Energy revision
Solutions

Question 1

The reaction between solutions of hydrochloric acid and sodium hydroxide can be represented by the following equation.

$HCl(aq) + NaOH(aq) \Rightarrow NaCl(aq) + H_2O(l)$ ΔH = .56.0 kJ mol^{-1}

80.0 mL of 2.00 M HCl, at 21.0 °C, is mixed with 20.0 mL of 2.50 M NaOH, also at 21.0 °C, in a well-insulated calorimeter. The calibration factor for the calorimeter and contents is 420 J C^{-1}

What is the final temperature, in °C, of the resultant solution in the calorimeter ?

Step 1 Find the limiting reactant

a) Find the mole of HCl => n = C x V => 2.00 x 0.0800 = 0.160

b) Find the mole of NaOH => n = C x V => 2.50 x 0.0200 = 0.0500

⇨ **The limiting reactant is NaOH**

Step 2 Find the heat energy released when 0.0500 mole of NaOH reacts

⇨ **0.0500 X 56.0 kJ = 2.80 kJ**

Step 3 Calculate the temperature change.

a) 2,800J / 400 J/C = 7.00 C

Step 4 Calculate the final temperature

21.0 = 7.00 = 28.0 C

Question 2

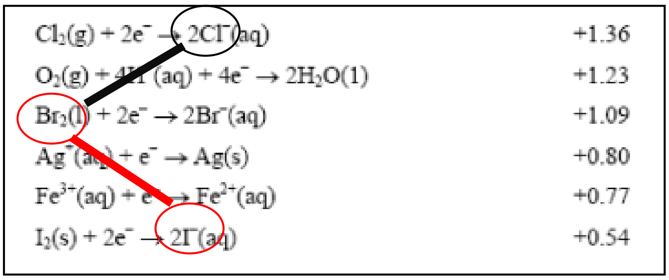
a) Which one of the following would be predicted to spontaneously oxidise aqueous iodide ions but not aqueous chloride ions? Explain

A. Au^{+} (aq)

B. Sn^{2+} (aq)

C. Fe^{2+} (aq)

D. Br_2(l)



$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	+1.23
$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$	+1.09
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	+0.54

Spontaneous reactions occur only between the reactants shown in red.

b) Two carbon electrodes are inserted into a solution of $MgCl_2$ and $Fe(NO_3)_2$. The electrodes are connected to the positive and negative terminals of an external power source.

i) Write the half cell reaction occurring at the anode.

The anode in electrolysis is the place where oxidation takes place and has positive polarity. It is the place where the strongest reductant present reacts

$Fe^{2+}(aq)$ => e^- + Fe^{3+}

ii) Write the half cell reaction occurring at the cathode.

The cathode in electrolysis is the place where reduction takes place and has negative polarity. It is the place where the strongest oxidant present reacts

$Fe^{2+}(aq)$ + $2e^-$ => Fe(s)

iii) Explain how the pH of the solution changes as the power source is switched on?
The pH remains unchanged

Question 3

The rechargeable nickel-cadmium cell is used to power small appliances such as portable computers. When the cell is being used, the electrode reactions are represented by the following equations.

$NiO_2(s)$ + $2H_2O(l)$ + 2e. => $Ni(OH)_2(s)$ + $2OH^-(aq)$
Cd(s) + $2OH^-(aq)$ => $Cd(OH)_2(s)$ + 2e.

a) When the cell is being **recharged** give the reaction that occurs at the:

When the battery is being recharged it acts as an electrolytic cell where oxidation occurs at the positive anode and reduction occurs at the negative cathode.

- Anode

$Ni(OH)_2(s)$ + $2OH^-(aq)$. => **$NiO_2(s)$ + $2H_2O(l)$ + 2e**

- Cathode

$Cd(OH)_2(s)$ + 2e. => **Cd(s) + $2OH^-(aq)$**

b) What happens to the pH of the electrolyte?

It remains unchanged as OH^- is created at the cathode and used at the anode in the same molar raito.

c) When the cell is being used what is the reaction occurring at the:

- anode.

Cd(s) + $2OH^-(aq)$ => **$Cd(OH)_2(s)$ + 2e**.

