Chemistry Unit 3+4 REDOX part 1 Test

Multiple choice -10 marks

1. In which of the following equations is the underlined species being oxidised?

(a) $\underline{Ca}^{2+}_{(aq)} CO^{2-}_{3\,(aq)} \rightarrow CaCO_{3(s)}$

(b) $Zn^{2+}_{(aq)} + \underline{Fe}_{(s)} \rightarrow Zn_{(s)} + Fe^{2+}_{(aq)}$

(c) $2\underline{H}^{+}_{(aq)} + Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + H_{2(g)}$

(d) $2I^{-}_{(aq)} + \underline{Br}_{2(g)} \rightarrow I_{2(aq)} + 2Br^{-}_{(aq)}$

2. Which of the following statements about oxidising and reducing agents is false?

(a) Bromine water can oxidise chloride ions to chlorine.

(b) Hydrogen peroxide solution is capable of spontaneous disproportionation.

(c) Group I metals are good reducing agents.

(d) Copper metal will react with a dilute silver nitrate solution.

3. The oxidation number for iodine in the iodate ion (IO^{3-}) is:

(a) +1

(b) -1

(c) +5

(d) -5

4. A student made the following observations:

(i) Clean metal A did not react with 1.0M B^{2+}

(ii) Clean metal B dissolved in 1.0M C^{2+} and crystals of C appeared

(iii) Clean metal C did not react with 1.0M A^{2+}

The order of strength **as a reducing agent** is

(a) A > B > C

(b) A > C > B

(c) B > C > A

(d) B > A > C

5. Which of the following is an example of an oxidation-reduction reaction?

(a) $2\ K_2CrO_4 + H_2SO_4 \rightarrow K_2Cr_2O_7 + K_2SO_4 + H_2O$

(b) $CaC_2 + 2\ H_2O \rightarrow Ca(OH)_2 + C_2H_2$

(c) $2\ Na + Cl_2 \rightarrow 2\ NaCl$

(d) $BaSO_3 + 2\ HCl \rightarrow BaCl_2 + H_2O + SO_2$

6. Which of the following statements about oxidation and reduction is FALSE?

(a) Oxidation and reduction occur simultaneously.

(b) The oxidising agent is reduced.

(c) More electrons are produced by the substance being oxidised than accepted by the substance being reduced.

(d) The reducing agent loses electrons in an oxidation-reduction reaction.

7. In which of the following reactions is the manganese containing species acting as a reducing agent?

(a) $MnO + Mg \longrightarrow Mn + MgO$

(b) $MnCl_2 + 2H_2O + Cl_2 \longrightarrow MnO_2 + 4Cl^- + 4H^+$

(c) $MnO_2 + 2Ag + 4H^+ \longrightarrow Mn^{2+} + 2Ag^+ + 2H_2O$

(d) $MnO_4^- + 5Fe^{2+} + 8H^+ \longrightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

8. Which one of the following reactions will be spontaneous under standard conditions?

a) $Cr_2O_7^{2-}{}_{(aq)} + 3H_2O_{2(l)} + 8H^+ \rightarrow 2Cr^{3+}{}_{(aq)} + H_2O_{(l)} + 3O_{2(g)}$

b) $3O_{2(g)} + 4Au_{(s)} + 12H^+ \rightarrow 4Au^{3+}{}_{(aq)} + 6H_2O_{(l)}$

c) $2Ag^+{}_{(aq)} + 2Br^- \rightarrow 2Ag_{(s)} + Br_{2(l)}$

d) $2Cl^-{}_{(aq)} + I_{2(s)} \rightarrow Cl_{2(g)} + 2I^-{}_{(aq)}$

9. Consider the incomplete chemical equation shown below.

$$Cr(s) + ClO_3^-(aq) + H^+(aq) \rightarrow Cr^{3+}(aq) + HClO_2(aq) + H_2O(l)$$

When this redox reaction is completed and balanced correctly (using whole numbers), the coefficient in front of $H^+(aq)$ will be;

(a) 1
(b) 3
(c) 6
(d) 9

10. Rank the following substances in order of increasing **nitrogen** oxidation number (i.e. from species with nitrogen in lowest oxidation state to highest oxidation state).

