Chemguide – answers

TRANSITION METALS: IRON

1. a) $N_2 + 3H_2 \rightleftharpoons 2NH_3$

   b) (i) $S_2O_8^{2-} + 2Fe^{2+} \longrightarrow 2SO_4^{2-} + 2Fe^{3+}$

   $2Fe^{3+} + 2I^- \longrightarrow 2Fe^{2+} + I_2$

   (ii) The uncatalysed reaction needs a reaction between two negative ions which would tend to repel each other. The two stages of the catalysed reaction both involve reactions between positive and negative ions. (Well done if you worked this out for yourself – you may of course have met it previously.)

2. a) (i) Very pale green solution gives a green precipitate which darkens on standing and becomes orange around the top.

   (ii) Yellow or orange solution produces an orange (or orange-brown) precipitate.

   b) The initial precipitate of iron(II) hydroxide is oxidised by the air to iron(III) hydroxide.

   c) (i) initially: $[Fe(H_2O)_6]^{2+}$ after addition: $[Fe(H_2O)_4(OH)_2]$

   (ii) initially: $[Fe(H_2O)_6]^{3+}$ after addition: $[Fe(H_2O)_3(OH)_3]$

3. a) (i) Very pale green solution gives a green precipitate

   (ii) Yellow or orange solution produces an orange (or orange-brown) precipitate and bubbles of a colourless gas.

   b) Hexaaquairon(III) ions are more acidic than hexaaquairon(II) ions. The hexaaquairon(III) ions are sufficiently acidic to react with carbonate ions to form carbon dioxide and the precipitate whose formula is in part (c) below. The hexaaquairon(II) ions aren't acidic enough to react with carbonate ions, and so you just get a precipitate of iron(II) carbonate.

   c) $[Fe(H_2O)_3(OH)_3]$

4. iron(III) ions (or better, hexaaquairon(III) ions). The blood red complex is $[Fe(SCN)(H_2O)_5]^{2+}$

5. You need to work out the electron-half-reactions for the things taking part in the reaction. Common to both of them is the oxidation of Fe^{2+} to Fe^{3+}. This is easy and you can just write it straight down.

   $Fe^{2+} \longrightarrow Fe^{3+} + e^-$

   a) Write down what you know. You know that MnO_4^- ions are reduced to Mn^{2+}.

   $MnO_4^- \longrightarrow Mn^{2+}$

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When you have reactions under acidic conditions, you can only add water, hydrogen ions and electrons to balance the half-equation. In this case, the oxygens must end up as water.

$$MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$$

Now balance the hydrogens:

$$MnO_4^- + 8H^+ \longrightarrow Mn^{2+} + 4H_2O$$

. . . and finally the charges:

$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$$

Now combine this with the iron half-equation, making sure that you transfer enough electrons.

$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$$

$$5 \times (Fe^{2+} \longrightarrow Fe^{3+} + e^-)$$

$$MnO_4^- + 8H^+ + 5Fe^{2+} \longrightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$$

Finally, check that everything balances including the charges – it is easy to make a mistake when you add the two equations together.

(Sorry if all this seems totally obvious to you! It isn't totally obvious to a lot of students.)

b) I am not going to repeat all this. The equation you should end up with is:

$$Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \longrightarrow 2Cr^{3+} + 7H_2O + 6Fe^{3+}$$

(If you got this wrong, a common mistake is not to balance the chromiums in the half-equation before you do anything else.)