



2020
Version

Chemistry AS 91393

C3.7 Redox

What is this NCEA Achievement Standard?

When a student achieves a standard, they gain a number of credits. Students must achieve a certain number of credits to gain an NCEA certificate (80 for Level 3)

The standard you will be assessed on is called **Chemistry 3.7 AS91393** Demonstrate understanding of oxidation-reduction processes

It will be internally (in Class) assessed as part of a **In-Class Examination** and will count towards **3 credits** for your Level 3 NCEA in Chemistry



What are the main steps required in this Internal Assessment?

AS91393 Demonstrate understanding of oxidation-reduction processes

Interpretation of evidence for Achieved

The student demonstrates an understanding of the oxidation-reduction processes involved in discharging and recharging of batteries.

- ☐ Can identify reactants and products /can write $\frac{1}{2}$ equations.
- ☐ Can identify what oxidant/reductant during charging and discharge
- ☐ Can identify oxidation number of the species involved
- ☐ Can link energy output during battery discharge and energy input during charging



Interpretation of evidence for Merit

The student demonstrates an in-depth understanding of the reduction-oxidations processes involved in discharging and recharging of batteries.

ACHIEVED PLUS

- ☐ Can write balanced half equations for the charging and discharging processes
- ☐ Can calculate cell potentials



Interpretation of evidence for Excellence

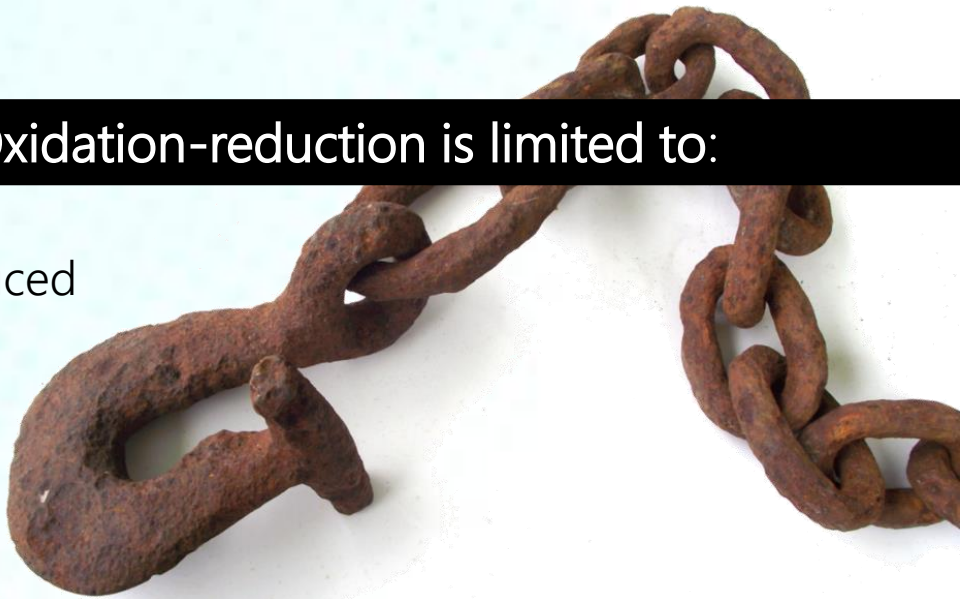
The student demonstrates a comprehensive understanding of the oxidation-reduction processes involved in discharging and recharging of batteries.

MERIT PLUS

- ☐ Can write fully balanced equations for the discharging and charging reactions
- ☐ Can write the cell expressions for both discharging and charging
- ☐ Can compares the charge and discharge processes in terms of spontaneity, products, and oxidant/reductant

In this Achievement Standard Oxidation-reduction is limited to:

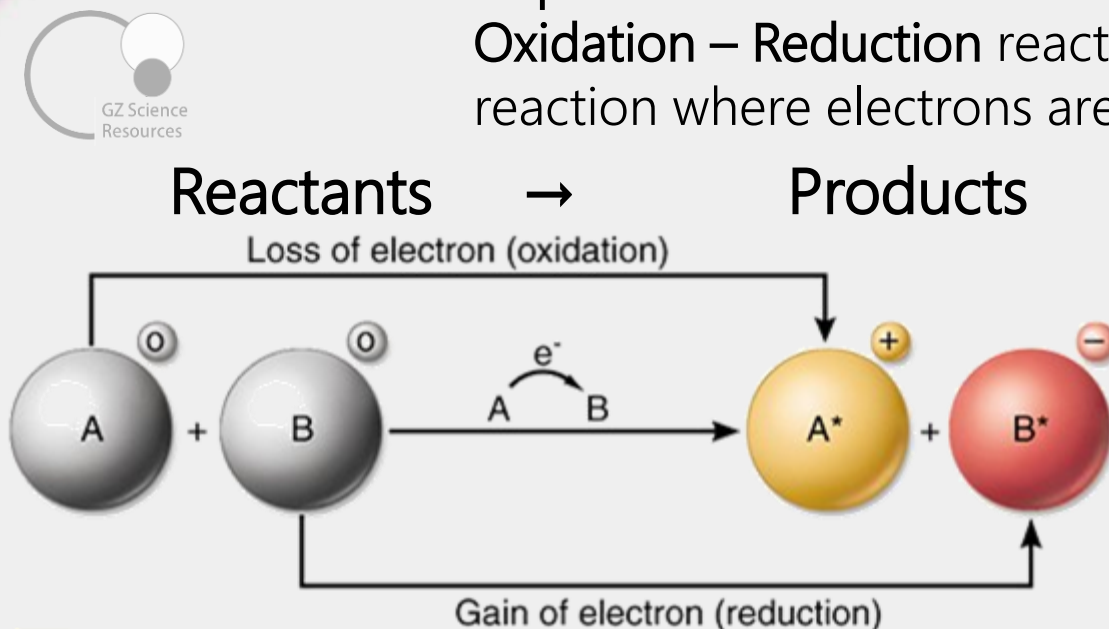
- ☐ identify the species oxidised and reduced
- ☐ identify oxidation numbers in relation to species
- ☐ write balanced half and full oxidation-reduction equations
- ☐ give a conventional cell diagrams (not required for assessment)
- ☐ calculate cell potentials using data provided
- ☐ make and explain links between the calculations and spontaneity of the reactions
- ☐ elaborate on the recharge process of batteries.
- ☐ justify why the recharge process is necessary in terms of amount of species
- ☐ compare and contrast the discharge and recharge processes in the battery



Redox Reactions - reactants & products

A **chemical reaction** is a process that produces a chemical change to one or more substances and will produce one or more **new substances**. Other observations may include a temperature change, a colour change or production of gas. Chemicals that are used in a chemical reaction are known as **reactants**. Those that are formed are known as **products**.

Oxidation – Reduction reactions are a specific type of reaction where electrons are transferred



A reactant and what product it changes into after the redox reaction is known as a **species** i.e. Cu changing to Cu^{2+} so Cu/Cu^{2+} is the species

RedOx terms

A redox reaction is where one reactant is oxidised and the other reactant is reduced.

Reduction and oxidation occur in pairs of reactants

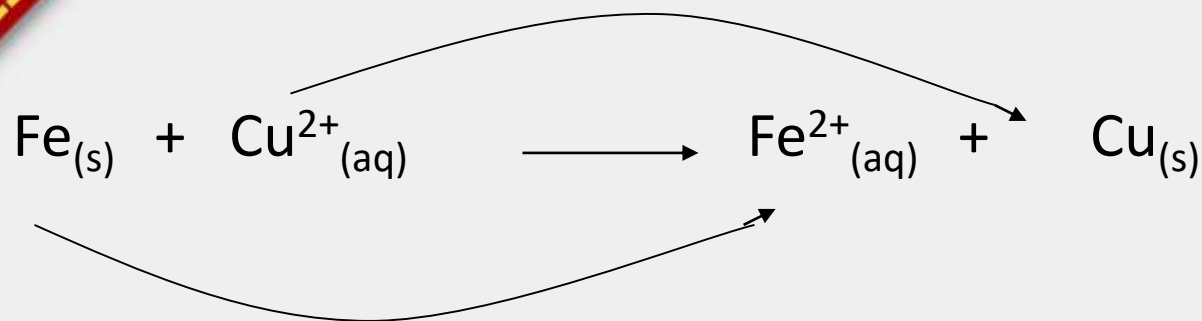
Oxidation of one reactant	Reduction of the other reactant
<input type="checkbox"/> loss of electrons and a	<input type="checkbox"/> gain of electrons
<input type="checkbox"/> loss of hydrogen and a	<input type="checkbox"/> gain of hydrogen
<input type="checkbox"/> gain of oxygen and a	<input type="checkbox"/> loss of oxygen

Oxidation numbers are used to determine what is oxidised and what is reduced in a reaction. These will be explained later

An Iron nail left in copper sulfate

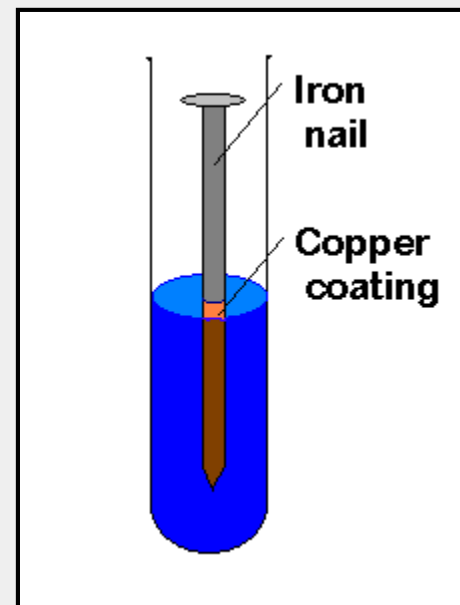
Copper is reduced – gained electrons

Oxidising agent (oxidant)



Iron is oxidised – lost electrons

Reducing Agent (reductant)



