

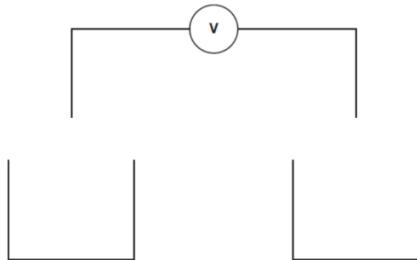
DM.a-n - Developing Metals | DM1-6 |

DM.Q Exam questions from past papers

Complete the diagram below that shows how the E_{cell}° value of an electrochemical cell using the following two half-reactions is measured under standard conditions.

half-reaction	E°/V
$\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$	+0.34
$\text{Fe}^{3+} + e^- \rightarrow \text{Fe}^{2+}$	+0.77

Label the electrodes and solutions used and state the conditions required.



Calculate the E_{cell}° value for the electrochemical cell in (i).

$$E_{\text{cell}}^{\circ} = \dots\dots\dots\dots\dots V [1]$$

Complete the diagram below that shows the E_{cell} value (5)?

Calculate the E_{cell} value (1) ?

Pt electrode for $\text{Fe}^{3+} / \text{Fe}^{2+}$ half cell ✓
in $\text{Fe}^{3+} / \text{Fe}^{2+}$ ✓
Cu electrode in Cu^{2+} ✓
salt bridge labelled and in solutions ✓
conditions: 1 mol dm⁻³ and 298K / 25C ✓

5

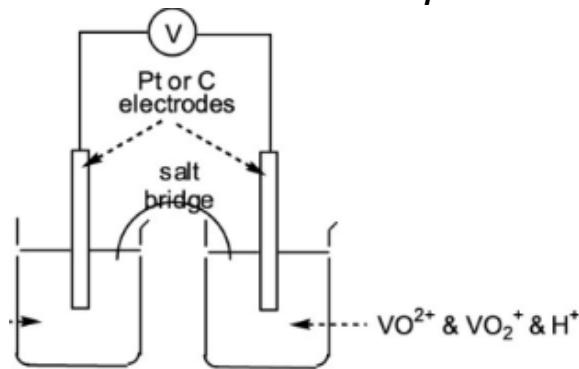
ALLOW CuSO_4 etc. instead of Cu^{2+}
ALLOW 1 molar / 1M

0.43 V ✓

| 1 | IGNORE any sign

- Pt electrode for $\text{Fe}^{3+} / \text{Fe}^{2+}$ half cell (1)
- In $\text{Fe}^{3+} / \text{Fe}^{2+}$ (1)
- Cu electrode in Cu^{2+} (1)
- Salt bridge labelled and in solutions (1)
- Conditions: 1mol dm⁻³ and 298K / 25C (1)
- 0.43V (1)

If a half cell was $\text{VO}_2^+ + 2\text{H}^+ + e^- \rightarrow \text{VO}^{2+} + \text{H}_2\text{O}$ then a label for H^+ would also be required



(iii) Suggest why rusting takes place faster in seawater than in rainwater.

..... [1]

Suggest why rusting takes place faster in seawater than in rainwater?

(iii) (Faster in salt water as) more (dissolved) ions (make it a better conductor) ✓

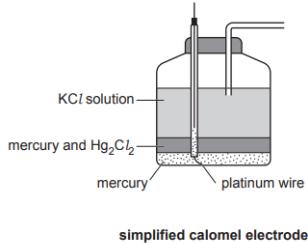
1

ALLOW (water acts as a) 'salt bridge' and
sea water contains a higher concentration
of ions
ALLOW (the salt) acts as an electrolyte

- Faster in sea water as more dissolved ions makes it a better conductor

(c)* A student investigating the rusting of iron is given a calomel electrode.

The diagram shows a simplified calomel electrode.



The student wants to use this electrode to measure the standard electrode potential of a $\text{Fe}^{2+}(\text{aq})|\text{Fe}(\text{s})$ half-cell.

Give instructions on how to do this, justifying the uses of the pieces of apparatus you name. You may add to the diagram above to illustrate your answer.

The student wants to use this electrode to measure the standard electrode potential of a Fe/Fe^{2+} half cell

Give instructions on how to do this justifying the uses of the pieces of apparatus you name.

[You may add to the diagram to illustrate ur answer]

Indicative scientific points may include:

Set up

- make sure reading on voltmeter is positive
- electrode/half-cell connected to positive terminal of voltmeter is the positive electrode
- discussion of how E_{cell} can be used to find electrode potential of iron half cell

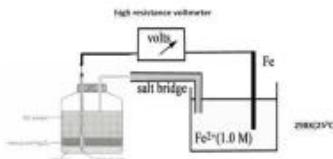
Description of the apparatus used and why

- high resistance voltmeter so negligible current is taken (so concs of ions stay the same)
- salt bridge to keep the charge in each beaker constant.
- both the above correctly connected to both cells
- iron electrode and $\text{Fe}^{2+}(\text{aq})$ solution in beaker so the reaction $\text{Fe}(\text{s}) \rightleftharpoons \text{Fe}^{2+}(\text{aq})$ can take place

Conditions/concentrations

- 298K/25°C
- 1.00 mol dm⁻³ $\text{Fe}^{2+}(\text{aq})$ solution as E° is changed with higher temp and conc

Note: some of the above points may be scored from a suitably labelled diagram, such as the one below



Set up:

- Make sure reading on voltmeter is positive
- Electrode/half cell connected to positive terminal of voltmeter is the positive electrode
- From knowing

Description of the apparatus used and why:

- High resistance voltmeter so negligible current is taken (so conc's of ions stay the same)
- Salt bridge to keep the charge in each half cell constant
- Draw complete electrochemical cell ;
- $\text{Fe}(\text{s})$ electrode

Conditions/Conc

- 298K/25C
- 1mol dm⁻³ in Fe/Fe²⁺

DM.a - Formulae, equations & amount of substance | DM1 |

What acid should be used for manganate redox titrations? Why?

Excess dilute sulfuric acid as other acids can set up alternative redox reactions (leading to inaccurate titres) as seen below

Table 34.1

redox half-cell	half-equation	E°/V
1	$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	-0.76
2	$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	-0.44
3	$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	-0.14
4	$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
5	$\frac{1}{2}O_2(g) + H_2O(l) + 2e^- \rightleftharpoons 2OH^-(aq)$	+0.40
6	$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
7	$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
8	$Cl_2(aq) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
9	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51

(d) A student carries out the analysis described in part (c) but uses hydrochloric acid instead of sulfuric acid to dissolve the sample of the alloy.

