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# Topic 4.9: Introduction to Oxidation-Reduction Reactions

Worksheet · AP Chemistry

UNIT 4 · CHEMICAL REACTIONS

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## How to Use This Worksheet

Work in order. Section A checks the definitions, the oxidation number rules and the agent names. Section B practises assigning oxidation numbers, writing half-reactions and balancing in acid, including a disproportionation. Section C combines balancing with redox titration calculations in exam style.

Reference facts used throughout. OIL RIG: Oxidation Is Loss, Reduction Is Gain (of electrons). Oxidation numbers: free element 0; monatomic ion equal to its charge; fluorine −1; oxygen −2 except in peroxides (−1) and with fluorine (positive); hydrogen +1 except in metal hydrides (−1); the numbers in a species sum to 0 for a neutral compound and to the charge for an ion. Molar masses: Zn 65.38 g/mol, Cu 63.55 g/mol. Carry unrounded values through each calculation and round only the final answer, to three significant figures.

Three tools recur in nearly every question. Assign oxidation numbers before and after to find what changed. Name agents by the partner's fate: the oxidising agent is itself reduced, the reducing agent is itself oxidised. Balance redox equations in halves, and let the charge check tell you how many electrons each half needs.

One habit protects every answer in Section C: never assume a 1 to 1 ratio. Permanganate takes five electrons where dichromate takes six and iodine takes two, so the mole ratios differ each time. Read the ratio from the balanced equation in front of you before touching the numbers.

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Do not look at the answer key until you have attempted every question.

## Section A: Recall and Concepts

**A1.** Define each term in one sentence. (a) Oxidation. (b) Reduction. (c) Oxidising agent. (d) Reducing agent. (e) Oxidation number. (f) Disproportionation.

**A2.** Give the oxidation number of the element named in each species. (a) S in $H_2SO_4$ (b) N in $NO_3^-$ (c) Mn in $MnO_4^-$ (d) O in $H_2O_2$ (e) H in NaH (f) Fe in $Fe_2O_3$ (g) C in $CO_2$ (h) S in $H_2S$

**A3.** For each change, state whether it is oxidation or reduction, and give the number of electrons transferred per particle. (a) $Zn \rightarrow Zn^{2+}$ (b) $Cu^{2+} \rightarrow Cu$ (c) $Fe^{2+} \rightarrow Fe^{3+}$ (d) $Cl_2 \rightarrow 2Cl^-$ (e) $2I^- \rightarrow I_2$ (f) $MnO_4^- \rightarrow Mn^{2+}$

**A4.** Consider $2Na + Cl_2 \rightarrow 2NaCl$. (a) Which species is oxidised? (b) Which species is reduced? (c) Name the oxidising agent. (d) Name the reducing agent. (e) State the oxidation number of chlorine before and after the reaction.

**A5.** Decide whether each reaction is redox, justifying your answer with oxidation numbers. (a) $CaCO_3 \rightarrow CaO + CO_2$ (b) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ (c) $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$ (d) $Mg + 2HCl \rightarrow MgCl_2 + H_2$ (e) $HCl + NaOH \rightarrow NaCl + H_2O$

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## Section B: Core Practice

**B1.** Show the sum equation and work out each oxidation number. (a) Cr in $Cr_2O_7^{2-}$ (b) Cl in $ClO_3^-$ (c) S in $SO_3^{2-}$ (d) P in $H_3PO_4$ (e) Mn in $MnO_2$ (f) N in $NH_3$

**B2.** Zinc metal is placed in copper(II) sulfate solution: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$. (a) Write the oxidation half-reaction. (b) Write the reduction half-reaction. (c) Name the oxidising agent and the reducing agent. (d) A 5.00 g piece of zinc reacts completely with excess $Cu^{2+}$. Calculate the mass of copper formed. (e) State the oxidation number of zinc and of copper before and after the reaction.

**B3.** Iron(II) ions are oxidised by acidified permanganate. (a) State the oxidation number of manganese in $MnO_4^-$ and in $Mn^{2+}$, and hence the number of electrons gained per $MnO_4^-$ ion. (b) State the number of electrons lost per $Fe^{2+}$ ion as it becomes $Fe^{3+}$. (c) Deduce the ratio $Fe^{2+}$ : $MnO_4^-$ in the balanced equation. (d) Write the two half-reactions. (e) Combine them into the full balanced equation, and check the total charge on each side.

