NAME ................................... Chemistry Class ...............................



The oxidation states of chlorine

Halogens

answers

**Topic 4B: The elements of Group 7 (halogens)**

9. understand reasons for the trends in melting and boiling temperatures,

Physical state at room temperature, and electronegativity for Group 7 elements

10. understand reasons for the trend in reactivity of Group 7 elements down the

group

11. understand the trend in reactivity of Group 7 elements in terms of the redox

reactions of Cl2, Br2 and I2 with halide ions in aqueous solution, followed by the

addition of an organic solvent

12. understand, in terms of changes in oxidation number*,* the following reactions of

the halogens:

i oxidation reactions with Group 1 and 2 metals

ii the disproportionation reaction of chlorine with water and the use of chlorine

in water treatment

iii the disproportionation reaction of chlorine with cold, dilute aqueous sodium

hydroxide to form bleach

iv the disproportionation reaction of chlorine with hot alkali

v reactions analogous to those specified above

13. understand the following reactions:

i solid Group 1 halides with concentrated sulfuric acid, to illustrate the trend in

reducing ability of the hydrogen halides

ii precipitation reactions of the aqueous anions Cl–, Br– and I– with aqueous

silver nitrate solution, followed by aqueous ammonia solution

iii hydrogen halides with ammonia and with water (to produce acids)

14. be able to make predictions about fluorine and astatine and their compounds, in

terms of knowledge of trends in halogen chemistry

**Topic 4C: Analysis of inorganic compounds**

15. know reactions, including ionic equations where appropriate, for identifying:

i carbonate ions, and hydrogencarbonate ions, using an aqueous

acid to form carbon dioxide

ii sulfate ions, using acidified barium chloride solution

iii ammonium ions, using sodium hydroxide solution and warming to form

ammonia









**The Halogens**

Facer Chapter 11 p217 – 230

**Department Website Factsheets**

|  |  |
| --- | --- |
| **14** | **Group 7** |
| 88 | Disproportionation |
| *111* | *Oxyacids Extension* |

**Websites**

<http://www.chemguide.co.uk/inorganic/group7menu.html#top>

A thorough coverage of the topic for AS and A level

<http://www.chemtopics.com/elements/halogen/halogen.htm>

Contains information on the properties of the elements and pictures together with information about their discovery and basic physical and chemical data.

<http://www.rsc.org/chemsoc/visualelements/pages/data/intro_groupvii_data.html>

Basic information on properties and reactivity

In this pack we will study various aspects of Halogen chemistry

Physical properties of halogens and solubility in water and organic solvents

 Oxidation reactions

 Disproportionation reactions

Reaction of hydrogen halides with ammonia and water

 Reactions of halides with conc. sulfuric acid

 Tests for halide ions and other anions

**Halogens**

The elements in group 7 of the periodic table are known as the [Halogens](http://en.wikipedia.org/wiki/Halogen) They are highly reactive and only exist as compounds in nature. They are able to **gain an electron** to form ions with a 1- charge and oxidation number -1 called the **halide ions**, but, as we shall see later they are also able to form more complex ‘**oxo’ ions** with various other oxidation numbers.

In addition, they form a variety of important **covalently bonded** chemical compounds including:-



Mustard gas

(Used during WW1)





Halothane

(anaesthetic)

Valium

(drug)



**Tetrabromobisphenol A (TBBPA)**

**Flame retardant**



See the articles on [Bromine](file:///H%3A%5Ctopic%202%5CHalogens%5CResources%5CBromine.pdf) and [Chlorine](file:///H%3A%5Ctopic%202%5CHalogens%5CResources%5Cchlorine.pdf)

**The Halogens – Physical Properties**

Complete the following table for the halogen elements.

Teacher demonstration & Chemistry Set DVD videos

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | **Fluorine** | **Chlorine** | **Bromine** | **Iodine** |
| Formula of molecule | F2 | Cl2 | Br2 | I2 |
| Appearance and state at room temp | Pale yellow gas | Green/yellow gas | Dark red/brown liquid | Shiny grey solid |
| Electron configuration | 1s22s22p5 | 1s22s22p63s23p5 | [Ar]4s23d104p5 | [Kr]5s24d105p5 |
| Bleaching ability  | Very strong | Strong | Weak | None |
| Test for the gas |  | Damp blue litmus paperTurns red then bleachedOr Damp starch iodide paperTurns blue | Damp starch + iodide paperTurns blueBubble through KI solution with starch | Starch solution turns blue/black |
| Solubility in water and appearance of solution | Reacts with water to form HF and O2 | Slight solubility but also reacts to form chlorine water HCl and HOClPale green soln. | Slight solubility but also reacts to form bromine water HBr and HOBrOrange solution | Almost insoluble producing a very pale brown solution.Soluble in KI solution 🡪 dark brown solution. |
| Trend in solubility in H2O of Cl2🡪I2 | X | Decreasing solubility 🡪 |
| Solubility in cyclohexane and appearance of solution |  | SolubleYellow/green solution | Solublered solution | Solublepurple solution |

**Solubility**

The solubility of the halogens decreases down the group. Iodine is almost insoluble in water, but is soluble in KI due to the reaction:-

I2(s) + I-(aq) **→** I3- (aq)

State, and explain the change in melting pts down .group 7

Increases, as the number of electrons increases the London forces also increase (IMF) requiring more energy to overcome.

State and explain the change in electronegativity down group 7.

Decreases, outer electrons are progressively more shielded from the nucleus as the atomic radius increases, also the distance between the bonding pair and nucleus increaces. (proton number increaes but this is more than offset by the other factors) So electrons in a covalent bond are attracted less to the halogen.