Using the data in Table 34.1, predict and explain the effect of this change of acid on the titre of potassium manganate(VII).

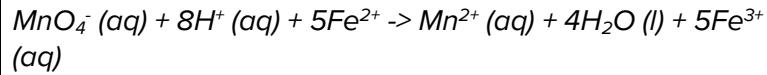
.....
.....
.....

[2]

The Cl_2/Cl (redox half-cell) electrode potential/ E° value is more negative than that for MnO_4^-/Mn^{2+} ✓	2	3.2 (x2)	ALLOW any clear indication of half-cell e.g. half cell 8 & 9
Cl^- are oxidised to Cl_2 (as well as Fe^{2+} to Fe^{3+}) so the titre would be larger ✓			ORA throughout

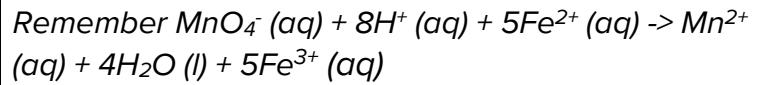
Why is an indicator not required for a manganate sulfuric acid redox titration and its colour change?

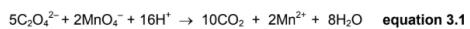
As manganate (MnO_4^-) is pink (the only coloured reagent) and so when it's reduced it produces Mn^{2+} which is colourless and the endpoint is when the pink reappears



Why must EXCESS sulfuric acid be used in a manganate sulfuric acid redox titration

Otherwise, the solution won't be acidic enough and MnO_2 will be produced instead of Mn^{2+}





The reaction above can be used to find the concentration of an ethanedioate solution.

The ethanedioate solution is titrated with a standard potassium manganate(VII) solution in acid conditions.

The reaction is slow at room temperature and needs to be heated above 60°C.

- i)* Describe how the titration would be carried out and explain how the end-point of the reaction can be recognised without an indicator.

indicative scientific points may include:

titration

- manganate(VII) (solution) in burette
- pipette known/stated volume of ethanedioate
- OR use graduated/volumetric pipette for ethanedioate.
- add sulfuric acid
- some indication that sulfuric acid is in excess
- warm/heat solution/ 60°C
- titrate (AW) until pink colour persists/ remains (AW)
- add (manganate(VII)) dropwise (AW) near end
- repeat for concordance (AW)

indicator

- MnO_4^{2-} /manganate(VII) is only coloured reagent
- colour change takes place during reaction from colourless to pale pink

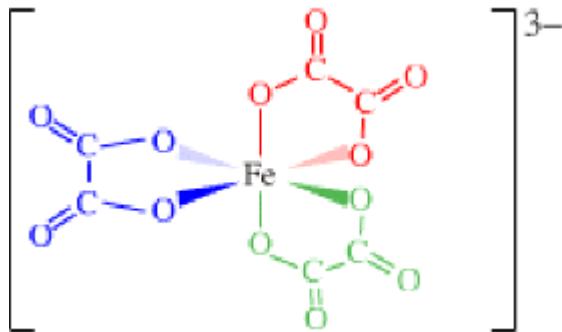
[6]

DM.b - Bonding & Structure | DM6 |

Define coordination number?

The number of coordinate bonds formed to a central metal ions.

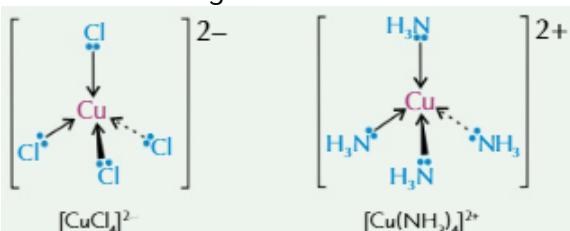
This won't always equal the number of ligands like here:



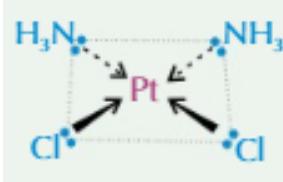
Coordinate bonds are the same thing as a dative covalent bond

What are the shape and bond angles of the complexes with the different coordinate numbers?

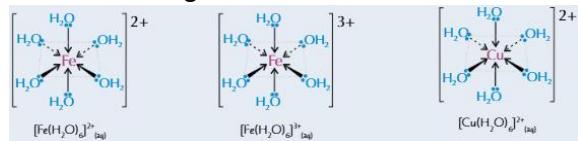
- Coordinate number 2 - Linear structure - 120 degrees
- Coordinate number 4 - Tetrahedral structure - 109.5 degrees



- Coordinate number 4 - Square planar structure - 90 degrees



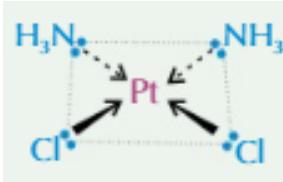
- Coordinate number 6 - Octahedral structure - 90 degrees



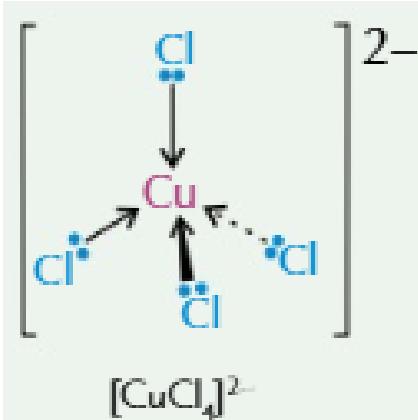
A question would state or imply that a structure with coordinate number 4 is either tetrahedral or square planar or both would be an acceptable answer

Complexes with other coordinate numbers not stated above are not required for this spec this includes 8

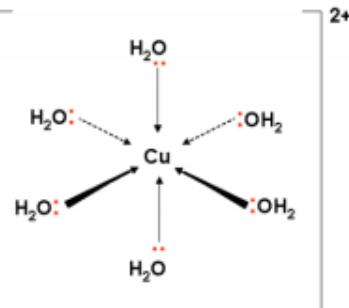
Draw the shape of the complex of Cisplatin $[\text{Pt}(\text{NH}_3)_2(\text{Cl})_2]$ as a square planar structure?



Draw the shape of the complex of $[\text{CuCl}_4]^{2-}$ as a tetrahedral structure?



