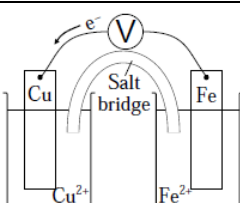
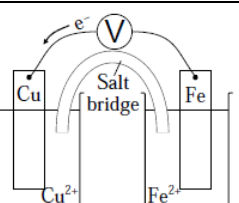








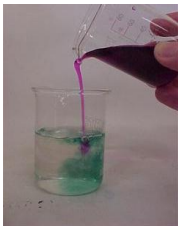

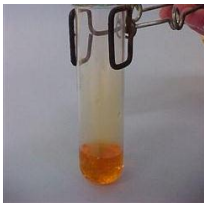
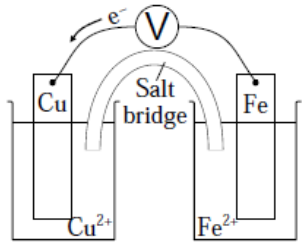
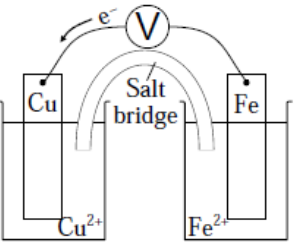
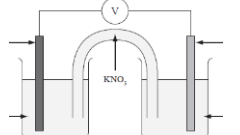


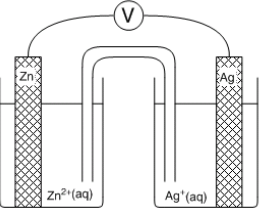
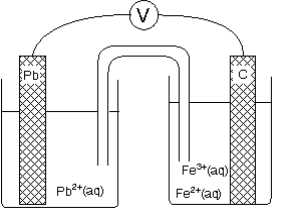
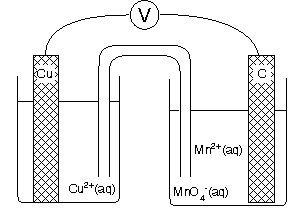
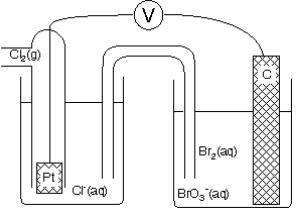
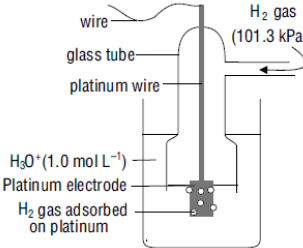


<ul style="list-style-type: none"> <li>Loss of electrons</li> <li>Increase in oxidation n<sup>o</sup>.</li> </ul>	<ul style="list-style-type: none"> <li>Gain of electrons</li> <li>Decrease in oxidation n<sup>o</sup>.</li> </ul>	direction of electron flow in external circuit	in cell diagrams*, the right-hand cell is written as ____.  *by convention
oxidation	reduction	from anode to cathode	reduction
<p>— , —    — , —</p> <p>order of species is written ...</p>	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 <b>C</b> tion occurs at the ..	Oxid <b>A</b> tion occurs at the..	$E_o(\text{RHE}) - E_o(\text{LHE})$	when written in format $\text{Cu}^{2+}, \text{Cu}$ the left species is __ , right species is __
<b>C</b> athode	<b>A</b> node	$E_o(\text{cell})$	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 n <sup>o</sup>	the species that causes another element to decrease in oxidation n <sup>o</sup>	OIL RIG Or LEO the lion says GER
oxidation number	an oxidising agent	a reducing agent	mnemonic for redox reactions