**B4.** Build the reduction half-reaction for dichromate in acidic solution, $Cr_2O_7^{2-} \rightarrow Cr^{3+}$, step by step. (a) Balance the chromium atoms. (b) Balance the oxygen atoms using $H_2O$. (c) Balance the hydrogen atoms using $H^+$. (d) Work out how many electrons must be added, and on which side. (e) Verify the final half-reaction by checking atoms and charge. (f) Combine your half-reaction with $Fe^{2+} \rightarrow Fe^{3+} + e^-$ to write the full balanced equation for dichromate oxidising iron(II).

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**B5.** Hydrogen peroxide decomposes: $2H_2O_2 \rightarrow 2H_2O + O_2$. (a) Give the oxidation number of oxygen in $H_2O_2$, in $H_2O$ and in $O_2$. (b) Explain why this reaction is a disproportionation. (c) Chlorine reacts with cold dilute alkali: $Cl_2 + 2OH^- \rightarrow Cl^- + ClO^- + H_2O$. Give the oxidation number of chlorine in $Cl_2$, in $Cl^-$ and in $ClO^-$, and state whether this is also a disproportionation. (d) In $OF_2$, oxygen has oxidation number +2. Explain why the usual −2 rule fails in this compound.

**B6.** Two more transfers. (a) In $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$, identify the species oxidised, the species reduced, and both agents. (b) How many electrons does each copper atom transfer? (c) In $2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$, how many electrons are lost per aluminium atom, how many gained per copper ion, and how many electrons are transferred in the equation as written? (d) In $Mg + 2H^+ \rightarrow Mg^{2+} + H_2$, name the oxidising agent and justify your choice. (e) In $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$, identify the species oxidised and reduced, and name both agents.

## Section C: Exam Style

**C1.** A 25.0 mL sample of an iron(II) solution is acidified and titrated with 0.0200 mol/L potassium permanganate. The endpoint, the first permanent purple tinge, comes at a titre of 24.0 mL. The equation is $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$. (a) Calculate the moles of $MnO_4^-$ used. (b) Calculate the moles of $Fe^{2+}$ in the sample. (c) Calculate the concentration of the iron(II) solution. (d) Explain why no indicator is needed. (e) Explain why the sample must be acidified for this equation, and its 5 to 1 ratio, to apply. (f) Calculate the mass of iron in the sample. (Molar mass of Fe: 55.85 g/mol.)

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**C2.** Acidified potassium dichromate also oxidises iron(II): $Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \rightarrow 2Cr^{3+} + 7H_2O + 6Fe^{3+}$. A 25.0 mL sample of an iron(II) solution needs 15.0 mL of 0.0200 mol/L dichromate solution to reach the endpoint. (a) Calculate the moles of $Cr_2O_7^{2-}$ used. (b) Calculate the moles of $Fe^{2+}$ in the sample. (c) Calculate the concentration of the iron(II) solution. (d) State the colour change at the endpoint region, and the oxidation number change for chromium that it signals. (e) A student uses a 1 to 1 ratio for part (b). By what factor is the student's answer wrong?

**C3.** Iodine solution is titrated with sodium thiosulfate: $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$. A 25.0 mL sample of iodine solution reacts exactly with 20.0 mL of 0.100 mol/L thiosulfate solution. (a) Write the oxidation half-reaction and the reduction half-reaction. (b) Name the oxidising agent and the reducing agent. (c) Calculate the moles of thiosulfate used. (d) Calculate the moles of $I_2$ in the sample. (e) Calculate the concentration of the iodine solution. (f) Starch solution is added near the endpoint. Describe the colour change seen at the endpoint and name the species responsible for the colour.

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## Answer Key

**A1.** (a) The loss of electrons by a species; its oxidation number increases. (b) The gain of electrons by a species; its oxidation number decreases. (c) The species that takes electrons from another species, causing oxidation; it is itself reduced. (d) The species that gives electrons to another species, causing reduction; it is itself oxidised. (e) A bookkeeping number assigned by fixed rules that tracks how electrons are distributed in a species; the values in a species sum to its overall charge. (f) A redox reaction in which the same species is both oxidised and reduced.

**A2.** (a) +6: 2(+1) + S + 4(−2) = 0. (b) +5: N + 3(−2) = −1. (c) +7: Mn + 4(−2) = −1. (d) −1: peroxide exception, 2(+1) + 2(O) = 0. (e) −1: metal hydride exception, (+1) + H = 0. (f) +3: 2(Fe) + 3(−2) = 0. (g) +4: C + 2(−2) = 0. (h) −2: 2(+1) + S = 0.

