2.6 The Periodic Table and Inorganic Chemistry: 1-hour Pack

Total Marks

AQA 7405 · Calculator allowed · about 60 minutes

Name: _____ Date: ____ / ____ / _____

Answer ALL questions. Show all your working.

1 A student reacts small samples of magnesium ribbon and calcium metal, in turn, with cold water.

(a) Write a balanced equation, including state symbols, for the reaction of calcium with cold water. **(2)**

(b) Describe the observations you would expect when comparing magnesium and calcium each added to cold water. **(2)**

(c) Explain why reactivity with water increases down Group 2 from magnesium to calcium. **(3)**

(d) Magnesium does not react with cold water but does react with steam. Write a balanced equation for the reaction of magnesium with steam, and explain why the product formed is different from the product formed when calcium reacts with cold water. **(3)**

(Total for Question 1 is 10 marks)

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2 This question uses practical techniques from the required practical qualitative test for halide ions. A student is given a colourless solution and asked to identify which halide ion it contains. The student adds a few drops of dilute nitric acid, followed by silver nitrate solution, $AgNO_3(aq)$, and then tests the solubility of any precipitate formed in dilute and concentrated ammonia solution.

(a) State the purpose of adding dilute nitric acid to the solution before the silver nitrate solution. **(1)**

(b) Complete the following, stating the colour of the precipitate formed with silver nitrate solution, and its solubility in ammonia, for chloride ions, bromide ions and iodide ions. **(3)**

(c) Write an ionic equation for the reaction between silver ions and bromide ions. **(1)**

(d) A student carries out this test on an unknown solution, X, and observes a cream precipitate that dissolves only when concentrated ammonia solution is added (it does not dissolve in dilute ammonia). Identify the halide ion present in X and justify your answer. **(2)**

(e) Explain why this silver nitrate/ammonia test alone would not distinguish between a solution containing only chloride ions and a solution containing a mixture of chloride and bromide ions. **(2)**

(Total for Question 2 is 9 marks)

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3 This question tests key definitions used throughout the periodic table and inorganic chemistry topic.

(a) Define periodicity. **(1)**

(b) Define disproportionation. **(1)**

(c) Define an amphoteric oxide. **(1)**

(d) State, in terms of oxidation number, what happens to a reducing agent during a redox reaction. **(1)**

(Total for Question 3 is 4 marks)

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4 This question is about halogen displacement and disproportionation reactions.

(a) Chlorine water is added to a colourless solution of potassium bromide. Write an ionic equation for the reaction that occurs and state the observation. **(3)**

(b) Explain, in terms of oxidising power, why chlorine displaces bromide ions from solution but iodine does not displace chloride ions from solution. **(2)**

(c) Chlorine gas is bubbled into cold, dilute sodium hydroxide solution. Write a balanced equation for this reaction, state one commercial use of the products formed and name the type of redox reaction that has taken place. **(4)**

(d) Explain what is meant by the term 'disproportionation', using oxidation numbers to show the changes in oxidation state of chlorine in the reaction in part (c). **(3)**

(Total for Question 4 is 12 marks)

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5 This question is about the reactions of solid sodium halides with concentrated sulfuric acid.

(a) Write an equation for the reaction of solid sodium chloride with concentrated sulfuric acid, and state the observation made. **(2)**

(b) With solid sodium bromide, the HBr initially formed is partly further oxidised by concentrated sulfuric acid, forming bromine and sulfur dioxide. Write an equation for this further reaction, and state the oxidation state of sulfur in H_2SO_4 and in SO_2. **(3)**

(c) With solid sodium iodide, concentrated sulfuric acid is reduced much further, all the way to hydrogen sulfide, H_2S. Write the overall equation for this reaction, and calculate the oxidation state of sulfur in H_2SO_4 and in H_2S. State what the change in oxidation state shows about iodide compared with bromide. **(4)**

(d) Explain why the reducing power of the halide ions increases down Group 7, from chloride to iodide. **(3)**

(Total for Question 5 is 12 marks)

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Mark scheme · 2.6 The Periodic Table and Inorganic Chemistry: 1-hour Pack

Question 1

(a) correct formulae: Ca, H_2O, $Ca(OH)_2$, H_2 **M1**

(a) $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$ fully balanced with correct state symbols cao **A1**