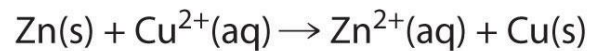
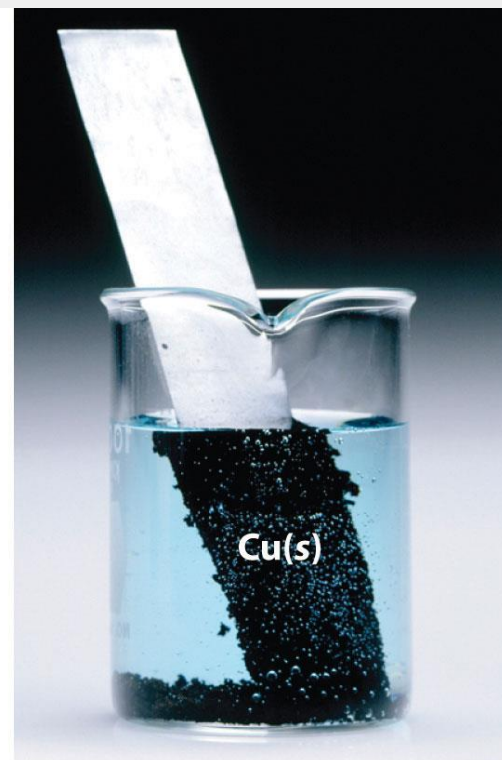
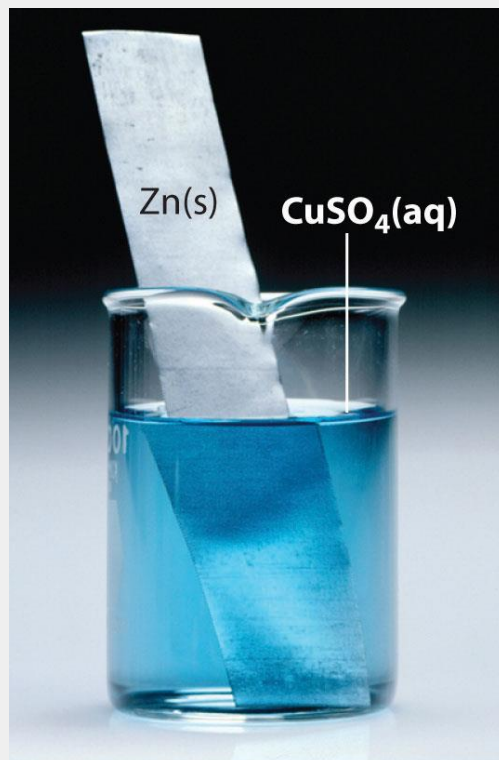
Background Knowledge



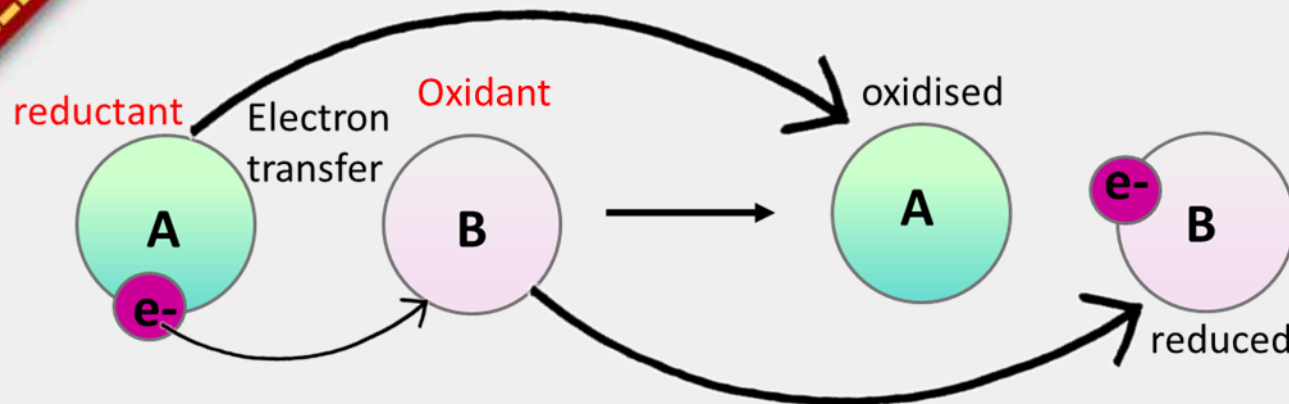
During electron transfer Redox reactions we often just write **ionic equations**.

For example the Cu^{2+} ions come from the CuSO_4 but only the Cu^{2+} is written into the equation. The SO_4^{2-} ions are **spectators** as they **play no part in the reaction**. They are also in solution and detached from the Cu^{2+} ions

Electron transfer



Summary of Terms



LEO (loss electrons oxidation) **A**

Reductant

Acts as a reducing agent to B

☐ is oxidised

☐ loses electrons

GER (gain electrons reduction) **B**

Oxidant

Acts as an oxidising agent to A

☐ is reduced

☐ gains electrons

Background Knowledge



Oxidation numbers can be used to predict whether a species – the reactant and its product – are undergoing oxidation or reduction.

The oxidation number is assigned to a **single atom only** and the corresponding atom in the product using a **set of rules**. If the oxidation number **increases** from reactant to product then **oxidation** has taken place. If the oxidation number **decreases** from reactant to product then **reduction** has taken place.

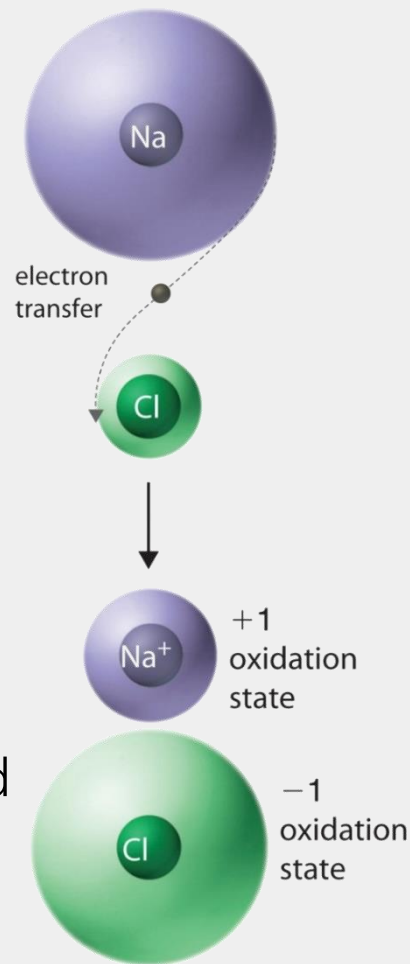
Oxidation Numbers

Reductant:

Loses e^-
Becomes oxidized

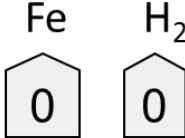
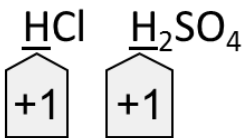
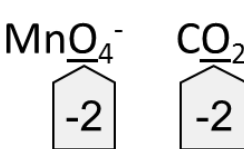


Oxidant:

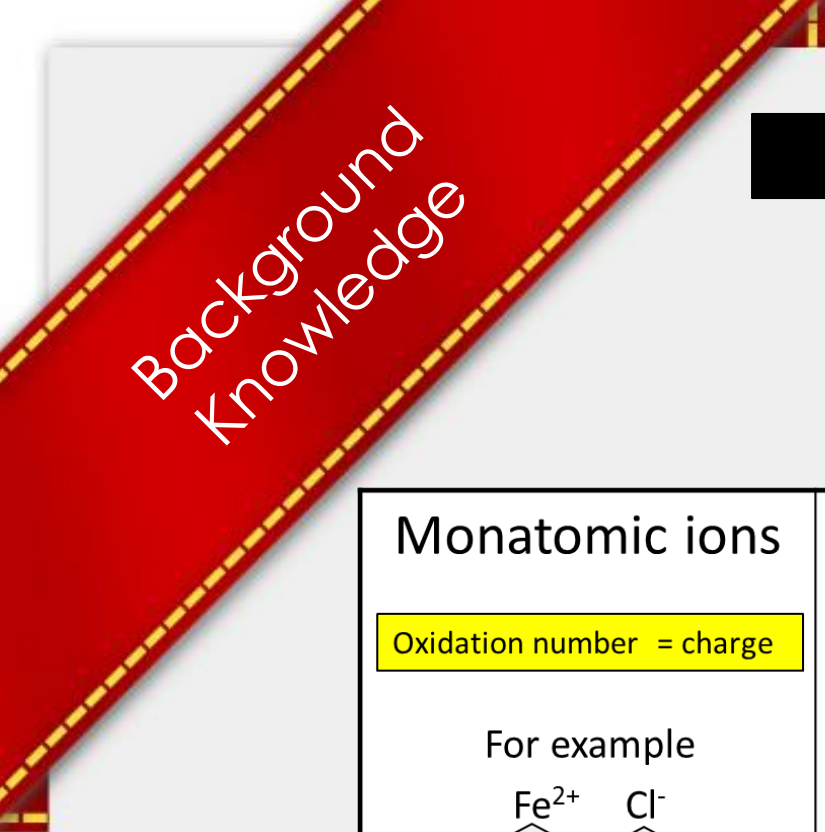
Gains e^-
Becomes reduced



Oxidation Numbers and Rules

The Oxidation Number (ON) gives the 'degree' of oxidation or reduction of an element.
They are assigned to a **INDIVIDUAL** atom using the following rules.

Elements	Hydrogen atom (not as element)	Oxygen atom (not as element)
Oxidation number = 0	Oxidation number = +1	Oxidation number = -2
For example Fe H ₂ 	For example HCl H ₂ SO ₄ 	For example MnO ₄ ⁻ CO ₂ 
	Except Hydrides Oxidation number = -1 For example LiH 	Except peroxides Oxidation number = -1 for example H ₂ O ₂ 

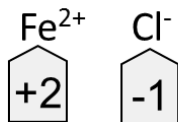


Oxidation Numbers and Rules

Monatomic ions

Oxidation number = charge

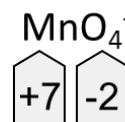
For example



Polyatomic ions

Sum of Oxidation number = charge

For example



Because

Total charge = -1

And

Oxygen = -2

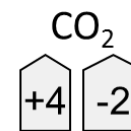
$$+7 + (4 \times -2) = -1$$



Molecules

Sum of Oxidation number = 0

For example



Because

Total charge = 0

And

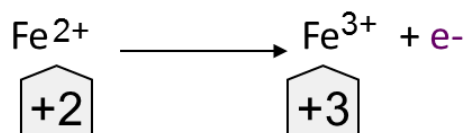
Oxygen = -2

$$+4 + (2 \times -2) = 0$$



Oxidation is a loss of electrons
and causes an **increase** in ON

Oxidation of Fe^{2+}



Fe has increased ON

(+2 to +3) caused by
a loss of electrons e^-

Reduction is a gain of electrons
and causes an **decrease** in ON

Reduction of MnO_4^-



Mn has decreased ON

(+7 to +2) caused by
a gain in electrons e^-

OXIDATION and **REDUCTION** always occur together. The electrons lost by one atom are gained by another atom.

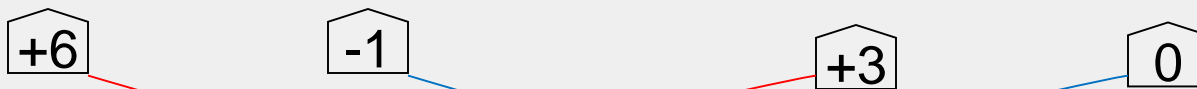
This is called a **REDOX** reaction.

