InkChemistry

ANSWER BOOK

Oxidation states and the rules for assigning them

Redox · Unit C07

AQA 3.1.7 · OCR A 2.1.5

19 questions · 65 marks

NAVY EDITION

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SECTION B · Standard, 3 to 5 marks

B1 OSR-Q2

B1 sulfur in H_2SO_4 is +6, since 2(+1) + S + 4(-2) = 0. **B1** nitrogen in NO_2 is +4, since N + 2(-2) = 0. **B1** chlorine in ClO_3^- is +5, since Cl + 3(-2) = -1. **B1** manganese in MnO_4^{2-} is +6, since Mn + 4(-2) = -2. [4]

B2 OSR-Q6

B1 oxidation is the loss of electrons and reduction is the gain of electrons; reject definitions in terms of oxygen alone. **B1** each magnesium atom loses two electrons to form Mg^{2+}, so magnesium is oxidised from 0 to +2. **B1** each chlorine atom gains one electron to form Cl^-, so chlorine is reduced from 0 to -1. **B1** oxidation and reduction both occur, with electrons passing from magnesium to chlorine, so the reaction is redox. [4]

B3 OSR-Q9

B1 Cu_2O is copper(I) oxide, since each copper is +1. **B1** $NaClO_3$ is sodium chlorate(V), since chlorine is +5. **B1** iron(III) sulfate is $Fe_2(SO_4)_3$. **B1** potassium manganate(VII) is $KMnO_4$. [4]

B4 OSR-Q16

B1 phosphorus is +5: 4(+1) + 2P + 7(-2) = 0, so 2P = +10. **B1** nitrogen is -2: 2N + 4(+1) = 0. **B1** chlorine is +7: 2Cl + 7(-2) = 0, chlorine bonded to oxygen being positive. **B1** hydrogen is -1: calcium is +2 by the Group 2 rule, which outranks the rule for hydrogen in a metal hydride. [4]

B5 OSR-Q17

B1 bromine rises from -1 in Br^- to 0 in Br_2, so bromide is oxidised. **B1** sulfur falls from +6 in H_2SO_4 to +4 in SO_2, so the acid is reduced. **B1** one element is oxidised and another reduced in the same reaction, so it is redox; sodium, hydrogen and oxygen do not change. **B1** sulfuric acid accepts the electrons, so it is the oxidising agent. [4]

B6 OSR-Q3

M1 four oxygen atoms at -2 give -8, so the three iron atoms together are +8. **A1** mean oxidation state = +8/3, which is +2.67 to three significant figures; accept $+2\frac{2}{3}$. **B1** the oxide holds one Fe^{2+} for every two Fe^{3+}, and 2 + 3 + 3 = 8, so the fraction is the average of two real oxidation states. [3]

B7 OSR-Q4

M1 moles of vanadium = 0.100 × 25.0/1000 = 2.50×10^{-3} mol. **M1** electrons per vanadium = $5.00 \times 10^{-3} \div 2.50 \times 10^{-3} = 2$. **A1** the oxidation state falls by 2, from +5 to +3; it is an integer with no unit. [3]

B8 OSR-Q5

B1 fluorine is the most electronegative element and is always -1 in its compounds, a rule of higher priority than the rule for oxygen. **B1** in OF_2 two fluorines contribute -2, so by the sum rule oxygen must be +2. **B1** in H_2O oxygen is more electronegative than hydrogen, hydrogen is +1, so the two hydrogens give +2 and oxygen is -2. [3]

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B9 OSR-Q7

B1 +1 is the mean found by treating NH_4NO_3 as one unit: $2N + 4(+1) + 3(-2) = 0$. **B1** the two nitrogen atoms are in different ions, NH_4^+ and NO_3^-, and have different oxidation states. **B1** in NH_4^+ nitrogen is -3 and in NO_3^- it is +5, so no nitrogen atom is at +1. [3]

B10 OSR-Q8

B1 Fe_2O_3 is the oxidising agent: iron falls from +3 to 0, so it is reduced; accept iron(III) oxide, reject 'iron'. **B1** Al is the reducing agent: aluminium rises from 0 to +3, so it is oxidised. **B1** oxygen stays at -2 in Fe_2O_3 and in Al_2O_3, so it takes no part in the electron transfer. [3]

B11 OSR-Q11

B1 lithium is +1 and aluminium is +3, so the four hydrogens share -4 and each is -1. **B1** this is a metal hydride: hydrogen is bonded to less electronegative metals, so the metal rules take priority over the rule for hydrogen. **B1** hydrogen at -1 is readily oxidised to 0 or +1, donating electrons to the species it reacts with, so the compound is a reducing agent. [3]