$\mathbf{N}O_3^-$ $\mathbf{N}_2O$ $HN\mathbf{O}_2$ $\mathbf{N}H_4^+$ $\mathbf{N}_2$

(a)	NH_4^+	<	N_2	<	N_2O	<	HNO_2	<	NO_3^-
(b)	NO_3^-	<	N_2O	<	HNO_2	<	N_2	<	NH_4^+
(c)	NH_4^+	<	HNO_2	<	N_2	<	NO_3^-	<	N_2O
(d)	N_2	<	NH_4^+	<	NO_3^-	<	N_2O	<	HNO_2

11. Consider the following reaction between cobalt metal and hydrochloric acid.

$$Co(s) + 2\,H^+(aq) \rightarrow Co^{2+}(aq) + H_2(g)$$

Which of the following statements is **correct**?

(a) Electrons are transferred from Co(s) to $H^+(aq)$.
(b) Electrons are transferred from $H^+(aq)$ to Co(s).
(c) Both Co(s) and $H^+(aq)$ will each gain and lose some electrons.
(d) Electrons are not transferred, as this is not a redox reaction.

12. In which of the following species is the oxidation state of sulfur the lowest?

(a) SO_3

(b) SO_3^{2-}

(c) $S_2O_4^{2-}$

(d) $S_2O_6^{2-}$

Questions 13 and 14 refer to the information below.

Corrosion occurs when metals are oxidised by coming into contact with oxygen. This process is increased in the presence of water or acidic and basic conditions. When iron is corroded it often forms rust.

13. Which of the following is **correct** in relation to the rusting of iron metal?

(a) Oxygen gas is the reducing agent / reductant
(b) Liquid water is the reducing agent / reductant
(c) The oxidation number of iron would decrease
(d) The presence of salt water would increase the rate of rusting

14. Using your table of standard reduction potentials, choose the metal that is **not** likely to corrode under standard conditions.

(a) Zn
(b) Ni
(c) Pb
(d) Au

15. Which of the following halogen displacement reactions would **not** occur under standard conditions?

(a) $Cl_2(aq) + 2\,Br^-(aq) \rightarrow 2\,Cl^-(aq) + Br_2(aq)$
(b) $I_2(aq) + 2\,Br^-(aq) \rightarrow 2\,I^-(aq) + Br_2(aq)$
(c) $Cl_2(aq) + 2\,I^-(aq) \rightarrow 2\,Cl^-(aq) + I_2(aq)$
(d) $Br_2(aq) + 2\,I^-(aq) \rightarrow 2\,Br^-(aq) + I_2(aq)$

Short answer - 36 marks

1. a. $Mg \rightarrow Mg^{2+}$ or MgO (½ each)
   b. $H_2O \rightarrow H_2$ (½ each)
   c. H_2O (1 mark)
   d. Mg (1 mark)

(4 marks)

2. Two beakers contained separate samples of zinc bromide solution, $ZnBr_2(aq)$. To one beaker a piece of tin metal, Sn(s), was added. To the second beaker a piece of magnesium metal, Mg(s), was added. In one beaker, a reaction took place, while in the other beaker no reaction was observed.

(a) Which of these metals (i.e. magnesium or tin) is the strongest reducing agent? Explain your answer.

(2 marks)

- **Magnesium**
- **E^0 value of oxidation of magnesium is 2.36 V compared to 0.14 V for tin, therefore a stronger tendency to be oxidised and thus a stronger reducing agent**

(b) Write a balanced chemical equation for the reaction that does occur, and explain why no reaction is observed in the other beaker.