State and explain the change in reactivity of the Group 7 elements down group 7

As most reactions with halogens result in the halogen gaining an electron. They become less reactive on descending. Reactivity decreases as atomic radius increases and more shielding means electrons attracted less strongly to the nucleus.

What is the problem in distinguishing between bromine and iodine solutions?

Depending on concentration both could appear a brown/orange colour.

How could you show that you have a solution of iodine? Give two methods.

* Add an organic solvent and shake. Halogens are more soluble in organic solvents, Iodine 🡪 purple colour in upper, organic layer, bromine 🡪 red organic layer
* Add starch solution and shake. Iodine 🡪 blue / black.

One of these methods could be used to separate bromine or iodine from aqueous solution:-

*Clarify by colouring the layers with* ***appropriate colours***

Shake and allow to settle

Add an organic solvent e.g. hexane

Separating funnel

Iodine moves into hexane layer

aqueous solution of iodine

Open tap and run off aqueous layer and discard

**Chemical reactions of the Halogens - with metals**

Most metals react with halogens to form a compound with ionic bonds.

**e.g The reaction of magnesium with chlorine**

Equation:.. Mg(s) + Cl2(g) 🡪 MgCl2(s)

What has been oxidised? Magnesium What has been reduced? chlorine

**Demonstration expt. to show the reactions of Aluminium with bromine and with iodine.**

(Classic Chemistry Experiment 77)

Chlorine, bromine and iodine are strong enough oxidising agents to oxidise Al to Al3+.

Describe your observations and write half ionic equations for the oxidation and reduction.

|  |  |  |
| --- | --- | --- |
|  | **aluminium + bromine** | **aluminium + iodine** |
| Observations | Bromine liquid dropped on Aluminium foil. Sparks and glowed redGot hot and a lot of bromine vapour evolved. A white solid remained. | Vigorous reaction, clouds of purple I2vapour.Solid glowed red-hotGrey solid AlI3 remained |
| Oxidation equation | Al(s) 🡪 Al3+ + 3e- | Al(s) 🡪 Al3+ + 3e- |
| Reduction equation | Br2 + 2e- 🡪 2Br- | I2 + 2e- 🡪 2I- |

****

**Reaction of halogens with iron(II) ions in solution**

We have seen in the redox pack that halogens are good oxidising agents as they readily **accept** electrons. In this experiment you will explore the ability of chlorine, bromine and iodine to act as oxidising agents. The Fe2+ can be oxidised to Fe3+, and the halogens can be reduced to halide ions.

* Place 2 cm3 of Fe2+ into a test tube and add aqueous chlorine until a change is observed.
* Record your observations in the table
* Repeat the experiment using aqueous bromine and then aqueous iodine

**NOTE: Fe2+ ions are pale GREEN. Fe3+ ions are RUST BROWN –** turn red with KSCN

|  |  |  |  |
| --- | --- | --- | --- |
|  | **Cl2 (aq)** | **Br2 (aq)** | **I2 (aq)** |
| Initial colour of halogen solution | Very pale green | orange | pale brown |
| Observation on adding Fe2+  | Solution became very pale yellow(thiocyanate 🡪 red ∴Fe3+ present) | Solution became very pale yellow(thiocyanate 🡪 red ∴Fe3+ present) | No change(thiocyanate 🡪 unchanged - no Fe3+ ) |
| Reduction half ionic equation(if any reaction) | Cl2 + 2e- 🡪 2Cl- | Br2 + 2e- 🡪 2Br- | None |
| Oxidation half ionic equation (if any reaction) | Fe2+🡪 Fe3+ + e- | Fe2+🡪 Fe3+ + e- | None |
| Overall Redox equation (if any) | Cl2 + 2Fe2+🡪  2Fe3+ + 2Cl- | Br2 + 2Fe2+🡪  2Fe3+ + 2Br- | None |

Additional test for Fe3+  = add potassium thiocyanate (KSCN) a red colour indicates Fe3+

**Redox reactions of the halogens and halides**

****

In this experiment you will explore the ability of halogens to oxidise one another, this will enable you to place them in order of oxidising ability.

Method

* Note the initial appearance of the solutions in Tables 1 and 2.
* Add chlorine water dropwise to 1 cm3 of potassium bromide solution in a test tube and mix, (the potassium ions are spectator ions).
* If you have any difficulty interpreting what is happening, add 1 cm3 of an organic solvent e.g. hexane or cyclohexane and shake.
* Repeat the experiment with chlorine water and potassium iodide.
* Record your observations in the Table 3 (next page).
* Replace chlorine water with bromine water and repeat experiment as indicated in Table 3.

Table 1 – Initial appearance of the halogen element in solution

|  |  |
| --- | --- |
| **Halogen element** in solution | Appearance |
| Chlorine water | Very pale green – almost colourless |
| Bromine water | Orange |
| Iodine (in KI) solution | orange / brown |

Table 2 – Initial appearance of potassium halide salts in aqueous solution

|  |  |
| --- | --- |
| **Halide salt** in solution | Appearance |
| Potassium chloride | Colourless solution |
| Potassium bromide | Colourless solution |
| Potassium iodide | Colourless solution |

**Table 3 Results of mixing halogen solutions with solutions of potassium halides.**

|  |  |  |  |
| --- | --- | --- | --- |
|  | Potassium chloridesolution | Potassium bromide solution | Potassium iodide solution |
| ChlorineVery pale green |  | Colourless solutions turn pale yellow √hexane 🡪 red | Colourless solutions turn orange brown√hexane 🡪 purple |
| Bromineorange colour | just a dilution of the orange colourX |  | turned dark orange,starch 🡪 blue/ black√ |
| Iodinebrown colour | just a dilution of the brown colour X | just a dilution of the brown colour X |  |

