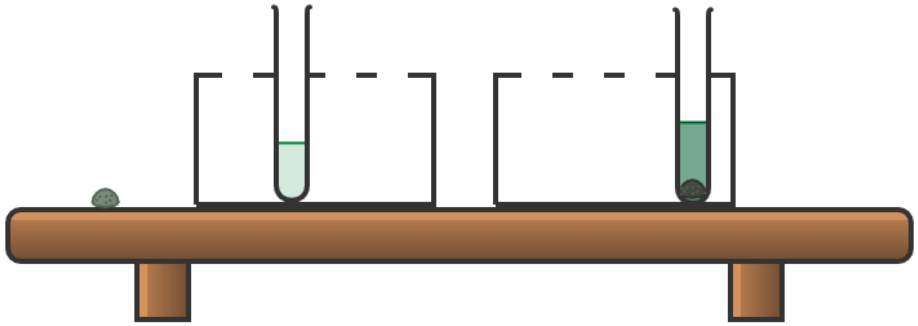
Practice questions - answers

$Cr(s) + Fe^{2+}(aq) \rightarrow Cr^{3+}(aq) + Fe(s)$ (unbalanced)
Step #1: What colour are the 4 species? Step #2: Balance the half equations Step #3: Identify which half reaction is oxidation and which is reduction – and WHY you knew this. Step #4: Combine the half equations to produce the overall balanced equation. Remember the number of e^- in each side MUST be the same so that they cancel out.

Write a description of what would be observed at the start and end of this reaction:

A piece of shiny light grey metal, Cr was added to a pale green solution, the colour being due to $Fe^{2+}(aq)$; a dark grey solid, Fe, was seen and the solution turned a dark green due to the formation of $Cr^{3+}(aq)$.



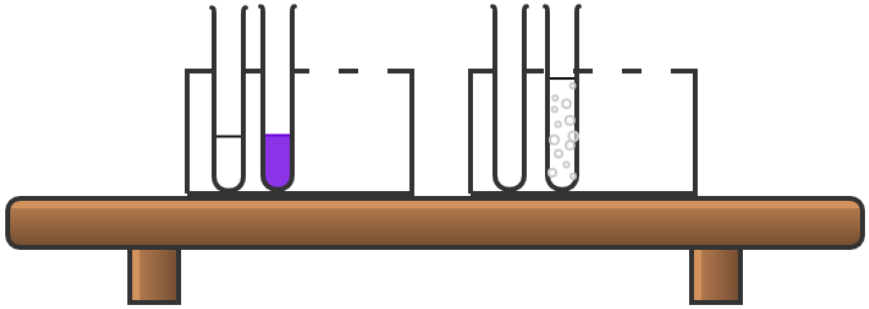
Unbalanced half equations:	$Fe^{2+} \rightarrow Fe$	$Cr \rightarrow Cr^{3+}$	Species Colours
Balanced half equation:	$Fe^{2+} + 2e^- \rightarrow Fe$	$Cr \rightarrow Cr^{3+} + 3e^-$	Fe^{2+} =pale green solution Fe = dark grey metal Cr = light grey metal Cr^{3+} = green solution
Oxidation / Reduction:	reduction	oxidation	
This is because….	Each Fe^{2+} ion has gained 2 electrons	Each Cr atom has lost3 electrons	
Overall equation:	$3Fe^{2+} + 2Cr \rightarrow 3Fe + 2Cr^{3+}$		

$HCOOH(aq) + MnO_4^-(aq) \rightarrow CO_2(g) + Mn^{2+}(aq)$ (unbalanced)

Step #1: What colour are the 4 species?
Step #2: Separate out the species and balance the half equations
Step #3: Identify which half reaction is oxidation and which is reduction – and WHY you knew this.
Step #4: Combine the half equations to produce the overall balanced equation. Remember the number of e^- in each side MUST be the same so that they cancel out.

Write a description of what would be observed at the start and during/end of this reaction:

Colourless methanoic acid is added to purple permanganate, the purple colour being due to $MnO_4^{2-}(aq)$; the colour disappears as colourless $Mn^{2+}(aq)$ ions are formed, and bubbles of a colourless gas are seen. The gas is carbondioxide



Unbalanced half equations:	$MnO_4^- \rightarrow Mn^{2+}$	$HCOOH \rightarrow CO_2$	Species Colours
Balanced half equation:	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	$HCOOH \rightarrow CO_2 + 2H^+ + 2e^-$	MnO_4^- = purple solution Mn^{2+} = colourless solution HCOOH = colourless solution CO_2 = colourless gas
Oxidation / Reduction:	reduction	oxidation	
This is because….	Each MnO_4^- ion gains 5 electrons	Each HCOOH molecule loses 2 electrons	
Overall equation:	$2MnO_4^- + 6H^+ + 5HCOOH \rightarrow 2Mn^{2+} + 8H_2O + 5CO_2$		