(2 marks)

- **$Mg(s) + Zn^{2+}(aq) \rightarrow Mg^{2+}(aq) + Zn(s)$**
  OR
- **$Mg(s) + ZnBr_2(aq) \rightarrow MgBr_2(aq) + Zn(s)$**

- **No reaction observed with tin metal as Sn^{2+} has a higher reduction potential than Zn^{2+} / Sn^{2+} is a stronger oxidant than Zn^{2+} / Zn is a stronger reductant than Sn**

Some chlorine water, $Cl_2(aq)$, was added to a separate third sample of zinc bromide solution.

(c) Explain what would occur when these solutions were mixed. Include in your answer the expected observations.

(2 marks)

- **a halogen displacement reaction would occur**
- **pale yellow and colourless solutions would mix to produce an orange solution**

3. Tin is a metallic element located in Group 14 of the periodic table. It is used to make many different alloys such as bronze and solder, as well as finding application in the plating of steel to produce 'tin cans' for storage.

A chemistry student had 1.0 mol L^{-1} solutions of the following four substances;

$Ni(NO_3)_2$ $Zn(NO_3)_2$ $Pb(NO_3)_2$ $Mg(NO_3)_2$

(a) Which of these solutions could **not** be stored in a tin container? Explain your answer using a relevant chemical equation.

(3 marks)

- **$Pb(NO_3)_2$**
- **metal displacement reaction would occur OR Pb^{2+} has higher E^0 value than Sn^{2+} and will cause oxidation of tin container**
- **$Pb^{2+}(aq) + Sn(s) \rightarrow Pb(s) + Sn^{2+}(aq)$**

4. In acidic conditions the chlorate ion $ClO_{3(aq)}^-$ undergoes disproportion to $Cl_{2(g)}$ and $ClO_{4(aq)}^-$. Write appropriate half equations for this redox, labelling them oxidation and reduction, and the overall equation. **Include states.**

(4 marks)

Reduction half equation $2ClO_3^- + 12H^+ + 10e^- \rightarrow Cl_2 + 6H_2O$ (1 mark)

Oxidation half equation $5X(ClO_3^- + H_2O \rightarrow ClO_4^- + 2H^+ + 2e^-)$ (1 mark)

Full equation $7ClO_{3(aq)}^- + 2H^+_{(aq)} \rightarrow Cl_{2(g)} + 5ClO_{4(aq)}^- + H_2O_{(l)}$ (2 mark)

5. Consider the reaction between hypophosphorous acid (H_3PO_2) which is added to potassium dichromate and produces Chromium (III) ions in solution and phosphoric acid.

(a) Write the skeletal equation for the reaction and identify what has been oxidised and reduced. Then from 1st principles balance the half equations and derive the full equation, be sure to include states for the final equation.

(6 marks)

(a) $H_3PO_2 + Cr_2O_7^{2-} \rightarrow H_3PO_4 + Cr^{3+}$

Oxidation $H_3PO_2 \rightarrow H_3PO_4$ (1 mark)

Reduction $Cr_2O_7^{2-} \rightarrow Cr^{3+}$ (1 mark)

$$2H_2O + H_3PO_2 \rightarrow H_3PO_4$$

$$2H_2O + H_3PO_2 \rightarrow H_3PO_4 + 4H^+ + 4e^- \quad \text{(1 mark)}$$

$$6e^- + 14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$$

X3 $2H_2O + H_3PO_2 \rightarrow H_3PO_4 + 4H^+ + 4e^-$ (1 mark)

X2 $6e^- + 14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$

$$6H_2O + 3H_3PO_2 \rightarrow 3H_3PO_4 + 12H^+ + 12e^-$$

$$12e^- + 28H^+ + 2Cr_2O_7^{2-} \rightarrow 4Cr^{3+} + 14H_2O$$

$$16H^+ + 3H_3PO_2 + 2Cr_2O_7^{2-} \rightarrow 3H_3PO_4 + 4Cr^{3+} + 8H_2O \quad \text{(1 mark)}$$

$$16H^+_{(aq)} + 3H_3PO_{2(aq)} + 2Cr_2O_7^{2-}{}_{(aq)} \rightarrow 3H_3PO_{4(aq)} + 4Cr^{3+}_{(aq)} + 8H_2O_{(l)} \quad \text{(1 mark)}$$

(b) **Detail a full observation** of the reaction.

