Chem Factsheet

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Chlorine and Chlorates

Chlorine	Chloride	Chlorate (I)
A group seven element. It exists as a diatomic molecule (Cl_2). It is a pale green gas at room temperature and pressure. It can be bubbled through water to produce chlorine water.	The monoatomic ion of chlorine. Forms through one atom of chlorine gaining one electron to produce Cl^-. Chloride ions are often found as common salts such as NaCl and KCl. These are formed through ionic bonding.	Chlorate (I) ion is a polyatomic ion consisting of oxygen and chlorine, ClO^-. The (I) included in the name represents the positive oxidation state of chlorine in this compound. i.e. Na(+1) + Cl(+1) + O(-2) = 0

Exam Hint:- *Be very careful when answering questions involving the species above e.g.*
- ***Chlorine*** *is an example of a halogen.*
- *One* ***Chloride*** *ion reacts with one silver ion to produce silver* ***chloride****, (AgCl).*
- *A* ***chlorate (I)*** *ion combines with a sodium ion to produce sodium* ***chlorate (I)****, (NaClO).*

If you do not use the correct terms you will not get the marks!

Chlorine as an Oxidising Agent

Chlorine is a strong oxidising agent.

An Oxidising agent is *a reagent that takes electrons (oxidises) from another species*

e.g. $Cl_2(aq) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(aq)$

The equation above can also be written as an **ionic** equation by removing the spectator ions.

i.e. $Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$

A spectator ion is *an ion that is not involved in a chemical reaction. It appears unchanged on the reactant and the product side*. Here the potassium ion (K^+) is the spectator ion.

Tip: This reaction is based on the general principle that a more reactive halogen will displace a less reactive halogen from its halide.

Uses of chlorine:
1. Bleach – a substance that removes colour by oxidising "coloured" molecules to colourless molecules.
2. To purify drinking water by killing bacteria by an oxidation process.
3. Killing bacteria in swimming pools.

The Reaction of Chlorine with Sodium Hydroxide

The compounds produced are dependent on the temperature and the concentration of the NaOH.

CONDITIONS 1:
Reagents: Dilute aqueous NaOH and chlorine gas.

Temperature: Room temperature (ca. 20°C)

The overall equation is:
$Cl_2(g) + 2NaOH(aq) \rightarrow NaCl(s) + NaClO(aq) + H_2O(l)$
The products are sodium chloride and sodium chlorate(I)

Omitting the sodium ions gives the following ionic equation for this reaction: $Cl_2 + 2OH^- \rightarrow Cl^- + ClO^- + H_2O$

This reaction is used industrially to produce bleach.

The reaction is an example of a **disproportionation** reaction.

A disproportionation reaction is *a reaction where the same species is simultaneously oxidised and reduced in the same reaction.*

In this reaction two chlorine atoms have undergone redox reactions.

One chlorine atom has been oxidised by losing one electron to form the chlorate(I) ion. $\frac{1}{2}Cl_2 + 2OH^- \rightarrow ClO^- + H_2O + e^-$

One chlorine atom has been reduced by gaining one electron to form a chloride ion. $\frac{1}{2}Cl_2 + e^- \rightarrow Cl^-$

CONDITIONS 2:
Reagents: Concentrated NaOH and chlorine gas

Temperature: Hot.

The overall equation is:

$3Cl_2(g) + 6NaOH(aq) \rightarrow NaClO_3(aq) + 5NaCl(aq) + 3H_2O(l)$

This again is an example of a **disproportionation** reaction.

Chlorines atom have been oxidised by losing electrons to form chlorate(V) ions.

$\frac{1}{2}Cl_2 + 6OH^- \rightarrow ClO_3^- + 3H_2O + 5e^-$

Chlorine atoms have been reduced by gaining electrons to form chloride ions. $\frac{1}{2}Cl_2 + e^- \rightarrow Cl^-$

The Purification of Drinking Water

Reagents: Chlorine gas and water

Temperature: Room temperature

$Cl_2(g) + H_2O(l) \rightleftharpoons HCl(aq) + HClO(aq)$ ———— (1)

This again is an example of a disproportionation reaction.

One Cl atom has been reduced to produce a Cl^- ion : $\frac{1}{2}Cl_2 + e^- \rightarrow Cl^-$

One Cl atom has been oxidised to chloric(I) acid (HClO) :
$\frac{1}{2}Cl_2 + H_2O \rightarrow HClO + H^+ + e^-$

The chloric (I) acid (HClO) slowly decomposes, the reactive oxygen atoms ([O]) killing the bacteria in water. $HClO \rightarrow HCl + [O]$

Chlorine in an Organic Solvent

Aqueous chlorine is a very pale green solution because of the presence of Cl_2 molecules in equilibrium (1). If it is added to an organic solvent (usually cyclohexane) two layers form, an inorganic layer and an organic layer. Due to the non polar nature of the chlorine molecules it dissolves more readily in the organic layer producing a more distinct pale yellow colour, and reducing the pale green aqueous colour

Bromine and iodine produce similarly enhanced colours when they preferentially dissolve in an organic solvent. Bromine goes from orange in aqueous to red in organic. Iodine changes from brown to purple.

