REDOX - AS91167

TEST YOURSELF 1

Question One

Pale green iron(II) sulfate solution is added drop by drop to acidified potassium permanganate solution. The solution turns from **purple** to **colourless**.

(a) Identify the ion responsible for the **purple** colour

(b) Identify the ion formed as the purple colour goes **colourless**

Question Two.

Chlorine gas, $Cl_2(g)$ is bubbled into a solution containing iron(II) ions, $Fe^{2+}(aq)$. The pale green solution changes to a pale orange colour.

Explain these observations.

Question Three

When sulfur dioxide gas (SO_2) is bubbled into a solution of acidified potassium dichromate solution, a colour change is observed. The *unbalanced* equation for this reaction is given below.

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

(a) Describe the colour change that would be expected when this reaction occurs.

(b) Explain these expected observations by referring to the species involved in the reaction.

Question Four

Complete the balanced half-equation for the reaction

$$Cr_2O_3 \rightarrow CrO_4^{2-}$$

Is this reaction oxidation or reduction?

Question Five

The halogens act as oxidizing agents in reactions. Aqueous chlorine $Cl_2(aq)$ can react with a solution containing iodide ions, $I^-(aq)$. Write balanced half-equations for the oxidation and reduction reactions that occur. Use the half equations to write a balanced equation for the overall oxidation-reduction reaction that occurs.

Question Six

Chlorine is a non-metal that forms compounds with many other elements.

Work out the oxidation number of chlorine in each of the following species.

$$\mathbf{Cl}_2 \quad H\mathbf{Cl} \quad \mathbf{Cl}_2O \quad \mathbf{Cl}O_3^- \quad H\mathbf{Cl}O_4$$

Question Seven

Determine the oxidation number of copper in malachite, $Cu_2(CO_3)(OH)_2$. Show how you worked out the oxidation number.

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Answers to TEST YOURSELF 1

Question One

(a) MnO_4^- / permanganate ion / manganate(VII) ion.

(b) Mn^{2+} / manganese (II) ion. (It's also called the manganous ion)

Question Two

Chlorine $Cl_2(g)$ oxidises the pale green $Fe^{2+}(aq)$ to $Fe^{3+}(aq)$ which is pale orange in solution.

Question Three

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

(a) Colour change from an orange to a green solution.

(b) Orange $Cr_2O_7^{2-}$ is reduced to green Cr^{3+} and the colourless SO_2 is oxidised to colourless sulfate, so over all, the colour is from an orange solution to a green solution.

Question Four

$Cr_2O_3 \rightarrow CrO_4^{2-}$

$Cr_2O_3 \rightarrow 2CrO_4^{2-}$

$5H_2O + Cr_2O_3 \rightarrow 2CrO_4^{2-}$

$5H_2O + Cr_2O_3 \rightarrow 2CrO_4^{2-} + 10H^+$

$5H_2O + Cr_2O_3 \rightarrow 2CrO_4^{2-} + 10H^+ + 6e^-$

This reaction is oxidation (loss of electrons).

Question Five

$2I^- \rightarrow I_2 + 2e^-$ [oxidation]

$Cl_2 + 2e^- \rightarrow 2Cl^-$ [reduction]

Overall: $Cl_2 + 2I^- \rightarrow 2Cl^- + I_2$

Question Six

0	-1	+1	+5	+7
$\mathbf{Cl}_2$	H**Cl**	$\mathbf{Cl}_2O$	$\mathbf{Cl}O_3^-$	H$\mathbf{Cl}O_4$

Question Seven

The charge on the CO_3^{2-} ions is 2- and on OH^- is 1-. Therefore together the anion charge must be 4 -. Since $Cu_2(CO_3)(OH)_2$ is neutral overall so the total cation charge is must be 4+. So each Cu ion has a 2+ charge and therefore an oxidation number of +2.