- cathode

$NiO_2(s)$ + $2H_2O(l)$ + 2e. => **$Ni(OH)_2(s)$ + $2OH^-(aq)$**

Question 4

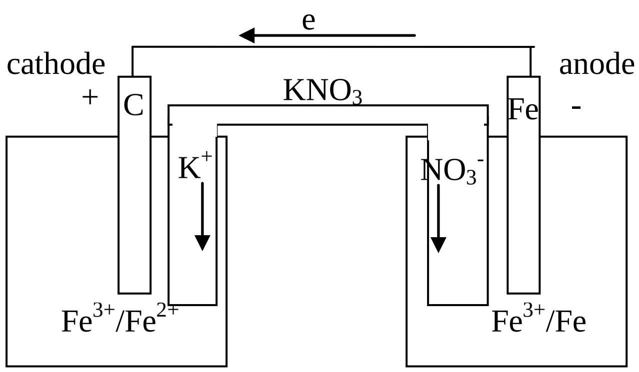
A galvanic cell consists of one half cell that is made up of an inert graphite electrode in a solution containing 1.0 M Fe^{2+} (aq) and 1.0 M Fe^{3+} (aq) at 25° C.

a) Select, from the electrochemical series provided, a second half cell so that:

- when connected will provide approximately 1.2 V ;

b) Draw the galvanic cell. Label the following.

i) anode and cathode
ii) polarity of each electrode
iii) oxidation and reduction reactions
iv) an appropriate electrolyte for the salt bridge
v) direction of electron flow
vi) direction of ion flow from salt bridge



Question 5

A copper disc is to be silver-plated in an electrolytic cell. The disc forms one electrode and a silver rod the other electrode. The electrolyte provides a source of Ag^+ (aq). The mass of silver to be deposited is 0.150 g.

a) If the current is held steady at 1.50 amps, calculate the time, in seconds, that it takes to complete the plating

Step 1 Calculate the mol of Ag deposited.
=> n_{Ag} = 0.150g/ 107.9 = 1.39 X 10^{-3}

Step 2 Calculate the mol of electrons needed
Ag^+(aq) + e => Ag(s)
n_e = 1.39 X 10^{-3}

Step 3 Calculate the charge needed in Coulombs
=> Q = 1.39 X 10^{-3} X 96500 = 134.2

Step 4 Calculate the time it takes for 1.50 amps to deliver this charge
Q =It
=> 134.5 = 1.50 t
=> t = 134.2 / 1.50 = 89.5 seconds

b) Write the equation to the reaction occurring at the:

- Anode
  $Ag(s) => Ag^{+}(aq) + e^{-}$

- Cathode

  $Ag^{+}(aq) + e^{-} => Ag(s)$

c) An identical disc is to be zinc-plated with a solution containing Zn^{2+} (aq) as the electrolyte using a current of 1.50 amps. Calculate the ratio of the time that is needed to plate the disc with 0.150 g of zinc to the time needed to plate the disc with 0.150 g of silver.

Step 1 Calculate the mol of Zn

⇨ **n_{zn} = 0.150 / 65.4 = 0.00229**

Step 2 Calculate the mol of electrons needed
Zn^{+2}(aq) +2 e => Zn(s)

⇨ **n_e = 2 X 0.002294 = 0.00459**

Step 3 Calculate the charge needed in Coulombs
=> Q = 0.00459 X 96500 = 443

Step 4 Calculate the time it take for 1.50 amps to deliver this charge
Q =It
=> 443 = 1.50 t
=> t = 443 / 1.50 = 295 seconds
Step 4 Find the ratio of Zn : Ag
295 : 89.5
3.3 : 1
or
10 : 3

Question 6

Two reactants "A" and "B" react according the equation shown below

A(g) + 2B(g) = > 2C(g) + 3D(g).

A calorimeter was used to record the energy given out when 6.00 grams of "B' reacts completely with excess "A".

A bomb calorimeter was calibrated by passing a current of 1.20 amps at 2.31 volts for 2.50 minutes through 100 mL of water at 25.2 ^{o}C. The temperature of the water increased to a maximum of 31.2 ^{o}C.

a) Calculate the calibration factor of the calorimeter.

Step 1 Calculate the amount of energy delivered.
=> E =vIt(seconds)
=> E = 2.31 X 1.20 X 2.50 X 60 = 415.8J

Step 2 Find the change in temperature
=> 31.2 - 25.2 = 6.00

Step 3 Calculate the calibration factor
=> 415.8 / 6.00 = 69.3 J/^{o}C

b) If the temperature of the water increased from 25.2 ^{o}C to 29.3 ^{o}C when 6.00g of reactant "B" reacted completely, find the ΔH of the reaction below if the molar mass of "B" is 46.3 g/mol..

A(g) + 2B(g) = > 2C(g) + 3D(g).

