Answers: Practice task

Part 1

O_2/H_2O 1.23V

Pb^{2+}/Pb -0.36V

At the cathode a layer of grey metal forms. Colourless $Pb^{2+}(aq)$ are reduced to grey Pb(s) which is the metal deposit. As colourless ions are removed from a colourless solution, there is no colour change in the solution.

$Pb^{2+} + 2e^- \rightarrow Pb$

Reduction is the gain of electrons. This is reduction as each Pb^{2+} ion gains 2 electrons *OR Reduction is the decrease on oxidation number. This is reduction as the oxidation number of Pb decreases from +2 in Pb^{2+} to 0 in Pb.*

At the anode the bubbles of colourless gas are oxygen. Colourless water is oxidised to colourless O_2 gas so again there is no colour change.

$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

Oxidation is the loss of electrons. This is oxidation as <u>each</u> H_2O molecule loses 2 electrons *OR Oxidation is the increase in oxidation number. It is oxidation as the oxidation number of O increases from -2 in H_2O to 0 in O_2.*

The overall equation is $2Pb^{2+} + 2H_2O \rightarrow 2Pb + O_2 + 4H^+$

E^ocell = E^o(red) – E^o(ox) = -0.36 – 1.23V = -1.59V.

Since this is a negative value the reaction is non spontaneous and would need a voltage of > 1.59V to make it take place.

O_2 is a better oxidising agent than Pb^{2+} (we know this as the O_2/H_2O has the more positive reduction potential than Pb^{2+}/Pb), so the spontaneous reaction would be the reduction of O_2 to H_2O and the oxidation of Pb to Pb^{2+}.

Part 2

Ag^+/Ag +0.80V

Pb^{2+}/Pb -0.36V

At the cathode colourless silver ions are reduced to a layer of silvery grey silver. The silver electrode would increase in size (as a layer of silver forms on it). Since $Ag^+(aq)$ ions are colourless, removal of these ions would not change the colour of the colourless solution. $Ag^+ + e^- \rightarrow Ag$. This is reduction as each Ag^+ ion gains one electron *OR this is reduction as the oxidation number of Ag decreases from +1 in Ag^+* to 0 in Ag.

At the anode the grey lead electrode gets smaller as Pb atoms are oxidised to Pb^{2+}ions. Since Pb^{2+} are colourless there is no change in colour to the colourless solution. $Pb \rightarrow Pb^{2+} + 2e^-$. This is oxidation as each Pb atom loses 2 electrons *OR this is oxidation as the oxidation number of Pb increases from 0 in Pb to +2 in Pb^{2+}.*

The overall equation is $2Ag^+ + Pb \rightarrow 2Ag + Pb^{2+}$.

E^ocell = E^o(red) – E^o(ox) = +0.80 – -0.36 = +1.16V

Since this is a positive value the reaction ($2Ag^+ + Pb \rightarrow 2Ag + Pb^{2+}$) is spontaneous.

Since Ag^+ is a better oxidising agent than Pb^{2+} (we know this as the Ag^+/Ag has the more positive reduction potential than Pb^{2+}/Pb) the spontaneous reaction would be the reduction of Ag^+ to Ag and the oxidation of Pb to Pb^{2+}.