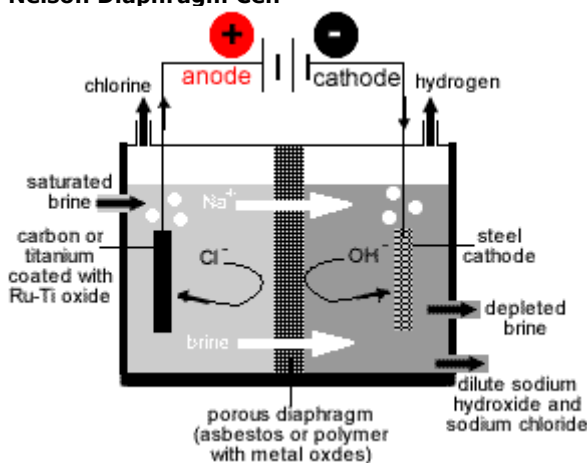- anti-bacterial agent in water

Electrolytic Cells for the Production of Sodium Hydroxide

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- Cathode reaction (reduction):
  $2H_2O_{(l)} + 2e -----> H_{2(g)} + 2OH^-_{(aq)}$
- Na^+ migrates across diaphragm to cathode compartment combining with OH^- to form NaOH.
- Overall cell reaction (showing Na^+ spectator ions):
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- Product contains sodium chloride and sodium hydroxide. $NaOH_{(s)}$ can be crystallised out.

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- o Ion-exchange membrane selectively allows Na^+ and water to flow to the cathode compartment but prevents products from moving between compartments.
- o Saturated brine enters the anode compartment where chlorine gas is formed.
- o Anode (positive electrode): titanium
- o Anode reaction (oxidation): $2Cl^-_{(aq)}$ -----> $Cl_{2(g)} + 2e$
- o Cathode (negative electrode): nickel
- o Cathode reaction (reduction): $2H_2O_{(l)} + 2e$ -----> $H_{2(g)} + 2OH^-_{(aq)}$
- o Na^+ migrates across the membrane to cathode compartment combining with OH^- to form NaOH.
- o Overall cell reaction (showing Na^+ spectator ions): $2H_2O_{(l)} + 2Cl^-_{(aq)} + 2Na^+_{(aq)}$ -----> $2Na^+_{(aq)} + 2OH^-_{(aq)} + H_{2(g)} + Cl_{2(g)}$
- o Product is concentrated sodium hydroxide.