Chloride Ion

A chloride ion is formed by one chlorine atom gaining one electron.

This reduction half equation can be written as: $\frac{1}{2}Cl_2 + e^- \rightarrow Cl^-$

A chloride ion can be identified using silver nitrate and nitric acid. Adding $AgNO_3$ and HNO_3 to a solution containing Cl^- produces a white precipitate.

$XCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + XNO_3(aq)$ where X = Na, K etc

The ionic equation for this reaction is: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

The silver chloride is the white precipitate that is seen.

The nitric acid is essential to remove any OH^- or CO_3^{2-} ions from the test solution. These produce interfering precipitates if they are not removed.

Practice Questions

1. Three unknown sodium halide solutions were found in a laboratory. The chemist needed to find out which solution contained chloride ions.
   (a) Describe a simple test the chemist could carry out to identify the solution containing the chloride ions.
   (b) Write the full equation of the reaction that has occurred in this test with state symbols.
   (c) Write the ionic equation that has occurred when performing this test with state symbols.

2. The introduction of chlorine to drinking water has meant that such diseases as cholera and other water-borne diseases have been minimised in such places as the United States and the United Kingdom.
   (a) Why is chlorine added to drinking water?
   (b) Write an equation to show what process has taken place when chlorine is added to drinking water.
   (c) State what type of reaction is occurring in the above reaction.
   (d) State what is being oxidised and reduced in terms of electron transfer.

3. (a) State what is meant by the term disproportionation reaction.
   (b) The reaction of chlorine with cold dilute aqueous sodium hydroxide is an example of a disproportionation reaction. Write a full equation including states symbols for this reaction.
   (c) Show, in terms of electron transfer, how this reaction is an example of a disproportionation reaction.
   (d) State a use for one of the products for the above reaction.
   (e) Another reaction occurs when chlorine is bubbled through hot concentrated sodium hydroxide. Write a full equation including state symbols for this reaction.
   (f) Write the chemical formula for the compound sodium chlorate (VII).

4. A solution of chlorine water was added to a solution of potassium bromide. The resultant mixture was shaken after an organic solvent had been added. A reaction occurs.
   (a) State what would be seen at the end of the experiment.
   (b) Write a full equation including state symbols for the reaction that takes place.
   (c) Write an ionic equation for the reaction that takes place.
   (d) What type of reaction has occurred?
   (e) Show what has been oxidised and reduced in the above equation.

Answers

1. (a) Acidify the sample with dilute nitric acid and then add a few drops of silver nitrate solution. Chloride is confirmed by a white precipitate.
   (b) $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
   (c) $Ag^+(aq) + Cl^- \rightarrow AgCl(s)$

2. (a) To kill bacteria to aid purification.
   (b) $Cl_2(g) + H_2O(l) \rightleftharpoons HCl(aq) + HClO(aq)$
   (c) Disproportionation.
   (d) Cl_2 reduced to Cl^- by gain of electrons ; Cl_2 oxidised to HClO by loss of electrons.

3. (a) A reaction where the same species is simultaneously oxidised and reduced in the same reaction.
   (b) $Cl_2(g) + 2NaOH(aq) \rightarrow NaCl(s) + NaClO(aq) + H_2O(l)$
   (c) Oxidation of chlorine : $\frac{1}{2}Cl_2 + 2OH^- \rightarrow ClO^- + H_2O + e-$
   Reduction of chlorine : $\frac{1}{2}Cl_2 + e^- \rightarrow Cl^-$
   (d) Bleach.
   (e) $3Cl_2(g) + 6NaOH(aq) \rightarrow NaClO_3(aq) + 5NaCl(aq) + 3H_2O(l)$
   (f) $NaClO_4$

4. (a) Two layers. An almost colourless aqueous layer and a red organic layer.
   (b) $2KBr(aq) + Cl_2(aq) \rightarrow Br_2(aq) + 2KCl(aq)$
   (c) $2Br^-(aq) + Cl_2(aq) \rightarrow Br_2(aq) + 2Cl^-(aq)$
   (d) Redox
   (e) $Br^-(aq)$ is oxidised to Br_2 by loss of electrons. $Cl_2(aq)$ is reduced to $Cl^-(aq)$ by gain of electrons.

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