Using Oxidation numbers to identify oxidants and reductants



What has been oxidised and what has been reduced?

STEP ONE – write the ON for each atom using rules (not oxygen or hydrogen)



Decrease - reduction

Increase - oxidation

STEP TWO – Identify the atom that has had its ON increased. It is **Oxidised**

I⁻ has increased **ON** (-1 to 0) so *I*⁻ is *Oxidised*. (the reductant)

STEP THREE – Identify the atom that has decreased ON. It is **reduced**.

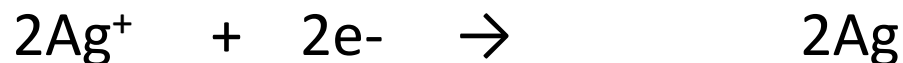
Cr has decreased **ON** (+6 to +3) so *Cr*₂*O*₇²⁻ is *Reduced*. (the oxidant)

Balancing half Redox equations

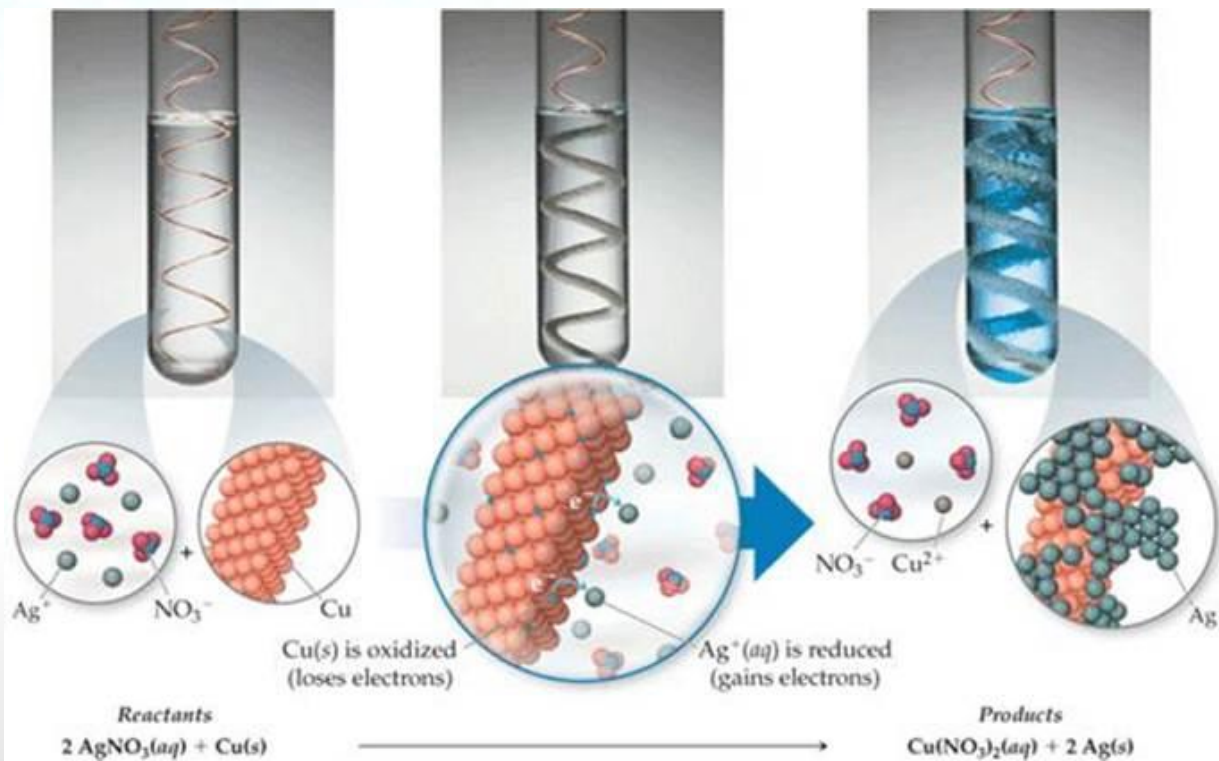
Oxidation half equation –
reductant is oxidised LEO



Reduction half equation –
oxidant is reduced GER



A balanced redox equation is broken into two half-equations, to show how electrons are transferred.



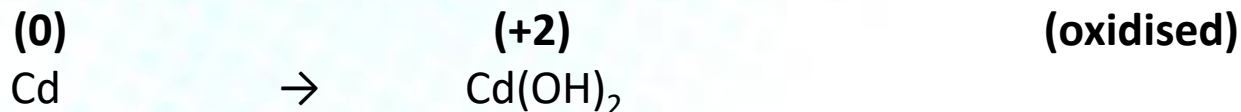
Balancing half Redox equations (part 1)

e.g. $\text{Cd} + \text{NiO}(\text{OH}) \rightarrow \text{Cd}(\text{OH})_2 + \text{Ni}(\text{OH})_2$ redox reaction in a discharging NiCad battery

1. Separate each half equation and balance each separately



2. Assign oxidation numbers and identify element oxidised or reduced.



3. Balance atom no. for element oxidised or reduced (other than oxygen and hydrogen)

4. Balance the Oxygen using H_2O



5. Add H^+ to balance the hydrogen



6. Add OH^- (in alkaline conditions) to cancel any H^+ [same amount on both sides] and cancel excess water



7. Balance charge by adding electrons (LHS on oxidants RHS on reductants)



8. Check balance of elements and charges

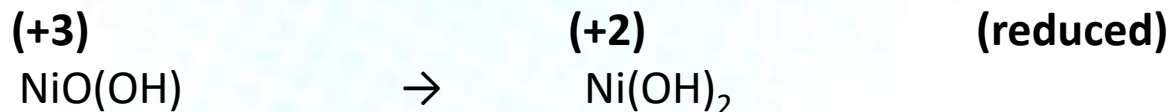
Balancing half Redox equations (part 2)

e.g. $\text{Cd} + \text{NiO(OH)} \rightarrow \text{Cd(OH)}_2 + \text{Ni(OH)}_2$ redox reaction in a discharging NiCad battery

1. Separate each half equation and balance each separately

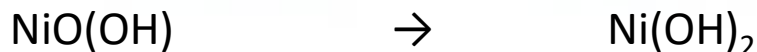


2. Assign oxidation numbers and identify element oxidised or reduced.



3. Balance atom no. for element oxidised or reduced (other than oxygen and hydrogen)

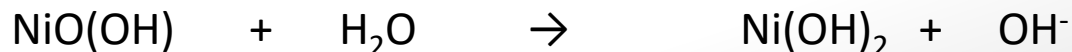
4. Balance the Oxygen using H_2O



5. Add H^+ to balance the hydrogen



6. Add OH^- (in alkaline conditions) to cancel any H^+ [same amount on both sides] and cancel excess water

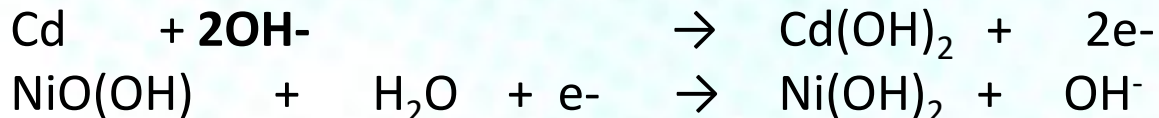


7. Balance charge by adding electrons (LHS on oxidants RHS on reductants)

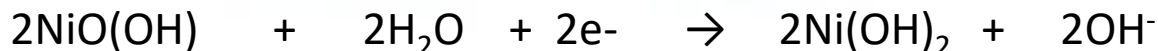


8. Check balance of elements and charges

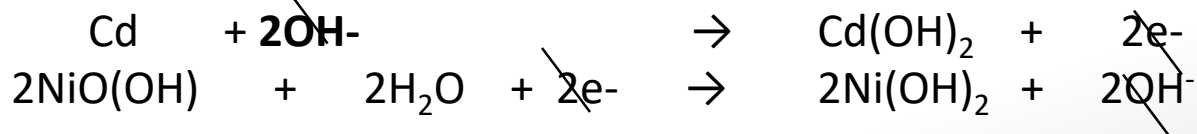
Joining half equations together



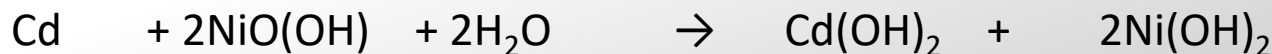
1. The two half equations must have electrons on opposite sides of the equation
2. Place the two equations one under the other
3. The electron numbers must equal each other – if not multiply one or both equations to the lowest common denominator (multiply every reactant/product)



4. Cancel out the electrons



5. Cancel out the same number of H^+ , OH^- and/or H_2O if present on both sides
6. Join the remainder together

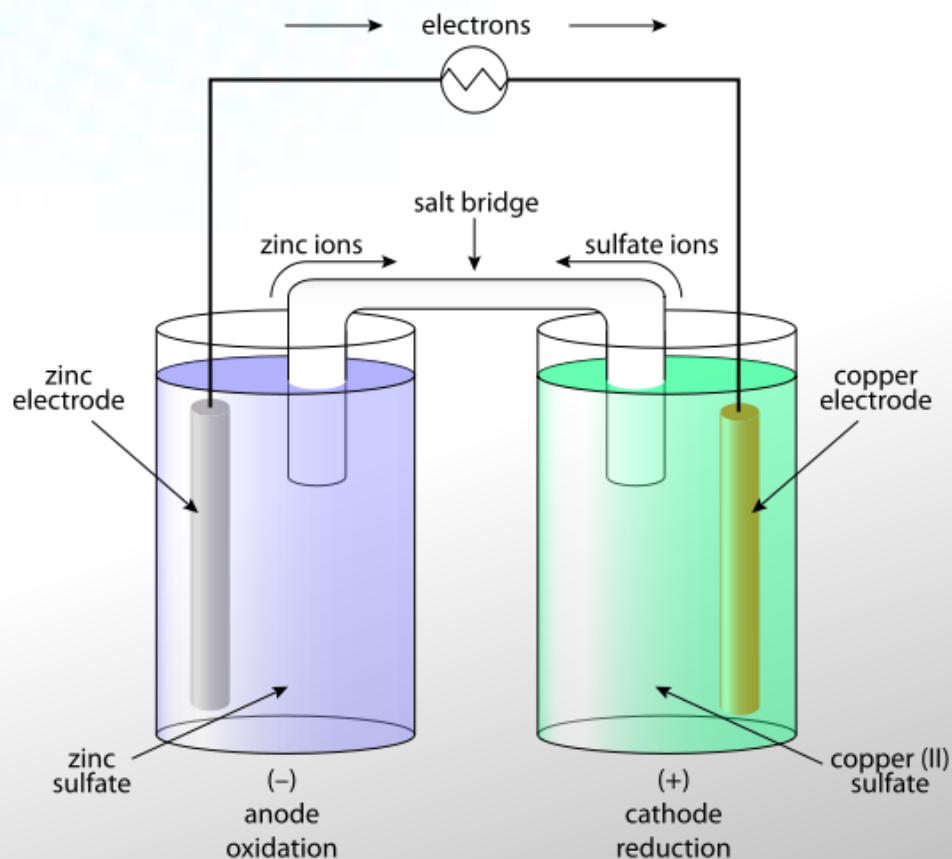


Electrochemical cells

Electrochemistry is the chemistry of reactions involving the transfer of electrons, which are redox reactions.

