CH1011 Tutorial 7 Answers

1. Assign the oxidation state of carbon in the following compounds: $C_{diamond}$ $NaHCO_3$

C diamond C = 0
element = 0 by definition

Na^+ HCO_3^- $\rightarrow$ HCO_3^- C = +IV
ionic cmpd
1-(charge) = (+1H) + (3x -2 O) + C
$\Rightarrow$ C = -1 + 6 -1 = +4

CH_2CF_2 0(charge) = C + (2x+1 H) + (2 x -1 F) $\Rightarrow$ C = 0
Covalent compound

2. The following two redox couple are combined: $E^o_{½}$ $Cr^{3+}_{(aq)}/Cr_{(s)}$ -0.74V $Fe^{3+}_{(aq)}/Fe^{2+}_{(aq)}$ 0.77V

- Work out and balance the spontaneous **redox reaction**.
- Calculate the standard cell voltage $\mathbf{E^o_{cell}}$.
- Identify the component that is the **oxidant** and the component that is the **reductant**.

The most negative is the anodic reaction the oxidation : Cr^{3+}/Cr -0.74V
The most positive is the cathodic reaction the reduction: Fe^{3+}/Fe^{2+} 0.77V

$Cr_{(s)} \rightarrow Cr^{3+}_{(aq)} + 3e^-$ **Cr reductant** (is oxidised)
$Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$ **$Fe^{3+}_{(aq)}$oxidant** (is reduced)

To balance the electrons we have to multiply 3x($Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$)

Combined to give the spontaneous voltaic cell:

$$\mathbf{Cr_{(s)} + 3Fe^{3+}_{(aq)} \rightarrow 3Fe^{2+}_{(aq)} + Cr^{3+}_{(aq)}}$$

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$ = 0.77 - -0.74 = **1.51V**

This verifies that the reaction is spontaneous as written as it is a positive voltage.

3. What is the cell reaction in the following cell?
Calculate the standard free energy change ΔG^o for the reaction:

Cu | $CuSO_4$ || KCl 0.100M | AgCl(s) | Ag(s)

Given that the potential of this cell at 298K is +0.030V and that E^o for the electrode reaction AgCl(s) + e $\rightarrow$ Ag(s) + Cl^- is + 0.223V, calculate the concentration of $CuSO_4$ in the cell.

$Cu(s) \rightarrow Cu^{2+}_{(aq)} + 2e^-$
$2AgCl(s) + 2e^- \rightarrow 2Ag(s) + 2Cl^-_{(aq)}$

Cell reaction: $\underline{Cu_{(s)} + 2AgCl_{(s)} \rightarrow 2Ag(s) + Cu^{2+}_{(aq)} + 2Cl^-_{(aq)}}$ (since E = positive)

$\Delta G^o = - nFE^o_{cell}$ (n = no of electrons transferred)

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$ = 0.223 - 0.340 = -0.117 V

$\Delta G^o = -2 \text{ x } 96{,}500 \text{ x} - 0.117 = 22581 \text{ J}$ = **22.6 kJ**

For the reaction $Cu(s) + 2AgCl(s) \rightarrow 2Ag(s) + Cu^{2+}{}_{(aq)} + 2Cl^-{}_{(aq)}$

$$E_{cell} = E^o{}_{cell} - \frac{RT}{n \text{ x } F} \ln\left(\frac{[Ag(s)] \text{x} [Cu^{2+}(aq)] \text{x} [Cl^-(aq)]^2}{[Cu(s)] \text{x} [AgCl(s)]^2}\right)$$

$$\therefore 0.030 = -0.117 - \frac{8.314 \text{ x } 298}{2 \text{ x } 96500} \ln\left\{\frac{1 \text{ x } [Cu^{2+}(aq)] \text{ x } 0.1^2}{1 \text{ x } 1^2}\right\}$$

$$\therefore 0.030 = -0.01284 \ln\{[Cu^{2+}] \text{ x } [0.01]\}$$

$\therefore 0.147 = -0.01284 \ln[Cu^{2+}{}_{(aq)}] - 0.01284 \ln[0.01]\}$

$\therefore 0.147 = -0.01284 \ln[Cu^{2+}{}_{(aq)}] + 0.0591$

$$\therefore -\left(\frac{0.147 - 0.0591}{0.01284}\right) = \ln[Cu^{2+}{}_{(aq)}]$$

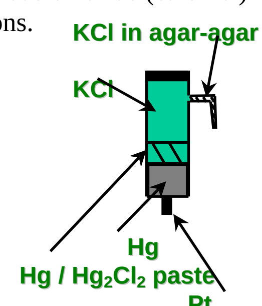
$\therefore -6.8458 = \ln[Cu^{2+}{}_{(aq)}]$

$\therefore \exp(-6.8458) = [Cu^{2+}{}_{(aq)}]$

Hence $[Cu^{2+}]$ = **$1.1 \text{ x } 10^{-3}$ mol dm^{-3}**

4. Outline the chief features of a **calomel electrode** and explain how it is used.

A calomel electrode is an example of a **metal-insoluble salt**-anion electrode or half-cell. It essentially consists of a mixture of mercury and mercurous chloride (calomel) in contact with a solution containing a fixed concentration of chloride ions.



$Hg_{(l)}\ Hg_2Cl_{2(l)}\ |\ Cl^-$

$Hg_2Cl_{2(l)} + 2e^- \rightarrow 2Hg(l) + 2Cl^-{}_{(aq)}$

A robust electrode which is used as the standard reference electrode in many electrochemical voltaic cells, it is referenced at 0.85V w.r.t. SHE.

5. What is the **membrane potential** and why is it important in cellular biochemistry?

The membrane potential $\mathbf{\Delta E = -RT/nF}$ $\boldsymbol{ln}\mathbf{([C_i] / [C_o])}$ written in Nernst Equation form.

n = charge on the ions
$[C_o]$ = concentration of ions outside the membrane (summed over H^+, Na^+, K^+, Ca^{2+})
$[C_i]$ = concentration of ions inside the membrane (summed over H^+, Na^+, K^+, Ca^{2+})

The membrane potential is an electrochemical potential (units mV)

Electrochemical potential is the combination of the two forces:

a. electrical potential which moves charged particles by electrical force.
b. chemical potential which moves particles down concentration gradients.

ΔE is the equilibrium electrical potential difference across the membrane where the electrical potential balances the chemical potential.

The cell manages to store useful energy (ΔG) which is instantly available by creating membrane potentials. Two important membrane potentials are the proton gradient used in ATP synthesis (mitochondria and chloroplast) and the Na^+/K^+ gradients used in nerve cell signal transmission.
These membrane potential potentials are created by active transport where the cell uses energy to pump ions across the cell membrane against a concentration gradient.