2.3.3 Halogens: displacement and disproportionation

Name: _____ Class: _____ Date: _____ **Total: 24 marks**

KEY WORDS disproportionation

Objective

Use oxidation-number and observation evidence while keeping reagent conditions explicit. All questions and marking guidance are original.

1 Key knowledge and worked examples

Cl_2, Br_2 and I_2 become weaker oxidising agents down Group 7: an incoming electron is less strongly attracted as radius and shielding increase. Larger electron clouds strengthen London forces and raise boiling temperature.

■ Worked explanation

$Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$. Chlorine gains electrons and is reduced, so it is the oxidising agent. Bromide is oxidised and is the reducing agent. In water bromine is orange/brown; iodine is brown. In a suitable non-polar organic solvent iodine is violet.

Cold dilute OH^- gives Cl^- and ClO^-. Hot concentrated OH^- gives Cl^- and $ClO_3{}^-$. In each case chlorine starts at 0 and is both reduced and oxidised. Never mix bleach with acid as an informal test.

2 Practice

2.1 Write the ionic equation for bromine displacing iodide and identify the reducing agent. [3]

2.2 State the physical states and characteristic colours of chlorine, bromine and iodine at room temperature. [3]

2.3 Explain why iodine has a higher boiling temperature but is a weaker oxidant than chlorine. [4]

3 Exam-style questions

3.1 Balance $Cl_2 + OH^- \rightarrow Cl^- + ClO_3{}^- + H_2O$ for hot concentrated alkali. Use oxidation numbers to explain the coefficient ratio. [5]

3.2 Explain the different main reactions of solid NaCl and NaBr with concentrated H_2SO_4. Give the redox equation involving HBr and name the sulfur product. [4]

3.3 Predict the relative oxidising powers of fluorine, chlorine and astatine using the Group 7 trend. Write equations for chlorine with water and hydrogen chloride with ammonia, then explain one benefit and one risk of chlorine water treatment. [5]

Solutions

Original marking guidance: M1 method, A1 accuracy, B1 explanation.

2.1 $Br_2 + 2I^- \rightarrow 2Br^- + I_2$. I^- loses electrons and is the reducing agent.

2.2 Chlorine: pale-green gas; bromine: red-brown liquid; iodine: grey-black solid. Iodine vapour is purple, not its ordinary solid appearance.

2.3 Iodine has more electrons/more polarisable cloud, stronger London forces. Its larger radius/more shielding weakens attraction for an incoming electron.

3.1 $3Cl_2 + 6OH^- \rightarrow 5Cl^- + ClO_3^- + 3H_2O$. Chlorine changes $0 \rightarrow -1$ and $0 \rightarrow +5$, so five atoms gain one electron each for every atom losing five.

3.2 NaCl releases HCl by acid–base reaction, with no reduction of sulfuric acid. Bromide/HBr can reduce the acid: $2HBr + H_2SO_4 \rightarrow Br_2 + SO_2 + 2H_2O$. SO_2 contains sulfur reduced from +6 to +4. This is a hazardous controlled demonstration.

3.3 $F_2 > Cl_2 > At_2$ from increasing radius/shielding and weaker attraction for an incoming electron down the group. $Cl_2 + H_2O \rightleftharpoons HCl + HClO$. $HCl + NH_3 \rightarrow NH_4Cl$. Disinfection reduces microbial disease; toxicity/harmful by-products require controlled dosing and monitoring.