

Chemistry

Unit 3

Area of Study 3 Test Answers:

Oxidation and reduction

Section 1: Multiple choice

(16 marks)

Question 1

- D** The oxidation number of silver and chlorine is unchanged. This is a precipitation reaction.

Question 2

- D** From the data given, Ga is more reactive than In because Ga donates electrons to In^{3+} ions.

Similarly, In is more reactive than Ge and Ge is more reactive than Pd. So in order of metal reactivity: $\text{Ga} > \text{In} > \text{Ge} > \text{Pd}$.

So Ga can donate electrons to Ge^{2+} and In can donate electrons to Pd^{2+} , but no reaction is expected between Pd and Ga^{3+} .

Question 3

- B** To be reduced, the oxidation number of sulfur must decrease.

ON of S in SO_2 is +4, in SO_3^{2-} it is +4, in H_2S it is -2, in HSO_4^- it is +6 and in Na_2SO_3 it is +4. So only H_2S could be a product of the reduction of SO_2 .

Question 4

- B** $0.340 - (-0.74) = 1.08 \text{ V}$

Question 5

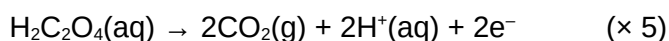
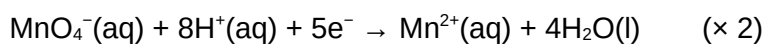
- A** Reduction occurs at the cathode, so the metal electrode at which Cu^{2+} ions are reduced is the cathode. Cu^{2+} ions are reduced at the surface of the Cu(s) electrode, which is therefore the cathode.

Question 6

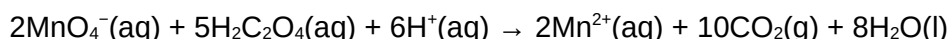
- A** In a galvanic cell, chemical energy of a spontaneous reaction is converted into electrical energy.

Question 7

- C** The half-equations for this reaction are as follows:



So the overall equation is:



Question 8

- C** Coating with paint or tin prevents oxygen from reacting with the iron so while the coating is intact, corrosion is prevented.

For corrosion protection, the iron needs to be made into the cathode of an electrolytic cell. Oxidation occurs at the anode so making the iron an anode would promote its oxidation.

Magnesium is more reactive than iron so a block of magnesium is preferentially oxidised and therefore protects the iron.

End of section 1

Section 2: Short answer

(21 marks)

* Indicates 1 mark

Question 9

(3 marks)

Compound	Oxidation number of N If x is the oxidation number of N, then:	
NaNO_3	$+1x + 3(-2) = 0$	so $x = +5^*$
HNO_2	$+1x + 2(-2) = 0$	so $x = +3^*$
N_2O_4	$+2x + 4(-2) = 0$	so $x = +4^*$

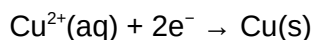
Question 10

(8 marks)

- a**
- i** $2\text{Fe}^{3+}(\text{aq}) + \text{H}_2(\text{g}) \rightarrow 2\text{Fe}^{2+}(\text{aq}) + 2\text{H}^+(\text{aq})^*$
 - ii** Standard reaction potentials give no indication of reaction rate. The reaction may be too slow to observe.*
- b**
- i** The resultant solution would turn brown.*
 - ii** The orange colour of $\text{Br}_2(\text{aq})$ would remain unchanged.*
 - iii** Br_2 is a stronger reducing agent than I_2 and so will accept electrons from I^- ions to produce the brown coloured $\text{I}_2(\text{aq})$.*

But Br_2 is a weaker reducing agent than Cl_2 so cannot attract electrons from Cl^- ions.*

- c** Since reduction occurs at the Cu electrode, the half-equation at that electrode is:



The $\text{Cu}^{2+}(\text{aq})$ is the oxidant.

$$\text{As } E^\circ_{\text{cell}} = E^\circ_{\text{oxidant}} - E^\circ_{\text{reductant}},$$

$$1.97 = 0.34 - E^\circ_{\text{reductant}}^*$$

$$E^\circ_{\text{reductant}} = 0.34 - 1.97 = -1.63 \text{ V}^*$$

Question 11

(6 marks)