(b) A colourless solution is added to an orange solution. Upon addition the solution turns a deep green.

(1 mark)

6. Tellurium (Te) is a rare, silver metalloid that can be used in solar panels and as a semiconducting material. It can be produced by reacting the mineral tellurite (TeO_2) with hypophosphorous acid (H_3PO_2). This produces tellurium metal and phosphorous acid (H_3PO_3).

Write the oxidation and reduction half-equations and the overall redox equation for this reaction, assuming acidic conditions. **No states required.**

(3 marks)

Oxidation half-equation	$\mathbf{H_3PO_2 + H_2O \rightarrow H_3PO_3 + 2H^+ + 2e^-}$
Reduction half-equation	$\mathbf{TeO_2 + 4H^+ + 4e^- \rightarrow Te + 2H_2O}$

Overall redox equation	$TeO_2 + 2H_3PO_2 \rightarrow Te + 2H_3PO_3$

7. Determine whether the following reactions represent SPONTANEOUS redox reactions or NOT. Be sure to justify your answer with working showing half equations with E^0 values, and the full equations with **phases** for all reactions. Where a reaction is not spontaneous you must state this as well, and show your working to justify this conclusion.

a. Tin filings added to dilute sulfuric acid.

b. Chlorine gas bubbled through a solution of calcium iodide.

c. Magnesium ribbon added to a solution of lead (II) sulfate.

d. Acidified potassium permanganate solution and hydrogen peroxide.

a. $Sn_{(s)} + 2H^+_{(aq)}$

$Sn \rightarrow Sn^{2+} + 2e^-$ $E° = +0.14V$ (best oxidation)

$2H^+ + 2e^- \rightarrow H_2$ $E° = +0.00V$ (best reduction)

$Sn_{(s)} + 2H^+_{(aq)} \rightarrow Sn^{2+}_{(aq)} + H_{2(g)}$ $E° = +0.14V$ (1)

Positive E° value, therefore reaction IS spontaneous. (1)

(2 marks total)

b. $Cl_{2(g)} + I^-_{(aq)}$

$2I^- \rightarrow I_2 + 2e^-$ $E° = -0.54V$ (best oxidation)

$Cl_2 + 2e^- \rightarrow 2Cl^-$ $E° = +1.36V$ (best reduction)

$2I^-_{(aq)} + Cl_{2(g)} \rightarrow 2Cl^-_{(aq)} + I_{2(aq)}$ $E° = +0.82V$ (1)

Positve E° value, therefore reaction is IS spontaneous. (1)

(2 marks total)

c. $Mg_{(s)} + Pb^{2+}_{(aq)}$

$Mg \rightarrow Mg^{2+} + 2e^-$ $E° = +2.36V$ (best oxidation)

$Pb^{2+} + 2e^- \rightarrow Pb$ $E° = -0.13V$ (best reduction)

$Mg_{(s)} + Pb^{2+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + Pb_{(s)}$ $E° = +2.20V$ (1)

Positive E° value, therefore reaction IS spontaneous. (1)

(2 marks total)

d. $MnO_4^-{}_{(aq)} + H_2O_{2(aq)}$

$H_2O_2 \rightarrow 2H^+ + O_2 + 2e^-$ $E° = -0.70V$ (best oxidation)

$\underline{MnO_4^- + 5e^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O \quad E° = +1.51V \quad \text{(best reduction)}}$

$5H_2O_{2(aq)} + 2MnO^-_{(aq)} + 6H^+ \rightarrow 5O_{2(g)} + 2Mn^{2+}_{(aq)} + 8H_2O_{(l)} \quad E° = +0.81V$
(1)

Positive E° value, therefore reaction IS spontaneous.
(1)

(2 marks total)

Total = 50 marks