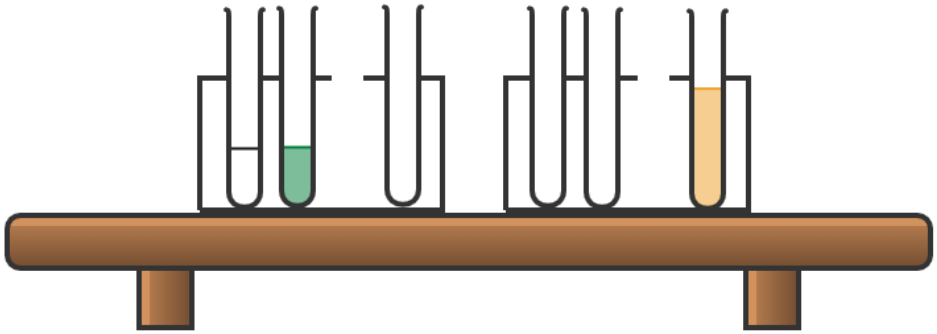
$H_2O_2(aq) + Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + H_2O(l)$ (unbalanced)

Step #1: What colour are the 4 species?
Step #2: Separate out the species and balance the half equations
Step #3: Identify which half reaction is oxidation and which is reduction – and WHY you knew this.
Step #4: Combine the half equations to produce the overall balanced equation. Remember the number of e^- in each side MUST be the same so that they cancel out.

Write a description of what would be observed at the start and end of this reaction:

Colourless hydrogen peroxide solution is mixed with a solution containing pale green $Fe^{2+}(aq)$. The solution turns a pale orange due to the formation of $Fe^{3+}(aq)$. The other product doesn't affect the colour as it is colourless water.

(Rather confusingly bubbles of colourless gas are also seen but this is due to the $Fe^{3+}(aq)$ ions catalysing the decomposition of the remaining H_2O_2)



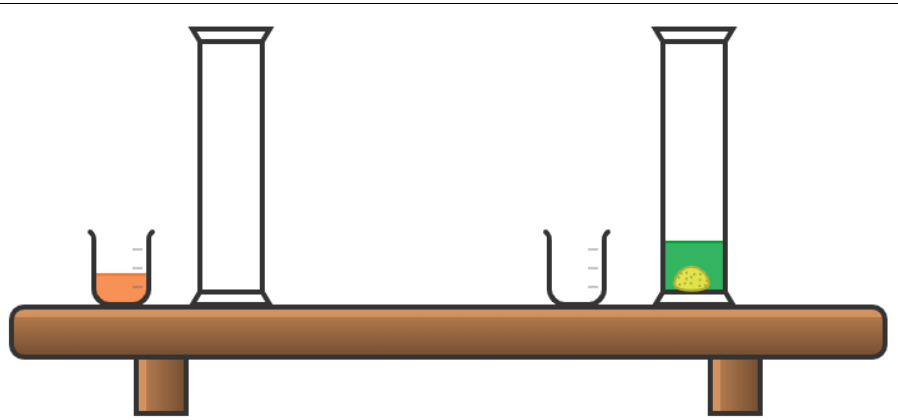
Unbalanced half equations:	$Fe^{2+} \rightarrow Fe^{3+}$	$H_2O_2 \rightarrow H_2O$	Species Colours
Balanced half equation:	$Fe^{2+} \rightarrow Fe^{3+} + e^-$	$2H^+ + H_2O_2 + 2e^- \rightarrow 2H_2O$	Fe^{2+}= pale green solution
Oxidation / Reduction:	oxidation	reduction	Fe^{3+} = pale orange solution
This is because….	Each Fe^{2+} ion loses one electron	Each H_2O_2 molecule gains 2 electrons	H_2O_2 = colourless solution
Overall equation:	$2H^+ + H_2O_2 + 2Fe^{2+} \rightarrow 2H_2O + 2Fe^{3+}$		H_2O = colourless liquid

A pinky-brown metal, Cu, was added to some colourless concentrated nitric acid (H^+ and NO_3^- are both colourless ions). The solution turned from colourless to blue, due to the formation of Cu^{2+} and a large amount of brown gas, NO_2, was released.