Spontaneous redox reactions occur in **Electrochemical cells**, which use the energy released from a chemical reaction to generate electric current. These are called Galvanic cells or batteries.

A voltmeter is connected to record voltage. A saltbridge filled with electrolyte (anion/cation solution) is used to complete a circuit so there is a flow of current.



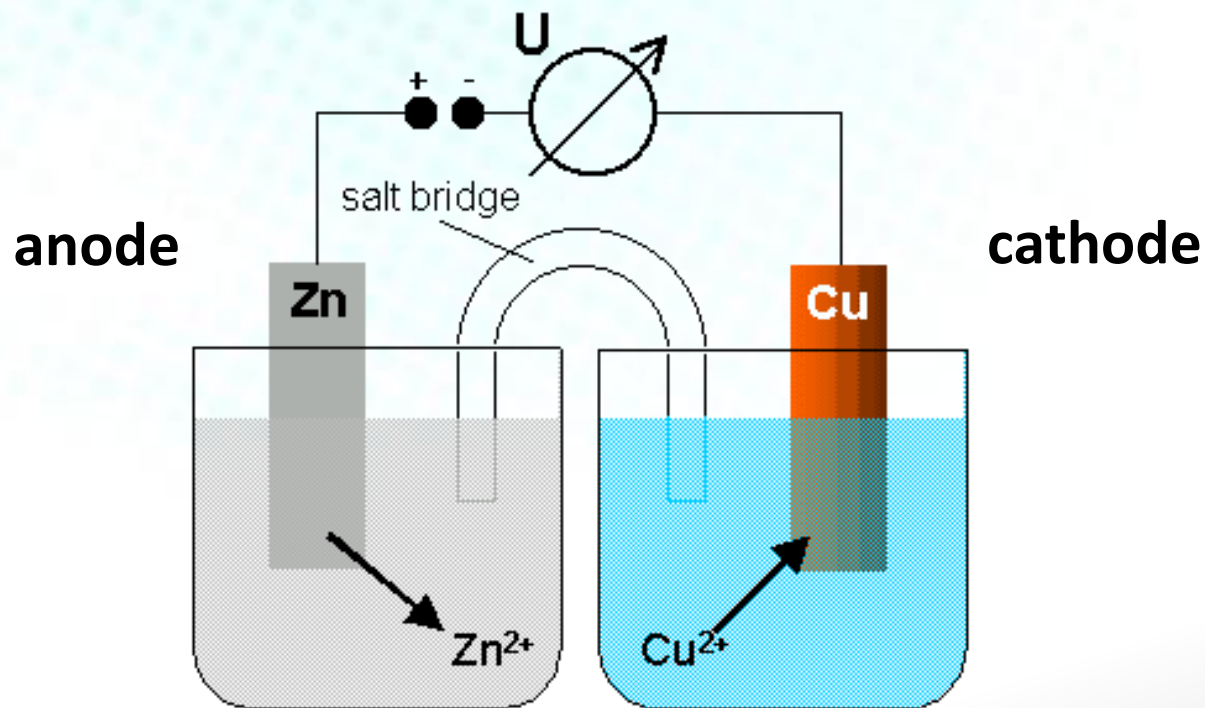
Galvanic Cells and Salt Bridges

Under normal conditions a redox reaction occurs spontaneously when an oxidising agent is in contact with a reducing agent. If the two half reactions are physically separated, the transfer of electrons is forced to take place through an **external metal wire**. As the reaction progresses a flow of electrons occurs.

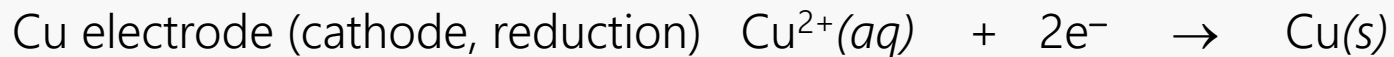
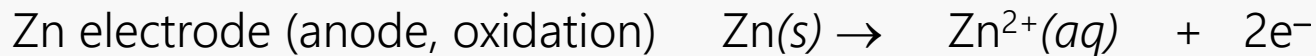


This only happens if there is a full circuit so that there is no net build-up of charge. To complete this circuit the separate solutions are connected using a **salt bridge** which allows ions to flow and transfer charge. Typically the salt bridge is a glass tube filled with a gel prepared using a strong electrolyte such as $\text{KNO}_3(\text{aq})$ (which contains ions that do not react with the electrodes or species in the solutions). The anions (NO_3^-) and cations (K^+) can move through the salt bridge so that charge does not build up in either cell as the redox reaction proceeds.

Galvanic Cells and Redox reactions

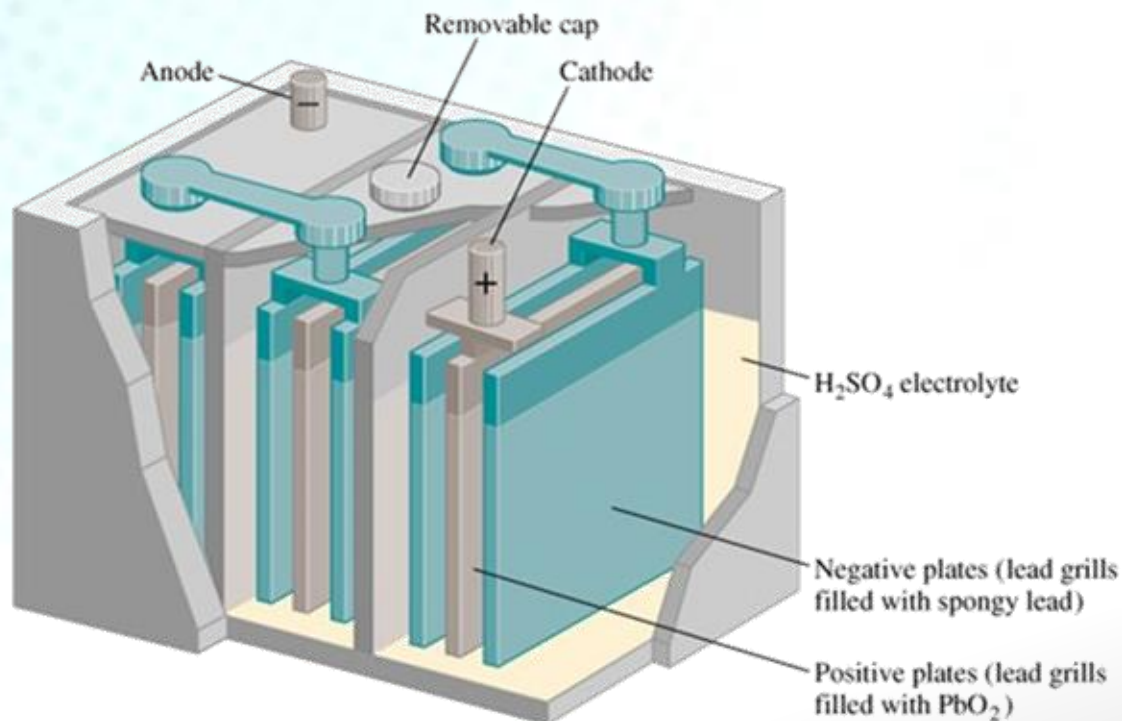


The oxidation and reduction reactions that occur at the electrodes are called **half-cell reactions**.



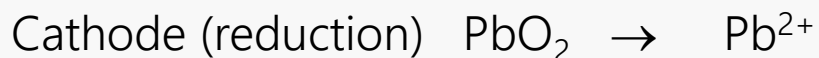
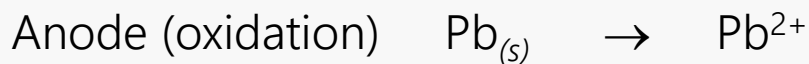
Galvanic Cells - Lead Acid battery example

This is the redox reaction that occurs when the battery is **discharging** – and the **energy produced** is used to power electrical systems (usually inside a vehicle)



The oxidation and reduction reactions that occur at the electrodes are called **half-cell reactions**.

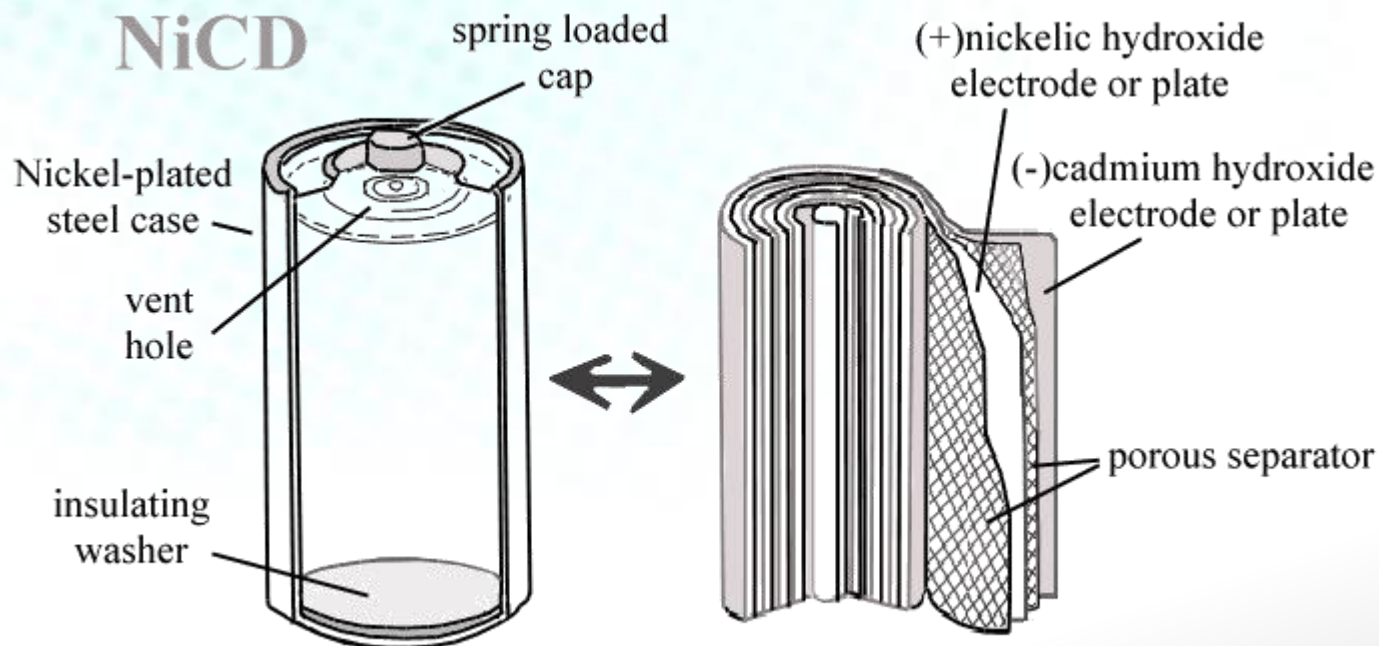
reductant



oxidant

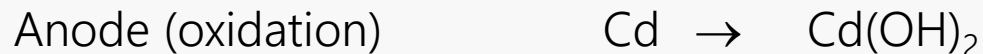
Galvanic Cells - NiCad Battery (nickel cadmium)

NiCad batteries are rechargeable batteries. The redox reaction shown is the spontaneous reaction when the battery is **discharging** and producing energy



The oxidation and reduction reactions that occur at the electrodes are called **half-cell reactions**.

