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 ($Cr_2O_7^{2-}$) has changed to a chromium ion (Cr^{3+}), according to the equation:

$$\underset{\text{orange}}{Cr_2O_7^{2-}} + 3SO_2 + 2H^+ \rightarrow \underset{\text{green}}{2Cr^{3+}} + H_2O + 3SO_4^{2-}$$

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.

$Fe^{3+} + 3e^-$	$\rightleftharpoons$	Fe	−0,06
$2H^+ + 2e^-$	$\rightleftharpoons$	$H_2(g)$	0,00
$S + 2H^+ + 2e^-$	$\rightleftharpoons$	$H_2S(g)$	+0,14
$Sn^{4+} + 2e^-$	$\rightleftharpoons$	Sn^{2+}	+0,15
$Cu^{2+} + e^-$	$\rightleftharpoons$	Cu^+	+0,16
$SO_4^{2-} + 4H^+ + 2e^-$	$\rightleftharpoons$	$SO_2(g) + 2H_2O$	+0,17
$Cu^{2+} + 2e^-$	$\rightleftharpoons$	Cu	+0,34
$2H_2O + O_2 + 4e^-$	$\rightleftharpoons$	$4OH^-$	+0,40
$SO_2 + 4H^+ + 4e^-$	$\rightleftharpoons$	$S + 2H_2O$	+0,45
$Cu^+ + e^-$	$\rightleftharpoons$	Cu	+0,52
$I_2 + 2e^-$	$\rightleftharpoons$	$2I^-$	+0,54
$O_2(g) + 2H^+ + 2e^-$	$\rightleftharpoons$	H_2O_2	+0,68
$Fe^{3+} + e^-$	$\rightleftharpoons$	Fe^{2+}	+0,77
$NO_3^- + 2H^+ + e^-$	$\rightleftharpoons$	$NO_2(g) + H_2O$	+0,80
$Ag^+ + e^-$	$\rightleftharpoons$	Ag	+0,80
$Hg^{2+} + 2e^-$	$\rightleftharpoons$	$Hg(\ell)$	+0,85
$NO_3^- + 4H^+ + 3e^-$	$\rightleftharpoons$	$NO(g) + 2H_2O$	+0,96
$Br_2(\ell) + 2e^-$	$\rightleftharpoons$	$2Br^-$	+1,07
$Pt^{2+} + 2e^-$	$\rightleftharpoons$	Pt	+1,20
$MnO_2 + 4H^+ + 2e^-$	$\rightleftharpoons$	$Mn^{2+} + 2H_2O$	+1,23
$O_2(g) + 4H^+ + 4e^-$	$\rightleftharpoons$	$2H_2O$	+1,23
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightleftharpoons$	$2Cr^{3+} + 7H_2O$	+1,33
$Cl_2(g) + 2e^-$	$\rightleftharpoons$	$2Cl^-$	+1,36
$MnO_4^- + 8H^+ + 5e^-$	$\rightleftharpoons$	$Mn^{2+} + 4H_2O$	+1,51
$H_2O_2 + 2H^+ + 2e^-$	$\rightleftharpoons$	$2H_2O$	+1,77
$Co^{3+} + e^-$	$\rightleftharpoons$	Co^{2+}	+1,81
$F_2(g) + 2e^-$	$\rightleftharpoons$	$2F^-$	+2,87

Oxi: $SO_2 + 2H_2O \rightarrow SO_4^{-2} + 4H^+ + 2e^-$ (x3)

Red: $Cr_2O_7^{-2} + 14H^+ + 6e^- \rightarrow 2Cr^{+3} + 7H_2O$

Balanced Redox:

$3SO_2 + 6H_2O + Cr_2O_7^{-2} + 14H^+ + 6e^- \rightarrow 3SO_4^{-2} + 12H^+ + 6e^- + 2Cr^{+3} + 7H_2O$

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

$2H^+$ $1H_2O$

$3SO_2 + \cancel{6H_2O} + Cr_2O_7^{-2} + \cancel{14H^+} + \cancel{6e^-} \rightarrow 3SO_4^{-2} + \cancel{12H^+} + \cancel{6e^-} + 2Cr^{+3} + \cancel{7}H_2O$

