

IB HL Redox 1

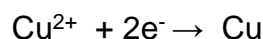
- 1 Define *oxidation* and *reduction* in terms of electrons.
- 2 Define *oxidation* and *reduction* in terms of loss/gain of hydrogen/oxygen.
- 3 Define *oxidation* and *reduction* in terms of oxidation state (oxidation number).
- 4 Determine the oxidation states of all elements in the following species:

P ₄	SO ₂	SO ₃	SO ₄ ²⁻	HCl	HClO ₃	ClO ₄ ⁻	KMnO ₄	C ₂ H ₅ Cl
Ca	CrO ₄ ²⁻	Cr ₂ O ₇ ²⁻	H ₂ O ₂	BaO ₂	LiH	LiAlH ₄	Na ₂ S ₂ O ₃	S ₄ O ₆ ²⁻

- 5 Name the following molecules/ions/compounds using oxidation states/numbers:

SO ₂	SO ₃	Cr ₂ O ₇ ²⁻	FeCl ₂	CuSO ₄	KMnO ₄	NaClO ₄	Cu ₂ O	Fe ₂ (SO ₄) ₃
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- 6 Explain whether each of the following involves oxidation, reduction or neither:

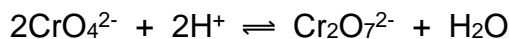
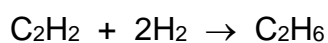
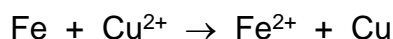
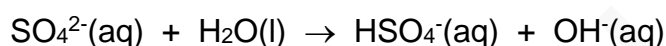
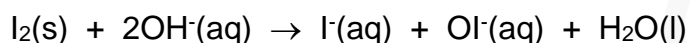


The conversion of Mn₂O₃ to MnO₂

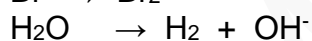
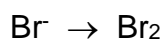
The conversion of CH₃CH₂OH to CH₃CHO

The conversion of CH₃CH₂CHO to CH₃CH₂COOH

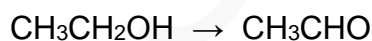
- 7 Explain which of the following equations represent Redox reactions



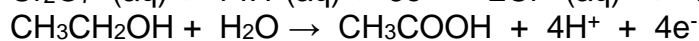
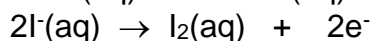
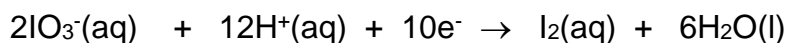
- 8 Balance the following half-equations in neutral solution:



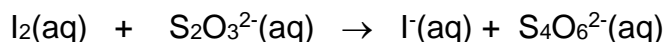
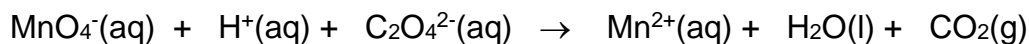
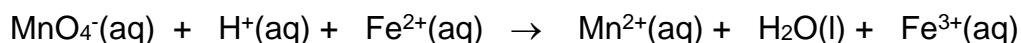
- 9 Balance the following half-equations in acidic solution:



- 10 Write balanced redox reactions from each pair of half equations



11 Balance the following redox equations:



12 Define the terms *oxidising agent* and *reducing agent*.

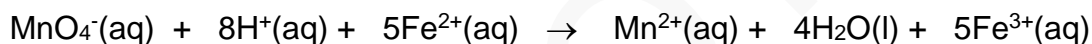
13 Identify the oxidising and reducing agents in the equations in 11.

14 Explain whether A or Q is more reactive given the equation $\text{A}(\text{s}) + \text{Q}^{2+}(\text{aq}) \rightarrow \text{A}^{2+}(\text{aq}) + \text{Q}(\text{s})$

15 Complete the following table by predicting whether a 'reaction' or 'no reaction' will occur in each case and write an ionic equation for each **reaction**.

	ZnSO ₄ (aq)	MgCl ₂ (aq)	CuSO ₄ (aq)
Zinc			
Magnesium			
copper			

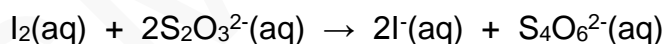
16 Solve redox titration problems involving iron(II) and manganate(VII)



(a) 25.00 cm³ of 1 mol dm⁻³ sulfuric acid was added to 25.00 cm³ of a solution of iron(II) sulfate. This was then titrated against 0.02000 mol dm⁻³ potassium manganate(VII). 22.50 cm³ of KMnO₄ was required to be added to the mixture to give the first permanent pink colour. Calculate the concentration of the original iron(II) sulfate solution.

