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VCE Chemistry ¾

Primary Cells & Faraday's Laws [1.10]

Homework Solutions

Homework Outline:

Compulsory Questions	Pg 2 – Pg 12
Supplementary Questions	Pg 13 – Pg 20

Section A: Compulsory Questions (32 Marks)

Sub-Section [1.10.1]: Identify Features of Primary Cells & how they Operate

Question 1 (3 marks)

Manganese metal is commonly used in batteries, especially in alkaline and Manganese-carbon cells for daily consumer use.

a. By referring to information provided in the Data Book, give one reason why Manganese is used as a reactant in these galvanic cells. (1 mark)

Manganese is a relatively strong reductant (-1.18V). This means that many reactants will have larger EMF values and act as a suitable oxidant, allowing for the galvanic cell to operate. (1)

b. When manganese comes into contact with water, an explosion is observed. Using the half equations, explain this observation. (2 marks)

Manganese metal and water will spontaneous react releasing heat. As hydrogen gas produced accumulates, the heat produced can cause ignition of this gas causing an explosion. (1) (2) – half eqns

Oxidation Reaction:

$$Mn_{(s)} \rightarrow Mn^{2+}_{(aq)} + 2e^-$$

Reduction reaction:

$$2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}$$

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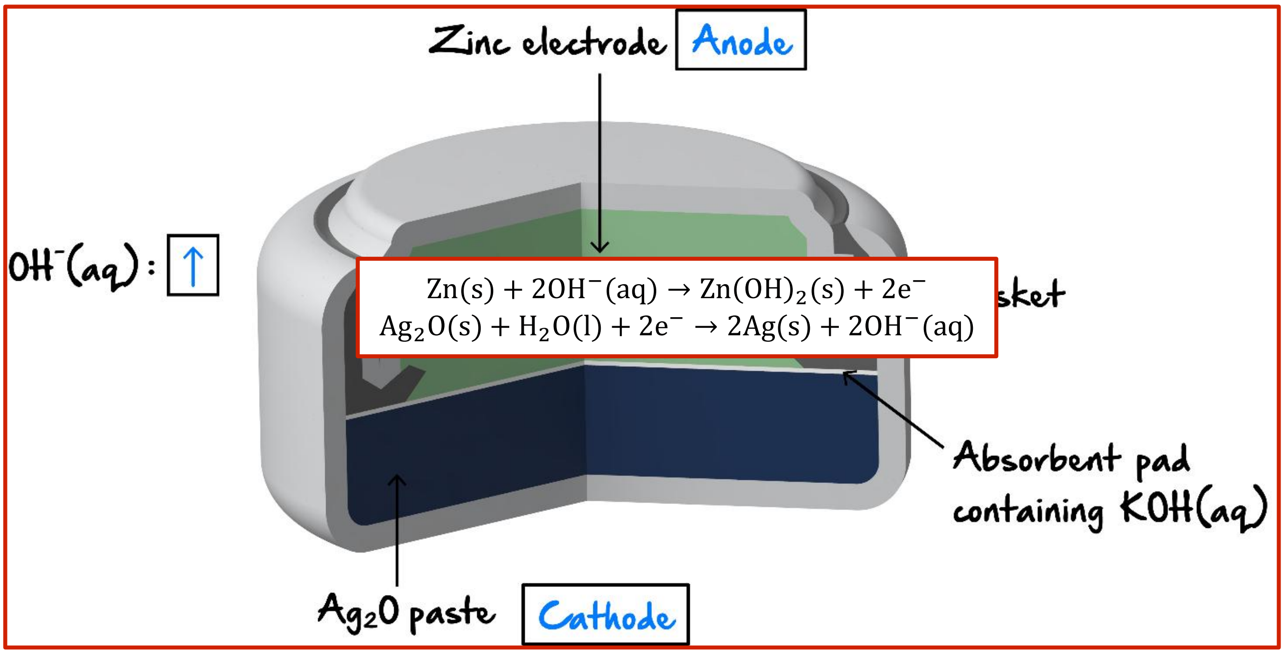
Question 2 (2 marks)

Zinc-silver oxide batteries are commonly used in watches due to their long battery life and durability.



a. Label the anode and the cathode in the boxes provided above. (1 mark)

b. Label the direction of movement of electrolyte in the cell in the box provided above. (1 mark)

