

More complex redox reactions

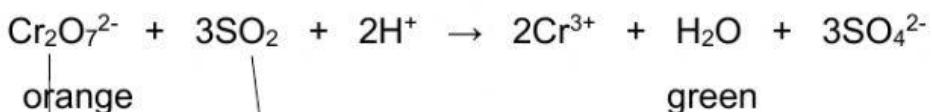
Instructions

- ✓ **Do not** include the phases
- ✓ **Write H⁺ instead of H+1**
- ✓ Always write the charge before the number eg Mg+2 not Mg2+
- ✓ Write the substances in the oxidation half reaction first in the redox reaction and then the substances in the reduction half reaction

Now sometimes the redox reactions are a little more 'complicated' and involve some of the compound ions oxidising or reducing.

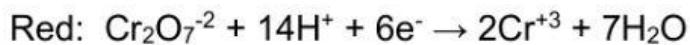
Eg

When SO_2 gas is bubbled through an acidified potassium dichromate solution, the colour of the solution changes from orange to green. The change in colour indicates that the dichromate ion ($\text{Cr}_2\text{O}_7^{2-}$) has changed to a chromium ion (Cr^{3+}), according to the equation:



Now you need to find one of these substances on the right of the table and another on the left. One substance still needs to reduce and one oxidise. Continue to use the C-rule.

$\text{Fe}^{3+} + 3\text{e}^-$	\rightleftharpoons	Fe	-0,06
$2\text{H}^+ + 2\text{e}^-$	\rightleftharpoons	$\text{H}_2(\text{g})$	0,00
$\text{S} + 2\text{H}^+ + 2\text{e}^-$	\rightleftharpoons	$\text{H}_2\text{S}(\text{g})$	+0,14
$\text{Sn}^{4+} + 2\text{e}^-$	\rightleftharpoons	Sn^{2+}	+0,15
$\text{Cu}^{2+} + \text{e}^-$	\rightleftharpoons	Cu^+	+0,18
$\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^-$	\rightleftharpoons	$\text{SO}_2(\text{g}) + 2\text{H}_2\text{O}$	+0,17
$\text{Cu}^+ + 2\text{e}^-$	\rightleftharpoons	Cu	+0,34
$2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^-$	\rightleftharpoons	4OH^-	+0,40
$\text{SO}_3^{2-} + 4\text{H}^+ + 4\text{e}^-$	\rightleftharpoons	$\text{S} + 2\text{H}_2\text{O}$	+0,45
$\text{Cu}^+ + \text{e}^-$	\rightleftharpoons	Cu	+0,52
$\text{I}_2 + 2\text{e}^-$	\rightleftharpoons	2I^-	+0,54
$\text{O}_2(\text{g}) + 2\text{H}^+ + 2\text{e}^-$	\rightleftharpoons	H_2O_2	+0,68
$\text{Fe}^{2+} + \text{e}^-$	\rightleftharpoons	Fe^{2+}	+0,77
$\text{NO}_3^- + 2\text{H}^+ + \text{e}^-$	\rightleftharpoons	$\text{NO}(\text{g}) + \text{H}_2\text{O}$	+0,80
$\text{Ag}^+ + \text{e}^-$	\rightleftharpoons	Ag	+0,80
$\text{Hg}^{2+} + 2\text{e}^-$	\rightleftharpoons	$\text{Hg}(\text{l})$	+0,85
$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^-$	\rightleftharpoons	$\text{NO}(\text{g}) + 2\text{H}_2\text{O}$	+0,96
$\text{Br}_2(\text{l}) + 2\text{e}^-$	\rightleftharpoons	2Br^-	+1,07
$\text{Pt}^{4+} + 2\text{e}^-$	\rightleftharpoons	Pt	+1,20
$\text{MnO}_4^- + 4\text{H}^+ + 2\text{e}^-$	\rightleftharpoons	$\text{Mn}^{2+} + 2\text{H}_2\text{O}$	+1,23
$\text{O}_2(\text{g}) + \text{H}^+ + \text{e}^-$	\rightleftharpoons	$2\text{H}_2\text{O}$	+1,23
$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^-$	\rightleftharpoons	$2\text{Cr}^{3+} + 7\text{H}_2\text{O}$	+1,33
$\text{Cl}_2(\text{g}) + 2\text{e}^-$	\rightleftharpoons	2Cl^-	+1,36
$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^-$	\rightleftharpoons	$\text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1,51
$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^-$	\rightleftharpoons	$2\text{H}_2\text{O}$	+1,77
$\text{Co}^{3+} + \text{e}^-$	\rightleftharpoons	Co^{2+}	+1,81
$\text{F}_2(\text{g}) + 2\text{e}^-$	\rightleftharpoons	2F^-	+2,87



Balanced Redox:



Now if anything appears on both sides of the reaction, cancel them out.



Does this reaction look like the original given redox reaction- yes it does

1.1 Write down the oxidation number of the Cr in $\text{Cr}_2\text{O}_7^{2-}$:

1.2 Was the Cr oxidised or reduced in the reaction:

1.3 Which substance acted as the reducing agent during the reaction:

Exercise 2:

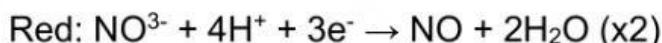
When $\text{Cu}_{(s)}$ is reacted with dilute nitric acid (HNO_3), the copper is oxidised to Cu^{2+} ions, and the nitrate ion NO_3^- is reduced to NO gas.

