## Writing Reduction – Oxidation Half Equations

**Example:** A pale green Fe(II) solution is reacted with a purple $MnO_4^-$ solution. On mixing the purple solution becomes colourless showing that permanganate has reduced to $Mn^{2+}$. $Fe^{2+}$ solution turns orange showing $Fe^{3+}$ has formed.

**1.** Write the reactants and products for one half reaction (oxidation of reduction):

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

**2.** Balance atoms that are not O or H (Mn is already balanced)

**3.** Balance O by adding $H_2O$

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

**4.** Balance the Hydrogens that have appeared with $H^+$

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

**5.** Balance charges by adding negatively charged **electrons** to the **most positive** side

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

(Reduction– ON↓, **Gain of electrons**)

**6.** Follow steps 1-5 for the other half equation;

1. $Fe^{2+} \rightarrow Fe^{3+}$
2. Already balanced
3. No Oxygens
4. No Hydrogens
5. $Fe^{2+} \rightarrow Fe^{3+} + e^-$ (Oxidation– ON↑, **Loss of electrons**)

**7.** Multiply each half equation by an appropriate number so as to give the same number of electrons in each equation ( so that you can cancel them all out in the final equation)

Reduction Half Equation $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

Oxidation Half Equation $(Fe^{2+} \rightarrow Fe^{3+} + e^-)$ x5

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**Balanced Redox Equation:** $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$