Step #1: Identify the species using the observations.
Step #2: Separate them out into two half equations and then balance these half equations
Step #3: Identify which half reaction is oxidation and which is reduction – and WHY you knew this.
Step #4: Combine the half equations to produce the overall balanced equation. Remember the number of e^- in each side MUST be the same so that they cancel out.

Unbalanced half equations:	$Cu \rightarrow Cu^{2+}$	$HNO_3 \rightarrow NO_2$	Species Colours
Balanced half equation:	$Cu \rightarrow Cu^{2+} + 2e^-$	$H^+ + HNO_3 + e^- \rightarrow NO_2 + H_2O$ Or $2H^+ + NO_3^- + e^- \rightarrow NO_2 + H_2O$	HNO_3 = colourless liquid Or NO_3^- = colourless solution Cu = pinky orange solid NO_2 = brown gas Cu^{2+} = blue solution
Oxidation / Reduction:	oxidation	reduction	
This is because….	Each Cu atom loses 2 electrons	Each HNO_3 gains one electron Or each NO_3^- gains one electron	
Overall equation:	$Cu + 4H^+ + 2NO_3^- \rightarrow Cu^{2+} + 2NO_2 + 2H_2O$		

Now for some harder ones……

<table>
<tr><td colspan="4">Acidified dichromate solution was mixed with hydrogen sulfide gas. A dark green solution formed and a yellow solid.</td></tr>
<tr><td>Orange dichromate solution $Cr_2O_7^{2-}(aq)$ is added to colourless $H_2S(g)$, hydrogen sulfide gas</td><td colspan="2">  </td><td>A green solution is seen due to $Cr^{3+}(aq)$ being formed and the yellow solid is sulfur, S(s)</td></tr>
<tr><td>Unbalanced half equation:</td><td>$Cr_2O_7^{2-} \rightarrow Cr^{3+}$</td><td>$H_2S \rightarrow S$</td><td>Species Colours</td></tr>
<tr><td>Balanced half equation:</td><td>$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$</td><td>$H_2S \rightarrow S + 2H^+ + 2e^-$</td><td rowspan="4">$Cr_2O_7^{2-}$= orange solution
H_2S = colourless gas
Cr^{3+} = green solution
S = yellow solid</td></tr>
<tr><td>Oxidation / Reduction:</td><td>reduction</td><td>oxidation</td></tr>
<tr><td>This is because….</td><td>Each $Cr_2O_7^{2-}$ ion gains six electrons</td><td>Each H_2S molecule loses 2 electrons</td></tr>
<tr><td>Overall equation:</td><td colspan="2">$Cr_2O_7^{2-} + 8H^+ + 3H_2S \rightarrow 2Cr^{3+} + 7H_2O + 3S$</td></tr>
</table>

SO_2 gas + dilute acidified $Cr_2O_7^{2-}$ solution react together to produce a dark green solution.

Hint: *remember that SO_2 and HSO_3^- both react in similar ways to each other and are converted into the sulfate ion SO_4^{2-} unless a yellow solid is observed – which is sulfur, S.*

<table>
<tr><td colspan="4">$SO_2 + Cr_2O_7^{2-} \rightarrow$ [you have to work it out] $SO_4^{2-} + Cr^{3+}$</td></tr>
<tr><td colspan="4">Expected observations. Link these to the species involved.

The colourless SO_2 gas reacts with the dichromate solution which is orange due to the $Cr_2O_7^{2-}$(aq) ions. The observed colour change is because colourless sulfate ions / SO_4^{2-}(aq) are formed as well as Cr^{3+}(aq) and these are green in solution.</td></tr>
<tr><td>Unbalanced half equation:</td><td>$Cr_2O_7^{2} \rightarrow Cr^{3+}$</td><td>$SO_2 \rightarrow SO_4^{2-}$</td><td>Species Colours</td></tr>
<tr><td>Balanced half equation:</td><td>$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$</td><td>$2H_2O + SO_2 \rightarrow SO_4^{2-} + 4H^+ + 2e^-$</td><td rowspan="4">$Cr_2O_7^{2-}$ = orange solution
SO_2 = colourless gas
SO_4^{2-} = colourless solution
Cr^{3+} = green solution</td></tr>
<tr><td>Oxidation / Reduction:</td><td>reduction</td><td>oxidation</td></tr>
<tr><td>This is because….</td><td>Each $Cr_2O_7^{2-}$ ion gains six electrons</td><td>Each SO_2 molecule loses 2 electrons</td></tr>
<tr><td>Overall equation:</td><td colspan="2">$3SO_2 + Cr_2O_7^{2-} + 2H^+ \rightarrow 2Cr^{3+} + H_2O + 3SO_4^{2-}$</td></tr>
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