Draw the shape of the complex of $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ as an octahedral structure? (3)



Six water molecules around the central copper. ✓

Water molecules oxygen atoms donating lone pairs
AND OR arrow from oxygen ✓

shape ✓

MP2 must show dative bonds i.e. arrowhead to Cu and start of arrow from O atoms

ALLOW any attempt (wedges, lines, dots) to show
correctly 3D nature of complex ion even if other than 6
ligands

Wedges and dashes should be adjacent to each other as
in diagram

DM.c-f - Redox | DM4 | DM5 | FOCUS ON THIS ONE ITS ON THE PRELEASE 2025 GUARANTEED TO COME UP ON Scientific Notice Paper 2

How can half-equations containing ^{12}O be balanced?

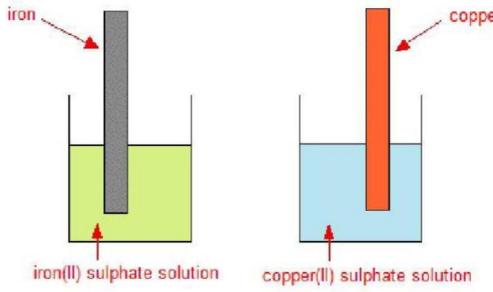
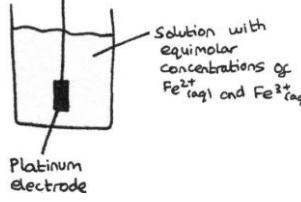
1. Check that the element that is changing oxidation state is initially balanced i.e. $\text{MnO}_4^{2-} > \text{Mn}^{2+}$ checking Mn atoms are balanced
2. Hydrogen ions and waters - balance atoms using H^+ and H_2O
3. Oxidation - balance charges using electrons
4. Check - check that atoms and charges balance.

This can be remembered as dodgy CHOC

You could also do the longer way by balancing the equation using oxidation states

What is the method for setting up electrochemical cells from textbook?

1. Construct the half-cell whose electrode potential is to be measured. For a metalion/metal half-cell (e.g., the Cu^{2+}/Cu half-cell), the electrode will be made from the metal which its ionic form in the solution. If the reaction involves two ions the same element in different oxidation states (e.g., $\text{Fe}^{3+}/\text{Fe}^{2+}$) the electrode should be either platinum or carbon (graphite) the solution will contain a

	<p>mixture of the two ions. Ensure cases all the solutions have a conc of 1M are at 298K</p> <ol style="list-style-type: none"> 2. Connect the half-cell to a standard hydrogen half-cell or other reference cell using a high-resistance voltmeter and salt bridge 3. Connect the two electrodes to a high resistance voltmeter and salt bridge [Figure 21] 4. Record the voltmeter reading - in the case of a measurement taken using a hydrogen half-cell this value is the required EMF E°_{cell}
What do simple half-cells usually consist of?	<p>A metal (acting as an electrode) and a solution of compound containing said metal. Eg, Cu and CuSO_4</p>  <p>Usually as if there was a CrO_7^{2-} to Cr^{3+} half reaction then a platinum/graphite electrode has to be used</p> <p><i>The electrode in this case is not an anode/cathode as this is not electrolysis its electrochemistry 'they are built differently' think of the meme :)</i></p>
When is a inert platinum electrode used in electrode potentials and why? (with example)	<ul style="list-style-type: none"> ● When the half equation doesn't contain a solid ● E.g. $\text{Fe}^{3+}(\text{aq}) + \text{e}^- = \text{Fe}^{2+}(\text{aq})$ 
What causes the transfer of charge through the solutions and the wire of an electrochemical cell (2)	<ul style="list-style-type: none"> ● The transfer of charge through the solution is caused by the ions (1) ● The transfer of charge through the wire is caused by electrons (1) <p><i>This is measured using a high-resistance voltmeter to stop current from flowing whilst preventing p.d. Preventing reaction. If this was replaced with a bulb, current will flow. Reactions will then occur at each electrode yet the voltage would fall to zero as the reactants are used up.</i></p>
What is the salt bridge made up of in	<ul style="list-style-type: none"> ● Usually filter paper with KNO_3 solution (1)

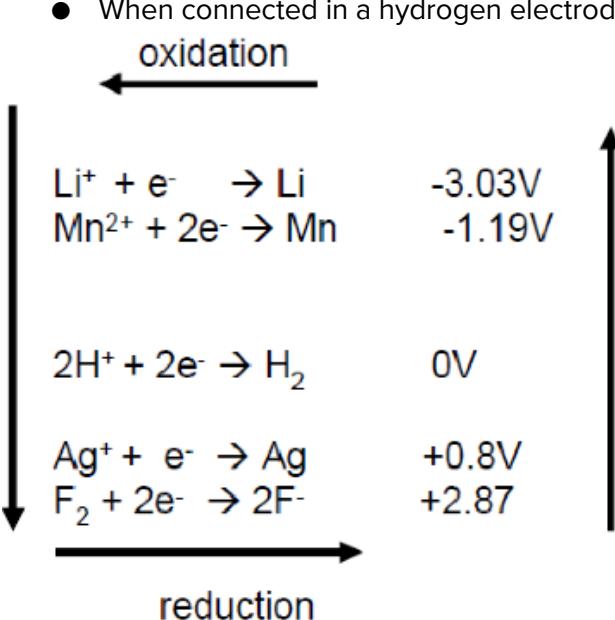
electrochemical cells and its purpose (structure and reason) (3)	<ul style="list-style-type: none"> Allows ions to flow between half cells (1) which completes the circuit (1) <p><i>Allows ions to form preventing imbalance (which would otherwise repel incoming metals).</i></p> <p><i>The salt should be unreactive. Eg, KCl will form complexes instead</i></p>						
Why isn't a wire used as the 'salt bridge'?	It would set up its own electrode system with solutions						
<p>Which direction do electrons flow in electrochemical cells?</p> <p>Table 2.1 shows the electrode potential data for the cell in Fig. 2.1.</p> <table border="1" data-bbox="203 691 584 803"> <thead> <tr> <th>Half-equation</th> <th>E°/V</th> </tr> </thead> <tbody> <tr> <td>$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s})$</td> <td>+0.34</td> </tr> <tr> <td>$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$</td> <td>+0.80</td> </tr> </tbody> </table> <p>State and explain which way electrons flow in the wire when the standard cell is delivering a current.</p> <p>.....</p> <p>..... [1]</p>	Half-equation	E°/V	$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	+0.34	$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+0.80	<p>From Cu to Ag AND Cu electrode potential is more negative/less positive ✓ 1 2.3 ALLOW Ag electrode is more positive/less negative ALLOW Cu is oxidised/loses electrons ALLOW Ag^+ ions are reduced/accept electrons</p> <ul style="list-style-type: none"> From Cu to Ag as Cu electrode potential is more negative (1) Or Ag electrode potential is more positive <p><i>From whichever electrode does oxidation to the reduction electrode</i></p>
Half-equation	E°/V						
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	+0.34						
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+0.80						
Define standard electrode potential of a half-cell (E°)?	<p>The voltage measured under standard conditions when the half-cell is connected to a standard hydrogen electrode</p> <p><i>This is by convention, we cannot measure an absolute potential for obvious reasons</i></p>						

What do standard electrode potentials tell us?

