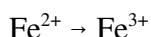


CHEMISTRY 135

BALANCING REDOX EQUATIONS

Writing balanced ionic equations for redox reactions is often rather difficult unless a step-wise approach is taken. This involves writing balanced half-equations and then combining them in such a way as to eliminate electrons. Take the example of the oxidation of $\text{Fe}^{2+}(\text{aq})$ to $\text{Fe}^{3+}(\text{aq})$ by the action of acidified potassium permanganate. The active ion in the potassium permanganate is $\text{MnO}_4^-(\text{aq})$ and it is reduced to Mn^{2+} .

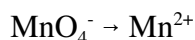
The half equation for the Fe^{2+} is easy to write. We start by writing:



and then balancing the charge in this half equation with electrons:



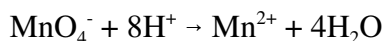
The half equation for the MnO_4^- is a little more difficult. First we write:



which is already balanced as far as the Mn goes, but not anything else. We then balance oxygen by adding water molecules:



We next balance hydrogen by adding hydrogen ions:



Finally we balance charge by adding electrons:



and this is our second balanced half equation.

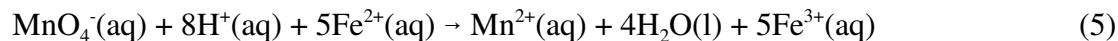
To construct the balanced ionic equation required, we equalise the number of electrons in equations (1) and (2) by multiplying equation (1) by 5, giving:



and then adding this new equation to (2):

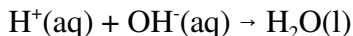


Electrons (and anything else that appears on both sides of the equation) may then be cancelled, giving:

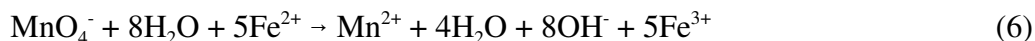


as the final balanced ionic equation. States must be included at this stage.

Note that the presence of hydrogen ions amongst the reactants tells us that the reaction can only proceed in the presence of acid. For a reaction which occurs in alkaline solution, OH⁻ ions must be added to both sides in the same number as the hydrogen ions. For the above equation (5), 8OH⁻ ions would be added to each side. Where OH⁻ and H⁺ ions occur together (on the left in the modified equation (5)) they are combined to give water molecules according to:



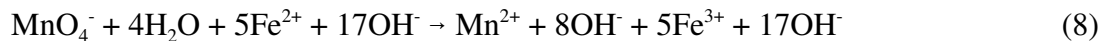
In this way equation (5) becomes:



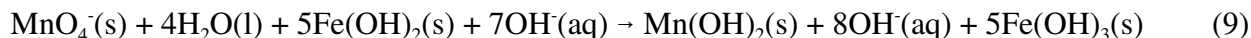
Some further adjustments are required to make a reasonable equation. 4 water molecules should be cancelled to give:



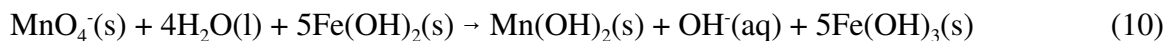
and it must be realised that Mn²⁺, Fe²⁺ and Fe³⁺ will be precipitated as Mn(OH)₂(s), Fe(OH)₂(s) and Fe(OH)₃(s) in alkaline solution, requiring the addition of 17OH⁻ to both sides, giving:



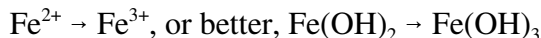
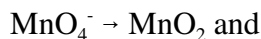
followed by a rearrangement to give:



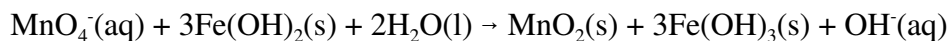
and then a cancellation:



Although equation (10) looks reasonable and is balanced, it may not actually take place to any significant extent. In practice, in alkaline solution MnO₄⁻ is usually reduced to MnO₂, so we should arrive at a more probable result by starting with:



The student is left to complete this exercise by a series of logical steps to give:



Exercises

- 1) a) Write ionic half-equations, including states for:
- the conversion of Fe^{2+} to Fe^{3+} .
 - the conversion of Al to Al^{3+} .
 - the conversion of Cu^{2+} to Cu.
 - the conversion of Cu to Cu^{2+} .
 - the conversion of Fe to Fe^{2+} .
 - the conversion of Ag to Ag^+ .
 - the conversion of Pb to Pb^{2+} .
 - the conversion of Mg to Mg^{2+} .
 - the conversion of NO_3^- in the presence of acid to NO.
 - the conversion of NO_3^- in the presence of acid to NO_2 .
 - the conversion of H_2 to H^+ .
 - the conversion of MnO_4^- in the presence of acid to Mn^{2+} and water.
 - the conversion of $\text{Cr}_2\text{O}_7^{2-}$ in the presence of acid to Cr^{3+} and
 - The addition of metallic aluminium to a copper(II) nitrate solution results in the precipitation of copper and the formation of Al^{3+} in solution.
 - When metallic copper is added to a dilute solution of nitric acid, gaseous NO is evolved and the solution becomes pale blue.
 - Potassium permanganate solution is decolorised by
 - acidified iron(II) sulphate solution.
- n) the conversion of H_2S to S and H^+ .
- o) the conversion of OH^- to O_2 and H_2O .
- b) In each of the above cases write down the appropriate E value.
- c) Classify each of the above half-equations as either an oxidation or a reduction reaction. Rank the oxidation equations in order of reducing power (since an oxidation process has the power to cause reduction elsewhere) and the reduction equations in order of oxidising power. In each case write the strongest first.
- 2) a) Write balanced overall ionic equations for the following reactions. You may combine some of the half equations you have written in question (1).
- b) In each of the above cases calculate the E^\ominus value for the ionic equation from the E^\ominus values for the half-equations. (Look them up in published tables.)