

## Answers: Practice task

### Part 1

$O_2/H_2O$  1.23V

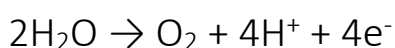
$Pb^{2+}/Pb$  -0.36V

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.



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.



Oxidation is the loss of electrons. This is oxidation as each  $H_2O$  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  $2Pb^{2+} + 2H_2O \rightarrow 2Pb + O_2 + 4H^+$

$$E^\circ_{cell} = E^\circ(\text{red}) - E^\circ(\text{ox}) = -0.36 - 1.23V = -1.59V.$$

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

$\text{Ag}^+/\text{Ag} +0.80\text{V}$

$\text{Pb}^{2+}/\text{Pb} -0.36\text{V}$

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  $\text{Ag}^+(\text{aq})$  ions are colourless, removal of these ions would not change the colour of the colourless solution.  $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$ . This is reduction as each  $\text{Ag}^+$  ion gains one electron *OR this is reduction as the oxidation number of Ag decreases from +1 in  $\text{Ag}^+$  to 0 in Ag.*

At the anode the grey lead electrode gets smaller as Pb atoms are oxidised to  $\text{Pb}^{2+}$  ions. Since  $\text{Pb}^{2+}$  are colourless there is no change in colour to the colourless solution.  $\text{Pb} \rightarrow \text{Pb}^{2+} + 2\text{e}^-$ . 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  $\text{Pb}^{2+}$ .*

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

$$E^\circ_{\text{cell}} = E^\circ(\text{red}) - E^\circ(\text{ox}) = +0.80 - -0.36 = +1.16\text{V}$$

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

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