Table 20: Standard electrode potentials
The values given are at 298 K.

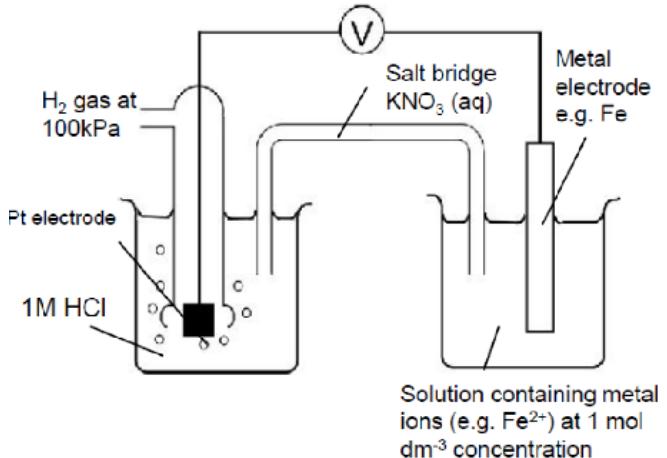
Half-reaction least oxidising	<i>Anion to cation is more likely oxidation</i>	E^\ominus/V
\downarrow		
$\text{Li}^+(\text{aq}) + \text{e}^- \rightarrow \text{Li}(\text{s})$	-3.04	
$\text{K}^+(\text{aq}) + \text{e}^- \rightarrow \text{K}(\text{s})$	-2.92	
$\text{Ca}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ca}(\text{s})$	-2.84	
$\text{Na}^+(\text{aq}) + \text{e}^- \rightarrow \text{Na}(\text{s})$	-2.71	
$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Mg}(\text{s})$	-2.36	
$\text{Al}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Al}(\text{s})$	-1.68	
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s})$	-0.76	
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe}(\text{s})$	-0.44	
$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ni}(\text{s})$	-0.26	
$\text{Sn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Sn}(\text{s})$	-0.14	
$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb}(\text{s})$	-0.13	
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00	
$\text{Sn}^{4+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Sn}^{2+}(\text{aq})$	+0.15	
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$	+0.34	
$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$	+0.40	
$\text{Cu}^{2+}(\text{aq}) + \text{e}^- \rightarrow \text{Cu}(\text{s})$	+0.52	
$\text{I}_2(\text{aq}) + 2\text{e}^- \rightarrow 2\text{I}^-(\text{aq})$	+0.54	
$\text{Fe}^{2+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$	+0.77	
$\text{Ag}^{+}(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$	+0.80	
$\text{Br}^{-}(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Br}^-(\text{aq})$	+1.07	
$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{aq})$	+1.23	
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	+1.36	
$\text{Cl}^{-}(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$	+1.36	
$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51	
\uparrow		
most oxidising	least reducing	

- The measure of the ability of a species to gain electrons. The more positive E^\ominus is, the stronger the ability.
- When connected in a hydrogen electrode system.



What is the standard hydrogen electrode (SHE), its equation with standard conditions (1mol dm⁻³, 298K and 1atm/100kPa)?

- A half cell using hydrogen
- Has the following equilibrium: $\text{H}_2(\text{g}) \rightleftharpoons 2\text{H}^+(\text{aq}) + 2\text{e}^-$



Details of the set-up of the hydrogen electrode are not required for this specification

Indirectly used by using a different standard electrode (which is calibrated against the SHE) since it's easier to use.

How do you work out the electrode potential of a cell?

- $E^\ominus_{\text{cell}} = E^\ominus_{\text{more positive}} - E^\ominus_{\text{more negative}}$
- The more positive E^\ominus at a half cell would always reduce

