CHEM1101 Worksheet 13 – 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: The Effect of Concentration on the Cell Potential

1. If $[H^+(aq)] = 1.0$ M then pH = 0.
2. At a lower $[H^+(aq)]$, the reaction is *less* favourable and so E will be less positive.

Model 2: The Nernst Equation

1. For reaction (1), $n = 2$. For reaction (2), $n = 2$.
2. $$Q = \frac{[H_2O][NAD^+]}{[O_2]^{1/2}[H^+][NADH]} = \frac{(1)(1)}{(1)(10^{-7})(1)} = 10^{+7}$$

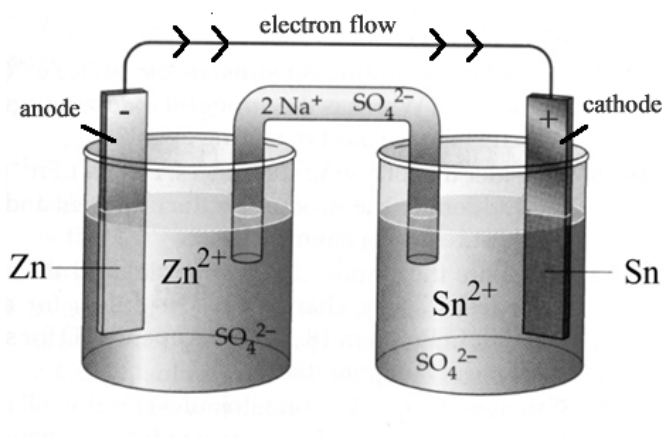
   With $T = 37$ °C and $n = 2$:

   $$E_{\text{cell}} = E^0_{\text{cell}} - \frac{RT}{nF}\ln Q = (+1.335\text{ V}) - \frac{(8.314\text{ J K}^{-1}\text{ mol}^{-1})(310.\text{ K})}{(2)(96485\text{ C mol}^{-1})}\ln(10^7) = 1.12\text{ V}$$

3. E_{cell} is less positive than E^0_{cell} as predicted.

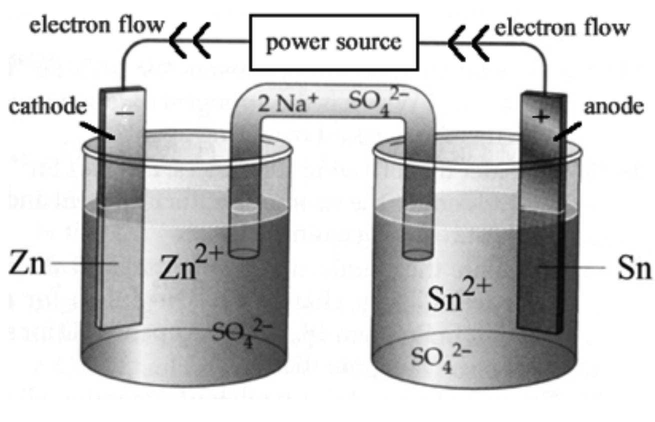
Model 3: Voltaic Cells

1. The zinc electrode will *lose* mass and the tin electrode will *gain* mass.
2. $Sn^{2+}(aq) + Zn(s) \rightarrow Sn(s) + Zn^{2+}(aq)$
3. Zn is being oxidised and Sn^{2+} is being reduced. The Zn electrode is the anode. The Sn electrode is the cathode.
4. In voltaic cells, electrons flow through the wire, from the zinc electrode towards the tin electrode.
5. In voltaic cells, 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).



Model 4: Electrolytic Cells

1. The zinc electrode will *gain* mass and the tin electrode will *lose* mass.
2. $Sn(s) + Zn^{2+}(aq) \rightarrow Sn^{2+}(aq) + Zn(s)$



3. Zn is being reduced and Sn^{2+} is being oxidised. The Zn electrode is the cathode. The Sn electrode is the anode.

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

5. The anode is positive and the cathode is negative. The power source pumps electrons to the cathode from the anode.

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

Model 5: Electrolysis of Water

1. $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$.
2. $F = (1.602 \times 10^{-19}\ C) \times (6.022 \times 10^{23}) = 96470\ C\ mol^{-1}$. This is equal to the tabulated value for Faraday's constant within the accurary of the data used here.
3. Number of moles of electrons = $I \times t / F = (10.0\ A) \times (2.00 \times 60 \times 60\ s) / (96485\ C\ mol^{-1}) = 0.746$ mol
4. From the half cell equation for the reduction of H_2O, $2e^-$ are required for each H_2. Therefore, 0.746 mol will produce ½ × 0.746 mol = 0.373 mol of H_2.
5. From Q1, half as much O_2 will be produced: 0.187 mol.

   Alternatively, from the half cell equation for the oxidation of H_2O, $4e^-$ are produced for each O_2. Therefore, 0.746 mol will have been produced by ¼ × 0.746 mol = 0.187 mol of O_2.

6. The reduction potential of water is -0.83 V so a cation with a *more* negative reduction potential should be used: Cr^{3+}, Al^{3+}, Mg^{2+}, Na^+, Ca^{2+} or Li^+,

   The oxidation potential of water is -1.23 V so an anion with a *more* negative oxidation potential should be used: Cl^- or SO_4^{2-}.

   Na_2SO_4 or K_2SO_4 are commonly used.

- How many hours does it take to form 10.0 L of O_2 measured at 99.8 kPa and 28 °C from water if a current of 1.3 A passes through the electrolysis cell?

Marks 3

10.0 L corresponds to 0.0100 m^3. Using $PV = nRT$, this corresponds to:

$$n = PV / RT$$
$$= (99.8 \times 10^3 \text{ Pa}) \times (0.0100 \text{ m}^3) / (8.314 \text{ Pa m}^3 \text{ mol}^{-1} \text{ K}^{-1}) \times ((28 + 273) \text{ K})$$
$$= 0.399 \text{ mol}$$

O_2(g) is formed by electrolysis of water according to the reaction:

$$2H_2O(l) \rightarrow 4H^+(aq) + O_2(g) + 4e^-$$

Hence, 4 × 0.399 mol = 1.60 mol of electrons are required. As the number of moles of electrons passed by a current *I* in a time *t* is:

number of moles of electrons = It / F

$$1.60 \text{ mol} = (1.3 \text{ A})t / (96485 \text{ C mol}^{-1})$$

$$t = 1.2 \times 10^5 \text{ s} = (1.2 \times 10^5 / 3600) \text{ hours} = 33 \text{ hours}$$

Answer: **33 hours**

- What mass of $PbSO_4$ is reduced at the cathode when a lead-acid storage battery is charged for 1.5 hours with a constant current of 10.0 A?

Marks 3

The total charge delivered is:

$$Q = It = (10.0) \times (1.5 \times 60 \times 60) = 54000 \text{ A}$$

This is equivalent to $\frac{Q}{F} = \frac{54000}{96485}$ = 0.56 moles of electrons.

The reduction reaction during recharging is a 2e$^-$ process:

$$PbSO_4 + H^+ + 2e^- \rightarrow Pb + HSO_4^-$$

Hence, $\frac{0.56}{2} = 0.28$ moles of $PbSO_4$ are reduced. The molar mass of $PbSO_4$ is (207.2 (Pb)) + (32.07 (S)) + (4 × 16.00 (O)) = 303.27. Hence, the mass reduced is given by:

mass of $PbSO_4$ reduced = 0.28 × 303.27 = 85 g