**Does chlorine (Cl2) oxidise Br- (aq)?.. yes . I- (aq)? ... yes**

Half equations for:

|  |  |
| --- | --- |
| Oxidation of bromide ions | Oxidation of iodide ions |
| 2Br-(aq) 🡪 Br2(aq) + 2e- | 2I-(aq) 🡪 I2(aq) + 2e-. |

Full redox equations (leaving out spectator ions) for:

|  |  |
| --- | --- |
| Oxidation of bromide ions by chlorine | Oxidation of iodide ions by chlorine |
| Cl2(aq)+2Br-(aq)🡪Br2(aq)+2Cl-(aq) | Cl2(aq)+2I-(aq)🡪I2(aq)+2Cl-(aq) |

**Does bromine (Br2) oxidise Cl- (aq)? ... no I- (aq)? .. yes**

Half equations for this reaction

|  |  |
| --- | --- |
| Oxidation half equation | Reduction half equation |
| 2I-(aq) 🡪 I2(aq) + 2e- | Br2(aq) + 2e-🡪 2Br-(aq) |

Full redox equation (including spectator ions)

Br2(aq) + 2KI(aq) 🡪I2(aq) + 2KBr(aq)

**Does iodine (I2) oxidise Cl- (aq)? ... no Br- (aq)? .. no**

**Put the halogens in order of their oxidising ability and explain the trend.**

 Cl2 > Br2 > I2 Chlorine is able to oxidise both bromide and iodide, it has fewest electon shells with least shielding and smallest radius so has the strongest ability to attract the electron from iodide or bromide.

W/S [Halogens 1](file:///H%3A%5Ctopic%202%5CHalogens%5CWorksheets%5CHalogens%201%20redox.doc)

**Disproportionation reactions**

Refer to Factsheet 88 (part)

**A disproportionation reaction is:** the simultaneous oxidation and reduction of an atom of the same element in an ion or molecule.

As we have seen in the Redox pack, chlorine, bromine and iodine can exist in a variety of oxidation states.

**** Revision exercise – Give the Oxidation number of Chlorine in the following compounds

Cl2 0 ClO3- +5 ClO- +1 Cl2O7  +7 Cl2O3 +3

**The reaction of Chlorine with water**

As we saw on p2 the halogens are only sparingly soluble in water. When we add chlorine to water some of it dissolves, but it also reacts to form a mixture of two acids

Cl2+ H2O 🡪 HCl + HOCl

Use oxidation numbers to show that this is a disproportionation reaction

Chlorine is simmultaneously oxidised from 0 in Cl2 to +1 in HOCl and reduced from 0 in Cl2 to -1 in HCl

Chlorine is often added to drinking water as it can kill the pathogens responsible for water-born diseases such as cholera and typhoid.

**The reactions of halogens with cold alkali**

All the halogens undergo a disproportionation reaction with **cold alkali**.

**Results**

|  |  |  |
| --- | --- | --- |
| Halogen | Initial appearance of aqueous solution | Observation on adding NaOH |
| Cl2 | green / yellow | turned colourless, negative test for Cl2 |
| Br2 | orange | turned colourless |
| I2 | red/brown | turned colourless |

All the halogens react with cold sodium hydroxide in a similar way to chlorine:-

 0 -1 +1

 Cl2 (g) + 2NaOH (aq) **→** NaCl (aq) + NaClO (aq) + H2O

Bleach is a solution of Cl2 in NaOH

What are the oxidations numbers of the chlorine species in:-

Cl2 0

NaCl -1

NaClO, sodium chlorate(I) +1

Use these oxidation numbers to show that this is a disproportionation reaction.

Chlorine, (oxidation state 0) has been reduced to Cl- (-1) in NaCl

Chlorine (0) has been oxidised to (+1) in NaClO

Simultaneous oxidation and reduction of atoms in a species is called disproportionation.

Write **ionic equations** for the disproportionation reactions of:- leave out spectator ions

bromine with cold alkali

 Br2 + 2OH- 🡪 Br - + OBr - + H2O

iodine with cold alkali

 I2 + 2OH- 🡪 I- + OI- + H2O

**Reactions of the halogens with hot alkali**

When the solution is heated or reacted with hot NaOH, the chlorate(I) ions, ClO-, themselves undergo disproportionation, so the reaction between chlorine and **hot alkali** is:-

3Cl2 (aq) + 6NaOH (aq) **→** 5NaCl (aq) + NaClO3 (aq) + 3H2O (l)

Ionic equation (leaving out Na+ spectator ions).

3Cl2 (aq) + 6OH- (aq) **→** 5Cl- (aq) + ClO-3 (aq) + 3H2O (l)

****

What are the oxidation numbers of chlorine in:-

Cl2 0

NaCl -1

NaClO3, sodium chlorate(V)? +5

Use these oxidation numbers to show that this is a disproportionation reaction

Chlorine (0) has been reduced to Cl- (-1) in NaCl

Chlorine (0) has been oxidised to (+5) in NaClO3

Simultaneous oxidation and reduction of atoms in a species is called disproportionation.

Write full equations for the disproportionation reactions of iodine and bromine with **hot sodium hydroxide** (including spectator ions).

