InkChemistry

ANSWER BOOK

Metal-aqua ions and the acidity of their solutions

Ions in aqueous solution · Unit C15

AQA 3.2.6

18 questions · 63 marks

NAVY EDITION

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SECTION A · Short, 1 to 2 marks

A1 MAI-Q14

B1 Na^+ has only a single charge and a larger radius than Al^{3+}, so its charge/size ratio is much lower. **B1** Na^+ does not polarise the O–H bonds of the surrounding water enough to release H^+, whereas Al^{3+} does. [2]

SECTION B · Standard, 3 to 5 marks

B1 MAI-Q15

B1 both solutions give a precipitate, green with iron(II) and brown with iron(III). **B1** only the iron(III) solution gives effervescence, of carbon dioxide. **B1** the iron(II) precipitate is the carbonate, $FeCO_3$. **B1** the iron(III) precipitate is the hydroxide, $[Fe(H_2O)_3(OH)_3]$, not a carbonate; reject $Fe_2(CO_3)_3$. **B1** $2[Fe(H_2O)_6]^{3+}(aq) + 3CO_3^{2-}(aq) \rightarrow 2[Fe(H_2O)_3(OH)_3](s) + 3CO_2(g) + 3H_2O(l)$. [5]

B2 MAI-Q1

B1 iron(III) nitrate contains $[Fe(H_2O)_6]^{3+}$ and its solution is yellow-brown; allow pale violet for the unhydrolysed ion. **B1** copper(II) sulfate contains $[Cu(H_2O)_6]^{2+}$, giving a blue solution. **B1** aluminium sulfate contains $[Al(H_2O)_6]^{3+}$, giving a colourless solution. **B1** Al^{3+} has no d electrons, so no electron can be promoted and no visible light is absorbed. [4]

B3 MAI-Q4

B1 the products are $[Al(H_2O)_5(OH)]^{2+}$ and H_3O^+, with the complex's charge falling from 3+ to 2+. **B1** both products are shown as (aq) and the equilibrium arrow is kept; allow $H^+(aq)$ only if the $H_2O(l)$ on the left is removed. **B1** the acid is $[Al(H_2O)_6]^{3+}$, which donates a proton from one of its water ligands. **B1** the base is the solvent water molecule, which accepts the proton and becomes H_3O^+. [4]

B4 MAI-Q3

B1 six water molecules act as ligands around the central Fe^{2+} ion. **B1** each water molecule forms a coordinate bond by donating a lone pair from its oxygen atom to the iron ion. **B1** the six oxygen atoms are arranged octahedrally around the iron, with bond angles of 90°. [3]

B5 MAI-Q6

B1 in water the Fe^{3+} ion is surrounded by six water ligands as $[Fe(H_2O)_6]^{3+}$. **B1** the metal ion attracts electron density from the oxygen atoms, polarising and weakening the O–H bonds of the water ligands. **B1** a water ligand loses H^+ to a solvent water molecule, giving $[Fe(H_2O)_5(OH)]^{2+}$ and H_3O^+, so the solution is acidic. [3]

B6 MAI-Q8

M1 $K_a = [H^+]^2 \div [Al(H_2O)_6^{3+}]$, taking $[H^+]$ equal to the concentration of $[Al(H_2O)_5(OH)]^{2+}$ and the aqua ion concentration as unchanged. **M1** $[H^+]^2 = 1.0 \times 10^{-5} \times 0.0500 = 5.0 \times 10^{-7}$, so $[H^+] = 7.07 \times 10^{-4}$ mol dm^{-3}. **A1** $pH = -\log_{10}(7.07 \times 10^{-4}) = 3.15$, to two decimal places. [3]

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B7 MAI-Q9 **M1** $[H^+]$ = 10 to the power -5.25 = 5.62×10^{-6} mol dm^{-3}. **M1** Ka = $[H^+]^2 \div [Fe(H_2O)_6^{2+}] = (5.62 \times 10^{-6})^2 \div 0.100$. **A1** Ka = 3.2×10^{-10} mol dm^{-3}, to two significant figures. [3]