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Question 3 (3 marks)

a. Which of the following statements is correct for both fuel cells and galvanic cells? (1 mark)

A. Use porous electrodes to increase the reaction surface area.

B. Require conductive electrodes.

C. Have direct energy conversions from electrical to chemical.

D. Have low energy efficiency.

b. Which of the following is an advantage that primary cells have over fuel cells? (1 mark)

A. Lower carbon emissions.

B. Greater voltage generated.

C. Lower start-up time.

D. Cheaper to produce.

c. Determine how a fuel cell differs from a primary cell. (1 mark)

A. Fuel cells require a continuous supply of reactants, whereas primary cells have a fixed amount of reactants.

B. Fuel cells store electrical energy, whereas primary cells generate it from chemical reactions.

C. Primary cells produce only water as a byproduct, whereas fuel cells produce CO_2.

D. Primary cells can be recharged multiple times, whereas fuel cells cannot.

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Sub-Section [1.10.2]: Apply Faraday's First & Second Law and $Q = It$ & $Q = n(e)F$ to Calculations

Question 4 (4 marks)

Answer the following questions regarding three separate galvanic cells.

a. In a galvanic cell, 310 C of electric charge passes through the circuit in 25 minutes. Calculate the current, in A, running through the cell. (1 mark)

$$I = \frac{Q}{t} = \frac{310}{25 \times 60} = 0.207\ A$$

b. In another galvanic cell, 2.56 A of current runs through the cell during a 15 minute period. Calculate the moles of electrons produced in the cell. (1 mark)

$$Q = It = 2.56 \times 15 \times 60 = 2304\ C$$

$$n(e^-) = \frac{2304}{96500} = 0.0239\ mol \cong 2.39 \times 10^{-2} mol$$

c. Calculate the moles of zinc (Zn) produced in a cell, when 2.51 A of current is running through the circuit for 35 minutes. The half equation for zinc has been provided below: (2 marks)

$$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$$

$$n(e^-) = \frac{It}{F} = \frac{2.51 \times 35 \times 60}{96500} = 0.054622\ mol$$

$$\therefore n(Zn) = \frac{0.054622}{2} = 0.0273\ mol \approx 2.7 \times 10^{-2}\ mol$$

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Question 5 (1 mark)

Kripa has set up a Copper-Nickel galvanic cell for a school experiment.

a. Write the half equations for the galvanic cell. (1 mark)

Reduction Half reaction:

$Cu^{2+}(aq) + 2e^{-} \rightleftharpoons Cu(s)$

Oxidation Half reaction:

$Ni_{(s)} \rightarrow Ni^{2+}_{(aq)} + 2e^{-}$

b. Kripa runs the cell for 17.5 minutes and finds that 6.75 A of current passed through the cell. Calculate the mass of metal deposited on the electrode.

$n(e^{-}) = \frac{It}{F}$

$= \frac{17.5 \times 60 \times 6.75}{96500}$

$= 0.07345\ mol$

$\therefore m(Cu) = \frac{0.07345}{2} \times 63.5$

$= 2.33g$

c. The setup is reset and the experiment is run again and Kripa notices 5.65 g of copper produced. Given that 5.45 A of current was passed through the cell, calculate the time (in seconds) for which the cell was running.

$n(Cu) = \frac{5.65}{63.5}$

$= 0.08898\ mol$

$\therefore n(e^{-}) = \frac{It}{F} = 0.08898 \times 2$

$\therefore t = \frac{0.08898 \times 2 \times 96500}{5.45}$

$= 3151\ s$

$\simeq 3.15 \times 10^{3}$ seconds

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Question 6 (2 marks)

Scott sets up a galvanic cell where chromium metal oxidises to form chromium (III) ions. Given that the change in mass of chromium was 4.55 g and the cell ran for 450 seconds, calculate the current running through the cell.

$$n(Cr) = \frac{4.55}{52}$$

$$= 0.0875\ mol$$

As $Cr_{(s)} \rightarrow Cr^{3+}_{(aq)} + 3e^-$,

$$\therefore n(e^-) = 3 \times 0.0875$$

$$= 0.2625\ mol$$

$$\therefore \frac{It}{F} = 0.2625$$

$$I = \frac{0.2625 \times 96500}{450}$$

$$\approx 56.3\ A$$

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Sub-Section [1.10.3]: Calculate the Charge of a Metal

Question 7 (1 mark)

Kevin runs a galvanic cell where 4.56 mol of iron is formed on the cathode. Given that the moles of electrons running through the cell is 2.28 mol, calculate the charge of the iron ions in the cell.

$$\frac{4.56}{2.28} = 2$$

Fe^{2+}

Question 8 (2 marks)

Shiven sets up a galvanic cell in the school laboratory and notes that 32.4 g of manganese metal deposits on the electrode. Given that the amount of electrons running through the cell is 1.18 mol, calculate the charge of the manganese ions in the cell.