<ul style="list-style-type: none"> <li>Balance atoms (that are <u>not</u> H or O)</li> <li>Balance O (add water)</li> <li>Balance H (add H<sup>+</sup>)</li> <li>Balance charge (add e<sup>-</sup> to more + side)</li> </ul>	<ul style="list-style-type: none"> <li>Balance atoms (that are <u>not</u> H or O)</li> <li>Balance O (add water)</li> <li>Balance H (add H<sup>+</sup>)</li> <li>Balance charge (add e<sup>-</sup> to more + side)</li> <li>Add OH<sup>-</sup> to both sides to cancel out H<sup>+</sup>'s -make H<sub>2</sub>O</li> </ul>	 <p>Reaction at RHE as drawn here</p>	 <p>Reaction at LHE as drawn here</p>
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...	<p>Standard reduction potentials E°</p> <p>MnO<sub>4</sub><sup>-</sup>, Mn<sup>2+</sup> = +1.51V</p> <p>Au<sup>3+</sup> / Au = +1.50 V</p> <p>Cl<sub>2</sub> / Cl<sup>-</sup> = +1.40 V</p> <p>Cu<sup>2+</sup> / CuI = +0.86 V</p> <p>Fe<sup>3+</sup> / Fe<sup>2+</sup> = +0.77 V</p> <p>I<sub>2</sub> / I<sup>-</sup> = +0.54 V</p> <p>Cu<sup>2+</sup> / Cu = +0.34 V</p> <p>Zn<sup>2+</sup> / Zn = -0.76V</p> <p>Fe<sup>2+</sup> / Fe = -0.47 V</p> <p>Al<sup>3+</sup> / Al = -1.66 V</p> <p>Mg<sup>2+</sup> / Mg = -2.36 V</p> <p>Na<sup>+</sup> / Na = -2.71 V</p> <p><b>strongest RA is...</b></p>
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 <sup>-1</sup> ; Pressure of gases 1.0 atm or 101.3 kPa	<p>Standard reduction potentials E°</p> <p>MnO<sub>4</sub><sup>-</sup>, Mn<sup>2+</sup> = +1.51V</p> <p>Au<sup>3+</sup> / Au = +1.50 V</p> <p>Cl<sub>2</sub> / Cl<sup>-</sup> = +1.40 V</p> <p>Cu<sup>2+</sup> / CuI = +0.86 V</p> <p>Fe<sup>3+</sup> / Fe<sup>2+</sup> = +0.77 V</p> <p>I<sub>2</sub> / I<sup>-</sup> = +0.54 V</p> <p>Cu<sup>2+</sup> / Cu = +0.34 V</p> <p>Zn<sup>2+</sup> / Zn = -0.76V</p> <p>Fe<sup>2+</sup> / Fe = -0.47 V</p> <p>Al<sup>3+</sup> / Al = -1.66 V</p> <p>Mg<sup>2+</sup> / Mg = -2.36 V</p> <p>Na<sup>+</sup> / Na = -2.71 V</p> <p><b>strongest OA ...</b></p>
101.3	hydrogen	standard conditions	MnO <sub>4</sub> <sup>-</sup>

			
blue solution	brown gas	orange-brown solution	pale green solution
copper(II), $\text{Cu}^{2+}$	$\text{NO}_2$ gas nitrogen dioxide	$\text{I}_2$ , iodine solution	$\text{Fe}^{2+}(\text{aq})$ iron (II) solution
			
pale orange solution	orange solution	blue/green solution	purple solution
$\text{Fe}^{3+}(\text{aq})$ iron (III) solution	dichromate, $\text{Cr}_2\text{O}_7^{2-}$	chromium(III) ion, $\text{Cr}^{3+}$	permanganate(VII), $\text{MnO}_4^-$
			In $\text{Zn}(\text{s}) \text{Zn}^{2+}$ the   represents the
green solution	colourless solution	(red)-orange solution	
manganate (VI), $\text{MnO}_4^{2-}$	$\text{Cl}^-$ , $\text{Br}^-$ , $\text{I}^-$ , $\text{Mn}^{2+}$ , $\text{IO}_3^-$ , $\text{BrO}_3^-$ , $\text{SO}_3^{2-}$ , $\text{SO}_4^{2-}$ , $\text{C}_2\text{O}_4^{2-}$ , $\text{S}_2\text{O}_3^{2-}$ , $\text{S}_4\text{O}_6^{2-}$ , $\text{H}^+$	Bromine solution, $\text{Br}_2$	phase boundary
When two half-cells are connected with a voltmeter we measure the _____	$\text{H}^+/\text{H}_2$ half-cell emf of zero is called the...	IUPAC convention $\text{Zn}(\text{s}) \text{Zn}^{2+}(\text{aq})  \text{Cu}^{2+}(\text{aq}) \text{Cu}(\text{s})$	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 ,

