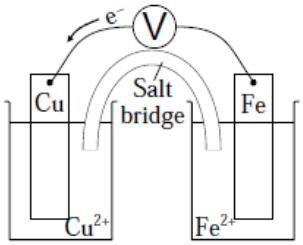
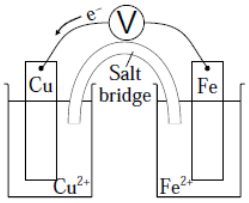
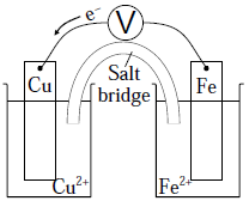
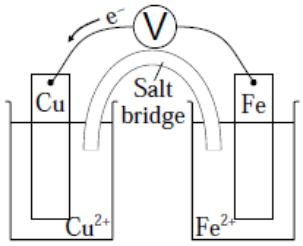
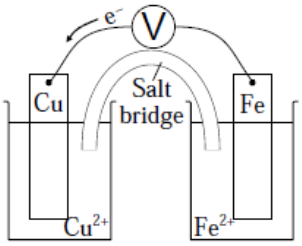
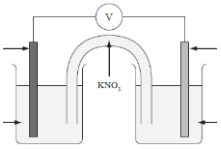
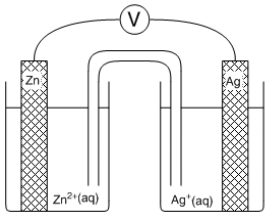
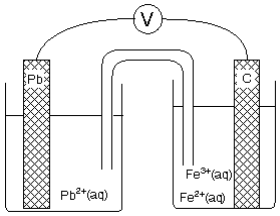
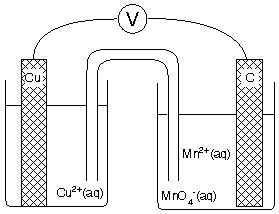
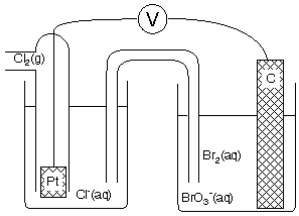
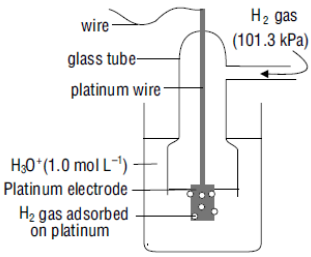
• Loss of electrons • Increase in oxidation nº.	• Gain of electrons • Decrease in oxidation nº.	direction of electron flow in external circuit	in cell diagrams*, the right-hand cell is written as ___. *by convention
oxidation	reduction	from anode to cathode	reduction
__ , __ || __ , __ order of species is written ...	suitable salt solution for making salt bridge ||	completes the circuit by allowing the movement of ions AND equalises the charge	
R , O || O , R	potassium nitrate	role of salt bridge	diagram of a cell (not the same as a cell diagram!!)
Redu**C**tion occurs at the ..	Oxid**A**ton occurs at the..	E_o(RHE) – E_o(LHE)	when written in format Cu^{2+},Cu the left species is __ , right species is __
Cathode	**A**node	$E_o$cell	oxidised form, reduced form
number assigned to an atom or ion to describe its relative state of oxidation or reduction	the species that causes another element to increase in oxidation nº	the species that causes another element to decrease in oxidation nº	OIL RIG Or LEO the lion says GER
oxidation number	an oxidising agent	a reducing agent	mnemonic for redox reactions