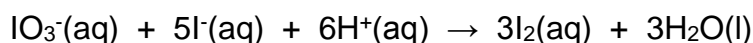
(b) Five iron tablets are dissolved in about 100 cm³ of 1.0 mol dm⁻³ sulfuric acid and the solution is made up to a total volume of 250 cm³ with distilled water. 25.0 cm³ of the iron(II) solution is titrated with 0.0100 mol dm⁻³ potassium manganate(VII) solution until the first permanent pink colour is seen. 11.60 cm³ of KMnO₄ was required. Calculate the mass of iron (in mg) in 1 tablet.

17 Solve redox titration problems involving sodium thiosulfate and iodine

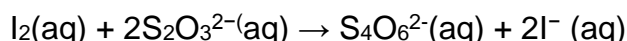
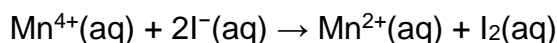
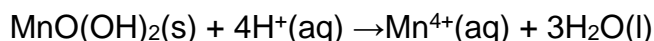
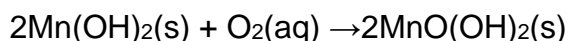


(a) 25.00 cm³ of iodine solution is pipetted into a conical flask. 0.100 mol dm⁻³ sodium thiosulfate solution is added until the mixture becomes pale yellow. Starch indicator is added and the titration continued until the reaction mixture becomes colourless. The total volume of sodium thiosulfate solution required was 21.50 cm³. Calculate the concentration of the iodine solution.

(b) x g of potassium iodate(V) is dissolved in distilled water and made up to a total volume of 250.0 cm³. 25.00 cm³ of this solution is pipetted into a conical flask and 20 cm³ of 1 mol dm⁻³ sulfuric acid added – this is excess. Excess potassium iodide solution is added to the reaction mixture and the liberated iodine titrated against sodium thiosulfate solution. 25.30 cm³ of 0.100 mol dm⁻³ sodium thiosulfate is required for the titration. Determine the value of x.



18 Calculate BOD from titrations using the Winkler Method



- (a) The Winkler method was used to measure the concentration of dissolved oxygen in a sample of water. Manganese(II) sulfate, sulfuric acid and potassium iodide were added to 50.0 cm³ of the water. The iodine that was formed was titrated against a sodium thiosulfate solution with a concentration of 2.00x10⁻³ mol dm⁻³. It was found that 17.40 cm³ of sodium thiosulfate was required for the titration.

Another sample of water from the same source was incubated for 5 days at 20 °C. At the end of the incubation the Winkler method was used to determine concentration of dissolved oxygen. It was found that the concentration of dissolved oxygen in the sample was 3.50 ppm.

Calculate the concentration of dissolved oxygen (in ppm) in the original sample and the biochemical oxygen demand.

- (b) The Winkler method was used to measure the concentration of dissolved oxygen in a sample of water. Manganese(II) sulfate, sulfuric acid and potassium iodide were added to 25.0 cm³ of the water. The iodine that was formed was titrated against a sodium thiosulfate solution with a concentration of 1.00x10⁻³ mol dm⁻³. It was found that 22.30 cm³ of sodium thiosulfate was required for the titration.

Another sample of water from the same source was incubated for 5 days at 20 °C. At the end of the incubation the Winkler method was used to determine concentration of dissolved oxygen. It was found that the concentration of dissolved oxygen in the sample was 2.80 ppm.

Calculate the concentration of dissolved oxygen in the original sample and the biochemical oxygen demand.

- 19 Explain the connection between biochemical oxygen demand and the degree of pollution of a water sample.
- 20 (a) Draw a labelled diagram showing a voltaic (Galvanic) cell made up of Zn/Zn²⁺ and Cu/Cu²⁺ half-cells. Labelling should include the anode, cathode, positive and negative electrodes and show the direction of electron flow in the external circuit and the direction of ion flow in the solutions.
- (b) Write half-equations for the reactions at each electrode
- (c) Write the overall redox reaction that occurs in the cell.
- (d) State whether the overall reaction that occurs is exothermic or endothermic.
- 21 Explain the purpose of the salt bridge in a voltaic (Galvanic) cell.
- 22 State the energy conversion that occurs in a voltaic cell.
- 23 Explain what is meant by the term *standard electrode potential* and *standard cell potential*
- 24 Describe, with the aid of a labelled diagram, the standard hydrogen electrode
- 25 Write the half-equation for the standard hydrogen electrode and state the value of the standard electrode potential.