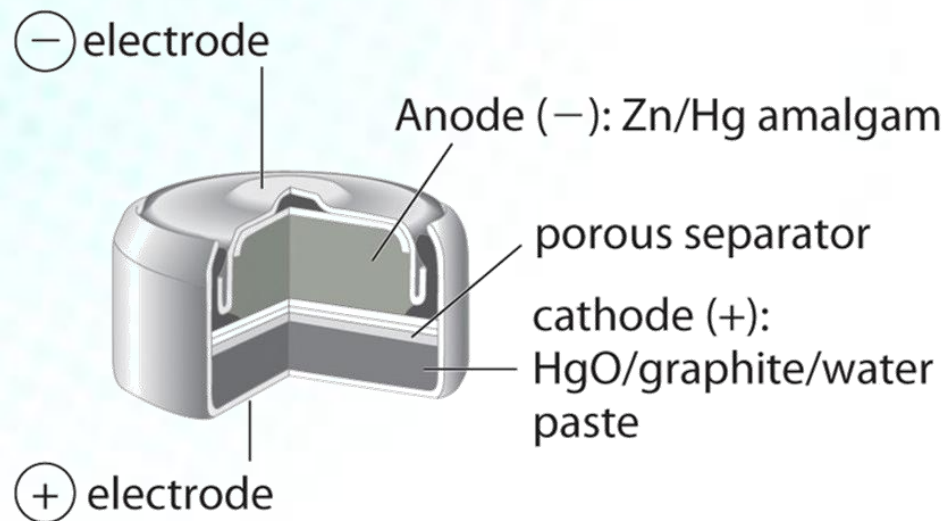
reductant



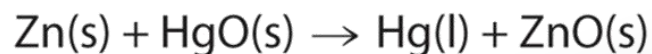
oxidant

Electrochemical Cells - Mercury Zinc Battery

This is the redox reaction that occurs when the battery is **discharging** – and the **energy produced** is used to power electrical systems (usually a small appliance or toy)

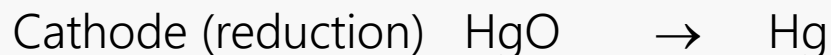
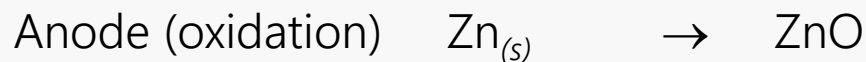


cell reaction:



The oxidation and reduction reactions that occur at the electrodes are called **half-cell reactions**.

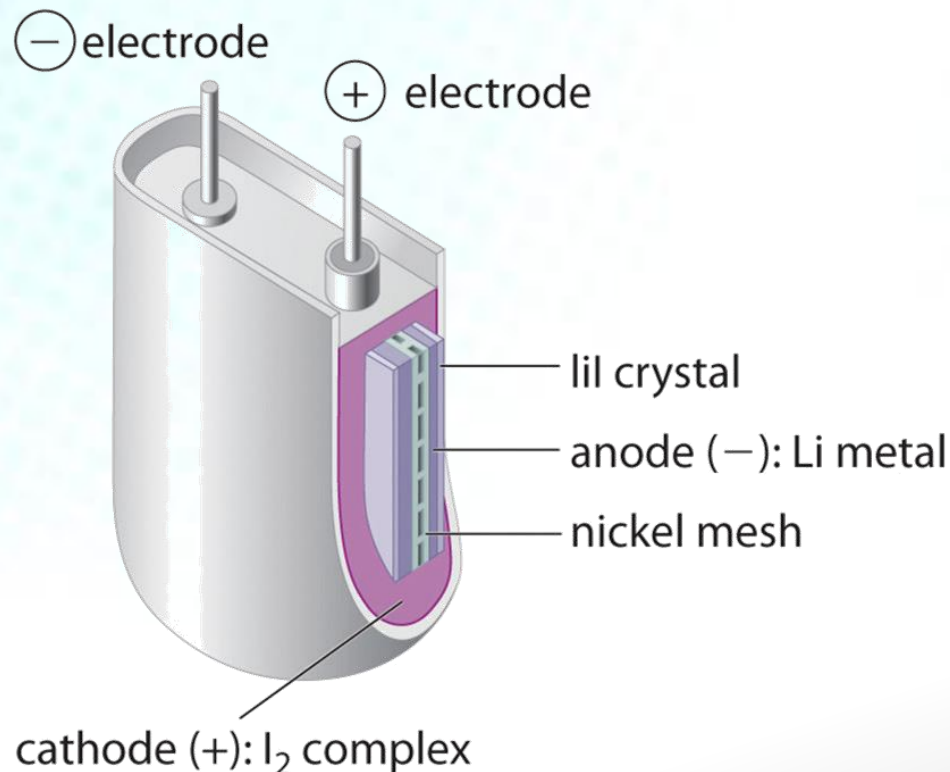
reductant



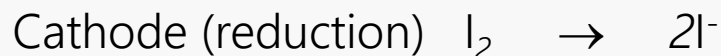
oxidant

Galvanic Cells - Lithium-iodine Battery

This is the redox reaction that occurs when the battery is **discharging** – and the **energy produced** is used to power electrical systems (used in pacemakers as they last up to 10 years and very reliable)



The oxidation and reduction reactions that occur at the electrodes are called **half-cell reactions**.



reductant

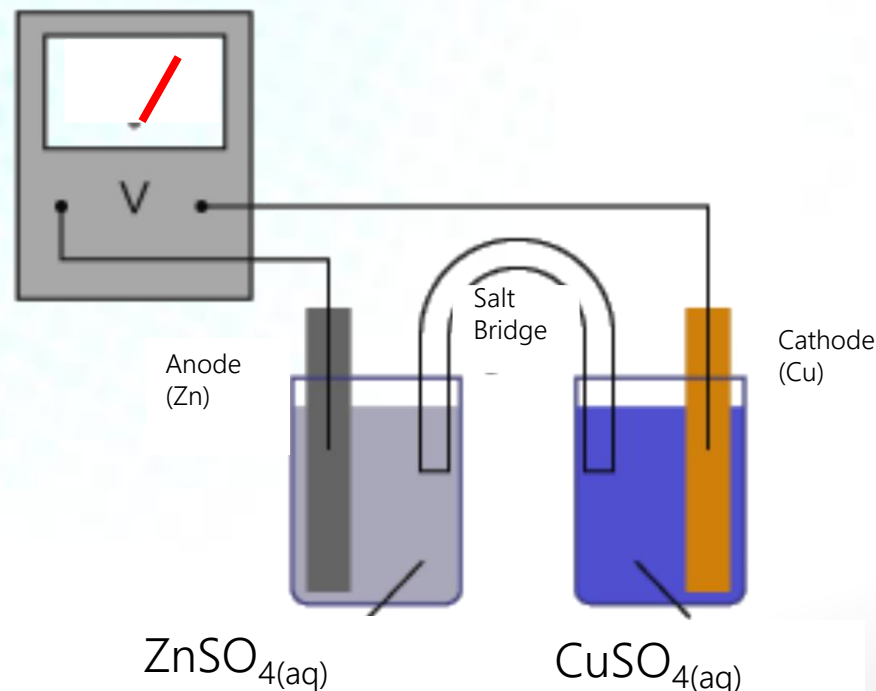
oxidant

Electromotive force

NOTE: Cell diagrams not required for assessment

The reduced and oxidised substances in each cell form a redox couple. The 2 couples in this cell (the Daniel cell) are $\text{Zn}^{2+}|\text{Zn}$ and $\text{Cu}^{2+}|\text{Cu}$. By convention, when writing redox couples, the oxidised form is always written first.

The fact that electrons flow from one electrode to the other indicates that there is a voltage difference between the two electrodes. This voltage difference is called the **electromotive force** or **emf** of the cell and can be measured by connecting a voltmeter between the two electrodes. The emf is therefore measured in volts and is referred to as the cell voltage or cell potential.



