Oxidation numbers can be used to determine whether an oxidation–reduction reaction has taken place. **An increase in oxidation number means that oxidation of the species has occurred, whereas a decrease in oxidation number means that reduction has occurred.** For example, consider the oxidation number of manganese as it changes from MnO_4^- to Mn^{2+}. The oxidation number of manganese in MnO_4^- is +7, but in Mn^{2+} manganese is in oxidation state +2. This means that when MnO_4^- is changed to Mn^{2+} a reduction reaction has taken place. This fits in with the more familiar definition that reduction is a gain of electrons, as shown in the ion–electron equation:

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

It can be seen that manganese in oxidation state +7 has been reduced and so MnO_4^- is acting as an oxidising agent when it reacts in such a manner. It is generally true that compounds containing metals in high oxidation states tend to be oxidising agents whereas compounds with metals in low oxidation states are often reducing agents.

Questions

1. Write an ion–electron equation for Fe^{2+} acting as:
   (i) an oxidising agent
   (ii) a reducing agent.

2. Work out the oxidation number of Cr in $Cr_2O_7^{2-}$ and decide whether the conversion of $Cr_2O_7^{2-}$ to Cr^{3+} is oxidation or reduction. Is the $Cr_2O_7^{2-}$ acting as an oxidising agent or a reducing agent in this reaction? Confirm your answer by writing the appropriate ion–electron equation.

3. Work out the oxidation number of chromium in $Cr_2O_7^{2-}$ and in CrO_4^{2-} and decide whether the conversion of $Cr_2O_7^{2-}$ to CrO_4^{2-} is oxidation or reduction.

4. The most common oxidation states of iron are +2 and +3. Using orbital box notation, draw out their respective electronic configurations and suggest which of the two ions is the more stable.

Page 57

Question 1

Write an ion electron equation for Fe^{2+} acting as:
(i) an oxidising agent (ii) a reducing agent.

Answers

(i) $Fe^{2+} + 2e^{-} \rightarrow Fe$ (ii) $Fe^{2+} \rightarrow Fe^{3+} + e^{-}$

Question 2

Work out the oxidation number of Cr in $Cr_2O_7^{2-}$ and decide whether the conversion of $Cr_2O_7^{2-}$ to Cr^{3+} is oxidation or reduction. Is the $Cr_2O_7^{2-}$ acting as an oxidising agent or a reducing agent in this reaction? Confirm your answer by writing the appropriate ion–electron equation.

Answer

In $Cr_2O_7^{2-}$, Cr has oxidation number +6 and so the conversion of $Cr_2O_7^{2-}$ to Cr^{3+} is reduction since there is a decrease in oxidation number. The $Cr_2O_7^{2-}$ is acting as an oxidising agent. The appropriate ion–electron equation is

$$Cr_2O_7^{2-} + 14H^{+} + 6e^{-} \rightarrow 2Cr^{3+} + 7H_2O$$

Question 3

Work out the oxidation number of chromium in $Cr_2O_7^{2-}$ and in CrO_4^{2-} and decide whether the conversion of $Cr_2O_7^{2-}$ to CrO_4^{2-} is oxidation or reduction.

Answer

Since the oxidation number of chromium in both $Cr_2O_7^{2-}$ and CrO_4^{2-} is +6 then the conversion of $Cr_2O_7^{2-}$ to CrO_4^{2-} is neither oxidation nor reduction.

Question 4

The most common oxidation states of iron are +2 and +3. Using orbital box notation, draw out their respective electronic configurations and suggest which of the two ions is the more stable.

Fe^{2+}	$[Ar]3d^6$	$\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$
Fe^{3+}	$[Ar]3d^5$	$\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

Answer

Fe^{3+} will be the more stable since there is a special stability associated with all the d orbitals being half filled. In Fe^{2+} only four of the d orbitals are half filled and the other is completely filled.

Naming complex ions

Complex ions and complexes are written and named according to IUPAC rules.

The formula of a complex ion should be enclosed within square brackets, although common complexes such as MnO_4^- are often written without brackets. The metal symbol is written first, then the negative ligands followed by the neutral ligands, e.g. $[Fe(OH)_2(H_2O)_4]^+$.

When naming the complex ion or molecule the ligands should be named first, in alphabetical order, followed by the name of the metal. If the ligand is a negative ion the name of which ends in -ide, the ending changes to 'o', e.g. chloride, Cl^-, becomes chloro, cyanide, CN^-, becomes cyano.

If the ligand is a negative ion the name of which ends in -ite, the final 'e' changes to 'o', e.g. nitrite, NO_2^-, changes to nitrito.

If the ligand is water it is named aqua, ammonia is named ammine and carbon monoxide is carbonyl.

If there is more than one particular ligand it is prefixed by di, tri, tetra, penta, or hexa, etc. as appropriate.

If the complex ion is an anion (negative ion) the suffix -ate is added to the name of the metal. Sometimes the Latin name for the metal is used in this context, e.g. ferrate not ironate and cuprate rather than copperate.

The oxidation state of the metal is given in Roman numerals after its name.

For example, the complex ion $[Ni(NH_3)_6]^{2+}$ is named hexaamminenickel(II) and the negative complex ion $[Fe(CN)_6]^{3-}$ is named hexacyanoferrate(III).

Question

Name the following complexes:

(a) $[CoCl_4]^{2-}$
(b) $[Ni(H_2O)_6]^{2+}$
(c) $[Fe(CN)_6]^{4-}$
(d) $[Ti(NH_3)_6]^{3+}$
(e) $[Ni(CN)_6]^{4-}$
(f) MnO_4^-
(g) $[PtCl_6]^{2-}$
(h) $Ni(CO)_4$
(i) $[Cu(NH_3)_4]^{2+}$

Page 61

Name the following complexes:

(a)	$[CoCl_4]^{2-}$	tetrachlorocobaltate(II)
(b)	$[Ni(H_2O)_6]^{2+}$	hexaaquanickel(II)
(c)	$[Fe(CN)_6]^{4-}$	hexacyanoferrate(II)
(d)	$[Ti(NH_3)_6]^{3+}$	hexaamminetitanium(III)
(e)	$[Ni(CN)_6]^{4-}$	hexacyanonickelate(II)
(f)	MnO_4^-	tetraoxomanganate(VII)
(g)	$[PtCl_6]^{2-}$	hexachloroplatinate(IV)
(h)	$Ni(CO)_4$	tetracarbonylnickel(0)
(i)	$[Cu(NH_3)_4]^{2+}$	tetraamminecopper(II)