B12 OSR-Q12

B1 in $S_2O_3^{2-}$ three oxygens give -6, so the two sulfurs share +4 and the mean is +2. **B1** in $S_4O_6^{2-}$ six oxygens give -12, so the four sulfurs share +10 and the mean is +2.5. **B1** the mean oxidation state of sulfur rises, so thiosulfate loses electrons to iodine, which is reduced from 0 to -1: thiosulfate is the electron donor, the reducing agent. [3]

B13 OSR-Q13

B1 all three agree for a reaction such as magnesium burning in oxygen, where magnesium gains oxygen, loses electrons and rises from 0 to +2. **B1** the oxygen definition fails for reactions with no oxygen, such as magnesium with chlorine, which the electron definition covers. **B1** the oxidation state definition extends the electron definition to covalent species such as SO_2, where electrons are shared rather than fully transferred. [3]

B14 OSR-Q14

B1 the claim is wrong: hydrogen is +1 in H_2O_2, in H^+ and in H_2O, so its oxidation state does not change. **B1** oxygen falls from -1 in the peroxide to -2 in water, so hydrogen peroxide is the species reduced. **B1** iodine rises from -1 to 0, so iodide is oxidised, and hydrogen peroxide is the oxidising agent. [3]

B15 OSR-Q15

B1 sulfur rises from -2 in H_2S to +4 in SO_2, a change of six, so sulfur is oxidised. **B1** oxygen falls from 0 in O_2 to -2 in SO_2 and in H_2O, so oxygen is reduced; hydrogen stays at +1. **B1** hydrogen sulfide donates the electrons, so H_2S is the reducing agent. [3]

B16 OSR-Q18

B1 with magnesium, $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$: magnesium rises from 0 to +2 and is oxidised. **B1** hydrogen falls from +1 in H^+ to 0 in H_2 and is reduced, so electrons pass from magnesium to hydrogen ions. **B1** with magnesium oxide, magnesium stays at +2, oxygen at -2 and hydrogen at +1, so no oxidation state changes and it is an acid–base reaction. [3]

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B17 OSR-Q19 **B1** chromium is +3: ammonia ligands count 0 and each chloride -1, so Cr + 2(-1) = +1. **B1** cobalt is +3: six neutral ammonia ligands add nothing, so the whole 3+ charge belongs to the cobalt. **B1** nickel is +2: four cyanide ligands give -4, so Ni - 4 = -2; the sign is required for each mark. [3]

B18 OSR-Q20 **B1** $Ba(s) + Cl_2(g) \rightarrow BaCl_2(s)$, with barium oxidised from 0 to +2 and chlorine reduced from 0 to -1. **B1** barium reacts more vigorously than magnesium. **B1** barium's outer electrons are further from the nucleus and more shielded, so its first and second ionisation energies are lower and it loses electrons more readily. [3]

SECTION C · Long, 6 marks or more

C1 OSR-Q21 **Indicative content** [6]

B1 manganese falls from +7 in $KMnO_4$ to +2 in $MnCl_2$, a gain of five electrons per manganese atom, so the manganate(VII) is reduced. **B1** ten of the sixteen chlorine atoms rise from -1 in HCl to 0 in Cl_2, losing one electron each, so those chloride ions are oxidised. **B1** the electrons balance: two manganese atoms gain ten electrons and ten chlorine atoms lose ten electrons. **B1** the other six chlorine atoms stay at -1 in KCl and $MnCl_2$, so that part of the HCl supplies H^+ and chloride ions as an acid would. **B1** potassium stays at +1, hydrogen at +1 and oxygen at -2 throughout, so none of them takes part in the electron transfer. **B1** $KMnO_4$ is therefore the oxidising agent, accepting electrons, and HCl is the reducing agent, donating them, as well as acting as the acid.

Levels

Level 3 (5–6 marks) all the oxidation states are assigned correctly, the reduction of manganese and the oxidation of chlorine are identified with the electron count balanced, the six chlorine atoms that stay at -1 are used to show the acid role, and the oxidising and reducing agents are named.

Level 2 (3–4 marks) manganese is shown to go from +7 to +2 and chlorine from -1 to 0; the electrons are not balanced, the six chlorines that stay at -1 are not used, or the agents are not named.

Level 1 (1–2 marks) one oxidation state change is right; HCl is often called the oxidising agent, and its role as an acid is not explained.

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