(a) Answer: $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

(b) magnesium reacts very slowly (if at all) with cold water, with few/no visible bubbles over a reasonable timescale oe **B1**

(b) calcium reacts noticeably/steadily with cold water, producing a stream of bubbles (hydrogen gas) and forming a cloudy white suspension/precipitate of calcium hydroxide oe **B1**

(b) Answer: Magnesium shows almost no reaction with cold water; calcium reacts steadily, fizzing (releasing hydrogen gas) and forming a cloudy white suspension of calcium hydroxide

(c) atomic radius increases down the group (extra electron shell for each successive element) oe **B1**

(c) shielding by inner shells increases, which more than compensates for the increase in nuclear charge oe **B1**

(c) so the outer two electrons are less strongly attracted to the nucleus / (first and second) ionisation energies decrease down the group, so the atoms lose their outer electrons more easily, making the metal more reactive oe **B1**

(c) Answer: Atomic radius and shielding increase down the group, outweighing the increase in nuclear charge; this lowers ionisation energies down the group, so the outer electrons are lost more easily, increasing reactivity

(d) $Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$ cao **B1**

(d) magnesium hydroxide, $Mg(OH)_2$, is (much less soluble/almost insoluble) and unstable at the high temperature of steam, so it decomposes to the oxide oe **B1**

(d) at the higher temperature of steam magnesium reacts fast enough to form the oxide directly, whereas at room temperature with cold water the reaction is far too slow to be observed oe **B1**

(d) Answer: $Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$; the reaction with steam occurs at a much higher temperature, forming the oxide (as any hydroxide formed is not stable/is far less soluble than $Ca(OH)_2$), whereas cold water reacts too slowly with Mg to be observed

Question 2

(a) to remove/react away any carbonate or hydroxide ions present, which would otherwise also form a precipitate with Ag^+ and give a false-positive/misleading result oe **B1**

(a) Answer: It removes carbonate or hydroxide ions that would otherwise also form a precipitate with silver ions, giving a false result

(b) chloride: white precipitate, dissolves in dilute ammonia solution **B1**

(b) bromide: cream precipitate, dissolves only in concentrated ammonia solution **B1**

(b) iodide: yellow precipitate, insoluble in ammonia (even concentrated) **B1**

(b) Answer: Cl^-: white precipitate, dissolves in dilute NH_3; Br^-: cream precipitate, dissolves only in concentrated NH_3; I^-: yellow precipitate, insoluble even in concentrated NH_3

(c) $Ag^+ + Br^- \rightarrow AgBr$ cao **B1**

(c) Answer: $Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$

(d) bromide ions, Br^- **B1**

(d) justification: a cream precipitate that is insoluble in dilute ammonia but dissolves in concentrated ammonia is characteristic of AgBr (unlike AgCl, which dissolves in dilute ammonia, or AgI, which does not dissolve even in concentrated ammonia) oe **B1**

(d) Answer: Bromide ions (Br^-); the observations (cream precipitate, insoluble in dilute ammonia but soluble in concentrated ammonia) uniquely match AgBr

(e) both AgCl and AgBr would be present as a mixed precipitate, and the overall colour/behaviour may be dominated by or hard to distinguish from a pure chloride result oe **B1**

(e) in particular, both AgCl and AgBr dissolve in (a suitable concentration of) ammonia, and small amounts of a second halide can be masked by the excess of the first, so the test cannot reliably confirm the absence of a second halide ion oe **B1**

(e) Answer: A mixture of AgCl and AgBr precipitates would form together, and since both dissolve in ammonia (AgCl in dilute, AgBr only in concentrated), a small amount of a second halide can be masked by an excess of the other, so the simple test cannot reliably distinguish a pure chloride solution from a chloride/bromide mixture

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Question 3

(a) a repeating (periodic) trend in the physical or chemical properties of elements when arranged in order of increasing atomic (proton) number, oe **B1**

(a) Answer: A repeating trend in element properties with increasing atomic number.

(b) a redox reaction in which the same element is simultaneously oxidised and reduced, oe **B1**

(b) Answer: A redox reaction in which the same element is both oxidised and reduced.

(c) an oxide that can react with both acids and bases (alkalis), oe **B1**

(c) Answer: An oxide that reacts with both acids and bases.

