<table>
<tr><th></th><th>1</th><th>2</th><th>3</th><th>4</th><th>5</th><th>6</th></tr>
<tr><td>1</td><td>redox</td><td>half equation</td><td>disproprotio nation</td><td rowspan="2">Make into a balanced half equation:
$MnO_4^- \rightarrow Mn^{2+}$</td><td>spectator ion</td><td>oxidation</td></tr>
<tr><td>2</td><td>oxidation state = 0</td><td>peroxide</td><td>oxidising agent</td><td>OILRIG</td><td rowspan="2">simultaneous oxidation and reduction of a single species to yield two different products</td></tr>
<tr><td>3</td><td>$SO_4^{2-}(aq)$</td><td rowspan="2">Combine the following half equations into a full redox equation:
$Zn \rightarrow Zn^{2+} + 2e^-$
$Ag^+ + e^- \rightarrow Ag$</td><td colspan="2">$Cl_2 + 2OH^- \longrightarrow ClO^- + Cl^- + H_2O$</td><td>Oxidation state of Cl in $ClO_3^-{}_{(aq)}$</td></tr>
<tr><td>4</td><td>Balance the following half equation:</td><td>electron donator</td><td>use e^-, H^+ and H_2O ONLY</td><td>Sulfate (VI) ions</td><td>ionic equation</td></tr>
<tr><td>5</td><td>oxidation state (oxidation number)</td><td>sum of the oxidation states for a neutral compound</td><td>reduction</td><td>Full redox equation</td><td>Copper (I) oxide or copper (II) oxide</td><td>$Na_{(s)} \rightarrow Na^+_{(aq)} + e^-$</td></tr>
<tr><td>6</td><td>$Cl_{2(g)} + 2e^- \rightarrow 2Cl^-_{(aq)}$</td><td><u>NOT</u> a half equation:
$2Mg + O_2 \rightarrow 2MgO$</td><td>oxidation state of O in H_2O_2</td><td>reducing agent</td><td>electron acceptor</td><td>oxidation state of Cr in $K_2Cr_2O_7$</td></tr>
</table>

DP Redox