- a**
- i** anode (because it is an oxidation reaction)*
 - ii** $\text{NiO}(\text{OH})(\text{s}) + \text{H}_2\text{O}(\text{l}) + \text{e}^- \rightarrow \text{Ni}(\text{OH})_2(\text{s}) + \text{OH}^-(\text{aq})^*$
 - iii** If x is the oxidation number of nickel in $\text{NiO}(\text{OH})(\text{s})$, then $x - 2 - 1 = 0$ and so $x = +3^*$
and in $\text{Ni}(\text{OH})_2$, $x - 2 = 0$, so $x = +2^*$
- b**
- i** $\text{Cd}(\text{OH})_2(\text{s}) + 2\text{e}^- \rightarrow \text{Cd}(\text{s}) + 2\text{OH}^-(\text{aq})^*$
 - ii** The products of the electrode reactions must remain on the electrodes.*

Question 12**(4 marks)**

- a** $\text{CH}_3\text{OH}(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{CO}_2(\text{g}) + 6\text{H}^+(\text{aq}) + 6\text{e}^-$ *
- b** $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$ *
- c** $2\text{CH}_3\text{OH}(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l})$ *
- d** The reactants are continuously pumped into a fuel cell, whereas in a galvanic cell the reactants are stored in the cell.*

End of section 2

Section 3: Extended answer

(23 marks)

* Indicates 1 mark

Question 13

(13 marks)

a platinum (or palladium or graphite)*

b The standard electrode potential of Cl^-/Cl_2 is 1.36.

The predicted cell potential is 1.49 V.

$$\text{As } E^\circ_{\text{cell}} = E^\circ_{\text{oxidant}} - E^\circ_{\text{reductant}}$$

$$1.49 = 1.36 - E^\circ(\text{M}^{x+}/\text{M}_{\text{half-cell}}) \text{ or } 1.49 = E^\circ(\text{M}^{x+}/\text{M}_{\text{half-cell}}) - 1.36^*$$

$$\text{So } E^\circ(\text{M}^{x+}/\text{M}_{\text{half-cell}}) = -0.13 \text{ V or } +2.85 \text{ V}^*$$

Because there are no metal ion/metal half-reactions with a standard reduction potential of +2.85 V, the standard reduction potential of M^{x+}/M must be -0.13 V and metal M is likely to be Pb.*

c anode* (Pb is oxidised to Pb^{2+})

d Electrons flow in the external circuit from M to the Cl_2/Cl^- half-cell.*

e i KNO_3 (or any other soluble ionic compound that does not interfere with the electrode reactions)*

ii In the salt bridge, K^+ ions move from left to right and NO_3^- ions move from right to left.*

f i $\text{Pb} \rightarrow \text{Pb}^{2+} + 2\text{e}^-$ *

ii $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$ *

iii $\text{Pb(s)} + \text{Cl}_2(\text{g}) \rightarrow \text{Pb}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq})$ *

g Set up a galvanic cell, under standard conditions, using a M^{x+}/M half-cell and a half-cell of the unknown metal and its ions.

Identify the direction of electron flow.*

The more reactive metal will form the anode of the cell and electrons flow through the external circuit from the anode to the cathode.*

Question 14**(10 marks)**

- a** **i** $\text{Ni}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Ni}(\text{s})^*$
- ii** The cathode* as the $\text{Ni}^{2+}(\text{aq})$ ions are being reduced.*
- b** **i** $\text{Ni}(\text{s}) \rightarrow \text{Ni}^{2+}(\text{aq}) + 2\text{e}^{-}$ *
- ii** It would decrease in size.*
- c** There would be no change in the green colour intensity if the solution was thoroughly mixed because the amount of $\text{Ni}^{2+}(\text{aq})$ ions produced at the anode is the same as that consumed at the cathode.*
- d** $n(\text{Ni}) = \frac{0.935}{58.69} = 0.01593^*$
- From equation in part **a(i)**, $n(\text{e}^{-}) = 2 \times 0.01593 = 0.0319$ (3 sf)*
- e** Like Ni^{2+} ions, Ag^{+} ions are stronger oxidising agents than water and so can be reduced from an aqueous solution.*
- Mg^{2+} ions are very weak oxidising agents so in an aqueous solution, water molecules would be reduced in preference to Mg^{2+} ions.*

End of answers