• Write the two half equations and hence balance the equation for the following redox reaction:

\[
\text{MnO}_2 + \text{NaCl} + \text{H}_2\text{SO}_4 \rightarrow \text{MnSO}_4 + \text{H}_2\text{O} + \text{Cl}_2 + \text{Na}_2\text{SO}_4
\]

### Working

The Mn is being reduced since it begins with oxidation number +4 (in MnO\(_2\)) and ends with oxidation number +2 (in MnSO\(_4\)). The reduction reaction is:

\[
\text{MnO}_2 + 4\text{H}^+ + 2e^- \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O}
\]

The Cl is being oxidised since it begins with oxidation number -1 (in Cl\(^-\)) and ends with oxidation number 0 (in Cl\(_2\)). The oxidation reaction is:

\[
2\text{Cl}^- \rightarrow \text{Cl}_2 + 2e^-
\]

These are then summed together to give the overall reaction.

### Balanced equation:

\[
\text{MnO}_2 + 2\text{Cl}^- + 4\text{H}^+ \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O} + \text{Cl}_2
\]

Which species is oxidised? **Cl\(^-\) is oxidised**

• In the electro-refining of Pt, what mass of Pt is deposited from a solution of PtCl\(_6\)^{2-} in 1.00 hour, by a current of 1.62 A?

PtCl\(_6\)^{2-} contains Pt(IV) and so 4 moles of electrons are required to reduced 1 mol.

The number of moles of electrons delivered by a current of 1.62 A in 1.00 hour is:

\[
\text{number of moles of electrons} = \frac{It}{F} = \frac{(1.62 \text{ C s}^{-1})(1.00 \times 60 \times 60 \text{ s})}{(96485 \text{ C mol}^{-1})} = 0.0604 \text{ mol}
\]

This will deposit:

\[
\text{number of moles of Pt deposit} = \frac{1}{4} \times 0.0604 \text{ mol} = 0.0151 \text{ mol}
\]

As the molar mass of Pt is 195.09 g mol\(^{-1}\), this corresponds to:

\[
\text{mass of Pt} = \text{number of moles} \times \text{molar mass} = (0.0151 \text{ mol}) \times (195.09 \text{ g mol}^{-1}) = 2.95 \text{ g}
\]

Answer: **2.95 g**