REDOX REACTIONS

Redox

occurs simultaneously

Comproporpination

$Ag^{2+} + Ag \rightarrow 2Ag^{+}$

(both reduced and oxidized)

Disproportionation

$2Cu^{+} \rightarrow Cu^{2+} + Cu$

(middle ground)

Electron transfer

half equation [Oxidation (loss) OIL
reduction (gain) RIG

$Na \rightarrow Na^{+} + e^{-}$ oxidation

$O_2 + 4e^{-} \rightarrow 2O^{2-}$ reduction

$4Na \rightarrow 4Na^{+} + 4e^{-}$

$4Na + O_2 + 4e^{-} \rightarrow 4Na^{+} + 4e^{-} + 2O^{2-}$

$4Na + O_2 \rightarrow 2Na_2O$

	Oxidation	Reduction
Hydrogen transfer	Loss	Gain
Oxygen transfer	Gain	Loss
Electron transfer	Loss	Gain

→ always applicable

Oxidation

$4Al + 3O_2 \rightarrow 2Al_2O_3$ gains oxygen (-) of electron

↑ oxidized

- gain of oxygen
- loss of hydrogen
- loss of e^-

oxidising agent

- oxidises others but is in turn reduced
- usually has higher oxidation numbers

KI
H_2O_2 K
NH_3
C
CO
H_2
H_2S
metals
SO_2

KI (aq) test

$2I^-(aq) \rightarrow I_2(aq) + 2e^-$

- Added to oxidising agent
- brown sol is formed when I^- is oxidised to I_2

Reduction

$2Al_2O_3 \rightarrow 4Al + 3O_2$ loss of O_2

$Ag^+(aq) + e^- \rightarrow Ag(s)$ (+) gain of electrons

$KMnO_4$
$K_2Cr_2O_7$
H_2O_2 ↖
O_2
HNO_3
H_2SO_4
Br_2
Cl_2
I_2

- loss of oxygen
- gain in hydrogen
- loss of e^-

reducing agent

- reduces others but is in turn oxidised
- usually has lower oxidation numbers

$K_2Cr_2O_7$ (aq) acidified

1) Add dilute H_2SO_4 acid

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

$KMnO_4$ (aq) acidified

$MnO_4^-(aq) + 5e^- + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l)$ (colourless)

Oxidation States

1) Oxidation state of atoms/monoatomic ion is equal to its charge

OS of H in H_2 = 0

2) Charge of a molecule / compound is equal to sum of oxidation states that make up the identity

$[Cu(H_2O)_6]^{2+}$

OS of Cu → +2

3) Some elements have a fixed oxidation state

▲ fractions are possible

Grp 1	(+1)
Grp 2	(+2)
F	(-1)
H	(+1) Except NaH (-1)
O	(-2) Except peroxides H_2O_2 (-1) / OF_2 (+2)
Cl	(-1) Except w O/F (+1)

↑ oxidation # = tend to oxidise others

↓ oxidation # reduces others

oxidation

OS ↑

reduction

OS ↓

H_2SO_4 $x + 4(-2) = -2$, $x = +6$

$Cr_2O_7^{2-}$ $2x + 7(-2) = -2$, $x = +6$

Halogen displacement

$Cl_2(g) + 2NaBr(aq) \rightarrow 2NaCl(aq) + Br(aq)$

0 → -1 (reduced); -1 → 0 (oxidised)

Cl is more electronegative (more likely to gain e^-)

Balancing half equations

1) # of atoms
2) charge

ionic equation

1) conservation of mass
   - # of atoms on LHS = RHS
2) Balance of charge
   - sum of charges on LHS = RHS
3) "Add e^-"
4) Add H^+ / OH^- depending on medium
5) Add H_2O to balance
6) Add state symbols

↳ Some half equations don't need to be balanced in charged medium

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

acidic medium

1) Work out change in oxidation state

$Cr_2O_7^{2-} \rightarrow Cr^{3+}$ (+6) (+3)

2) Balance amt of element so charge matches

$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$

3) Add e^-

$Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+}$ (oxygen added to)

4) Add H^+

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

5) Add H_2O

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

6) State symbols

$Cr_2O_7^{2-}(aq) + 6e^- + 14H^+(aq) \rightarrow 2Cr^{3+}(aq) + H_2O(l)$

alkaline medium

$ClO^- \rightarrow Cl^-$ (+1) (-1) → match Cl charge

$ClO^- + 2e^- \rightarrow Cl^-$ (-1) (-2) (-1) → balance total charge

$ClO^- + 2e^- \rightarrow Cl^- + 2OH^-$ (-1) (-2) (-1) (-2) → balance equation

$ClO^-(aq) + 2e^- + H_2O(l) \rightarrow Cl^-(aq) + 2OH^-(aq)$

Combining

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

$SO_3^{2-} + 2OH^- \rightarrow SO_4^{2-} + 2e^- + H_2O$

LCM: 2 x 5 = 10 → make coeff of e^- the same as gain in e^- = loss in e^-

$2MnO_4^- + 10e^- + 16H^+ \rightarrow 2Mn^{2+} + 8H_2O$

$5SO_3^{2-} + 10OH^- \rightarrow 5SO_4^{2-} + 10e^- + 5H_2O$

$2MnO_4^- + 10e^- + 16H^+ + 5SO_3^{2-} + 10OH^- \rightarrow 2Mn^{2+} + 13H_2O + 5SO_4^{2-} + 10e^-$

combine to H_2O

$2MnO_4^-(aq) + 6H^+(aq) + 5SO_4^{2-}(aq) \rightarrow 2Mn^{2+}(aq) + 3H_2O(l) + 5SO_4^{2-}(aq)$