• Balance atoms (that are <u>not</u> H or O) • Balance O (add water) • Balance H (add H^+) • Balance charge (add e^- to more + side)	• Balance atoms (that are <u>not</u> H or O) • Balance O (add water) • Balance H (add H^+) • Balance charge (add e^- to more + side) • Add OH^- to both sides to cancel out H^+'s –make H_2O	 Reaction at RHE as drawn here	 Reaction at LHE as drawn here
balancing half equations – acidic conditions	balancing half equations – alkaline conditions	oxidation	reduction
a chemical reaction which results in the loss of electrons from a chemical species	a chemical reaction which results in a chemical species gaining electrons	a chemical species that has the ability to take electrons from other chemicals – i.e. it causes oxidation	a chemical species that has the ability to give electrons to another species – i.e. it causes reduction
oxidation	reduction	oxidant	reductant
the process by which ionic compounds are split into their atoms using electric currents	oxygen, permanganate, dichromate, iron(III), halogens, bromate, iodate… are all …	hydrogen, thiosulfate, iron(II), halides, zinc, oxalate, sulfur dioxide… are all…	Standard reduction potentials $E°$ MnO_4^-, Mn^{2+} = +1.51V Au^{3+} / Au = +1.50 V Cl_2 / Cl^- = +1.40 V Cu^{2+} / CuI = +0.86 V Fe^{3+} / Fe^{2+} = +0.77 V I_2 / I^- = +0.54 V Cu^{2+}/Cu = +0.34 V Zn^{2+} / Zn = –0.76V Fe^{2+}/Fe = –0.47 V Al^{3+} / Al = –1.66 V Mg^{2+} / Mg = –2.36 V Na^+ / Na = –2.71 V strongest RA is…
electrolysis	oxidising agents	reducing agents	Na
Standard E° potentials are measured at a pressure of ___ kPa.	Standard E° potentials are measured with respect to the standard ___ half-cell.	elements in pure form; Temp. 25 °C or 298 K; concentrations 1.0 mol L^{-1}; Pressure of gases 1.0 atm or 101.3 kPa	Standard reduction potentials $E°$ MnO_4^-, Mn^{2+} = +1.51V Au^{3+} / Au = +1.50 V Cl_2 / Cl^- = +1.40 V Cu^{2+} / CuI = +0.86 V Fe^{3+} / Fe^{2+} = +0.77 V I_2 / I^- = +0.54 V Cu^{2+}/Cu = +0.34 V Zn^{2+} / Zn = –0.76V Fe^{2+}/Fe = –0.47 V Al^{3+} / Al = –1.66 V Mg^{2+} / Mg = –2.36 V Na^+ / Na = –2.71 V strongest OA …
101.3	hydrogen	standard conditions	MnO_4^-

 blue solution	 brown gas	 orange-brown solution	 pale green solution
copper(II), Cu^{2+}	NO_2 gas nitrogen dioxide	I_2, iodine solution	$Fe^{2+}(aq)$ iron (II) solution
 pale orange solution	 orange solution	 blue/green solution	 purple solution
$Fe^{3+}(aq)$ iron (III) solution	dichromate, $Cr_2O_7^{2-}$	chromium(III) ion, Cr^{3+}	permanganate(VII), MnO_4^-
 green solution	 colourless solution	 (red)-orange solution	In $Zn(s)\|Zn^{2+}$ the | represents the
manganate (VI), MnO_4^{2-}	Cl^-, Br^-, I^-, Mn^{2+}, IO_3^-, BrO_3^-, SO_3^{2-}, SO_4^{2-}, $C_2O_4^{2-}$, $S_2O_3^{2-}$, $S_4O_6^{2-}$, H^+	Bromine solution, Br_2	phase boundary
When two half-cells are connected with a voltmeter we measure the ____	H^+/H_2 half-cell emf of zero is called the...	IUPAC convention $Zn(s)\|Zn^{2+}(aq)\|\|Cu^{2+}(aq)\|Cu(s)$	if two species are both present in a single solution, we separate them with a ____
electromotive force (emf of the cell)	standard half-cell	'cell diagram'	Comma ,