**A3.** (a) Oxidation: the atom loses 2 electrons, number rising 0 to +2. (b) Reduction: the ion gains 2 electrons, number falling +2 to 0. (c) Oxidation: 1 electron lost, +2 to +3. (d) Reduction: each chlorine atom gains 1 electron, 0 to −1, so 2 electrons per $Cl_2$ molecule. (e) Oxidation: each iodide loses 1 electron, −1 to 0, so 2 electrons per $I_2$ formed. (f) Reduction: 5 electrons gained, manganese falling +7 to +2.

**A4.** (a) Sodium, Na: it loses electrons, 0 to +1. (b) Chlorine, $Cl_2$: it gains electrons, 0 to −1. (c) $Cl_2$, the species that is reduced. (d) Na, the species that is oxidised. (e) 0 in $Cl_2$, −1 in NaCl.

**A5.** (a) Not redox: Ca is +2, C is +4 and O is −2 on both sides; no number changes. (b) Redox: carbon rises from −4 in $CH_4$ to +4 in $CO_2$, and oxygen falls from 0 in $O_2$ to −2 in the products. (c) Not redox: Ag stays +1, Na +1, Cl −1 and N +5 throughout; the ions only swap partners. (d) Redox: magnesium rises 0 to +2 and hydrogen falls +1 to 0. (e) Not redox: H is +1, Cl −1, Na +1 and O −2 on both sides; a proton moves, but no oxidation number changes.

**B1.** (a) 2Cr + 7(−2) = −2, so 2Cr = +12 and Cr = +6. (b) Cl + 3(−2) = −1, so Cl = +5. (c) S + 3(−2) = −2, so S = +4. (d) 3(+1) + P + 4(−2) = 0, so P = +5. (e) Mn + 2(−2) = 0, so Mn = +4. (f) N + 3(+1) = 0, so N = −3.

**B2.** (a) $Zn \rightarrow Zn^{2+} + 2e^-$. (b) $Cu^{2+} + 2e^- \rightarrow Cu$. (c) Oxidising agent: $Cu^{2+}$, because it gains the electrons and is reduced. Reducing agent: Zn, because it loses the electrons and is oxidised. (d) Moles of Zn = 5.00 ÷ 65.38 = 0.0765 mol. The equation runs 1 to 1, so moles of Cu = 0.0765 mol. Mass of Cu = 0.0765 × 63.55 = 4.86 g. (e) Zinc: 0 in Zn metal, +2 in $Zn^{2+}$. Copper: +2 in $Cu^{2+}$, 0 in Cu metal.

**B3.** (a) Manganese is +7 in $MnO_4^-$ (Mn + 4(−2) = −1) and +2 in $Mn^{2+}$, a fall of 5, so each $MnO_4^-$ gains 5 electrons. (b) Each $Fe^{2+}$ loses 1 electron. (c) Electrons must balance: 5 $Fe^{2+}$ supply the 5 electrons one $MnO_4^-$ takes, so the ratio $Fe^{2+}$ : $MnO_4^-$ is 5 : 1. (d) Oxidation: $Fe^{2+} \rightarrow Fe^{3+} + e^-$. Reduction: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$. (e) $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$. Charge check: left, −1 + 8 + 5(+2) = +17; right, +2 + 5(+3) = +17. Atoms also match: Mn 1, O 4, H 8, Fe 5 on each side.

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**B4.** (a) $Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$. (b) Add $7H_2O$ to the right: $Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$. (c) Add $14H^+$ to the left: $Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$. (d) Charge before electrons: left $-2 + 14 = +12$; right $2(+3) = +6$. The left must fall by 6, so add $6e^-$ to the left. (e) $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$. Atoms: Cr 2, O 7, H 14 on each side. Charge: $-2 + 14 - 6 = +6$ on the left, +6 on the right. Balanced. (f) The iron(II) half gives 1 electron, so multiply it by 6 and add: $Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \rightarrow 2Cr^{3+} + 7H_2O + 6Fe^{3+}$. The 6 electrons cancel, and the charge is +24 on each side.