- 26 Explain, with the aid of a diagram, how the standard electrode potential for Cu^{2+}/Cu could be measured
- 27 State the significance of the fact that the standard electrode potential for a Cu^{2+}/Cu half-cell is positive but that for a Zn^{2+}/Zn half-cell is negative.
- 28 For each of the following cells
- calculate the cell potential,
 - write an overall equation for the reaction that goes on in the cell
 - state which is the positive and which the negative electrode
 - identify the anode and cathode
 - state which way electrons flow in the external circuit
 - state which way ions flow in the salt bridge

$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Ni}(\text{s})$	$\text{Ag}^{+}(\text{aq}) + \text{e}^{-} \rightleftharpoons \text{Ag}(\text{s})$
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Zn}(\text{s})$	$\text{Sn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Sn}(\text{s})$
$\text{I}_2(\text{s}) + 2\text{e}^{-} \rightleftharpoons 2\text{I}^{-}(\text{aq})$	$\text{Fe}^{3+}(\text{aq}) + \text{e}^{-} \rightleftharpoons \text{Fe}^{2+}(\text{aq})$
$\text{Cl}_2(\text{g}) + 2\text{e}^{-} \rightleftharpoons 2\text{Cl}^{-}(\text{aq})$	$\text{Cl}_2(\text{g}) + 2\text{e}^{-} \rightleftharpoons 2\text{Cl}^{-}(\text{aq})$

- 29 For each of the cells in 28 write the cell diagram using the cell diagram convention
- 30 Explain using electrode potentials below whether the reactions below will be spontaneous or not
- $$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^{+}(\text{aq}) + 6\text{Fe}^{2+}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l}) + 6\text{Fe}^{3+}(\text{aq})$$
- $$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^{+}(\text{aq}) + 6\text{Br}^{-}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l}) + 3\text{Br}_2(\text{l})$$
- $$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^{+}(\text{aq}) + 6\text{F}^{-}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l}) + 3\text{F}_2(\text{g})$$
- 31 Given the electrode potentials below select the strongest reducing agent and the strongest oxidising agent.
- $$\text{U}^{4+} + \text{e}^{-} \rightleftharpoons \text{U}^{3+} \quad E^{\circ} = -0.61 \text{ V}$$
- $$\text{Eu}^{3+} + \text{e}^{-} \rightleftharpoons \text{Eu}^{2+} \quad E^{\circ} = -0.43 \text{ V}$$
- $$\text{Ho}^{3+} + \text{e}^{-} \rightleftharpoons \text{Ho} \quad E^{\circ} = -2.32 \text{ V}$$
- $$\text{Ce}^{4+} + \text{e}^{-} \rightleftharpoons \text{Ce}^{3+} \quad E^{\circ} = +1.61 \text{ V}$$
- $$\text{Pu}^{4+} + \text{e}^{-} \rightleftharpoons \text{Pu}^{3+} \quad E^{\circ} = +0.97 \text{ V}$$
- 32 Using the values in 31 explain whether the following are true or false:
- Eu^{3+} will oxidise Pu^{3+}
 - U^{3+} will reduce Ce^{4+}
 - Pu^{4+} will oxidise U^{3+}
- 33 For each of the following reactions calculate the cell potential and ΔG and predict whether the reaction is spontaneous.
- $\text{Mn}(\text{s}) + \text{Sn}^{2+}(\text{aq}) \rightarrow \text{Mn}^{2+}(\text{aq}) + \text{Sn}(\text{s})$
 - $\frac{1}{2}\text{Br}_2(\text{l}) + \text{Fe}^{2+}(\text{aq}) \rightarrow \text{Br}^{-}(\text{aq}) + \text{Fe}^{3+}(\text{aq})$
 - $\text{Br}_2(\text{l}) + 2\text{Fe}^{2+}(\text{aq}) \rightarrow 2\text{Br}^{-}(\text{aq}) + 2\text{Fe}^{3+}(\text{aq})$
 - $\text{MnO}_4^{-}(\text{aq}) + 8\text{H}^{+}(\text{aq}) + 5\text{Fe}^{2+}(\text{aq}) \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l}) + 5\text{Fe}^{3+}(\text{aq})$
 - $2\text{MnO}_4^{-}(\text{aq}) + 16\text{H}^{+}(\text{aq}) + 10\text{F}^{-}(\text{aq}) \rightarrow 2\text{Mn}^{2+}(\text{aq}) + 8\text{H}_2\text{O}(\text{l}) + 5\text{F}_2(\text{g})$