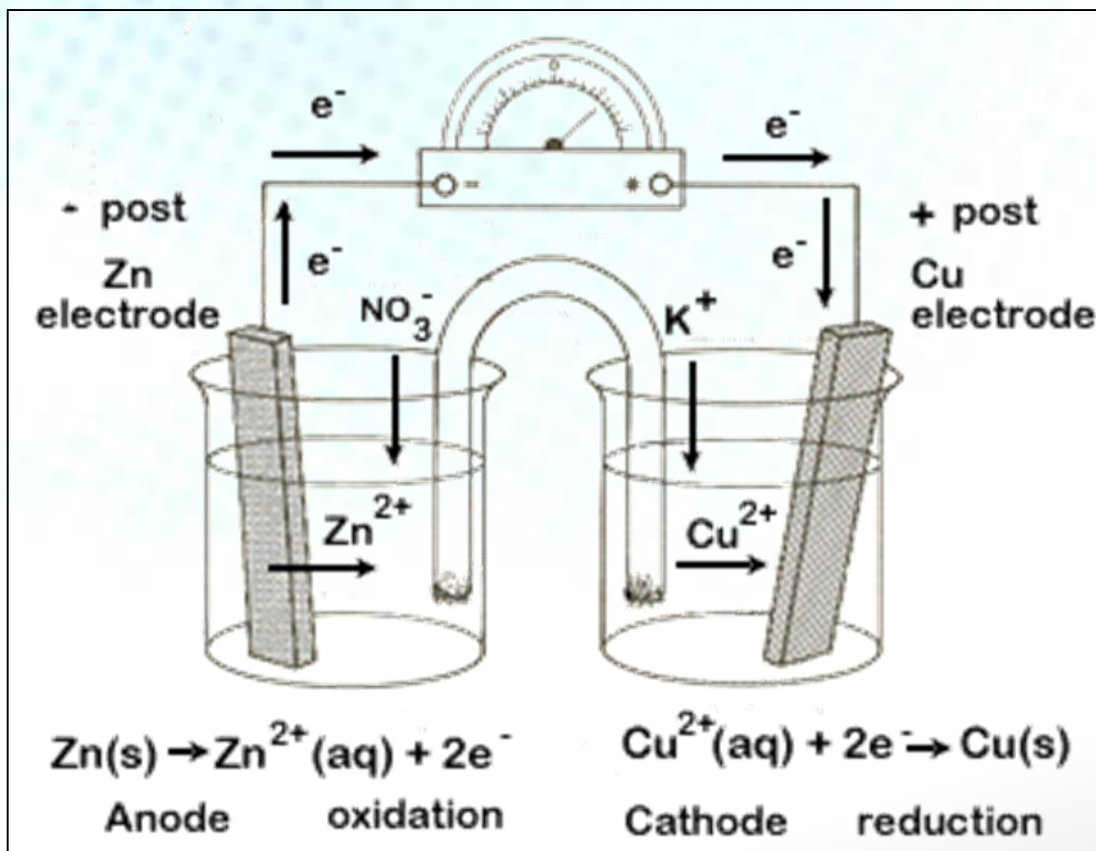
A high cell potential shows that the cell reaction has a high tendency to generate a current of electrons. Obviously the size of this voltage depends on the particular solutions and electrodes used, but it also depends on the concentration of ions and the temperature at which the cell operates.

Electrochemical cells Summary of terms

AAO

n n x
o i d
d o d
e n a
t i
o n

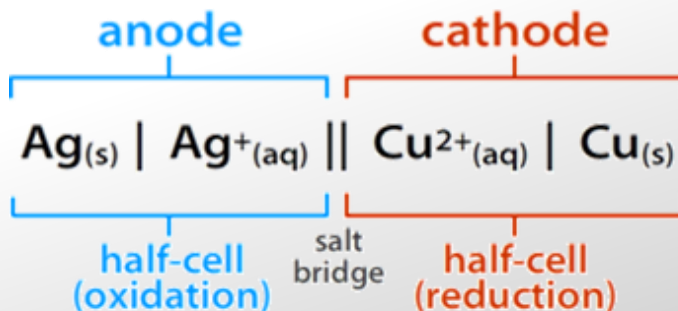
LEO



CCR

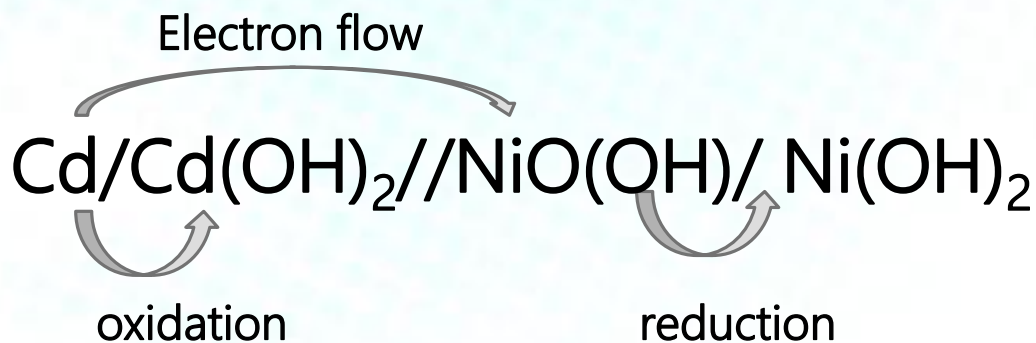
a a e
t t d
h i u
o o c
d n t
e i
o n

GER



NOTE: Cell diagrams not required for assessment

Galvanic cells can be represented using **cell diagrams**. This is a type of short hand notation that follows a standard IUPAC convention. For the NiCad battery cell the standard cell diagram is:



The vertical lines represent phase boundaries and $||$ represents the salt bridge.

The cathode (reduction reaction) is always shown on the right hand side and the anode (oxidation) on the left in a standard cell diagram.

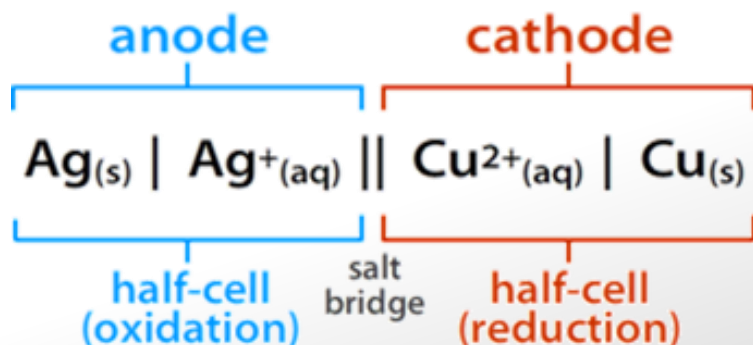
The electrons thus move from left of the $||$ to right in the standard cell diagram, representing a spontaneous redox reaction. The electrodes are always written in at the beginning and end of a cell diagram.

In each half cell the reactant appears first, followed by the product.

An **inert electrode** must be used in cells in which both species in a redox couple are in aqueous solution (MnO_4^- and Mn^{2+}). The inert electrodes are commonly either platinum, Pt(s) or graphite, C(s) electrodes. Since the two species in the redox couple are in solution, they are separated by a comma rather than a vertical line.



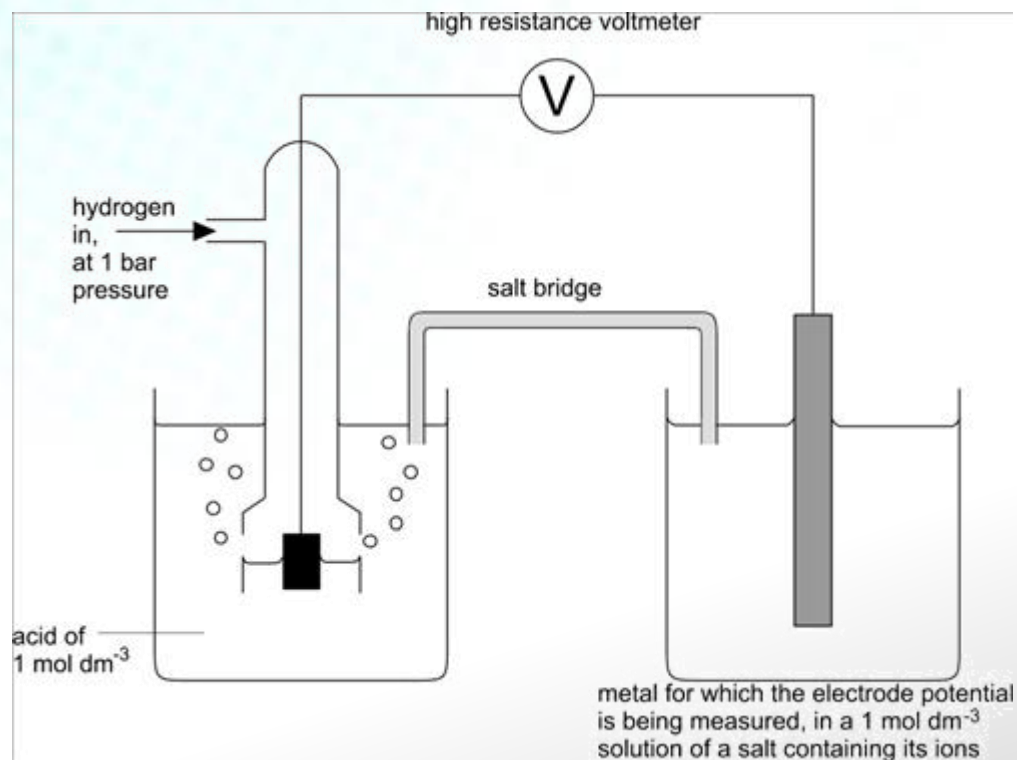
The cell diagram shows two half cells linked. Each half cell consists of the oxidant, the reductant and the electrode (which may be the oxidant or reductant). The two half cells above are $\text{Pb(s)}|\text{Pb}^{2+}(\text{aq})$ and $\text{PbO}_2(\text{s})|\text{Pb}^{2+}(\text{aq})|\text{PbO}_2(\text{s})$.



If one of the reactants is a suitable electrode, such as copper or zinc, then that will be the outside substance

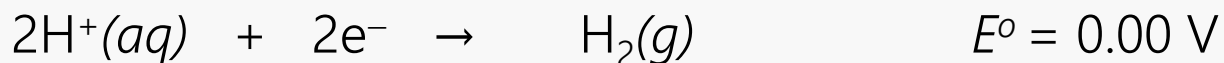
Standard electrode potential

The overall cell voltage is the sum of the electric potential at each electrode. If one of the electrode potentials is known, and the overall cell voltage is measured, then the potential of the other electrode can be calculated by subtraction. Clearly it is best if all electrode potentials are measured relative to a particular electrode. In this way, a scale of relative values can be established. The **standard hydrogen electrode** (SHE) is used as the standard reference electrode, and it has arbitrarily been given a value of **0.00 V**.



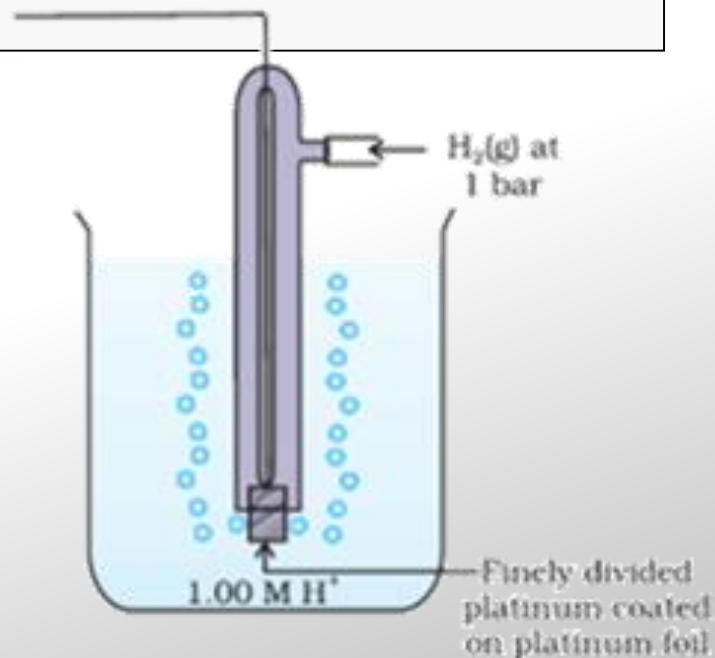
Standard conditions

Under **standard conditions** (when the pressure of hydrogen gas is 1 atm, and the concentration of acid is 1 mol L⁻¹) the potential for this standard Hydrogen electrode reduction reaction is assigned a value of **zero**.