$n(Mn) = \frac{32.4}{54.9}$

$= 0.5902\ mol$

$\therefore \frac{n(e^-)}{n(Mn)} = \frac{1.18}{0.5902}$

$\approx 2 \quad \therefore Mn^{2+}$

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Question 9 (3 marks)

Justin runs a galvanic cell for 3.0 minutes and finds that 0.113 g of chromium metal has been deposited at the cathode. Given that 3.5 A of current was running through the cell, determine the charge of the chromium ion.

$n(e^-) = \frac{It}{F}$

$n(Cr) = \frac{0.113}{52}$

$= 0.00217\ mol$

$n(e^-) = \frac{3.5 \times 180}{96500}$

$= 0.006528\ mol$

Let 'x' represent the charge of Cr

$\therefore n(e) = x \times n(Cr)$

$x = \frac{0.006528}{0.00217}$

$x \approx +3$

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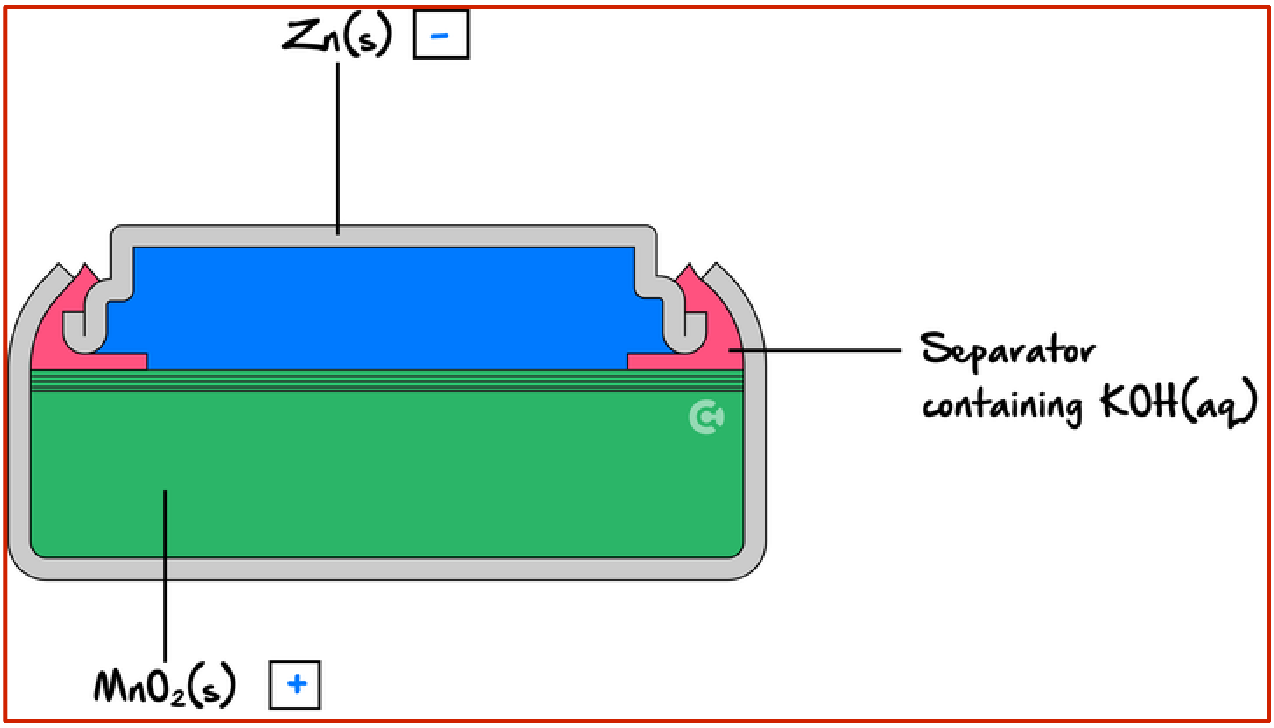
Sub-Section: The 'Final Boss'

Question 10 (10 marks)

A diagram of a Manganese-Oxide battery has been shown below.