B8 MAI-Q10 **M1** $[H^+]^2$ = Ka × $[Cu(H_2O)_6^{2+}] = 3.2 \times 10^{-8} \times 0.250 = 8.0 \times 10^{-9}$. **M1** $[H^+]$ = square root of $8.0 \times 10^{-9} = 8.94 \times 10^{-5}$ mol dm^{-3}. **A1** pH = $-\log_{10}(8.94 \times 10^{-5})$ = 4.05, to two decimal places. [3]

B9 MAI-Q11 **B1** $[Al(H_2O)_6]^{3+}$ is the stronger acid because its higher charge/size ratio polarises the O–H bonds more, so its solution holds more H_3O^+. **B1** that H_3O^+ reacts with carbonate ions to give carbon dioxide, and the aluminium hydroxide precipitates. **B1** $[Cu(H_2O)_6]^{2+}$ is too weak an acid to release carbon dioxide, so the carbonate ions precipitate copper(II) carbonate instead. [3]

B10 MAI-Q13 **B1** the yellow-brown colour is largely due to the hydrolysed complex $[Fe(H_2O)_5(OH)]^{2+}$. **B1** adding acid raises the concentration of H_3O^+, so the hydrolysis equilibrium shifts to the left, by Le Chatelier's principle. **B1** the solution then contains mainly $[Fe(H_2O)_6]^{3+}$, which is pale violet. [3]

B11 MAI-Q16 **B1** the approximation gives $[H^+]$ = square root of $(6.3 \times 10^{-3} \times 0.0500) = 1.8 \times 10^{-2}$ mol dm^{-3} and a pH of 1.75. **B1** that $[H^+]$ is about 35% of 0.0500, so the aqua ion concentration has not stayed close to its starting value. **B1** the approximation is therefore poor here; it overestimates $[H^+]$, and the true pH is higher than 1.75. [3]

B12 MAI-Q19 **M1** the concentration of $[Cu(H_2O)_6]^{2+}$ after dilution is $0.200 \times 10.0 \div 100.0$ = 0.0200 mol dm^{-3}. **M1** $[H^+]^2 = 3.2 \times 10^{-8} \times 0.0200 = 6.4 \times 10^{-10}$, so $[H^+] = 2.53 \times 10^{-5}$ mol dm^{-3}. **A1** pH = $-\log_{10}(2.53 \times 10^{-5})$ = 4.60, to two decimal places. [3]

B13 MAI-Q20 **B1** hydrochloric acid is fully ionised, so $[H^+]$ falls tenfold and the pH rises by $\log_{10}10 = 1$. **B1** for the aqua ion $[H^+] = \sqrt{(Ka \times c)}$, so a tenfold fall in c lowers $[H^+]$ only by $\sqrt{10}$, about 3.16, a pH rise of 0.5. **B1** on dilution the hydrolysis equilibrium shifts to the right, so a larger fraction of the aqua ions is hydrolysed and $[H^+]$ falls by less than the concentration does. [3]

B14 MAI-Q21 **B1** Mg^{2+} has a higher charge and a smaller radius than Na^+, so its charge/size ratio is high enough to polarise the O–H bonds of its water ligands a little. **B1** $[Mg(H_2O)_6]^{2+}$ therefore hydrolyses slightly, releasing a little H_3O^+, so the solution is very slightly acidic. **B1** Mg^{2+} has a lower charge and a larger radius than Al^{3+}, so it polarises the O–H bonds far less and the hydrolysis lies much further to the left than for $[Al(H_2O)_6]^{3+}$. [3]