	<table border="1"> <thead> <tr> <th>Redox system</th><th>E^\ominus/V</th></tr> </thead> <tbody> <tr> <td>$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu(s)}$</td><td>+0.34</td></tr> <tr> <td>$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag(s)}$</td><td>+0.80</td></tr> </tbody> </table> <p>Could also think about it as $E_{\text{cell}} = E_{\text{reduction}} - E_{\text{oxidation}}$</p>	Redox system	E^\ominus/V	$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu(s)}$	+0.34	$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag(s)}$	+0.80
Redox system	E^\ominus/V						
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu(s)}$	+0.34						
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag(s)}$	+0.80						
<p>How can electrode potentials be used to predict the feasibility of reactions and why might a reaction not occur</p> <p>For instance if we had a mixture only of MnO_4^- and Zn or $\text{Mg}^{2+}(\text{aq})$ and Zn(s)?</p> $\begin{array}{l} \text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Mg(s)} & -2.36 \\ \text{Al}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Al(s)} & -1.68 \\ \text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn(s)} & -0.76 \\ \hline \text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O(l)} & +1.51 \end{array}$	<ul style="list-style-type: none"> ● Zn(s) and MnO_4^- would react as MnO_4^- can be reduced by Zn ● Zinc will lose electrons (oxidised) that will be transferred to the MnO_4^- ● Zn^{2+}/Zn has a more negative E^\ominus than $\text{MnO}_4^-/\text{Mn}^{2+}$ <p>Whilst</p> <ul style="list-style-type: none"> ● Mg^{2+} and Zn would not react as Mg^{2+}/Mg has a more negative E^\ominus than Zn^{2+}/Zn so Zn(s) would not be oxidised by Mg^{2+} 						
<p>What's wrong with using electrode potentials to determine the feasibility of whether a reaction occurs? 2 reasons</p>	<ul style="list-style-type: none"> ● The reaction may have a high activation enthalpy ● Conditions may not be standard <p><i>The rate of reaction may be slow that it appears not to be happening may be another reason</i></p>						
<p>How does increasing the concentration of Zn^{2+} affect cell potential E_{cell}? (2)</p> $\begin{array}{ll} \text{Zn}_{(\text{s})} \rightleftharpoons \text{Zn}^{2+}(\text{aq}) + 2\text{e}^- & E^\star = -0.76 \text{ V} \\ \text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}_{(\text{s})} & E^\star = +0.34 \text{ V} \end{array} \quad \text{Zn}_{(\text{s})} + \text{Cu}^{2+}(\text{aq}) \rightleftharpoons \text{Zn}^{2+}(\text{aq}) + \text{Cu}_{(\text{s})} \quad E_{\text{cell}} = +1.10 \text{ V}$	<ul style="list-style-type: none"> ● The POE shifts to the left which reduces the loss of electrons ● Therefore the E^\ominus becomes less negative (1) and so the cell potential decreases (1) <p><i>By Le Chatlier's, increasing [reactants] increases E^\ominus</i></p>						
<p>How does increasing the concentration of Cu^{2+} affect cell potential E_{cell}? (2)</p> $\begin{array}{ll} \text{Zn}_{(\text{s})} \rightleftharpoons \text{Zn}^{2+}(\text{aq}) + 2\text{e}^- & E^\star = -0.76 \text{ V} \\ \text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}_{(\text{s})} & E^\star = +0.34 \text{ V} \end{array} \quad \text{Zn}_{(\text{s})} + \text{Cu}^{2+}(\text{aq}) \rightleftharpoons \text{Zn}^{2+}(\text{aq}) + \text{Cu}_{(\text{s})} \quad E_{\text{cell}} = +1.10 \text{ V}$	<ul style="list-style-type: none"> ● The POE shifts to the right which increases the gain of electrons ● Therefore the E^\ominus becomes more positive (1) and so the E_{cell} increases (1) <p><i>By Le Chatlier's, increasing [reactants] increases E^\ominus</i></p>						
<p>How does temperature affect cell potential E_{cell}?</p>	<p>Most cells are exothermic in the forward direction. Using Le Chatelier's increasing temperature decreases E_{cell}</p>						
<p>The Prediction could be Wrong if the Conditions are Not Standard</p>							
<ol style="list-style-type: none"> 1) Changing the concentration (or temperature) of the solution can cause the electrode potential to change. 2) For example the zinc/copper cell has these half equations in equilibrium... $\begin{array}{ll} \text{Zn}_{(\text{s})} \rightleftharpoons \text{Zn}^{2+}(\text{aq}) + 2\text{e}^- & E^\star = -0.76 \text{ V} \\ \text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}_{(\text{s})} & E^\star = +0.34 \text{ V} \end{array} \quad \text{Zn}_{(\text{s})} + \text{Cu}^{2+}(\text{aq}) \rightleftharpoons \text{Zn}^{2+}(\text{aq}) + \text{Cu}_{(\text{s})} \quad E_{\text{cell}} = +1.10 \text{ V}$							
<ol style="list-style-type: none"> 3) ...if you increase the concentration of Zn^{2+}, the equilibrium will shift to the left, reducing the ease of electron loss. The electrode potential of Zn/Zn^{2+} becomes less negative and the whole cell potential will be lower. 4) ...if you increase the concentration of Cu^{2+}, the equilibrium will shift to the right, increasing the ease of electron gain. The electrode potential of Cu^{2+}/Cu becomes more positive and the whole cell potential will be higher. 							

What are the 3 stages of rusting [equations]?
With colours of ppt and state symbols

1. $\frac{1}{2}O_2(g) + H_2O(l) + 2e^- \rightarrow 2OH^-(aq)$
 $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
2. $Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$ (green rust)
3. $Fe^{2+} \rightarrow Fe^{3+} + e^-$ (think about electron config) then
 $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$ or $Fe_2O_3 \cdot xH_2O(s)$ would also occur to produce orange rust

Stage 3 occurs due to Fe^{3+} being more stable than Fe^{2+} as Fe^{3+} has a partially filled d subshell which is more stable than a full or empty d subshell

Rusting is a type of corrosion

How does having alkaline conditions affect the process of rusting?

- 1) There are two half-equations involved:
 $Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s) \quad E^\circ = -0.44 V$
 $2H_2O(l) + O_{2(g)} + 4e^- \rightleftharpoons 4OH^-(aq) \quad E^\circ = +1.23 V$
So the overall reaction is: $2H_2O(l) + O_{2(g)} + 2Fe(s) \rightarrow 2Fe^{2+}(aq) + 4OH^-(aq) \quad E^\circ = +1.67 V$
- 2) The $Fe^{2+}(aq)$ and $OH^-(aq)$ ions produced combine to form a precipitate of iron(II) hydroxide, $Fe(OH)_2$.
 $Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$
- 3) The $Fe(OH)_2$ is further oxidised to $Fe(OH)_3$ by oxygen.
 $2H_2O(l) + O_{2(g)} + 4Fe(OH)_2(s) \rightarrow 4Fe(OH)_3(s)$
- 4) Iron(III) hydroxide gradually turns into hydrated iron(III) oxide, $Fe_2O_3 \cdot xH_2O$ — this is rust.

- Increase in $[OH^-]$ makes the POE for $0.5O_2 + H_2O + 2e^- \rightleftharpoons 2OH^-$ shift to the left to reduce $[OH^-]$ which produces electrons
- Then the POE for $Fe^{2+} + 2e^- \rightleftharpoons Fe$ shifts to the right due to the extra electrons produced
- Therefore the $[Fe^{2+}]$ decreases which produces less rust