Alina, curious about the cell, researches about it and finds some information about the half equations, which have been provided below.

$$2MnO_2(s) + H_2O(s) + 2e^- \rightarrow Mn_2O_3(s) + 2OH^-(aq)$$

$$Zn(s) + 2OH^-(aq) \rightarrow ZnO(s) + 2e^- + H_2O(s)$$

a. Referring to the diagram cell above:

i. Label the positive and negative electrodes in the boxes provided above. (1 mark)

ii. Explain the purpose of KOH(aq) in the cell. (1 mark)

The KOH acts as an electrolyte, allowing for the balancing of charge around each half cell due to the movement of ions.

b. Experimenting with the battery, Alina finds that 2.35 g of Mn_2O_3 is produced when the cell is running for 15.5 minutes.

i. Find the amount of electrons, in moles, which have passed through the circuit. (2 marks)

$$n(Mn_2O_3) = \frac{2.35}{54.9 \times 2 + 16 \times 3}$$

$$= 0.01489 \text{ mol}$$

$$n(e^-) = 2 \times 0.01489$$

$$= 0.0298\, mol \cong 2.98 \times 10^{-2}\, mol$$

ii. Find the current that ran through the circuit whilst the battery cell was operating. (2 marks)

$$n(e^-) = 0.0298\, mol$$

$$I = \frac{0.0298 \times 96500}{15.5 \times 60}$$

$$I = 3.09\, A$$

iii. Find the amount (in mole) of hydroxide consumed whilst the cell was running. (1 mark)

0 mole.
Whilst hydroxide is consumed in one half equation, an equal amount is produced in the other half equation, meaning that there is none consumed overall. (1)

c. Manganese (Mn) can also have other charges other than Mn(II). In another setup, Alina finds that 3650 C of charge runs through the cell and 0.519 g of manganese metal is produced. Find the charge of manganese in this new cell. (3 marks)

$n(e^-) = \frac{3650}{96500} = 0.0378\ mol$

$n(Mn) = \frac{0.519}{54.9} = 0.00945\ mol$

Let 'x' represent the charge of manganese:

$\frac{0.0378}{0.00945} = 4$

$\therefore Mn^{4+}$

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Section B: Supplementary Questions (25 Marks)

Sub-Section [1.10.1]: Identify Features of Primary Cells & how they Operate

Question 11 (3 marks)

Sodium metal is commonly used in batteries, especially in alkaline and Sodium-carbon cells for daily consumer use.

a. By referring to information provided in the Data Book, give one reason why Sodium is used as a reactant in these galvanic cells. (1 mark)

Sodium is a relatively strong reductant (-2.71V). This means that many reactants will have larger EMF values and act as a suitable oxidant, allowing for the galvanic cell to operate. (1)

b. When Sodium comes into contact with water, an explosion is observed. Using the half equations, explain this observation. (2 marks)

Sodium metal and water will spontaneously react releasing heat. As hydrogen gas produced accumulates, the heat produced can cause ignition of this gas causing an explosion. (1) (2) – half eqns