<p>Left hand electrode  <math>Zn(s) Zn^{2+}(aq)  </math></p>	<p>Right hand electrode  <math>  Cu^{2+}(aq) Cu(s)</math></p>	<p><math>Pt(s) H_2(g) H^+(aq)  </math></p>	<p><math>  Cl^-(aq) Cl_2(g) Pt(s)</math></p>																
<p>reduced   oxidised   </p>	<p>   oxidised   reduced</p>	<p>Inert Pt electrode placed to the outside</p>	<p>Inert Pt electrode placed to the outside</p>																
 <p>Observation @ LHE</p>	 <p>Observation @ RHE</p>	 <p><math>E^\circ (Fe^{2+} / Fe) = -0.44 V</math>  <math>E^\circ (Zn^{2+} / Zn) = -0.76 V</math>                  What emf could be made from this cell?</p>	<p><math>E^\circ (MnO_4^- / Mn^{2+}) = +1.51 V</math>  <math>E^\circ (Cl_2 / Cl^-) = +1.36 V</math>  <math>E^\circ (Fe^{3+} / Fe^{2+}) = +0.77 V</math>  <math>E^\circ (SO_4^{2-} / SO_2) = +0.20 V</math>  <math>E^\circ (H^+ / H_2) = +0.00 V</math>                  Can <math>Fe^{3+}</math> oxidise <math>Cl^-</math> to chlorine <math>Cl_2</math>?</p>																
<p>Copper plated electrode, colour of solution fades</p>	<p>Fe electrode gets thinner, solution becomes more green (pale)</p>	<p><math>E^\circ(cell) = +0.32V</math></p>	<p>No (spontaneous reaction is <math>Cl_2</math> oxidises <math>Fe^{2+}</math> to <math>Fe^{3+}</math>)</p>																
<p><math>E^\circ (H_2O_2 / H_2O) = +1.77 V</math>  <math>E^\circ (O_2 / H_2O_2) = +0.68 V</math>                  spontaneous reaction is.....</p>	<p><math>E^\circ (H_2O_2 / H_2O) = +1.77 V</math>  <math>E^\circ (O_2 / H_2O_2) = +0.68 V</math>                  strongest OA is</p>	<p><math>K_2MnO_4</math>  <math>MnO_2</math>  <math>MnO_4^-</math>                  oxidation numbers of Mn are</p>	<p><math>E^\circ (Fe^{2+} / Fe) = -0.44 V</math>  <math>E^\circ (Cu^{2+} / Cu) = 0.34 V</math>                  standard cell diagram is...</p>																
<p><math>H_2O_2 \rightarrow O_2 + H_2O</math>                  (unbalanced)</p>	<p><math>H_2O_2</math></p>	<p>+6, +4, +7</p>	<p><math>Fe   Fe^{2+}    Cu^{2+}   Cu</math></p>																
<table border="1" data-bbox="102 1682 416 1816"> <tr> <td></td> <td>Ga(s)</td> <td>Fe(s)</td> <td>Zn(s)</td> </tr> <tr> <td><math>Ga^{3+}(aq)</math></td> <td>-</td> <td>x</td> <td>✓</td> </tr> <tr> <td><math>Fe^{2+}(aq)</math></td> <td>✓</td> <td>-</td> <td>✓</td> </tr> <tr> <td><math>Zn^{2+}(aq)</math></td> <td>x</td> <td>x</td> <td>-</td> </tr> </table> <p>strongest to weakest reductant</p>		Ga(s)	Fe(s)	Zn(s)	$Ga^{3+}(aq)$	-	x	✓	$Fe^{2+}(aq)$	✓	-	✓	$Zn^{2+}(aq)$	x	x	-	<p>When looking at 2 std redox potentials reduction (L → R) occurs in the more ___ of the pair</p>	<p>When looking at 2 std redox potentials oxidation (L ← R) occurs in the more ___ of the pair</p>	<p>RED CAT                    AN OX  </p>
	Ga(s)	Fe(s)	Zn(s)																
$Ga^{3+}(aq)$	-	x	✓																
$Fe^{2+}(aq)$	✓	-	✓																
$Zn^{2+}(aq)$	x	x	-																
<p><math>Zn &gt; Ga &gt; Fe</math>                  as zinc can reduce both <math>Ga^{3+}</math> and <math>Fe^{2+}</math></p>	<p>positive</p>	<p>negative</p>	<p>reduction @cathode                  oxidation @anode</p>																