The superscript ° denotes standard state conditions. When the hydrogen electrode acts as a cathode, H⁺ ions are reduced, whereas when it acts as an anode, H₂ gas is oxidised.

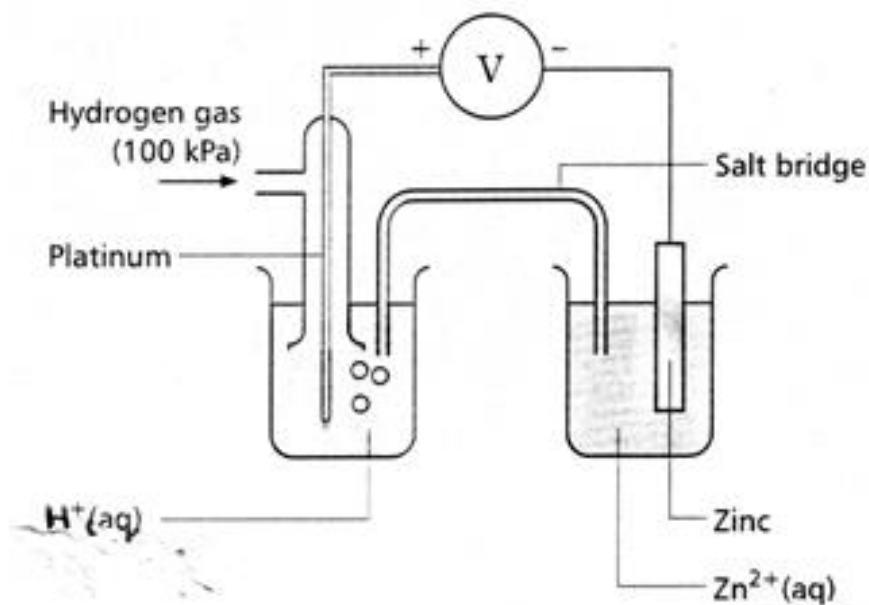
In order to measure the potential of any other redox couple they are measured against this standard hydrogen electrode (SHE)



Standard electrode (reduction) potential

For any redox couple, the standard electrode (reduction) potential is the voltage obtained under standard conditions when that half-cell is connected to the standard hydrogen electrode (value 0v).

For example, the electrode potential of a $\text{Zn}^{2+}|\text{Zn}$ electrode can be measured by connecting it to a hydrogen electrode.



NOTE: Cell diagrams not required for assessment

Experimentally, the more positive terminal is always where reduction is occurring in a spontaneous reaction. In example (a) reduction occurs in the hydrogen electrode (positive electrode) while oxidation occurs in the $\text{Zn}^{2+}|\text{Zn}$ compartment (negative electrode). The cell diagram for this electrochemical cell is

Flow of electrons



Flow of electrons



Standard reduction potential

Using the standard reduction potentials for many half reactions have been measured under standard conditions (at 25 °C). **Standard reduction potentials are provided in the Internal Assessment.**

The table can be used to decide the relative strength of species as oxidants or reductants. The species on the left in the couple with the most positive reduction potential, will be the strongest oxidising agent or oxidant. E.g it is $F_2(g)$ (**NOT** F_2 / F^-). This means F_2 has the greatest tendency to gain electrons. As the electrode potential decreases, the strength as an oxidant decreases.

Conversely the strongest reducing agent or reductant would have the least positive (or most negative) e.g. $Li(s)$. This means Li has the greatest tendency to lose electrons.

Electrode	Electrode reactions	E° volts
$Li^+ Li$ ↑	$Li^+(aq) + e^- \longrightarrow Li(s)$	-3.05
$K^+ K$	$K^+(aq) + e^- \longrightarrow K(s)$	-2.93
$Mg^{2+} Mg$	$Mg^{2+}(aq) + 2e^- \longrightarrow Mg(s)$	-2.37

More positive the standard reduction Potential the more likely to **Gain electrons** (be reduced)

Common Redox couples (in batteries)

	Redox couple	Standard reduction potential (V)
1	$\text{PbO}_2/\text{Pb}^{2+}$	1.69
2	$\text{MnO}_2/\text{Mn}^{3+}$	0.74
3	$\text{NiO(OH)}/\text{Ni(OH)}_2$	0.48
4	HgO/Hg	0.098
5	I_2/I^-	0.54
6	Pb^{2+}/Pb	-0.36
7	Zn^{2+}/Zn	-0.76
8	$\text{Cd(OH)}_2/\text{Cd}$	-0.82
9	Li^+/Li	-3.10

When 2 couples are placed together, because they are all shown as reduction reactions, the lower value couple will be reversed into an oxidation reaction (the charge will stay the same on the SRP)

All of these couples show reduction from left to right. i.e redox couple 1. PbO_2 is reduced to Pb^{2+} . If redox couple 6. was placed with 1. then it would have a lower reduction potential and therefore be reduced. Pb is therefore oxidised to Pb^{2+} (the order of the couple is reversed)

Using reduction potentials to determine E°_{cell}

In any electrochemical cell, the standard cell potential (voltage), E°_{cell} , is the difference between the reduction potentials of the two redox couples involved. The couple with the **more positive reduction potential** will be the **reduction half-cell (cathode)**. This means that the E°_{cell} for any combination of electrodes can be predicted using the relationship

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{(reduction half-cell)}} - E^{\circ}_{\text{(oxidation half-cell)}}$$

OR
$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{(cathode)}} - E^{\circ}_{\text{(anode)}}$$

OR
$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{(RHE)}} - E^{\circ}_{\text{(LHE)}}$$

Do not forget the
units are V (volts)

(where RHE is the right hand electrode and LHE is the left hand electrode in the **standard** cell diagram).

Predicting whether a reaction will occur

It is possible to use E° values to predict whether a reaction will occur. This simply involves identifying which species must be reduced and which species must be oxidised if the reaction is to proceed spontaneously. The appropriate reduction potentials are then substituted into the equation.

$$E^\circ_{\text{cell}} = E^\circ_{(\text{cathode/red})} - E^\circ_{(\text{anode/ox})}$$

where $E^\circ_{(\text{cathode})}$ is the reduction potential for the half cell where reduction occurs and $E^\circ_{(\text{anode})}$ is the reduction potential for the half cell where oxidation occurs. If the E°_{cell} calculated is positive, then the reaction will occur spontaneously. Conversely, a negative cell potential means the reaction will not proceed.

This E°_{cell} is positive therefore this redox reaction will occur spontaneously

Consider the lead acid battery cell $\text{Pb(s)} \mid \text{Pb}^{2+}(\text{aq}) \parallel \text{PbO}_2, \text{Pb}^{2+} \mid \text{PbO}_2(\text{s})$

Reduction reaction is $\text{PbO}_2 + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{Pb}^{2+} + 2\text{H}_2\text{O}$ $E^\circ(\text{PbO}_2/\text{Pb}^{2+}) = +1.69\text{V}$

Oxidation reaction is $\text{Pb}_{(\text{s})} \rightarrow \text{Pb}^{2+} + 2\text{e}^-$ $E^\circ(\text{Pb}^{2+}/\text{Pb}) = -0.36\text{V}$

$$E^\circ_{\text{cell}} = E^\circ(\text{PbO}_2/\text{Pb}^{2+}) - E^\circ(\text{Pb}^{2+}/\text{Pb}) = +1.69 - (-0.36) \text{ V} = \mathbf{+2.05\text{V}}$$

electrode

The acid in the battery is concentrated and there are 6 sets of cells so the battery normally produces 12V

Charging Batteries - non-spontaneous Redox reactions

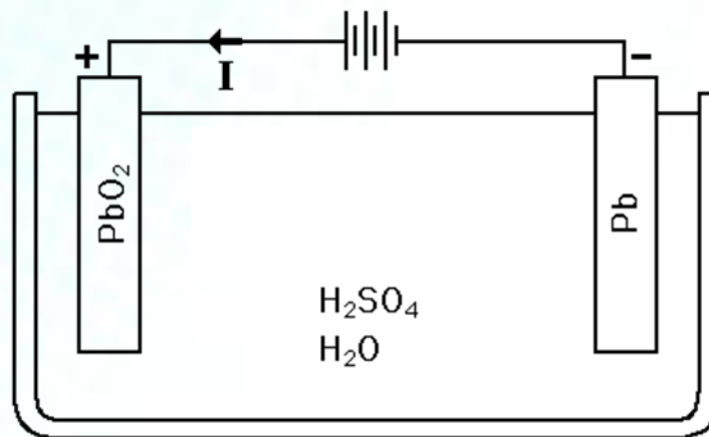
Eventually if the discharging of a battery continues (while supplying energy to the vehicle or appliance) the **reactants will "run out"** as they are changed into products during the redox reaction.

Some types of batteries can be **charged** – this involved supplying an **external source of energy** to power a **reverse** of the discharging reaction. The built up products will then be changed back into the original reactants to enable the battery to be discharged once more.