Iodine: 3I2 (aq) + 6NaOH (aq) **→** 5Nal (aq) + NalO3 (aq) + 3H2O (l)

Bromine: 3Br2 (aq) + 6NaOH (aq) **→** 5NaBr (aq) + NaBrO3 (aq) + 3H2O (l)

Write **ionic equations** for the disproportionation reactions of iodine and bromine with **hot alkali** (leaving out spectator ions)**.**

Iodine: 3I2 (aq) + 6OH- (aq) **→** 5l- (aq) + lO3- (aq) + 3H2O (l)

Bromine: 3Br2 (aq) + 6OH- (aq) **→** 5Br- (aq) + BrO3- (aq) + 3H2O (l)

**Halides**

So far we have looked at the halogens, the elements, but the halogens are so reactive that they do not exist in nature. The halogens are more often found as the halide ions Cl-, Br-, and I-.

In this section we will study the halides and their reactions with

1. Water
2. Ammonia
3. Conc. sulfuric acid
4. Silver nitrate solution.

**HCl and water**

Hydrogen chloride gas is very soluble in water, reacting with it to produce hydrochloric acid. The familiar steamy fumes of hydrogen chloride in moist air are caused by the hydrogen chloride reacting with water vapour in the air to produce a fog of concentrated hydrochloric acid.

HCl(g) + H2O(l) **→** HCl (aq) **→ H+(aq) + Cl- (aq)**

**HCl and ammonia demonstration**

Your teacher will demonstrate the reaction between conc. ammonia and conc. HCl.

Observations

Hydrogen chloride gas forms **steamy fumes in moist air** due to HCl dissolving in water droplets and forming HCl(aq).

HCl(g) forms **a white smoke with NH3(g)**, this is solid NH4Cl particles

Equation

NH3(g) + HCl(g) **→** NH4Cl (s)

Note the state symbols, the two gases react to form a solid.

Equation for HBr reacting with ammonia

NH3(g) + HBr(g) **→** NH4Br (s)

Equation for HI reacting with ammonia

NH3(g) + Hl(g) **→** NH4l (s)

**Reactions of the halides with concentrated sulphuric (VI) acid**

Two types of reaction can take place when metal halides (e.g.NaCl, NaBr, NaI) react with conc H2SO4

* **Redox reactions forming Cl2, Br2 or I2**
* **Displacement reactions forming HCl, HBr or HI**

**Redox Reactions**

We already know that the ability of the halogens Cl2, Br2 and I2 to **gain electrons** varies. The halogen which gains electrons most easily is chlorine This is, therefore the strongest oxidising agent.

It follows that the halide ions Cl-, Br- and I- will have differing ability to **lose electrons.**  Which halide ion will lose electrons most easily? iodide. This is, therefore the strongest reducing agent.

Concentrated sulphuric acid contains the sulfate (VI) ion of formula/charge SO42- . The S in this ion has oxidation number +6

**This can be reduced by the different halide ions to the following sulphur species**

|  |  |  |  |
| --- | --- | --- | --- |
| **Name** | Sulfur dioxide gas | Sulfur solid | Hydrogen sulfide gas |
| **Formula**  | SO2 | S | H2S |
| **Oxidation No of S** | +4 | 0 | -2 |

**Reactions of conc. H2SO4 and potassium halides –** demonstration / practical

CORROSIVE

* Put a spatula-full of KCl salt in a test tube.
* **Very carefully** add 2/3 drops of conc. H2SO4 to the tube (standing in a test tube rack.)
* Test the gases evolved:- with conc. NH3 on a glass rod, acidified dichromate paper, damp blue litmus paper and by bubbling through silver nitrate solution.
* Repeat with KBr and KI in place of KCl.

**Results**

|  |  |  |
| --- | --- | --- |
|  | Observations | Gases evolved |
| KCl  | Vigorous became hot (exothermic)Steamy fumes litmus 🡪 red. Acidic gasFumes produced white smoke when tested with NH3 and white ppt with AgNO3 | HCl |
| KBr  | Vigorous became hot (exothermic)Steamy fumes litmus 🡪 red. Acidic gasFumes produced white smoke when tested with NH3Orange vapour Gas turned orange dichromate(VI) paper or solution green. | HBr Br2 and SO2Cr(VI) has been reduced to Cr(III) by SO2gas |
| KI  | Very Vigorous became very hot Steamy fumes 🡪 white smoke + NH3Brown black depositsAcidified dichromate 🡪 green.Lead ethanoate 🡪 black | HI(g) I2 , SO2(g) , H2S(g)  |

**Displacement**

The equation for the displacement reaction of conc. sulphuric acid with sodium chloride

H2SO4 + NaCl **→** NaHSO4 (s) + HCl(g)

**NOTE:**  There is **no change in the oxidation numbers** of any species.

This is the **only reaction** for a metal **chloride** with concentrated sulphuric acid.

**Redox**

With two of the halides a **redox** reaction takes place as well as the straightforward displacement of one acid by another.

**Bromide**:-

KBr + H2SO4 **→** KHSO4 + HBr displacement

2HBr + H2SO4 **→** Br2 + 2H2O + SO2  redox

**Iodide**:-

KI + H2SO4 **→** KHSO4 + HI displacement

8HI + H2SO4 **→** H2S + 4H2­O + 4I2 redox

Give the oxidation numbers for the sulphur species in the following.

SO42- +6

SO2 +4

S0

H2S -2

Use these oxidation numbers to show that the reactions of HBr or HI with H2SO4 are redox reactions.

HBr and HI are reducing agents the Br- and I- ions reduce H2SO4 (+6) to SO2(+4) or H2S(-2)

The Br- and I- ions are oxidised:- HBr(-1) oxidised to Br2(0) and HI oxidised to I2(0)

Are HBr / HI oxidising or reducing agents? reducing agents

Which is the stronger agent? Explain your answer.