Step 1 Find the mol of "B"
=> n_B = 6.00 / 46.3 = 0.1296

Step 2 Find the amount of energy given out by 0.1296 mol of "B"
=> Energy = 69.3 X (29.3 - 25.2) = 284.13 J

Step 3 Find the energy per mol of "B"
=> 284.13 / 0.1296 = 2.192 kJ

Step 4 Find the ΔH of the reaction
=> Since the reaction indicates two mol of "B" then the ΔH is 2.192 X 2 = 4.384 kJ

- ⇨ **ΔH = - 4.384 kJ (negative sign indicates energy given out)**

Question 7

a) Fuel cell is constructed using ethanol and oxygen as the main reactants. Ethanol and oxygen react to form acetic acid according to the chemical equation below.

$$CH_3CH_2OH(l) + O_2(g) => CH_3COOH(aq) + H_2O(l)$$

i) Give the half equation that occurs at the anode
$CH_3CH_2OH(l) + H_2O(l) => CH_3COOH(aq) + 4H^+(aq) + 4e^-$

ii) Give the half equation that occurs at the cathode
$4e^- + 4H^+(aq) + O_2(g) => 2H_2O(l)$

iii) A set of fuel cells produce 230.0 grams of acetic acid every minute. What charge is produced by the fuel cells over a 24 hour period?

Step 1 Calculate the mol of acetic acid produced every minute.
=> 230.0 / 60.0 = 3.83

Step 2 Calculate the mol of acetic acid produced in 24 hours.
=> 3.83 X 24 X 60 = 5515 mol

Step 3 Calculate the mol of electrons every 24 hours
=> 5515 X 4 = 22060

Step 4 Calculate the charge generated over a 24 hour period.
=> 96500 X 22060 = 2.13 X 10^9 Q

8) The reaction between hydrogen and oxygen is the basis of energy production in a number of fuel cells

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \qquad \Delta H = -571.6 \text{ kJ mol}^{-1}$$

An alkaline electrolyte is used in a particular hydrogen/oxygen fuel cell.
b) Write a balanced half-equation for the reaction occurring at the

i. anode

$2OH^-(aq) + H_2 => 2H_2O(l) + 2e^-$

ii. cathode.

$O_2(g) + 2H_2O(l)) + 4e^- => + 4OH^-(aq)$

iii. 16.00 grams of oxygen reacted with excess hydrogen and the energy generated was used to heat 80.00 grams of water at 1.20 ^oC. Assuming no energy is lost what is the final temperature of the water.

Step 1 Calculate the oxygen gas reacting.
=> 16.00 /32.00 = 0.500 mol
Step 2 Calculate the amount of energy released.
=> For every mol of oxygen reacting 571.6 kJ of heat energy is released.
=> So for 0.500 mol of oxygen we get (0.500 X 571.6) = 285.8 kJ
Step 3 Calculate the rise in temperature
=> 285.8 kJ / 2.18 kJ/^{o}C = 131.1 ^{o}C
Step 4 Calculate the final temperature
=> 131.1 + 1.20 = 132.30 ^{o}C

iv. Given the equations below

a) 2A(g) + B(g) => C(s) + D(l) ΔH= -34.5 kJ/mol
b) E(l) + G(s) => 2C(s) + 2D(l) ΔH= -144.5 kJ/mol

Calculate the ΔH of the equation below

4A(g) + 2B(g) => E(l) + G(s)

Step 1 multiply a) by 2

c) 4A(g) + 2B(g) => 2C(s) +2D(l) ΔH= -69.0 kJ/mol

Step 2 reverse equation b)

d) 2C(s) + 2D(l) => E(l) + G(s) ΔH= +144.5 kJ/mol

Step 3 Add equations c) and d)

4A(g) + 2B(g) => E(l) + G(s) ΔH= +75.5 kJ/mol

Physical constants

$F = 96\,500$ C mol^{-1}
$R = 8.31$ J K^{-1} mol^{-1}
1 atm = 101 325 Pa = 760 mmHg
0°C = 273 K
Molar volume at STP = 22.4 L mol^{-1}
Avogadro constant = 6.02×10^{23} mol^{-1}

Ideal gas equation
$pV = nRT$

The electrochemical series

	$E°$ in volt
$F_2(g) + 2e^- \rightarrow 2F^-(aq)$	+2.87
$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightarrow 2H_2O(l)$	+1.77
$Au^+(aq) + e^- \rightarrow Au(s)$	+1.68
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(1)$	+1.23
$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$	+1.09
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.34
$S(s) + 2H^+(aq) + 2e^- \rightarrow H_2S(g)$	+0.14
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	–0.13
$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$	–0.14
$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$	–0.23
$Co^{2+}(aq) + 2e^- \rightarrow Co(s)$	–0.28
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	–0.44
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	–0.76
$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$	–0.83
$Mn^{2+}(aq) + 2e^- \rightarrow Mn(s)$	–1.03
$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$	–1.67
$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$	–2.34
$Na^+(aq) + e^- \rightarrow Na(s)$	–2.71
$Ca^{2+}(aq) + 2e^- \rightarrow Ca(s)$	–2.87
$K^+(aq) + e^- \rightarrow K(s)$	–2.93
$Li^+(aq) + e^- \rightarrow Li(s)$	–3.02