$$Na(s) \rightarrow Na^{+}(aq) + e^{-}$$

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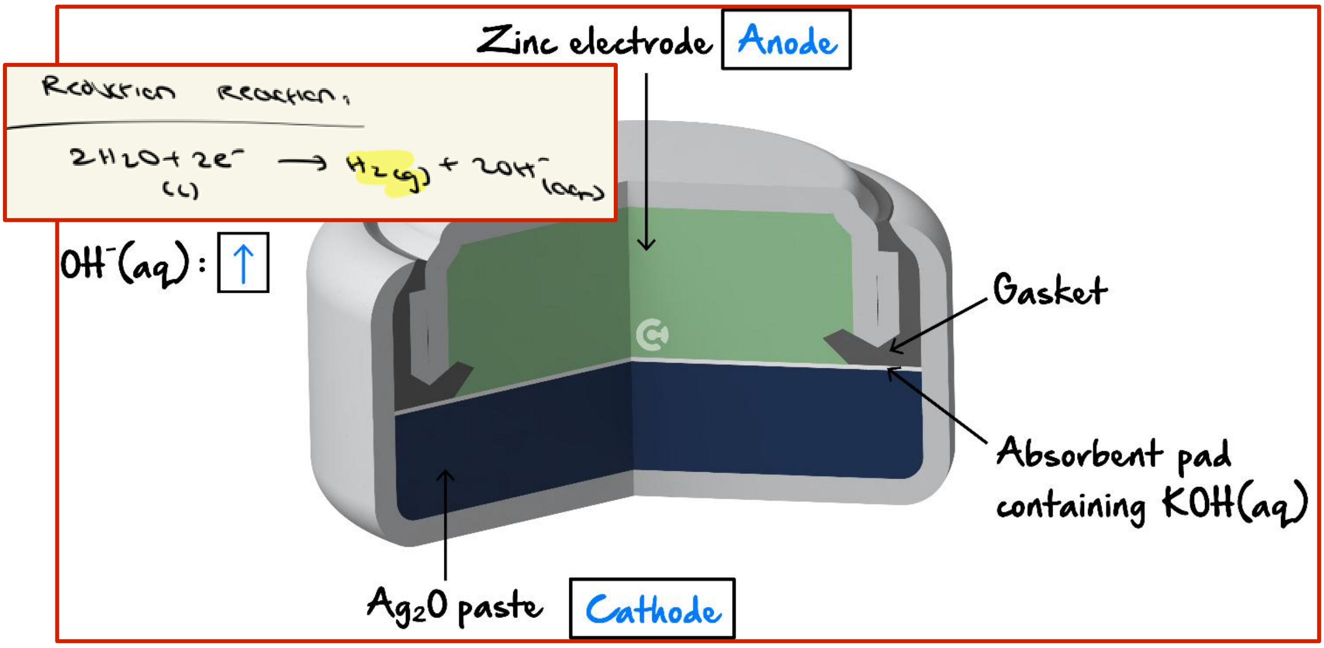
Question 12 (2 marks)

Zinc-Mercury Oxide batteries are commonly used commercially due to their long battery life and durability. A diagram of the cell has been provided below.



The half equations for the cell have also been provided below:

$$Zn(s) + 2OH^-(aq) \rightarrow ZnO(s) + H_2O(l) + 2e^-$$

$$HgO(s) + H_2O(l) + 2e^- \rightarrow Hg(l) + 2OH^-(aq)$$

a. Label the anode and the cathode in the boxes provided above. (1 mark)

b. Label the direction of movement of electrolyte in the cell in the box provided above. (1 mark)

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Question 13 (2 marks)

a. What is a key feature of primary cells that allows for their commercial usage over fuel cells? (1 mark)

A. Greater efficiency in the production of energy.

B. Cheaper electrodes that reduce the overall cost of the battery.

C. Separation of reactants into two half-cells is cheaper than a constant supply of reactants.

D. Less specific electrodes that are easier to source.

b. Which of the following outlines the properties required for electrodes in primary cells? (1 mark)

A. Porous, inert, catalytic, conduct electricity.

B. Reactive, catalytic, porous.

C. Conductive of electrons.

D. Porous, inert, conductive.

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Sub-Section [1.10.2]: Apply Faraday's First & Second Law and $Q = It$ & $Q = n(e)F$ to Calculations

Question 14 (4 marks)

Answer the following questions regarding three separate galvanic cells.

a. In a galvanic cell, 250 C of electric charge passes through the circuit in 20 minutes. Calculate the current, in A, running through the cell. (1 mark)

$I = \frac{Q}{t}$

$= \frac{250}{20 \times 60}$

$= 0.208\ A$

b. In another galvanic cell, 1.46 A of current runs through the cell during a 10.0 minute period. Calculate the moles of electrons produced in the cell. (1 mark)

$Q = 1.46 \times 10 \times 60$

$= 876\ C$

$n(e^-) = \frac{876}{96500}$

$= 0.00908\ mol$

$\approx 9.08 \times 10^{-3}\ mol$

c. Calculate the moles of zinc (Zn) produced in a cell, when 5.42 A of current is running through the circuit for 35 minutes. The half equation for zinc has been provided below: (2 marks)

$$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$$

$n(e^-) = \frac{It}{F}$

$= \frac{5.42 \times 35 \times 60}{96500}$

$= 0.117948\ mol$

$\therefore n(Zn) = \frac{0.117948}{2}$

$= 0.05897\ mol$

$\approx 5.90 \times 10^{-2}\ mol$

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Question 15 (3 marks)