<p>Test for chlorine gas... with starch-iodide paper; colour change is...</p>	<p><math>\text{Cu(s)} \text{Cu}^{2+}(\text{aq})  \text{Cr}_2\text{O}_7^{2-}(\text{aq}) \text{Cr}^{3+}(\text{aq})</math></p> <p><math>E^\circ_{\text{cell}} = +ve</math></p> <p>What would you see in the LHE?</p>	<p><math>\text{Cu(s)} \text{Cu}^{2+}(\text{aq})  \text{Cr}_2\text{O}_7^{2-}(\text{aq}) \text{Cr}^{3+}(\text{aq})</math></p> <p><math>E^\circ_{\text{cell}} = +ve</math></p> <p>What would you see in the RHE?</p>	<p>a salt bridge provides a path for ___ to move from one half-cell to the other</p>
<p>starch-iodide paper turns blue-black</p>	<p>Cu electrode shrinks, solution goes darker blue</p>	<p>solution goes from orange to blue/green</p>	<p>ions</p>
 <p>draw (write) the cell diagram</p>	 <p>draw (write) the cell diagram</p>	 <p>draw (write) the cell diagram</p>	 <p>draw (write) the cell diagram</p>
<p><math>\text{Zn} \text{Zn}^{2+}  \text{Ag}^+ \text{Ag}</math></p>	<p><math>\text{Pb} \text{Pb}^{2+}  \text{Fe}^{3+}, \text{Fe}^{2+} \text{C}</math></p>	<p><math>\text{Cu(s)} \text{Cu}^{2+}(\text{aq})  \text{MnO}_4^-(\text{aq}), \text{Mn}^{2+}(\text{aq}) \text{C(s)}</math></p>	<p><math>\text{Pt(s)} \text{Cl}^-(\text{aq}) \text{Cl}_2(\text{g})  \text{BrO}_3^-(\text{aq}), \text{Br}_2(\text{aq}) \text{C(s)}</math></p>
<p>platinum or graphite is used as an ___ electrode for many half-cells</p>	<p>an oxidising agent oxidises something else. Oxidation is loss of electrons (OIL RIG). An oxidising agent takes electrons from that other substance...</p>		<p><math>E^\circ(\text{Cr}_2\text{O}_7^{2-}, \text{Cr}^{3+}) = 1.33 \text{ V}</math>  <math>E^\circ(\text{Hg}^{2+}/\text{Hg}) = 0.85 \text{ V}</math>  <math>E^\circ(\text{Fe}^{3+}, \text{Fe}^{2+}) = 0.77 \text{ V}</math>  <math>E^\circ(\text{I}_2, \text{I}^-) = 0.54 \text{ V}</math>  <math>E^\circ(\text{S}/\text{S}^{2-}) = -0.48 \text{ V}</math></p> <p>What's the strongest reductant?</p>
<p>inert</p>	<p>...so an oxidising agent must gain electrons.</p>	<p>standard hydrogen electrode</p>	<p><math>\text{S}^{2-}</math></p>