HI(g) can reduce S in H2SO4 from +6 to -2 in H2S

Whereas HBr only reduces S in H2SO4 from +6 to +4 in SO2.

Explain why iodides react differently from chlorides.

(hydrogen) iodide is more easily oxidized / loses electrons more easily than (hydrogen) chloridebecause it has a larger radius (than chloride).

Try Facer page 230 questions 1-10

****See worksheets [Halogens 2](file:///%5C%5Cgodalming.ac.uk%5Cdfs%5CUsers%5CStaff%5Cgda%5CUnit2%5CGroup%207%5CWorksheets%5CHalogens%202%20reactions.005.doc) [Halogens 3](file:///%5C%5Cgodalming.ac.uk%5Cdfs%5CUsers%5CStaff%5Cgda%5CUnit2%5CGroup%207%5CWorksheets%5CHalogens%203%20reactions.doc) and [Halogens 4](file:///%5C%5Cgodalming.ac.uk%5Cdfs%5CUsers%5CStaff%5Cgda%5CUnit2%5CGroup%207%5CWorksheets%5CHalogens%204%20tests.doc)

**Tests for ions**

You have studied the flame tests for certain cations in the Group 2 pack. In addition to these you need to know tests for the anions, CO32-, SO42-, Cl-, Br-, I-, and the cation NH4+.



**Tests for halide ions Cl-, Br-, I-.**

**Method**

1. Put approx 1 cm3 of a solution of a chloride (e.g.sodium chloride), in a test tube.
2. Add approx 1 cm3 of dilute nitric acid followed by 1 cm3 silver nitrate solution and mix thoroughly. Record the appearance in row 1. Then divide into 2 separate test tubes.
3. To the first tube add dilute ammonia solution dropwise until in excess. Mix well. If there is no change add a few drops of concentrated ammonia (CARE) in the fume cupboard. Cork the tube and shake carefully. Record observations in row 2 of the table below.
4. Leave the second tube in strong light for 10 mins. Record its change of appearance in row 3 of the table below.
5. Repeat steps 1-4 using a bromide solution in place of chloride.
6. Repeat steps 1-4 using an iodide solution in place of chloride,

|  |  |  |  |
| --- | --- | --- | --- |
| Test | **Chloride** | **Bromide** | **Iodide** |
| Add Ag+(aq), acidified with HNO3  | White ppt. | Cream ppt. | Yellow ppt. |
| Add excess dilute NH3 to the above mixture. | Ppt. dissolved with dil.NH3 | Ppt. dissolved in conc. NH3 | Ppt. did not dissolve |
| Leave the precipitate in the light for 10 mins. | White ppt. turned lilac | Cream ppt. turned grey | Yellow ppt. unchanged |

Write ionic equations for the reactions between silver ions and halide ions, including the states of matter.

**Chloride**: Ag+(aq) + Cl-(aq) 🡪 AgCl(s)

**Bromide**: Ag+(aq) + Br-(aq) 🡪 AgBr(s)

**Iodide** Ag+(aq) + I-(aq) 🡪 AgI(s)

Suggest why dilute nitric acid was added to the silver nitrate solution. Think what other precipitates it might dissolve. Prevents precipitation of silver hydroxide, oxide or carbonate which would interfere with the observations.

How does adding NH3 allow you to identify the precipitate?

It allows confirmation of precipitate as AgCl is soluble in both dil and conc NH3, AgBr is only soluble in conc. NH3 and Agl is insoluble in NH3**Summary - Distinguishing AgCl , AgBr and AgI precipitates**

From the experiment above you should have observed:

|  |  |  |  |
| --- | --- | --- | --- |
| **Precipitate** | **Colour** | **Solubility in NH3(aq)** | **Effect of light** |
| AgCl | White | Soluble in dilute NH3(aq) | Decomposes to silver (goes lilac) |
| AgBr | Cream | Soluble in conc.NH3(aq) | Decomposes slowly to silver(grey) |
| AgI | Yellow | Not soluble in NH3 (aq) | No change |

**Explanation:**

* Silver chloride and silver bromide react with ammonia solution to form a **complex ion**

 e.g. AgCl(s) + 2NH3(aq) 🡪 [Ag(NH3)2]+(aq) + Cl-(aq)

* Silver chloride and silver bromide **decompose** on exposure to **light**

 e.g 2AgCl(s) 🡪 2Ag(s) + Cl2(g)

These reactions help distinguish between the halide ions.

See worksheets [Halogens 2](file:///%5C%5Cgodalming.ac.uk%5Cdfs%5CUsers%5CStaff%5Cgda%5CUnit2%5CGroup%207%5CWorksheets%5CHalogens%202%20reactions.005.doc) [Halogens 3](file:///%5C%5Cgodalming.ac.uk%5Cdfs%5CUsers%5CStaff%5Cgda%5CUnit2%5CGroup%207%5CWorksheets%5CHalogens%203%20reactions.doc) and [Halogens 4](file:///%5C%5Cgodalming.ac.uk%5Cdfs%5CUsers%5CStaff%5Cgda%5CUnit2%5CGroup%207%5CWorksheets%5CHalogens%204%20tests.doc)

**Test for CO32- or HCO3- - Addition of dilute hydrochloric acid**

**Method**

1. Place approx. 2 cm3 of a solution of soluble carbonate (e.g. sodium carbonate) in a test tube.
2. Add an equal volume of dilute HCl to the test tube and record your observations.
3. Repeat steps 1 and 2 using a hydrogen carbonate solution in place of the carbonate.

|  |  |  |
| --- | --- | --- |
|  | **CO32-** | **HCO3-** |
| Observation + HCl | Effervescence and colourless solution produced  | On heating / boiling the solution without acid a gas is produced. |

What gas was evolved? CO2.How would you test for the presence of this gas?

