Reaction of Chlorine with Water and NaOH

Disproportionation

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Reaction of chlorine with water: DISPROPORTIONATION

Cl_2 reacts with water.

$$Cl_2\ (g) + H_2O\ (l) \rightleftharpoons HCl\ (aq) + HClO\ (aq)$$

0 → -1 (reduction)

0 → +1 (oxidation)

This equilibrium mixture is called chlorine water

Chlorine is both simultaneously reducing and oxidising. Disproportionation is the name for the reaction where an element simultaneously oxidises and reduces.

If some universal indicator is added to the solution it will first turn red due to the acidity of both reaction products. It will then turn colourless as the HClO bleaches the colour

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Reaction with water in sunlight

If the chlorine is bubble through water in the **presence of bright sunlight** a different reaction occurs

$$2Cl_2 + 2H_2O \rightarrow 4H^+ + 4Cl^- + O_2$$

The same reaction occurs to the equilibrium mixture of chlorine water. The greenish colour of chlorine water fades as the Cl_2 reacts and a colourless gas (O_2) is produced

Chlorine is used in water treatment to kill bacteria. It has been used to treat drinking water and the water in swimming pools

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Reaction of chlorine with NaOH: DISPROPORTIONATION

Reaction with cold alkali:

Cl_2 (*and* Br_2, I_2) in aqueous solutions will react with cold sodium hydroxide. The colour of the halogen solution will fade to colourless

$$Cl_2\,(aq) + 2\,NaOH(aq) \rightarrow NaCl\,(aq) + NaClO(aq) + H_2O\,(l)$$

0 → -1 (reduction)

0 → +1 (oxidation)

Chlorine is both reducing and oxidising. Disproportionation is the name for the reaction where an element simultaneously oxidises and reduces.

The mixture of NaCl and NaClO is used as Bleach

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Variable oxidation numbers of halogens

Fluorine: Can only have two different oxidation numbers, F_2 = 0 , and as a fluoride ion F^- = -1

Chlorine, Bromine and Iodine can form several oxidation states.

Oxidation number chart for the halogens

The minimum ox. no. for the halogens is –1. This is achieved by gaining one electron to complete the outer shell. Halogen atoms would never gain two electrons.

The maximum ox. no. is +7. This is achieved by losing 7 electrons (in theory). Halogen atoms would never lose more than 7 electrons.

Oxidation number chart for the halogens

+7 : $KClO_4$ potassium chlorate (VII)
+5 : $KClO_3$ potassium chlorate (V)
+3 : $KClO_2$ potassium chlorate (III)
+1 : KClO potassium chlorate (I)
0 : Cl_2 chlorine
-1 : KCl potassium chloride

When the halogen has a high oxidation number, they have a tendency to reduce easily to lower oxidation states. All the halates from (I up to VII) can act as oxidising agents.

Note: All halide and halate salts with group 1 or 2 metals are white crystalline solids.

Ionic equation for reaction with Br_2 and cold NaOH:

Br_2 (aq) + $2OH^-$ (aq) → Br^- (aq) + BrO^- (aq) + H_2O (l)

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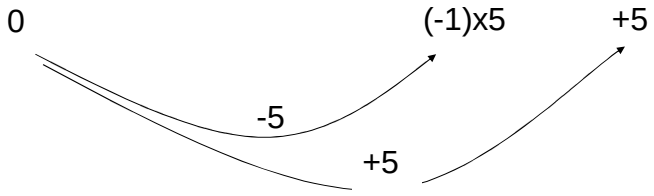
Reaction with hot alkali.

With hot alkali disproportionation also occurs but the halogen that is oxidised goes to a higher oxidation state.

$3Cl_2$ (aq) + 6 NaOH(aq) → 5 NaCl (aq) + $NaClO_3$ (aq) + $3H_2O$ (l)



Iodine & Bromine do a similar reaction

$3I_2$ (aq) + $6OH^-$ (aq) → 5 I^- (aq) + IO_3^- (aq) + $3H_2O$ (l)

extra

Oxidising power of the halates

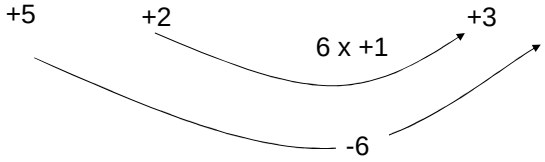
The halates tend to reduce when they react, and are therefore good oxiding agents. The higher the oxidation number of the halate, the stonger the oxidising power.

<u>Reacting the halates with Fe^{2+} solutions.</u>

ClO^- (aq)+ $2Fe^{2+}$(aq) +$2H^+$(aq)→ $2Fe^{3+}$ (aq) +Cl^-(aq) +H_2O(l)

observation: change from green (iron II) to brown (iron III)

$BrO_3^-{}_{(aq)}$ + $6Fe^{2+}{}_{(aq)}$ + $6H^+{}_{(aq)}$ → $6Fe^{3+}{}_{(a\ q)}$ + $Br^-{}_{(aq)}$ + $3H_2O$ $_{(l)}$



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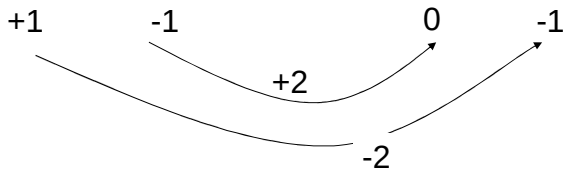
Reacting the halates with (iodide) I^- solutions.

$ClO^-(aq) + 2I^-(aq) + 2H^+(aq) \rightarrow I_2(aq) + Cl^-(aq) + H_2O(l)$



$BrO_3^-(aq) + 6I^-(aq) + 6H^+(aq) \rightarrow 3I_2(aq) + Br^-(aq) + 3H_2O(l)$

$IO_3^-(aq) + 5I^-(aq) + 6H^+(aq) \rightarrow 3I_2(aq) + 3H_2O(l)$

extra