<b>OXIDANT</b>  $O_2$ is reduced to....	<b>OXIDANT</b>  $Cl_2$ (pale green) is reduced to....	<b>OXIDANT</b>  $I_2$ (orange-brown) is reduced to....	<b>OXIDANT</b>  $Fe^{3+}$ (pale orange) is reduced to....
Oxide ion $O^{2-}$	Chloride ion $Cl^-$	Iodide ion $I^-$	Iron(II) ion $Fe^{2+}$ (pale green)
<b>OXIDANT</b>  $H^+$ (with metals) is reduced to....	<b>OXIDANT</b>  $MnO_4^-$ (in acidic conditions) is reduced to....	<b>OXIDANT</b>  $MnO_4^-$ (in basic conditions) is reduced to....	<b>OXIDANT</b>  $MnO_4^-$ (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) $MnO_2$
<b>OXIDANT</b>  $Cu^{2+}$ (blue) is reduced to....	<b>OXIDANT</b>  $Cr_2O_7^{2-}/H^+$ (orange) is reduced to....	<b>OXIDANT</b>  $OCI^-$ is reduced to....	<b>OXIDANT</b>  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)
<b>OXIDANT</b>  $IO_3^-$ is reduced to....	<b>OXIDANT</b>  $MnO_2$ is reduced to....	Oxidation number of underlined element $Cu\underline{S}O_4$	Oxidation number of underlined element $S_4\underline{O}_6^{2-}$
Iodine $I_2$ (orange brown)	Manganese(II) ion $Mn^{2+}$	<b>+6</b>	<b>+2.5</b>

<b>REDUCTANT</b>  Metal e.g. Mg is oxidised to....	<b>REDUCTANT</b>  C (black) is oxidised to....	<b>REDUCTANT</b>  CO is oxidised to....	<b>REDUCTANT</b>  Fe <sup>2+</sup> (pale green) is oxidised to....
Metal ion e.g. Mg <sup>2+</sup>	CO or CO <sub>2</sub> gas	CO <sub>2</sub> gas	Iron(III) ion Fe <sup>3+</sup> (pale orange)
<b>REDUCTANT</b>  Br <sup>-</sup> is oxidised to....	<b>REDUCTANT</b>  I <sup>-</sup> is oxidised to....	<b>REDUCTANT</b>  H <sub>2</sub> S is oxidised to....	<b>REDUCTANT</b>  SO <sub>2</sub> is oxidised to....
Bromine Br <sub>2</sub> (red-orange)	Iodine I <sub>2</sub> (orange-brown)	Sulfur S (yellow)	Sulfate ion SO <sub>4</sub> <sup>2-</sup>
<b>REDUCTANT</b>  SO <sub>3</sub> <sup>2-</sup> is oxidised to....	<b>REDUCTANT</b>  S <sub>2</sub> O <sub>3</sub> <sup>2-</sup> is oxidised to....	<b>REDUCTANT</b>  H <sub>2</sub> O <sub>2</sub> is oxidised to....	<b>REDUCTANT</b>  H <sub>2</sub> C <sub>2</sub> O <sub>4</sub> is oxidised to....
Sulfate ion SO <sub>4</sub> <sup>2-</sup>	Tetrathionate ion S <sub>4</sub> O <sub>6</sub> <sup>2-</sup>	Oxygen gas O <sub>2</sub>	Carbon dioxide gas CO <sub>2</sub>
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 <sub>2</sub> , O <sub>3</sub> and S <sub>8</sub> is equal to....	The oxidation number of a simple (monatomic) ion e.g. Na <sup>+</sup> , Mg <sup>2+</sup> , Cl <sup>-</sup> is equal to ....	The SUM of the oxidation numbers of a polyatomic ion e.g. SO <sub>4</sub> <sup>2-</sup> , MnO <sub>4</sub> <sup>-</sup> 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 <u>Cl</u> O <sup>-</sup>	Oxidation number of underlined element <u>Cr</u> O <sub>4</sub> <sup>2-</sup>	Oxidation number of underlined element <u>C</u> O <sub>2</sub>	Oxidation number of underlined element <u>N</u> <sub>2</sub>
+1	+6	+4	0
Oxidation number of underlined element <u>S</u> O <sub>3</sub>	Oxidation number of underlined element H <sub>2<u>S</u></sub>	Oxidation number of underlined element Na <u>H</u>	Oxidation number of underlined element K <u>I</u> O <sub>3</sub>
+6	-2	-1 (NaH is a metal hydride)	+5