Bubble through limewater which would give a white ppt./turn cloudy

Give the reaction including state symbols for the reaction between the carbonate and the acid

Na2CO3(aq) + 2HCl(aq) 🡪 2NaCl(aq) + H2O(l) + CO2(g)

Give an ionic equation for the reaction between carbonate and hydrogen ions

CO32-(aq) + 2H+(aq) 🡪 H2O(l) + CO2(g)

Give the reaction including state symbols for the reaction between the hydrogencarbonate and the acid

NaHCO3(aq) + HCl(aq) 🡪 NaCl(aq) + H2O(l) + CO2(g)

Give an ionic equation for the reaction between the hydrogencarbonate ion and hydrogen ions

HCO3-(aq) + H+(aq) 🡪 H2O(l) + CO2(g)

**Test for SO42- - Addition of HCl followed by BARIUM CHLORIDE soln.**

**Method**

1. To a 2 cm3 sample of a soluble sulfate (e.g. sodium sulfate) add 1 cm3 of dilute hydrochloric acid.
2. The add barium chloride solution dropwise and record your observations below:

|  |  |  |
| --- | --- | --- |
| Anion | Observation on addition of dil. HCl | Observation on addition of BaCl2 soln. |
| **SO42-** | Colourless solution produced  | Immediate thick white ppt |

Give the equation for the reaction between barium chloride and sodium sulfate, including state symbols.

Na2SO4(aq) + BaCl2(aq) 🡪 BaSO4(s) + 2NaCl(aq)

Give an ionic equation for the reaction between sulfate and barium ions

SO42-(aq) + Ba2+(aq) 🡪 BaSO4(s)

**Note**. Addition of barium chloride to a carbonate will give a white precipitate of barium carbonate. However, if HCl is added to the anion solution before BaCl2(aq) then only the sulfate(VI) SO42- will form a precipitate.

**Test for the ammonium ion NH4+ - addition of sodium hydroxide**

**Method**

* Add dilute NaOH to your sample and heat gently
* Test gas given off with damp red and blue litmus paper

**Results**

|  |  |  |  |
| --- | --- | --- | --- |
|  | Observations | Damp Red Litmus | Damp Blue litmus |
| **NH4+**  | Gas given off, smell of NH3 | Turned blue | No change |

The addition of NaOH causes the following reaction

NH4+ + OH- 🡪 NH3 + H2O

The warming releases the ammonia gas, which is detected by the litmus paper. This is a example of an acid/base reaction which you will study next year.

**Summary Table – Tests for ions LEARN THESE!**

****Fill in the table for each ion, giving details of the test and observations. You will need to learn these and the **Tests for gases** (see p 11) for Unit 2 and the practical assessment.

|  |  |  |
| --- | --- | --- |
| Ion  | Test | Observations |
| **Li+** | Flame testClean nichrome wire in HCl and use non-luminous flameDip in HCl then solid.Hold at edge of non-luminous flame | Red |
| **Na+** | Persistent yellow orange |
| **K+** | Lilac |
| **Ba2+** | Apple green |
| **Sr2+** | red  |
| **Ca2+** | Yellow-red |
| **Mg2+** | To a solution add dil.NaOH  | White ppt of Mg(OH)2(s)Mg2+(aq) + 2OH-(aq) 🡪 Mg(OH)2(s) |
| **NH4+** | To a solution add dil.NaOH, warm, and test gas with red litmusOR Heat the solid | NH4+(aq) + OH-(aq) 🡪 NH3(g) + H2O(l)NH3 alkaline gasNH4+ sublimes producing a white solid further up the test tube. |
| **Cl-** | 1)Acidify with dil HNO3Add AgNO3 dropwise To the ppt. add dil NH3 followed by conc. NH32)Add a few drops of conc H2SO4 to the solid sample 3) To 2cm3 of the sample add chlorine water until no further changes. | 1)White ppt of AgCl, Soluble in NH32) Steamy fumes HCl |
| **Br-** | 1) Cream ppt of AgBr, Soluble in conc. NH32) Steamy fumes of HBr + Orange fumes of Br23) Solution turns orange  |
| **I-** | 1) Yellow ppt of AgI, insoluble in conc. NH32) Steamy fumes of HI + Black solid I23) Solution turns brown |
| **SO42-** | Add BaCl2(aq) dropwise followed by HCl(aq) | White ppt produced BaSO4.Ba2+(aq) + SO42-(aq) 🡪 BaSO4(s)Unchanged with HCl |
| **CO32-** | Add acid, test gas by bubbling through limewater. | Effervescence 🡪 colourless solutionGas turns limewater milky |
| **HCO32-** | Heat aqueous solution | Effervescence, lime water 🡪 milky |