(d) its oxidation number increases (it is oxidised), oe **B1**

(d) Answer: Its oxidation number increases.

Question 4

(a) correct species: Cl_2, Br^-, Cl^-, Br_2 **M1**

(a) $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$ fully balanced cao **A1**

(a) solution turns (from colourless) to orange oe **B1**

(a) Answer: $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$; the solution turns orange

(b) oxidising power (ability to gain electrons/be reduced) decreases down Group 7, so chlorine is a stronger oxidising agent than bromine (and iodine) oe **B1**

(b) chlorine, as the stronger oxidising agent, can oxidise (remove electrons from) Br^-; but iodine is a weaker oxidising agent than chlorine, so it cannot oxidise Cl^- (remove electrons from the already more strongly held chloride ion) oe **B1**

(b) Answer: Oxidising power decreases down Group 7, so Cl_2 (a stronger oxidising agent) can oxidise Br^- to Br_2, but I_2 (a weaker oxidising agent than Cl_2) cannot oxidise Cl^- to Cl_2

(c) correct species: Cl_2, NaOH, NaCl, NaOCl (or NaClO), H_2O **M1**

(c) $Cl_2 + 2NaOH \rightarrow NaCl + NaOCl + H_2O$ fully balanced cao **A1**

(c) commercial use: household bleach / disinfectant oe **B1**

(c) this is an example of a disproportionation reaction **B1**

(c) Answer: $Cl_2 + 2NaOH \rightarrow NaCl + NaOCl + H_2O$; the mixture (containing sodium chlorate(I), NaOCl) is used as household bleach/disinfectant; this is a disproportionation reaction

(d) disproportionation is a reaction in which the same element is simultaneously oxidised and reduced oe **B1**

(d) chlorine in Cl_2 has oxidation state 0, and is reduced to -1 in Cl^- (NaCl) ft from (c) **B1**

(d) chlorine in Cl_2 (oxidation state 0) is also oxidised to +1 in ClO^- (NaOCl) ft from (c) **B1**

(d) Answer: Disproportionation is when the same element is simultaneously oxidised and reduced; here Cl (0 in Cl_2) is reduced to -1 (in Cl^-) and oxidised to +1 (in ClO^-) in the same reaction

Question 5

(a) $NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$ cao **B1**

(a) steamy/misty white fumes (of HCl gas) observed oe **B1**

(a) Answer: $NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$; steamy white fumes of HCl gas are observed

(b) $2HBr + H_2SO_4 \rightarrow Br_2 + SO_2 + 2H_2O$ cao **B1**

(b) sulfur in H_2SO_4 = +6 **B1**

(b) sulfur in SO_2 = +4 **B1**

(b) Answer: $2HBr + H_2SO_4 \rightarrow Br_2 + SO_2 + 2H_2O$; S in H_2SO_4 = +6, S in SO_2 = +4

(c) $8HI + H_2SO_4 \rightarrow 4I_2 + H_2S + 4H_2O$ (accept $8NaI + 9H_2SO_4 \rightarrow 8NaHSO_4 + 4I_2 + H_2S + 4H_2O$, or $8H^+ + 8I^- + H_2SO_4 \rightarrow 4I_2 + H_2S + 4H_2O$) oe **M1**

(c) sulfur in H_2SO_4 = +6 **A1**

(c) sulfur in H_2S = -2 **A1**

(c) this is a change of 8 in oxidation state (from +6 to -2), showing iodide is a much stronger reducing agent than bromide (which only reduces S to +4 in SO_2) oe **B1**

(c) Answer: $8HI + H_2SO_4 \rightarrow 4I_2 + H_2S + 4H_2O$; S in H_2SO_4 = +6, S in H_2S = -2

(d) reducing power depends on how readily the halide ion loses an electron (is oxidised) oe **B1**

(d) ionic radius increases down the group ($Cl^- < Br^- < I^-$), so the outer electron to be lost is further from the nucleus and more shielded oe **B1**

(d) this electron is less strongly attracted/held, so it is lost more easily down the group, meaning I^- loses an electron more readily than Cl^- and is therefore a stronger reducing agent oe **B1**

(d) Answer: Ionic radius increases down the group, so the outer electron of I^- is further from the nucleus and more shielded than in Cl^-; it is therefore lost more easily, making I^- a stronger reducing agent than Cl^-

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