high $[OH^-]$ pushes first equilibrium to left (AW) ✓

electrons (formed) push second equilibrium to left ✓

so less Fe^{2+} (and hence rust) formed ✓

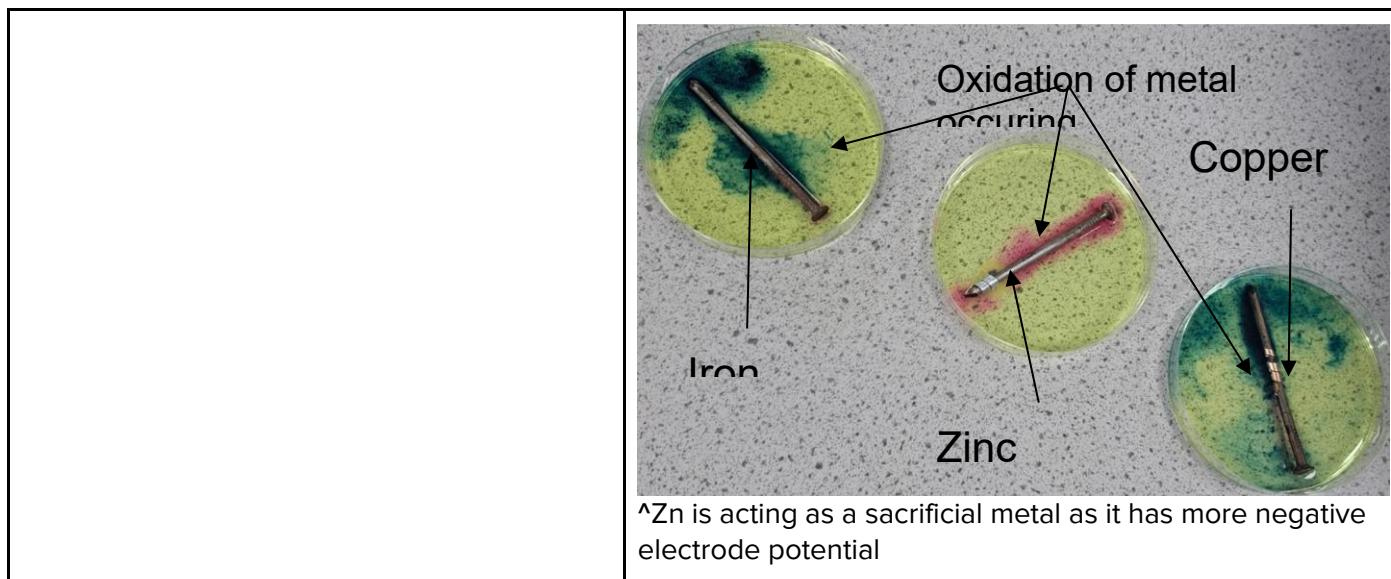
What are the preventive methods against rusting?

Barrier methods:

- Lubricating
- Painting
- Coating with a polymer
- Oiling/greasing

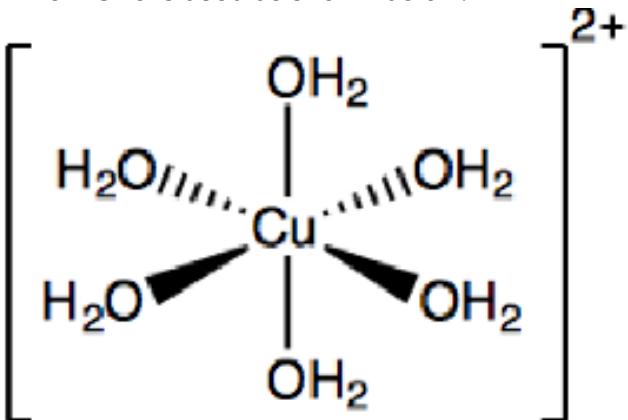
Sacrificial method:

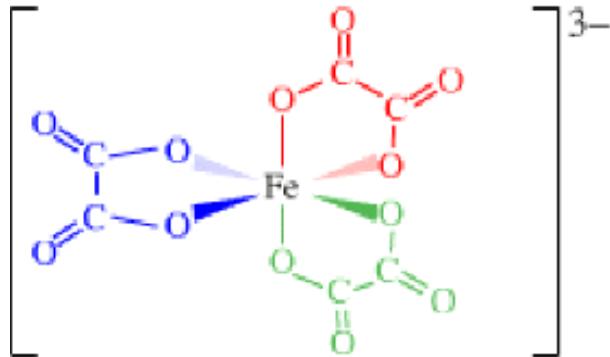
- Using a sacrificial metal, e.g zinc, and so Zn would be oxidised instead of Fe
- Galvanising by spraying zinc onto an object or using blocks of zinc like on a ship



DM.g-m - Inorganic chemistry & the periodic table | DM1 | DM2 | DM3 | DM6 |

Define a transition metal?	A d block element with at least one ion with a partially filled d subshell E.g, Cu has [Ar] 3d ¹⁰ 4s ¹ where Cu ²⁺ has [Ar] 3d ⁹ 4s ⁰ <i>For this spec when referring to the d block elements its only from scandium (3d1) to zinc (3d10) unless a question states otherwise</i>
Why aren't zinc and scandium transition elements?	<ul style="list-style-type: none"> ● Zinc can only form 2+ ion which has a complete d-subshell <ul style="list-style-type: none"> ○ Zn has [Ar] 3d¹⁰ 4s² whereas Zn²⁺ has [Ar] 3d¹⁰ 4s⁰ ● Scandium can only form 3+ ions which has an empty d-subshell <ul style="list-style-type: none"> ○ Sc has [Ar] 3d¹4s² whereas Sc³⁺ has [Ar] 3d⁰4s⁰
What are the characteristics of all transition elements?	<ul style="list-style-type: none"> ● Form coloured compounds (coloured ions) ● Variable oxidation states (Eg, Fe (II) and Fe (III)) ● Catalysts ● Form complex ions (in solution) <p><i>They form coloured ions due to their electronic configuration</i></p>
What is always true for when a d block element becomes an ion?	Whenever d block elements form ions the 4s electrons are lost FIRST
Which transition metals have anomalous	<ul style="list-style-type: none"> ● Copper - [Ar] 3d¹⁰ 4s¹ and chromium [Ar] 3d⁵ 4s¹

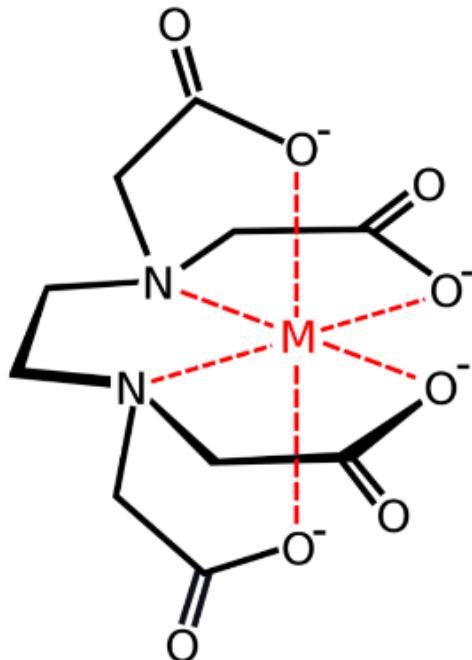
electron structures and why?	<ul style="list-style-type: none"> They have a half built/full built 3d subshells is lower in energy than a full built 4s subshell
Define ligand?	<p>An atom, ion, or molecule that uses its lone pair of electrons to form a coordinate bond to a central metal ion</p> <p><i>This bonds occurs through the ligand lone pairs of electrons</i></p>
Define complex/complex ion?	A central metal ion surrounded by ligands
How are (aq) complex ions written?	<ul style="list-style-type: none"> When they contain ligands that aren't water or hydroxide ions, you need to include all water ligands. Eg, $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ cannot be rewritten whereas $[\text{Cu}(\text{OH})_2(\text{H}_2\text{O})_4]^{2+}$ can be written as $\text{Cu}(\text{OH})_2$
Define monodentate ligand? (with examples of both ligands and complex ions)	<ul style="list-style-type: none"> Ligand that forms 1 coordinate bond to a central metal ion Examples include: H_2O, NH_3 and Cl^-, CN^- One is used as shown below:  <p>$[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ in this case could also be written as $\text{Cu}_{(\text{aq})}$</p> <p><i>In this case, the water ligands have no charge so the charge stays as 2+ from Cu^{2+}</i></p>
Define bidentate ligand? (with examples of both ligands and complex ions)	<ul style="list-style-type: none"> Ligands that forms 2 coordinate bonds to a central metal ion Examples include: $\text{C}_2\text{O}_4^{2-}$ (ethanedioate), $\text{C}_2\text{H}_4(\text{NH}_3)_2$ (ethene-1,2-diamine) This is used as shown below with Fe (III)



