Marks

- Give the oxidation number of carbon in each of the following. **2**

CF_2Cl_2(g)	**+4**
$Na_2C_2O_4$(s)	**+3**
HCO_3^-(aq)	**+4**
C(s)	**0**

- Consider a voltaic cell that uses the following half-reactions: **3**

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$$

$$Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$$

Write a balanced equation for the overall reaction.

The two cell potentials are +1.51 V (MnO_4^-/Mn^{2+}) and +0.15 V (Sn^{4+}/Sn^{2+}). The least positive (Sn^{4+}/Sn^{2+}) is reversed so that it is the oxidation reaction. Combining the half cells in this way gives, after balancing:

$$\mathbf{2MnO_4^-(aq) + 16H^+(aq) + 5Sn^{2+}(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5Sn^{4+}(aq)}$$

Which species is the oxidising agent?	**MnO_4^-(aq)**
Which species is the reducing agent?	**Sn^{2+}(aq)**

Calculate the standard cell potential. (Refer to the table of standard reduction potentials.)

As noted above, the Sn^{4+}/Sn^{2+} half cell is reversed so that its oxidation potential is -0.15 V. Hence, the standard cell potential is:

$$\mathbf{E^o_{cell} = E^o_{red} + E^o_{ox} = (+1.51) + (-0.15) = +1.38\,V}$$

Answer: **+1.38 V**