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B15 MAI-Q22 **M1** $[Cu(H_2O)_6]^{2+}(aq) + H_2O(l) \rightleftharpoons [Cu(H_2O)_5(OH)]^{+}(aq) + H_3O^{+}(aq)$, so $K_a = [H^+]^2 \div 0.100$ with the two weak-acid assumptions. **M1** $[H^+] = \sqrt{(3.2 \times 10^{-8} \times 0.100)} = \sqrt{(3.2 \times 10^{-9})} = 5.66 \times 10^{-5}$ mol dm^{-3}. **A1** pH = $-\log_{10}(5.66 \times 10^{-5}) = 4.25$; the solution is only weakly acidic because a 2+ ion polarises its water ligands slightly. [3]

SECTION C · Long, 6 marks or more

C1 MAI-Q7 **Indicative content** [6]

B1 $[Fe(H_2O)_6]^{3+}(aq) + H_2O(l) \rightleftharpoons [Fe(H_2O)_5(OH)]^{2+}(aq) + H_3O^{+}(aq)$, the charge on the complex falling by one as a proton is lost. **B1** $[Fe(H_2O)_6]^{2+}(aq) + H_2O(l) \rightleftharpoons [Fe(H_2O)_5(OH)]^{+}(aq) + H_3O^{+}(aq)$, the same hydrolysis for the 2+ ion, to a lesser extent. **B1** Fe^{3+} has a higher charge and a smaller radius than Fe^{2+}, so it has the higher charge density; accept charge/size ratio. **B1** Fe^{3+} therefore attracts the lone pairs of its water ligands more strongly, so the O–H bonds in those ligands are polarised more. **B1** the O–H bonds are weakened more, so H^+ is released more readily and the hydrolysis equilibrium lies further to the right. **B1** $[H_3O^+]$ is higher in the iron(III) solution, so its pH is lower: about 1.7, against about 5 for the iron(II) solution.

Levels

Level 3 (5–6 marks) both hydrolysis equations are correct with state symbols, the higher charge density of Fe^{3+} is explained through its charge and radius, the stronger polarisation and weakening of the O–H bonds is linked to the position of the equilibrium, and the lower pH is stated with approximate values.

Level 2 (3–4 marks) one hydrolysis equation is right and the higher charge density of Fe^{3+} explains the stronger polarisation of the water ligands; the second equation, the shift in the equilibrium or the pH comparison is not there.

Level 1 (1–2 marks) Fe^{3+} is said to be more acidic, often because it 'has more H^+' to release; no correct equation and no mention of the O–H bonds in the ligands.

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C2 MAI-Q17 **Indicative content** [6]

B1 a 3+ ion gives effervescence of carbon dioxide and a hydroxide precipitate; a 2+ ion gives a carbonate precipitate and no gas. **B1** for the 3+ ion, $2[M(H_2O)_6]^{3+}(aq) + 3CO_3^{2-}(aq) \rightarrow 2[M(H_2O)_3(OH)_3](s) + 3CO_2(g) + 3H_2O(l)$, balanced with state symbols. **B1** for the 2+ ion, $[M(H_2O)_6]^{2+}(aq) + CO_3^{2-}(aq) \rightarrow MCO_3(s) + 6H_2O(l)$, the metal carbonate precipitating with no gas. **B1** the 3+ ion has a higher charge/size ratio, so it polarises the O–H bonds of its water ligands more strongly. **B1** H^+ is released more readily, so the 3+ solution is acidic enough to react with carbonate and release carbon dioxide. **B1** the 2+ ion is too weak an acid to release carbon dioxide, so its metal carbonate precipitates; the gas can be confirmed by limewater turning milky.

Levels

Level 3 (5–6 marks) both observations are given and matched to the right charge, both equations are correct and balanced, and the difference is explained through the charge/size ratio polarising the water ligands so that only the 3+ solution is acidic enough to release carbon dioxide from carbonate.

Level 2 (3–4 marks) effervescence with the 3+ ion and a carbonate precipitate with the 2+ ion are described and one equation is right; the other equation, or the explanation through polarisation and acidity, is thin.

Level 1 (1–2 marks) one observation is given, typically carbon dioxide from both ions or a carbonate precipitate from the 3+ ion; no correct equation and no reason for the difference.

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