

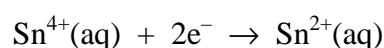
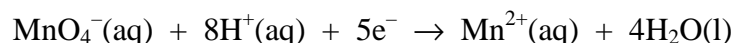
**Marks**  
**2**

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

CF <sub>2</sub> Cl <sub>2</sub> (g)	+4
Na <sub>2</sub> C <sub>2</sub> O <sub>4</sub> (s)	+3
HCO <sub>3</sub> <sup>-</sup> (aq)	+4
C(s)	0

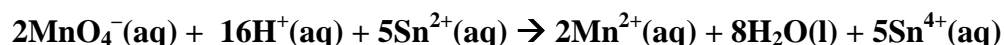
**3**

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



Write a balanced equation for the overall reaction.

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



Which species is the oxidising agent?

**MnO<sub>4</sub><sup>-</sup>(aq)**

Which species is the reducing agent?

**Sn<sup>2+</sup>(aq)**

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

**As noted above, the Sn<sup>4+</sup>/Sn<sup>2+</sup> half cell is reversed so that its oxidation potential is -0.15 V. Hence, the standard cell potential is:**

$$E_{\text{cell}}^{\circ} = E_{\text{red}}^{\circ} + E_{\text{ox}}^{\circ} = (+1.51) + (-0.15) = +1.38 \text{ V}$$

**Answer: +1.38 V**