|  |  |
| --- | --- |
| Gas | Test & Observations |
| Oxygen | * Put a glowing spill into a test tube of the gas.
* It will relight
 |
| Carbon dioxide | * Bubble the gas through a solution of .limewater
* This will go cloudy / form a white ppt
 |
| Ammonia | * Smelly gas. Damp red litmus goes blue
* With HCl gas **white smoke** is produced
 |
| Nitrogen dioxide | * This is a brown coloured gas
* Damp blue litmus goes red
 |
| Hydrogen | * Put a lighted spill into a test tube of the gas.
* It will give a squeaky pop
 |
| Hydrogen chloride | * Steamy fumes. Damp blue litmus goes red
* With NH3 gas **white smoke** is produced
* When bubbled into AgNO3(aq) **a white precipitate** forms
 |
| Hydrogen bromide | * Steamy fumes. Damp blue litmus goes red
* With NH3 gas  **white smoke** is produced
* When bubbled into AgNO3(aq)  **a cream coloured precipitate**
 |
| Hydrogen iodide | * Steamy fumes. Damp blue litmus goes red
* With NH3 gas  **white smoke** is produced
* When bubbled into AgNO3(aq)  **a yellow precipitate** forms
 |
| Chlorine | * Smells of bleach
* Damp blue litmus paper goes red then bleached
* Damp starch/iodide paper 🡪 blue
* Turns KBr (aq) from colourless to yellow / orange
 |
| Bromine | * Vapour is orange in colour.
* Turns KI (aq) from colourless to brown
 |
| Iodine | * Vapour is purple in colour. Solid is gray
* Turns starch soln. from white to blue/black
 |
| Water vapour  | * Anhydrous copper(II)sulfate goes from white to blue
 |
| Sulfur dioxide | * Acidified potassium dichromate solution goes from orange to green/blue
 |
| Hydrogen sulfide | * Rotten egg smell.
* Lead(II) nitrate solution goes from colourless to black
 |

**Revision page.**

****[See AS Inorganic Dry Labs](file:///H%3A%5Ctopic%202%5CHalogens%5CAS%20inorganic%20dry%20Labs.rtf)

**Try Facer page 230 questions 8 and 13**

In the space below, include:-

* Physical properties of the halogen elements
* Redox reactions of halogens with metals and ions
* Reactions of potassium halides with conc. H2SO4
* Reaction of halides with silver nitrate solution
* Reactions of hydrogen halides as acids and with ammonia.
* Tests for ions
* Tests for gases.

|  |
| --- |
| **TO DO and CHECK: Godalming online quiz ‘group 7’. https://online.godalming.ac.uk/mod/quiz/view.php?id=726** |



**Multiple Choice Exam Questions**

Multiple choice p25 Q12-14

**1.** When concentrated sulfuric acid is added to solid sodium bromide, bromine is produced. When concentrated sulfuric acid is added to solid sodium chloride, **no** chlorine is produced.

 The reason for this difference is

**A** sulfuric acid is a strong acid.

**B** hydrogen chloride is a weak acid.

**C** the chloride ion is a weaker reducing agent than the bromide ion.

**D** bromine is less volatile than chlorine.

(Total 1 mark)

**2.** Which of the following statements about the elements in Group 7 is **incorrect**?

**A** They all show variable oxidation states in their compounds.

**B** They all form acidic hydrides.

**C** Electronegativity decreases as the group is descended.

**D** They all exist as diatomic molecules.

(Total 1 mark)

**3.** What are the products, other than water, when chlorine is passed through cold, dilute aqueous sodium hydroxide solution?

**A** NaCl and NaClO

**B** NaClO and NaClO3

**C** NaCl and NaClO3

**D** NaClO and NaClO4

(Total 1 mark)

Do not write in the margin

**4.** What would be the colour of the solution when iodine is dissolved in a hydrocarbon solvent?

Multiple choice p26 - 27 Q15-18

**A** Grey

**B** Brown

**C** Yellow

**D** Purple

(Total 1 mark)

**5.** Which of the following is **not** a true statement about hydrogen iodide?

**A** It forms steamy fumes in moist air.

**B** It dissolves in water to form an acidic solution.

**C** It forms a cream precipitate with silver nitrate solution.

**D** It forms dense white smoke with ammonia.

(Total 1 mark)

**6.** Chemical reactions may involve

**A** oxidation

**B** reduction

**C** no change in oxidation number

**D** disproportionation

 Which of the terms above best describes what happens to the **chlorine** in the following reactions?

(a) Cl2(g) + H2O(l) → HCl(aq) + HOCl(aq)

**A**

**B**

**C**

**D**

(1)

(b) Cl2(g) + 2Na(s) → 2NaCl(s)

**A**

**B**

**C**

**D**

(1)

(c) NaCl(s) + H2SO4(l) → HCl(g) + NaHSO4(s)

**A**

**B**

**C**

**D**

 **(1)**

**Total 3 marks**

Do not write in the margin

**Structured Questions**

1. Chlorine was used in swimming pools as a bactericide.

 The amount of chlorine present can be determined by adding excess potassium iodide solution to a known volume of swimming pool water. This reacts to form iodine:

 Cl2(aq) + 2I–(aq)  I2(aq) + 2Cl–(aq)

 The amount of iodine formed is then found by titration with sodium thiosulfate solution of known concentration.

 The ionic equation for the reaction between iodine and sodium thiosulfate in aqueous solution is

 I2(aq) + 2S2O32–(aq)  S4O62–(aq) + 2I–(aq)

 A student carried out the determination of chlorine in a sample of swimming pool water.
A record of the measurements obtained is given below:

Volume of water sample tested = 1000 cm3

Final reading of burette = 16.3 cm3

Initial reading of burette = 7 cm3

Volume added from burette = 9.3 cm3

 Concentration of sodium thiosulfate solution = 0.00500 mol dm–1

(a) (i) The record of measurements reveals faults both in the procedure and the recording of measurements. State one fault in each of these.

Procedure Only one titration carried out/ no check on accuracy of titration OR 1000 cm3 volume to large to fit in titration flask

Recording of measurements Did not record burette readings to 0.05 cm3
/1 decimal place / sufficient precision / recording only one significant figure in a titration reading.