2.1 write the oxidation number of N in NO_3^{-1}

Now look for reactions on the table:

$S + 2H^+ + 2e^- \rightleftharpoons H_2S(g)$	$\Delta E^\circ = 0.14$
$Sn^{4+} + 2e^- \rightleftharpoons Sn^{2+}$	$\Delta E^\circ = 0.15$
$Cu^{2+} + e^- \rightleftharpoons Cu^+$	$\Delta E^\circ = 0.16$
$SO_3^{2-} + 4H^+ + 2e^- \rightleftharpoons SO_2(g) + 2H_2O$	$\Delta E^\circ = 0.17$
$Cu^{2+} + 2e^- \rightleftharpoons Cu$	$\Delta E^\circ = 0.34$
$2H_2O + O_2 + 4e^- \rightleftharpoons 4OH^-$	$\Delta E^\circ = 0.40$
$SO_2 + 4H^+ + 4e^- \rightleftharpoons S + 2H_2O$	$\Delta E^\circ = 0.45$
$Cu^+ + e^- \rightleftharpoons Cu$	$\Delta E^\circ = 0.52$
$I_2 + 2e^- \rightleftharpoons 2I^-$	$\Delta E^\circ = 0.54$
$O_2(g) + 2H^+ + 2e^- \rightleftharpoons H_2O_2$	$\Delta E^\circ = 0.68$
$Fe^{2+} + e^- \rightleftharpoons Fe^{3+}$	$\Delta E^\circ = 0.77$
$NO_3^- + 2H^+ + e^- \rightleftharpoons NO_2(g) + H_2O$	$\Delta E^\circ = 0.80$
$Ag^+ + e^- \rightleftharpoons Ag$	$\Delta E^\circ = 0.80$
$UO_2^{2+} + 2e^- \rightleftharpoons UO_2^{+}$	$\Delta E^\circ = 0.86$
$NO_3^- + 4H^+ + 3e^- \rightleftharpoons NO(g) + 2H_2O$	$\Delta E^\circ = 0.96$
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-$	$\Delta E^\circ = 1.07$
$Pt^{4+} + 2e^- \rightleftharpoons Pt^{2+}$	$\Delta E^\circ = 1.20$
$MnO_4^- + 4H^+ + 2e^- \rightleftharpoons Mn^{2+} + 2H_2O$	$\Delta E^\circ = 1.23$
$O_2(g) + 4H^+ + 4e^- \rightleftharpoons 2H_2O$	$\Delta E^\circ = 1.23$
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightleftharpoons 2Cr^{3+} + 7H_2O$	$\Delta E^\circ = 1.33$
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-$	$\Delta E^\circ = 1.36$

Now there are 2 reactions with NO_3^- , but the one in red, creates NO_2 , not NO as stated in the question.

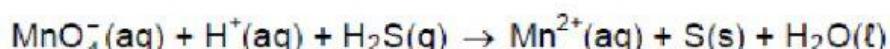


Exercise 3:

Instructions

- ✓ **Do not include the phases**
- ✓ **Write H^+ instead of $H+1$**
- ✓ Always write the charge before the number eg $Mg+2$ not $Mg2+$
- ✓ Write the substances in the oxidation half reaction first in the redox reaction and then the substances in the reduction half reaction

The reaction between permanganate ions (MnO_4^-) and hydrogen sulphide (H_2S) is given below.



Write the



3.3 By what number must the oxidation half reaction be multiplied to balance the reaction:

3.4 By what number must the reduction half reaction be multiplied to balance the reaction:

3.5 Balanced redox reaction (after any substances appearing on both sides have been cancelled out):



3.6 What is the oxidation number of Mn in the permanganate ion:

3.7 Write down the formula of the substance that undergoes oxidation :

3.8 Write down the formula of the oxidising agent:

Exercise 4

The unbalanced redox reaction , in which SO_2 is involved, is shown below.

Reaction 1: $\text{SO}_2(\text{g}) + \text{H}_2\text{S}(\text{g}) \rightarrow \text{S}(\text{s}) + \text{H}_2\text{O}(\text{l})$

4.1 oxi: $\rightarrow + +$

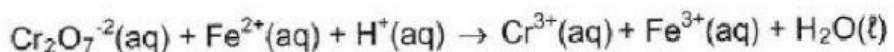
4.2 red: $+ + \rightarrow +$

4.3 balanced redox: (after any substances appearing on both sides have been cancelled out):



Exercise 5

The reaction between dichromate ions ($\text{Cr}_2\text{O}_7^{2-}$) and iron(II) ions (Fe^{2+}) in an acidic medium is given below.



5.1 oxi: $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) \rightarrow \text{Cr}^{3+}(\text{aq}) + \text{H}_2\text{O}(l)$

5.2 red: $\text{Fe}^{2+}(\text{aq}) + \text{H}^+(\text{aq}) \rightarrow \text{Fe}^{3+}(\text{aq}) + \text{H}_2\text{O}(l)$

5.3 balanced redox: (after any substances appearing on both sides have been cancelled out):