An electrochemical cell that undergoes a redox reaction when electrical energy is applied is called an ***electrolytic cell***

The discharging oxidation reaction will become a reduction reaction during charging

The discharging reduction reaction will become an oxidation reaction during charging



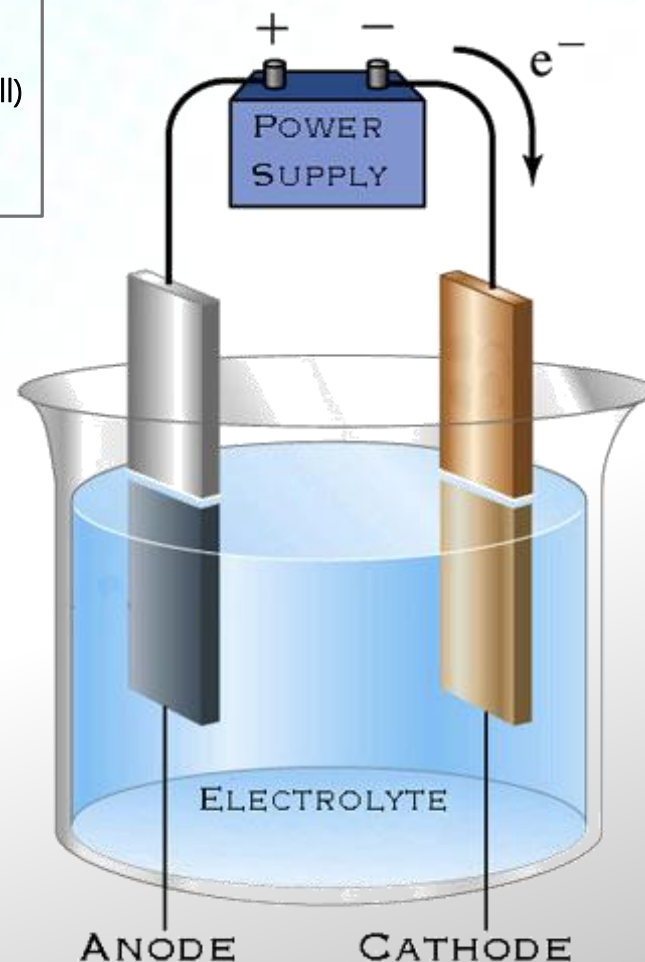
With energy from the charging battery, the lead sulfate is broken down and with oxygen from ionized water, lead oxide is deposited on the positive electrode and lead is deposited on the negative electrode

E°_{cell} in Charging Batteries - non-spontaneous Redox reactions

$$\text{Charged } E^{\circ}_{\text{cell}} = E^{\circ}_{(\text{reduction half-cell})} - E^{\circ}_{(\text{oxidation half-cell})}$$

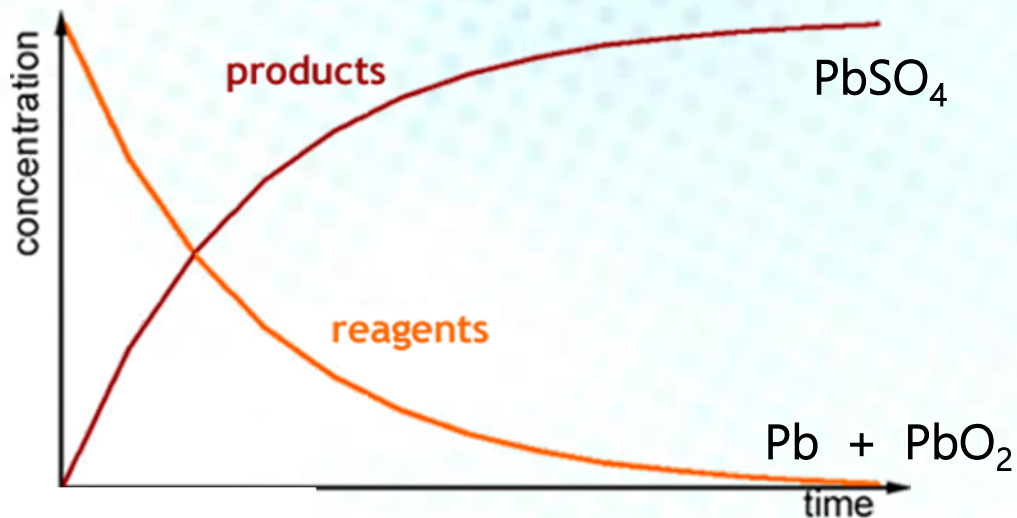
$$= \text{lowest reduction potential} - \text{highest reduction potential}$$

The E°_{cell} for the charging battery "swaps around" the reduction potentials to give a **negative value** – which indicates the redox reaction is not spontaneous



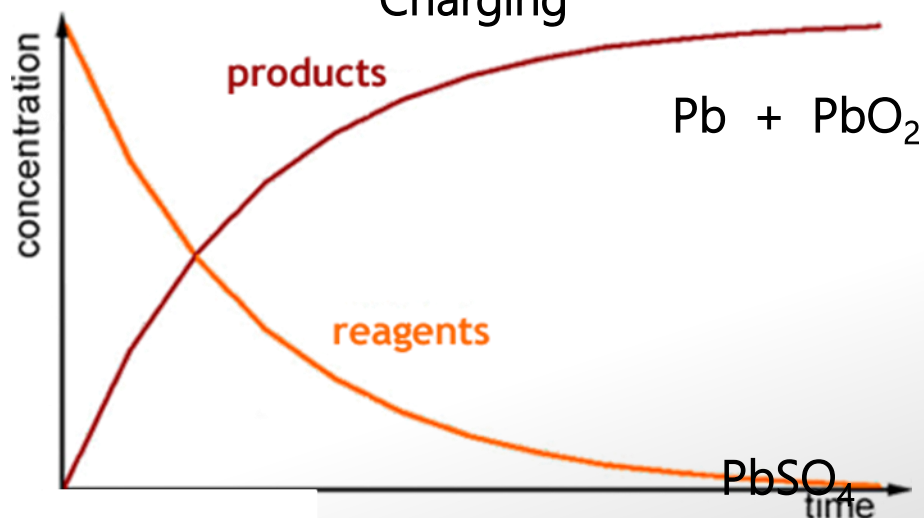
Reactants and Products during charging and discharging

Discharging



During **discharge** of a battery the amount of reactants (both the oxidant and reductant) will be decreased and the products formed increased. In the case of the lead-acid battery the Pb and PbO_2 will be decreased (the anode and cathode respectively) and the solid PbSO_4 will increase.

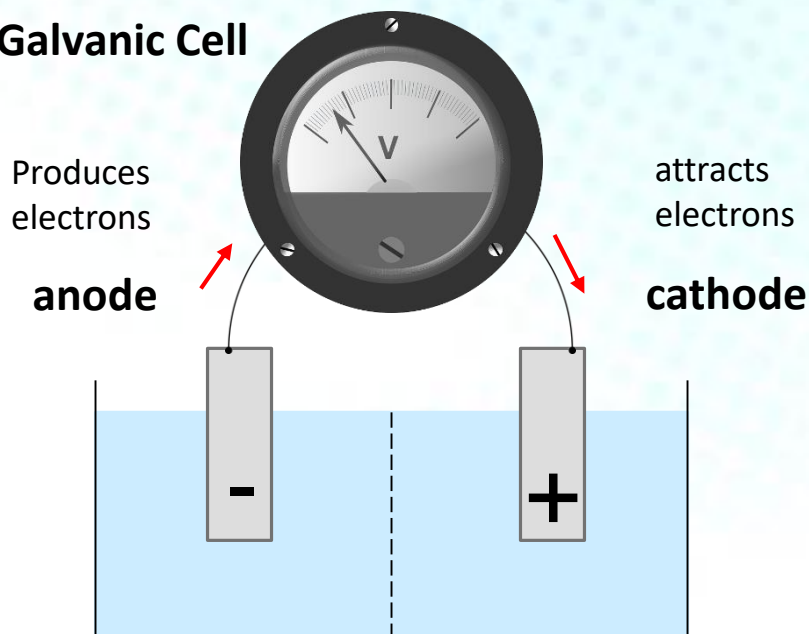
Charging



During **charging** of a battery the products from the discharging are now the reactants. In the case of the lead-acid battery the amount of PbSO_4 will be decreased and deposited back on the anode and cathode as Pb and PbO_2 respectively

Summary of charging and discharging a battery

Galvanic Cell



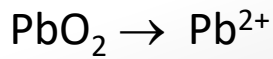
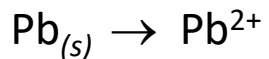
Oxidation

reductant

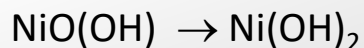
Reduction

oxidant

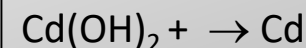
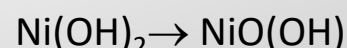
Discharging Battery where energy is released by spontaneous redox reaction and converted to electrical energy



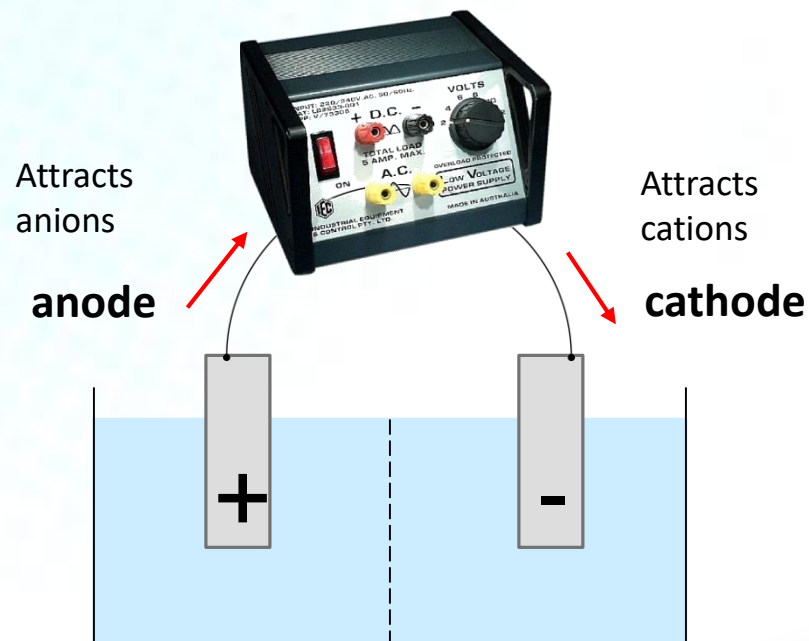
Lead
acid
battery



NiCad
battery



Electrolytic Cell



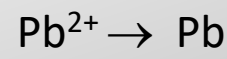
Oxidation

reductant

Reduction

oxidant

Charging Battery where energy is used to drive non-spontaneous redox reaction



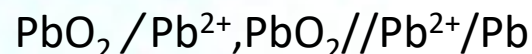
Summary of charging and discharging a battery

Discharging Battery where energy is released by spontaneous redox reaction and converted to electrical energy



$$E^\circ_{\text{cell}} = E^\circ_{(\text{red})} - E^\circ_{(\text{ox})}$$

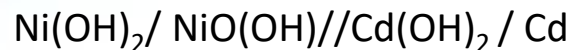
Charging Battery where energy is used to drive non-spontaneous redox reaction



$$E^\circ_{\text{cell}} = E^\circ_{(\text{red})} - E^\circ_{(\text{ox})}$$



$$E^\circ_{\text{cell}} = E^\circ_{(\text{red})} - E^\circ_{(\text{ox})}$$



$$E^\circ_{\text{cell}} = E^\circ_{(\text{red})} - E^\circ_{(\text{ox})}$$



NOTE: Cell diagrams not required for assessment