(2)

(ii) Calculate the number of moles of sodium thiosulfate used in the titration.

 4.65 × 10–5 / 4.7 × 10–5 / 0.0000465 / 0.000047 (mol)

**(1)**

W02 C 1 04

(iii) Use your answer to (ii) to calculate the number of moles of iodine which reacted.

 2.3 × 10–5 / 0.000023
OR candidates answer to (ii) divided by 2

(1)

iv) Deduce the concentration of chlorine, in mol dm-3, in the swimming pool water.

 2.3 × 10–5 / 0.000023 mol dm–3
OR candidates answer to (iii)

(1)

(b) The disinfecting action of chlorine in swimming pools is due to the presence of chloric(I) acid, HClO, formed by the reaction of chlorine with water.

 In many swimming pools, chemicals other than chlorine are used to form chloric(I) acid. This is partly because the use of chlorine gas causes much more corrosion of metal parts in swimming pools than does chloric(I) acid.

 Compounds used to chlorinate swimming pool water in this way include calcium chlorate(I) and chlorine dioxide.

(i) State and explain the type of reaction that occurs when chlorine attacks a metal, using the example of iron.

**Redox** as chlorine removes/gains electrons from the metal
(and is reduced) **(1)**And metal gives/loses electrons to the chlorine (and is oxidised) **(1)**

 Redox is essential in order to score both marks
The gain / loss of electrons can be awarded from two ionic
half equations**.**

(2)

(ii) Suggest **one** other reason why the use of chlorine is undesirable in swimming pools.

Chlorine is (highly) toxic/poisonous/irritant
OR chlorine has an unpleasant smell **(1)**.

(1)

 (iii) Give the formula for calcium chlorate(I).

Ca(ClO)2..

(1)

(iv) Chlorine dioxide, ClO2, undergoes a disproportionation reaction when it reacts with water.

 4ClO2 + 2H2O  HClO + 3HClO3

 Explain, in terms of oxidation numbers, why this is a disproportionation reaction.

Cl is oxidised from +4 (in ClO2) to +5 (in HClO3) **(1)**and is reduced (from +4) to +1 (in HClO) **(1)**..

(2)

**(Total 11 mark)**

2 (a) Hydrogen chloride can be made from sodium chloride and concentrated sulphuric acid. Write a balanced chemical equation to represent this reaction.

NaCl + H2SO4  NaHSO4 + HCl

(1)

 (b) (i) How would you confirm that a solution said to be HCl(aq) contained chloride ions?

add silver nitrate (solution) / correct formula AgN03 (aq) **(1)**
white ppt /solid **(1)**
soluble in dilute ammonia /ammonia solution **(1)**

(3)

(ii) Hydrogen chloride is soluble in water. Explain why the solution is acidic.

dissociates /reacts/lionises/changes into ions (as it dissolves) **(1)**
forming H+ ions / H3O+ ions/ donates a proton to water.
This makes the solution an acid **(1)**.

(2)

(c) (i) Give a chemical test for chlorine, stating what you would do and what you would see.

damp litmus paper **(1)** bleached **(1)** or
damp starch-iodide paper **(1)** goes blue **(1)**
Displacement acceptable.

(2)

(ii) Hydrogen chloride can be oxidised to chlorine by lead(IV) oxide, PbO2. Write the oxidation numbers of lead and of chlorine in the boxes provided.



 +4 -1 +2 -1 0 (2)

(d) Sodium iodide reacts with concentrated sulphuric acid to give iodine, not hydrogen iodide. Explain why iodides react differently from chlorides in this case.

(hydrogen) iodide is more easily oxidized / loses electrons more
easily than (hydrogen) chloride **(1)**because larger (than chloride) **(1)***Could argue from the reducing power of the iodide / chloride
for the first mark*

(2)

(Total 12 marks)

**3.** This question is about the manufacture of bromine from bromide ions found in seawater.

(a) In the first step, chlorine gas is bubbled into acidified seawater. This converts the bromide ions to bromine. The low pH prevents hydrolysis of the liberated bromine.

(i) Complete and balance the equation for the hydrolysis of bromine with water which is a disproportionation reaction.

 Br2(aq) + H2O(l)  2H+(aq) + Br–(aq) + BrO–(aq)
formulae **(1)**balancing (ignore state symbols) **(1)**

(2)

 (1)

(iv) Write the ionic equation, including state symbols, for the reaction of chlorine gas with bromide ions.

Cl2(g) + 2Br–(aq) → 2Cl–(aq) + Br2(aq)
formulae **(1)**balancing and state symbols **(1)**

(2)

(b) In the second step, air is blown through the reaction mixture to remove the bromine as a vapour which is then mixed with sulphur dioxide gas and water vapour. The unbalanced equation for this reaction is

 Br2 + H2O + SO2 → H+ + Br− + SO42–

(i) Identify the elements which are oxidised and reduced and give their oxidation numbers.

Element **oxidised** S

Oxidation number initial +4 final +6

Element **reduced** Br2

Oxidation number initial 0 final -1

(2)

(ii) Using this information, or otherwise, balance the equation.

Br2 + 2H2O + SO2 → 4H+ + 2Br– + SO42–

(Total 6 marks)

Halogens

|  |  |
| --- | --- |
| **Task** | Mark |
| Notes | /10 |
| Revision Notes /Summary page | /10 |
| Godalming online quiz Group 7 | /10 |
| Structured exam questions  | /39 |
| Overall Grade for this work | A B C D E U |

Student comments

Tutor comments

Tutor signature Date

Student targets for **next pack**