**B5.** (a) $H_2O_2$: $-1$ (peroxide). $H_2O$: $-2$. $O_2$: 0. (b) The same element, oxygen, in the same reactant moves in both directions at once: some atoms fall from $-1$ to $-2$ and are reduced, others rise from $-1$ to 0 and are oxidised. One species doing both jobs is disproportionation. (c) $Cl_2$: 0. $Cl^-$: $-1$. $ClO^-$: +1, since Cl + $(-2)$ = $-1$. Yes: chlorine starts at 0 and splits to $-1$ and +1, so it is both reduced and oxidised in the same reaction. (d) Fluorine is more electronegative than oxygen and always takes $-1$. With two fluorines at $-1$, oxygen must be +2 for the sum to reach 0: O + $2(-1)$ = 0. The $-2$ rule assumes oxygen is the most electronegative partner present, which is false here.

**B6.** (a) Oxidised: Cu, 0 to +2. Reduced: $Ag^+$, +1 to 0. Oxidising agent: $Ag^+$. Reducing agent: Cu. (b) 2 electrons per copper atom, matching the two $Ag^+$ ions each taking one. (c) Each aluminium atom loses 3 electrons; each copper ion gains 2; in the equation as written, $2 \times 3 = 6$ electrons are transferred, equal to $3 \times 2$ on the gaining side. (d) $H^+$: it gains electrons to form $H_2$ (falling from +1 to 0) and is therefore reduced, which makes it the oxidising agent; Mg is the reducing agent. (e) Oxidised: $Br^-$, rising from $-1$ to 0 as $Br_2$. Reduced: $Cl_2$, falling from 0 to $-1$. Oxidising agent: $Cl_2$. Reducing agent: $Br^-$.

**C1.** (a) Moles of $MnO_4^-$ = $0.0200 \times (24.0 \div 1000) = 0.0200 \times 0.0240 = 4.80 \times 10^{-4}$ mol. (b) The ratio $Fe^{2+}$ : $MnO_4^-$ is 5 : 1, so moles of $Fe^{2+} = 5 \times 4.80 \times 10^{-4} = 2.40 \times 10^{-3}$ mol. (c) Concentration = $2.40 \times 10^{-3} \div 0.0250 = 0.0960$ mol/L. (d) $MnO_4^-$ is purple and its product $Mn^{2+}$ is effectively colourless, so the titrant is its own indicator: the endpoint is the first permanent purple tinge from one excess drop. (e) $H^+$ is a reactant in the equation, consumed in an 8 to 1 ratio to $MnO_4^-$. Without acid in good excess the reduction does not follow this equation to $Mn^{2+}$, so the 5 to 1 ratio would not hold. (f) Mass of Fe = $2.40 \times 10^{-3} \times 55.85 = 0.134$ g.

**C2.** (a) Moles of $Cr_2O_7^{2-}$ = $0.0200 \times 0.0150 = 3.00 \times 10^{-4}$ mol. (b) The ratio $Fe^{2+}$ : $Cr_2O_7^{2-}$ is 6 : 1, so moles of $Fe^{2+} = 6 \times 3.00 \times 10^{-4} = 1.80 \times 10^{-3}$ mol. (c) Concentration = $1.80 \times 10^{-3} \div 0.0250 = 0.0720$ mol/L. (d) The orange colour of $Cr_2O_7^{2-}$ fades as green $Cr^{3+}$ forms; chromium falls from +6 to +3. (e) Six times too small: the student reports $3.00 \times 10^{-4}$ mol instead of $1.80 \times 10^{-3}$ mol, an answer wrong by a factor of 6.

**C3.** (a) Oxidation: $2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-$. Reduction: $I_2 + 2e^- \rightarrow 2I^-$. (b) Oxidising agent: $I_2$, which gains the electrons. Reducing agent: $S_2O_3^{2-}$, which loses them. (c) Moles of thiosulfate = $0.100 \times 0.0200 = 2.00 \times 10^{-3}$ mol. (d) The ratio $I_2$ : $S_2O_3^{2-}$ is 1 : 2, so moles of $I_2 = 2.00 \times 10^{-3} \div 2 = 1.00 \times 10^{-3}$ mol. (e) Concentration = $1.00 \times 10^{-3} \div 0.0250 = 0.0400$ mol/L. (f) The blue-black colour of the starch

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and iodine complex disappears, leaving the solution colourless, at the drop where the last of the $I_2$ is reduced to $I^-$; the colour is caused by $I_2$ held in the starch.

**Continue Revising**

Topic 4.9 Revision Notes

Topic 4.9 Quiz: Introduction to Oxidation-Reduction Reactions

Previous Worksheet: Topic 4.8 Introduction to Acid-Base Reactions

Next Worksheet: Topic 5.1 Reaction Rates

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