Joanne has set up a Copper-Nickel galvanic cell for a school experiment.

a. Write the half equations for the galvanic cell. (1 mark)

Oxidation Half reaction:

$Cu^{2+}(aq) + 2e^{-} \rightleftharpoons Cu(s)$

Reduction Half reaction:

$Ni_{(s)} \rightarrow Ni^{2+}_{(aq)} + 2e^{-}$

b. Joanne runs the cell for 19.5 minutes and finds that 8.75 A of current passed through the cell. Calculate the mass of metal deposited on the electrode. (2 marks)

$n(e^{-}) = \frac{19.5 \times 60 \times 8.75}{96500}$

$= 0.106088\ mol$

$\therefore m(Cu) = \frac{0.106088}{2} \times 63.5$

$= 3.37 g$

c. The setup is reset and the experiment is run again and Joanne notices 4.55 g of copper produced. Given that 2.45 A of current was passed through the cell, calculate the time (in seconds) for which the cell was running for.

$n(Cu) = \frac{4.55}{63.5}$

$= 0.07165\ mol$

$n(e^{-}) = \frac{It}{F} = 0.07165 \times 2$

$\therefore t = \frac{0.07165 \times 2 \times 96500}{5.45}$

$= 2537.33\ s$

$\approx 2.54 \times 10^{3}$ seconds

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Question 16 (2 marks)

Scott sets up a galvanic cell where chromium metal oxidises to form chromium (III) ions. Given that the change in mass of chromium was 2.55 g and the cell ran for 350 seconds, calculate the current running through the cell.

$$n(Cr) = \frac{2.55}{52}$$

$$= 0.049\ mol$$

As $Cr_{(s)} \rightarrow Cr^{3+}_{(aq)} + 3e^-$,

$\therefore\ n(e^-) = 3 \times 0.049$

$$= 0.147\ mol$$

$$\therefore \frac{It}{F} = 0.147$$

$$I = \frac{0.147 \times 96500}{350}$$

$$= 40.5\ A$$

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Sub-Section [1.10.3]: Calculate the Charge of a Metal

Question 17 (1 mark)

Kevin runs a galvanic cell where 9.54 *mol* of iron is formed on the cathode. Given that the moles of electrons running through the cell is 3.18 *mol*, calculate the charge of the iron ions in the cell.

$$\frac{9.54}{3.18} = 3$$

Fe^{3+}

Question 18 (2 marks)

Shiven sets up a galvanic cell in the school laboratory and notes that 27.6 *g* of manganese metal deposits on the electrode. Given that the amount of electrons running through the cell is *mol*, calculate the charge of the manganese ions in the cell.

$$n(Mn) = \frac{27.6}{54.9} = 0.5027\ mol$$

$$\therefore \frac{n(e^-)}{n(Mn)} = \frac{1.01}{0.5027} \approx 2$$

$\therefore Mn^{2+}$

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Question 19 (3 marks)

A solution of titanium ions is reduced at a cathode, whereby it is found that a current of 2.75 A is produced over 9.75 hours. It is found that 15.9 g of titanium metal is deposited at the cathode. Find the charge of the titanium ions.

$Q = It = 2.75 \times 9.75 \times 60 \times 60$

$= 96525\ C$

$n(e^-) = \frac{Q}{F} = \frac{96525\ C}{96500} = 1.00\ mol$

$n(Ti) = \frac{m}{Mr} = \frac{15.9\ g}{47.9} = 0.332\ mol$

$n(Ti) : n(e^-)$

$0.332 : 1.00$

$1 : 3$

$\therefore Ti^{3+}$

Question 20 (3 marks)

A solution of iodine ions is reduced at a cathode, whereby it is found that a current of 5.80 A is produced over 4.04 hours. It is found that 22.2 g of iodine metal is deposited at the cathode. Find the charge of the iodine ions.

$Q = It = 5.80\ A \times 4.04 \times 60 \times 60$

$= 84355\ C$

$n(e^-) = \frac{Q}{F} = \frac{84355\ C}{96500} = 0.874\ mol$

$n(I) = \frac{m}{Mr} = \frac{22.2\ g}{126.9} = 0.175\ mol$

$n(I) : n(e^-)$

$\frac{0.175}{0.175} : \frac{0.874}{0.175}$

$1 : 4.99$

$\therefore I^{5+}$

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