Left hand electrode $Zn(s)\|Zn^{2+}(aq)\|\|$	Right hand electrode $\|\|Cu^{2+}(aq)\|Cu(s)$	$Pt(s)\|H_2(g)\|H^+(aq)\|\|$	$\|\|Cl^-(aq)\|Cl_2(g)\|Pt(s)$
reduced | oxidised ||	|| oxidised | reduced	Inert Pt electrode placed to the outside	Inert Pt electrode placed to the outside
 Observation @ LHE	 Observation @ RHE	 $E^\circ\ (Fe^{2+}/Fe) = -0.44\ V$ $E^\circ\ (Zn^{2+}/Zn) = -0.76\ V$ What emf could be made from this cell?	$E^\circ\ (MnO_4^-/Mn^{2+}) = +1.51\ V$ $E^\circ\ (Cl_2/Cl^-) = +1.36\ V$ $E^\circ\ (Fe^{3+}/Fe^{2+}) = +0.77\ V$ $E^\circ\ (SO_4^{2-}/SO_2) = +0.20\ V$ $E^\circ\ (H^+/H_2) = +0.00\ V$ Can Fe^{3+} oxidise Cl^- to chlorine Cl_2?
Copper plated electrode, colour of solution fades	Fe electrode gets thinner, solution becomes more green (pale)	$E^o(cell) = +0.32V$	No (spontaneous reaction is Cl_2 oxidises Fe^{2+} to Fe^{3+})
$E^\circ\ (H_2O_2/H_2O) = +1.77\ V$ $E^\circ\ (O_2/H_2O_2) = +0.68\ V$ spontaneous reaction is.....	$E^\circ\ (H_2O_2/H_2O) = +1.77\ V$ $E^\circ\ (O_2/H_2O_2) = +0.68\ V$ strongest OA is	$K_2\underline{Mn}O_4$ $\underline{Mn}O_2$ $\underline{Mn}O_4^-$ oxidation numbers of Mn are	$E^\circ\ (Fe^{2+}/Fe) = -0.44\ V$ $E^\circ\ (Cu^{2+}/Cu) = 0.34\ V$ standard cell diagram is...
$H_2O_2 \rightarrow O_2 + H_2O$ (unbalanced)	H_2O_2	+6, +4, +7	$Fe \mid Fe^{2+} \mid\mid Cu^{2+} \mid Cu$
 strongest to weakest reductant	When looking at 2 std redox potentials reduction (L → R) occurs in the more ___ of the pair	When looking at 2 std redox potentials oxidation (L ← R) occurs in the more ___ of the pair	RED CAT  AN OX 
Zn > Ga > Fe as zinc can reduce both Ga^{3+} and Fe^{2+}	positive	negative	<u>red</u>uction @<u>cat</u>hode <u>ox</u>idation @<u>an</u>ode

	Ga(s)	Fe(s)	Zn(s)
$Ga^{3+}(aq)$	–	✘	✔
$Fe^{2+}(aq)$	✔	–	✔
$Zn^{2+}(aq)$	✘	✘	–

Test for chlorine gas… with starch-iodide paper; colour change is….	$Cu(s)\|Cu^{2+}(aq)\|\|Cr_2O_7^{2-}(aq)\|Cr^{3+}(aq)$ $E^o cell = +ve$ What would you see in the LHE?	$Cu(s)\|Cu^{2+}(aq)\|\|Cr_2O_7^{2-}(aq)\|Cr^{3+}(aq)$ $E^o cell = +ve$ What would you see in the RHE?	a salt bridge provides a path for ___ to move from one half-cell to the other
starch-iodide paper turns blue-black	Cu electrode shrinks, solution goes darker blue	solution goes from orange to blue/green	ions
 draw (write) the cell diagram	 draw (write) the cell diagram	 draw (write) the cell diagram	 draw (write) the cell diagram
$Zn\|Zn^{2+}\|\|Ag^{+}\|Ag$	$Pb\|Pb^{2+}\|\|Fe^{3+},Fe^{2+}\|C$	$Cu(s)\|Cu^{2+}(aq)\|\|MnO_4^-(aq), Mn^{2+}(aq)\|C(s)$	$Pt(s)\|Cl^-(aq)\|Cl_2(g)\|\|BrO_3^-(aq), Br_2(aq)\|C(s)$
platinum or graphite is used as an ____ electrode for many half-cells	an oxidising agent oxidises something else. Oxidation is loss of electrons (OIL RIG). An oxidising agent takes electrons from that other substance…		$E^o (Cr_2O_7^{2-},Cr^{3+})=1.33\ V$ $E^o (Hg^{2+}/Hg)= 0.85V$ $E^o (Fe^{3+},Fe^{2+})= 0.77V$ $E^o (I_2,I^-)= 0.54V$ $E^o (S/S^{2-})= -0.48V$ What's the strongest reductant?
inert	…so an oxidising agent must gain electrons.	standard hydrogen electrode	S^{2-}

