CHEM1612 Worksheet 10 – Answers to Critical Thinking Questions

The worksheets are available in the tutorials and form an integral part of the learning outcomes and experience for this unit.

Model 1: Reduction Potentials

1. Oxidising agents are themselves reduced – the strongest oxidising agent is the most easily reduced. This is Ag^+(aq) as it has the most positive $E_{red}°$ value (it has the strongest attraction to electrons).

2. Reaction (1) will remain a reduction. Reaction (2) will reverse to become an oxidation, as Ag^+(aq) is the strongest oxidising agent.

3. Reaction (3) will remain a reduction. Reaction (4) will reverse to become an oxidation as Zn^{2+}(aq) is the stronger reducing agent. It does not matter that they are both negative as it is the *difference* between the two E^0_{red} values which determines the reaction.

Model 2: Voltaic Cells

1. The Zn/Zn^{2+} half reaction is proceeding as an oxidation as it has a *lower* E^0_{red} value than that for Sn^{2+}/Sn. When we flip a reduction to an oxidation, we reverse the sign of the potential.

2. The zinc electrode will *lose* mass and the tin electrode will *gain* mass.

3. Oxidation (always) takes place at the anode. Reduction (always) takes place at the cathode

4. (b) Electrons flow through the wire, from the zinc electrode towards the tin electrode.

5. The anode is negative and the cathode is positive.

6. SO_4^{2-}(aq) moves into the zinc half cell (as cations are being made in the oxidation reaction in this cell). Na^+(aq) moves into the tin half cell (as cations are being lost in this cell).

7. Cathode - reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$. Anode - oxidation: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
E^0_{cell} = [0.80 +(-0.34)] V = +0.46 V

8. Cathode - reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$. Anode - oxidation: $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$
E^0_{cell} = [0.34 + 0.14] V = +0.48 V

9. Cathode - reduction: $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$. Anode -oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
E^0_{cell} = [-0.14 + 0.76] V = +0.62 V

10. Couple the cells with (i) the most positive and (ii) the least positive (or most negative) reduction potentials. The latter is reversed to become the oxidation reaction.

    $Ag^+(aq) + e^- \rightarrow Ag(s)$ and $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$; E^0_{cell} = +1.56 V

11. (a) $NAD^+ + HCOO^- \rightarrow NADH + CO_2$ $E°$ = (-0.105 + 0.20) V = +0.10 V
    NAD is reduced.

    (b) $O_2 + 2H^+ + 2NADH \rightarrow H_2O + 2NAD^+$ $E°$ = (+0.82 + 0.105) V = +0.925 V
    NAD is oxidised.