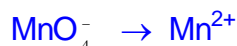


## Writing Reduction – Oxidation Half Equations

**Example:** A pale green Fe(II) solution is reacted with a purple  $\text{MnO}_4^-$  solution. On mixing the purple solution becomes colourless showing that permanganate has reduced to  $\text{Mn}^{2+}$ .  $\text{Fe}^{2+}$  solution turns orange showing  $\text{Fe}^{3+}$  has formed.

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

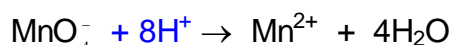


2. Balance atoms that are not O or H (Mn is already balanced)

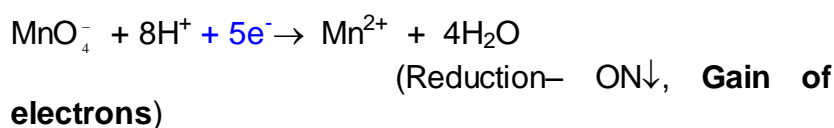
3. Balance O by adding  $\text{H}_2\text{O}$



4. Balance the Hydrogens that have appeared with  $\text{H}^+$



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

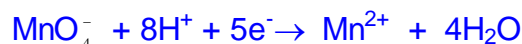


6. Follow steps 1-5 for the other half equation;

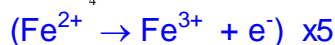
1.  $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+}$
2. Already balanced
3. No Oxygens
4. No Hydrogens
5.  $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$  (Oxidation–  $\text{O} \uparrow$ , **Loss of electrons**)

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



Oxidation Half Equation



**Balanced Redox Equation:**