OXIDANT O_2 is reduced to….	OXIDANT Cl_2 (pale green) is reduced to….	OXIDANT I_2 (orange-brown) is reduced to….	OXIDANT Fe^{3+} (pale orange) is reduced to….
Oxide ion O^{2-}	Chloride ion Cl^-	Iodide ion I^-	Iron(II) ion Fe^{2+} (pale green)
OXIDANT H^+ (with metals) is reduced to….	OXIDANT MnO_4^- (in acidic conditions) is reduced to….	OXIDANT MnO_4^- (in basic conditions) is reduced to….	OXIDANT MnO_4^- (in neutral conditions) is reduced to….
Hydrogen gas, H_2	Manganese(II) ion Mn^{2+}	Manganate(VI) ion MnO_4^{2-} (green)	Manganese dioxide (brown ppt) MnO_2
OXIDANT Cu^{2+} (blue) is reduced to….	OXIDANT $Cr_2O_7^{2-}/H^+$ (orange) is reduced to….	OXIDANT OCl^- is reduced to….	OXIDANT Conc. HNO_3 (NO_3^- ion) is reduced to….
Copper (metal) Cu (pinky orange)	Chromium(III) ion Cr^{3+} (greeny blue)	Chloride Cl^-	Nitrogen dioxide gas NO_2 (brown)
OXIDANT IO_3^- is reduced to….	OXIDANT MnO_2 is reduced to….	Oxidation number of underlined element $Cu\underline{S}O_4$	Oxidation number of underlined element $S_4O_6^{2-}$
Iodine I_2 (orange brown)	Manganese(II) ion Mn^{2+}	+6	+2.5

REDUCTANT Metal e.g. Mg is oxidised to….	REDUCTANT C (black) is oxidised to….	REDUCTANT CO is oxidised to….	REDUCTANT Fe^{2+}(pale green) is oxidised to….
Metal ion e.g. Mg^{2+}	CO or CO_2 gas	CO_2 gas	Iron(III) ion Fe^{3+} (pale orange)
REDUCTANT Br^- is oxidised to….	REDUCTANT I^- is oxidised to….	REDUCTANT H_2S is oxidised to….	REDUCTANT SO_2 is oxidised to….
Bromine Br_2 (red-orange)	Iodine I_2 (orange-brown)	Sulfur S (yellow)	Sulfate ion SO_4^{2-}
REDUCTANT SO_3^{2-} is oxidised to….	REDUCTANT $S_2O_3^{2-}$ is oxidised to….	REDUCTANT H_2O_2 is oxidised to….	REDUCTANT $H_2C_2O_4$ is oxidised to….
Sulfate ion SO_4^{2-}	Tetrathionate ion $S_4O_6^{2-}$	Oxygen gas O_2	Carbon dioxide gas CO_2
Oxidising agents (oxidants) are themselves	Reducing agents (reductants) are themselves	Oxidation is a loss of … (Leo the lion / OIL)	Reduction is a gain of … (says GER / RIG)
reduced	oxidised	electrons	electrons

The oxidation number of any free, uncombined element e.g. Cu, Pb, H_2, O_3 and S_8 is equal to….	**The oxidation number of a simple (monatomic) ion e.g. Na^{+}, Mg^{2+}, Cl^{-} is equal to ….**	**The SUM of the oxidation numbers of a polyatomic ion e.g. SO_4^{2-}, MnO_4^{-} is equal to….**	**In compounds the sum of the oxidation numbers of all atoms is equal to….**
Zero, 0	**the charge on that ion**	**the charge on that ion**	**Zero, 0**
The ox. number of oxygen (in compounds/ions) is __ , except in peroxides where it is __.	**The ox. number of hydrogen (in compounds/ions) is __ , except in metal hydrides where it is __.**	**If its oxidation number increases, the element has been….**	**If its oxidation number decreases, the element has been….**
-2 -1	**+1 -1**	**oxidised**	**reduced**
Oxidation number of underlined element **$\underline{Cl}O^{-}$**	**Oxidation number of underlined element** **$\underline{Cr}O_4^{2-}$**	**Oxidation number of underlined element** **$\underline{C}O_2$**	**Oxidation number of underlined element** **$\underline{N}_2$**
+1	**+6**	**+4**	**0**
Oxidation number of underlined element **$\underline{S}O_3$**	**Oxidation number of underlined element** **$H_2\underline{S}$**	**Oxidation number of underlined element** **$Na\underline{H}$**	**Oxidation number of underlined element** **$K\underline{I}O_3$**
+6	**-2**	**-1** **(NaH is a metal hydride)**	**+5**