Practise Assessment Answers

Question One

Electrolysis requires energy to occur. That is why there is a battery attached to the electrodes. A non-spontaneous reaction is occurring.

Molten Al2O3 is made up of Al3+ ions and O2- ions.

Al3+ moves to the negative electrode (cathode), because the positive ion is attracted to the negative electrode. Al forms when the Al3+ is reduced.

Al3+ + 3 e 🡪 Al

This is a reduction reaction (Al3+ gains electrons, oxidation number of Al decreases from +3 (Al3+) to 0 (Al))

Al is a grey molten metal.

O2- moves to the positive electrode (anode), because the negative ion is attracted to the positive electrode. O2 gas forms when the O2- is oxidised.

2 O2- 🡪 O2 + 4 e

This is a oxidation reaction (O2- loses electrons, oxidation number of O2- increases from -2 (O2-) to 0 (O))

The O2 produced is seen as bubbles of gas.

Question Two

(a) Right electrode = copper metal, right solution = Cu2+(aq)

Left electrode = graphite OR platinum, left solution = S2O32-/S4O62- mixture

This can be reversed so long as the answer in b in also reversed.

(b) from left beaker to right beaker (S2O32- is oxidised to S4O62-, losing electrins which flow to the copper half cell)

Question Three

(a) reduction: Br2 + 2e → 2 Br-, oxidation: Fe2+ → Fe3+ + e,

overall: Br2 + 2 Fe2+ → 2 Br- + 2 Fe3+

(b) C l Fe2+, Fe3+ ll Br2 , Br- l C Pt is also acceptable as an electrode

(c) *E*cell = *E*red – *E*ox = 1.09 – 0.77 = +0.32V

Question Four

Aluminium is the more easily oxidised metal.

The *E*o value for Cl2/Cl- is larger than both metals so both can be oxidised by Cl2. However, the *E*o for Al3+/Al is lower than that of Cu2+/Cu indicating that Al3+ is a weaker oxidant than Cu2+. This means that aluminium attracts its outermost electrons less strongly than copper and is therefore more easily oxidized.

Question Five

(a) The C6H12O6 reaction is reveresed so that it is an oxidation reaction, the other equations is multiplied by 6 so that the electrons can cancel each other out.

C6H12O6 + 6 H2O + 6 O2 + 12 H2O→ 6 CO2 + 24 H+ + 24 OH-

The OH- and H+ combine make water (H2O) and the 18 water on the left side cancel 18 of the waters on the right side leaving 6 on the right side.

C6H12O6 + 6O2 → 6CO2 + 6H2O

(b) Graphite could be used for both electrodes. It is a robust conducting solid which is unreactive towards the reagents in the cells

(c) Glucose → O2/OH-

The *E*o value for O2 is larger than that for glucose so O2 is a stronger oxidant than glucose. Electrons therefore flow from the glucose (which is oxidised at the anode) to O2 (which is reduced at the cathode)

(d) *E*cell = *E*red – *E*ox = 0.40 – (-0.01) = +0.41V

Question Six

(a) Reaction 1 circled with the reason thatbecause that electrode is where oxidation occurs which is loss of electrons or gain in oxidation number.

(b)Lead plate:Pb: is oxidised /changes: to PbSO4 / Pb2+: Lead plates shrink (lose mass)

Lead oxide plate:Pb2+ / PbO2: is reduced / changes to Pb2+ / PbSO4: PbO2 coating shrinks.

Sulfuric acid:[H2SO4] decreases as: H+ used up / SO42– used up / H2O produced

(c) When the cell is recharged, energy is supplied (from the recharger) to cause a non-spontaneous reaction (*Eo* < 0) to occur. As a result the cell components are remade as the reverse reactions are carried out. The overall reaction is simply reversed, and the anode (on discharge) becomes the cathode (when charging).

As current is reversed back through the cells, the lead sulfate ppt has to be turned back to Pb on the lead electrodes: Pb2+ + 2e → Pb

and to PbO2 on the PbO2 electrodes: Pb +2H2O → PbO2 +4H+ +4e

As this happens the sulfuric acid concentration increases again as the recharging reaction produces both hydrogen ions and sulfate ions. This makes the cell ready to use as a battery again.