$3SO_2 + Cr_2O_7^{-2} + 2H^+ \rightarrow 3SO_4^{-2\,-} + 2Cr^{+3} + H_2O$

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

1.1 Write down the oxidation number of the Cr in $Cr_2O_7^{2-}$:

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 $Cu_{(s)}$ 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:

$S + 2H^+ + 2e^-$	$\rightleftharpoons$	$H_2S(g)$	+ 0,14
$Sn^{4+} + 2e^-$	$\rightleftharpoons$	Sn^{2+}	+ 0,15
$Cu^{2+} + e^-$	$\rightleftharpoons$	Cu^+	+ 0,16
[illegible]	$\rightleftharpoons$	[illegible]	[illegible]
$Cu^{2+} + 2e^-$	$\rightleftharpoons$	Cu	+ 0,34
[illegible]	$\rightleftharpoons$	[illegible]	[illegible]
$SO_2 + 4H^+ + 4e^-$	$\rightleftharpoons$	$S + 2H_2O$	+ 0,45
$Cu^+ + e^-$	$\rightleftharpoons$	Cu	+ 0,52
$I_2 + 2e^-$	$\rightleftharpoons$	$2I^-$	+ 0,54
$O_2(g) + 2H^+ + 2e^-$	$\rightleftharpoons$	H_2O_2	+ 0,68
[illegible]	$\rightleftharpoons$	[illegible]	[illegible]
$NO_3^- + 2H^+ + e^-$	$\rightleftharpoons$	$NO_2(g) + H_2O$	+ 0,80
$Ag^+ + e^-$	$\rightleftharpoons$	Ag	+ 0,80
[illegible]	$\rightleftharpoons$	[illegible]	[illegible]
$NO_3^- + 4H^+ + 3e^-$	$\rightleftharpoons$	$NO(g) + 2H_2O$	+ 0,96
$Br_2(\ell) + 2e^-$	$\rightleftharpoons$	$2Br^-$	+ 1,07
$Pt^{2+} + 2e^-$	$\rightleftharpoons$	Pt	+ 1,20
$MnO_2 + 4H^+ + 2e^-$	$\rightleftharpoons$	$Mn^{2+} + 2H_2O$	+ 1,23
$O_2(g) + 4H^+ + 4e^-$	$\rightleftharpoons$	$2H_2O$	+ 1,23
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightleftharpoons$	$2Cr^{3+} + 7H_2O$	+ 1,33
$C\ell_2(g) + 2e^-$	$\rightleftharpoons$	$2C\ell^-$	+ 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.

Oxi: $Cu \rightarrow Cu^{+2} + 2e^-$ (x3)

Red: $NO^{3-} + 4H^+ + 3e^- \rightarrow NO + 2H_2O$ (x2)

Redox: $3Cu + 2NO^{3-} + \cancel{6e^-} \rightarrow 3Cu^{+2} + 2NO + 4H_2O + \cancel{6e^-}$

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.

$$MnO_4^-(aq) + H^+(aq) + H_2S(g) \rightarrow Mn^{2+}(aq) + S(s) + H_2O(\ell)$$

Write the

3.1 Oxidation half reaction: $\rightarrow$ + +

3.2 Reduction half reaction: + + $\rightarrow$ +

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):

$+ \quad + \quad \rightarrow \quad + \quad +$

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: $SO_2(g) + H_2S(g) \rightarrow S(s) + H_2O(\ell)$

4.1 oxi: $\rightarrow \quad + \quad +$

4.2 red: $+ \quad + \quad \rightarrow \quad +$

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

$+ \quad \rightarrow \quad +$

Exercise 5

The reaction between dichromate ions ($Cr_2O_7^{-2}$) and iron(II) ions (Fe^{2+}) in an acidic medium is given below.

$$Cr_2O_7^{-2}(aq) + Fe^{2+}(aq) + H^{+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq) + H_2O(\ell)$$

5.1 oxi: $\qquad \rightarrow \qquad + \qquad$

5.2 red: $\qquad + \qquad + \qquad \rightarrow \qquad + \qquad$

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

$$\qquad + \qquad + \qquad \rightarrow \qquad + \qquad + \qquad$$