Each ligand reduces the charge by 2. You have 3 hence 3- Charge

Define polydentate ligand (with an example)

- Ligands that can form **more than one coordinate bond** to a central metal ion
- Examples include: EDTA⁴⁻ (forming 6 coordinate bonds)



In this case, it would be a hexadentate but don't worry about it

Don't say it is a ligand that can form more than 2 coordinate bonds or any specific number of coordinate bonds as its **wrong from Mark Scheme see below**

(polydentate ligand can form more than one bond / donate at least two lone pairs to (central) atom / ion
OR has more than one atom with a lone pair which are used to bond AW
✓

DO NOT ALLOW 'many / multiple / several lone pairs', 'more than 2 lone pairs etc.' AW
ALLOW 'has more than 1 attachment site to the central (central) atom / ion'

Define Ligand substitution (Ligand exchange reaction)?

When one ligand is swapped for another ligand

	<p><i>Also known as ligand displacement</i></p>															
Describe how iron(II) reacts with hydroxide ions and ammonia (with colours and equations)?	<ul style="list-style-type: none"> ● $\text{Fe}^{2+} \text{(aq)} + 2\text{OH}^- \text{(aq)} \rightarrow \text{Fe(OH)}_2 \text{(s)}$ ● $\text{Fe}^{2+} \text{(aq)} + 2\text{NH}_3 \text{(aq)} \rightarrow \text{Fe(OH)}_2 \text{(s)} + 2\text{NH}_4^+ \text{(aq)}$ <ul style="list-style-type: none"> ○ Via $\text{NH}_3 \text{(aq)} + \text{H}_2\text{O} \rightleftharpoons \text{OH}^- \text{(aq)} + \text{NH}_4^+ \text{(aq)}$ ● From pale green solution to green ppt <p><i>Depending on the conditions Fe^{2+} can be oxidised to Fe^{3+} so it would make an orange ppt instead so look out for that like down here</i></p> <p>17 Excess ammonia solution is added separately to $\text{Cu}^{2+} \text{(aq)}$ and $\text{Fe}^{2+} \text{(aq)}$. The solutions are left in the air.</p> <p>Which row of the table represents the colour changes seen?</p> <table border="1"> <thead> <tr> <th></th> <th>$\text{Cu}^{2+} \text{(aq)}$</th> <th>$\text{Fe}^{2+} \text{(aq)}$</th> </tr> </thead> <tbody> <tr> <td>A</td> <td>Pale blue precipitate \rightarrow dark blue solution</td> <td>Dark green precipitate \rightarrow orange precipitate</td> </tr> <tr> <td>B</td> <td>Green/blue solution \rightarrow dark blue solution</td> <td>Green precipitate remaining green</td> </tr> <tr> <td>C</td> <td>Pale blue precipitate \rightarrow green/blue solution</td> <td>Orange precipitate \rightarrow dark green precipitate</td> </tr> <tr> <td>D</td> <td>Pale blue precipitate \rightarrow dark blue solution</td> <td>Orange precipitate \rightarrow dark green precipitate</td> </tr> </tbody> </table> <p>Your answer <input type="text"/> [1]</p> <p>Answer is A btw</p> <p><i>Fe(II) DOES NOT FORM COMPLEXES WITH AMMONIA</i></p>		$\text{Cu}^{2+} \text{(aq)}$	$\text{Fe}^{2+} \text{(aq)}$	A	Pale blue precipitate \rightarrow dark blue solution	Dark green precipitate \rightarrow orange precipitate	B	Green/blue solution \rightarrow dark blue solution	Green precipitate remaining green	C	Pale blue precipitate \rightarrow green/blue solution	Orange precipitate \rightarrow dark green precipitate	D	Pale blue precipitate \rightarrow dark blue solution	Orange precipitate \rightarrow dark green precipitate
	$\text{Cu}^{2+} \text{(aq)}$	$\text{Fe}^{2+} \text{(aq)}$														
A	Pale blue precipitate \rightarrow dark blue solution	Dark green precipitate \rightarrow orange precipitate														
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Describe how iron(III) reacts with hydroxide ions and ammonia (with colours and equations)	<ul style="list-style-type: none"> ● $\text{Fe}^{3+} \text{(aq)} + 3\text{OH}^- \text{(aq)} \rightarrow \text{Fe(OH)}_3 \text{(s)}$ ● $\text{Fe}^{3+} \text{(aq)} + 3\text{NH}_3 \text{(aq)} \rightarrow \text{Fe(OH)}_3 \text{(s)} + 3\text{NH}_4^+$ <ul style="list-style-type: none"> ○ Via $\text{NH}_3 \text{(aq)} + \text{H}_2\text{O} \rightleftharpoons \text{OH}^- \text{(aq)} + \text{NH}_4^+ \text{(aq)}$ ● Goes from a yellow solution to an orange ppt <p><i>Fe(III) DOES NOT FORM COMPLEXES WITH AMMONIA</i></p>															
Describe the reaction of copper (II) with hydroxide ions (with colours and equations)	<ul style="list-style-type: none"> ● $\text{Cu}^{2+} \text{(aq)} + 2\text{OH}^- \text{(aq)} \rightarrow \text{Cu(OH)}_2 \text{(s)}$ ● Goes from pale blue solution to a blue ppt (since no charge) <p>$[\text{Cu}(\text{H}_2\text{O})_6]^{2+} \text{(aq)} + 2\text{OH}^- \text{(aq)} \rightarrow [\text{Cu}(\text{H}_2\text{O})_4(\text{OH})_2] \text{(s)} + 2\text{H}_2\text{O}$</p>															
Describe the reaction of copper (II) with ammonia and excess ammonia (with colours and equations and the properties of ammonia for each reaction)?	<ul style="list-style-type: none"> ● $[\text{Cu}(\text{H}_2\text{O})_6]^{2+} \text{(aq)} + 2\text{NH}_3 \text{(aq)} \rightarrow \text{Cu(OH)}_2 \text{(s)} + 2\text{NH}_4^+$ (aq) Occurs from: <ul style="list-style-type: none"> ○ $\text{NH}_3 \text{(aq)} + \text{H}_2\text{O} \rightleftharpoons \text{OH}^- \text{(aq)} + \text{NH}_4^+ \text{(aq)}$ ○ $\text{Cu}^{2+} \text{(aq)} + 2\text{OH}^- \text{(aq)} \rightarrow \text{Cu(OH)}_2 \text{(s)}$ ● Goes from pale blue solution to a blue ppt ● During this process, Ammonia acts as a base ● In excess NH_3, $\text{Cu(OH)}_2 \text{(s)}$ reacts further to form $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+} \text{(aq)}$ ● Which is from the blue ppt to a dark blue solution as ppt dissolves ● During this process, Ammonia acts as a ligand <p><i>This occurs via ligand substitution (ligand exchange reaction)</i></p>															
How does Homogeneous and Heterogeneous catalytic activity of a	<ul style="list-style-type: none"> ● Homogeneous - They use their variable oxidation states 															

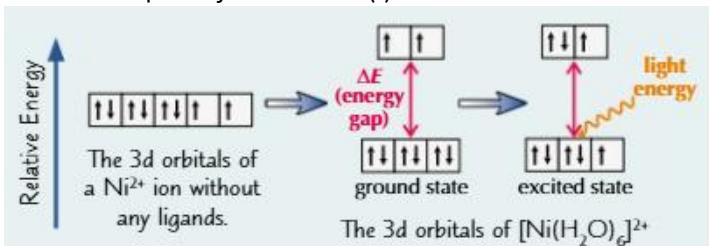
transition metal differ from the standard?

- Heterogeneous - During adsorption they use their (3)d and (4)s electrons of the atoms on the catalyst surface to form weak bonds to reactants

The majority of the time it would be the 3d or 4s electrons but it would differ if they don't use 4th-period transition metals

Why are complex ions coloured? (3)

- The d orbital (or d subshell) are split by ligands (1)
- Electrons get excited and move to a higher energy level when they absorb visible light via $\Delta E = h\nu$ (1)
- Colour transmitted is complementary to the frequency absorbed (1)

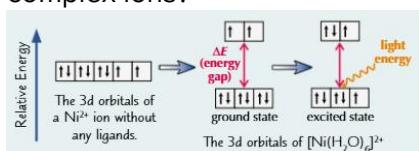


Don't accept $E=h\nu$ need the delta says on MS

Therefore different complex ions/different ligands would have a different sized gap and different colours as they would absorb a different frequency of colour and transmit a different complementary colour

I've only really seen them mention d orbital splitting in past papers not d subshell

What are 3 factors that would affect the gap size when ligands split d subshells/orbitals in complex ions?



- Different ligands
- Different central metal ion
- Different oxidation state (of central metal ion)
- Different coordinate number (from CGP guide)

What are the 3 pairs of complementary colours?

- Red and green
- Blue and orange
- Violet and yellow



So when a compound absorbs red wavelengths, it transmits (or reflects) the complementary colour green and so green is what seen

Do not allow 'emit' light

DM.n - Energy & Matter | DM3 |

What is the method for finding the concentration of an unknown solution using colorimetry/visible spectrophotometer?
[Including fine detail] and Mnemonic
[Mrs Marsden Script]

Setting up colorimeter:

- **Choose** a suitable filter / wavelength of colorimeter, showing greatest absorbance/complementary colour *
- **Zero** colorimeter using a cuvette of water *

Calibration:

- **Make** up 5 or more **standard solutions** of coloured compound/transition metal for which the concentration is known [Minimum of 5 different concentrations, use a burette to take accurate volume readings, use concentrations more and less than the unknown]
- **Measure** the absorbance for each **known** concentration of diluted standard solutions [ensure each sample is in a clean cuvette]
- **Plot** a calibration curve of absorbance against concentration and draw a line of best fit

Measuring unknown:

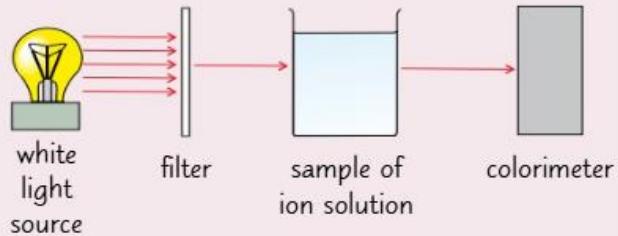
- **Measure** the absorbance of **unknown** sample and use calibration curve to determine the concentration of unknown solution *

* May be achieved by drawing a diagram

Remember this as CZMMPM

Choosing Zero Measurements Mass Produces

Moles



Remember control variables which could be keeping temperature constant or use the same total volume of solution in each cuvette bestie 😊

Visible Spectroscopy can be used to Find Concentrations of Ions too

Visible absorption spectroscopy uses a beam of **monochromatic light** — light of a **single colour**.

White light is a **mixture** of all the different wavelengths of visible light.

It can be split into monochromatic beams using a **filter**.

The machine you use in visible spectroscopy is called a **visible spectrophotometer**. Fancy.

- 1) The beam of monochromatic light is passed through a **solution** of the complex. A detector measures the **intensity of light** before and after it's passed through the solution. From this you can calculate **absorbance** — a measure of how much light the solution has absorbed.
- 2) Different frequencies of light are passed through the solution to produce a **visible absorption spectrum** (a graph of frequency or **wavelength** against absorbance). The **peaks** in the graph tell you which colours of light the complex absorbs most **strongly**.
- 3) Different complexes absorb different colours, so produce **characteristic visible absorption spectra**. This allows you to **identify** the complexes. You can then use a **calibration curve** to convert absorbance data into concentration data.

Example: This is the visible **absorption** spectrum of $[\text{Co}(\text{NH}_3)_6]^{3+}$, a yellow-orange complex.

