Topic 4 - Cell Stoichiometry - pgs 652-656

- In the production of elements, the refining of metals, and electroplating, the quantity of electricity that passes through a cell determines the masses of substances that react or are produced at the electrodes

- We know from oxidation and reduction half-reactions, a specific number of electrons are lost or gained

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

- This reduction half reaction tells us that 2 moles of electrons must be gained to produce 1 mole of zinc.

- This gives us our mole ratio that we can use in stoichiometry

- In order for us to be able to use electrons as part of our stoihiometric calculations, we have to be able to determine the number of moles or electrons that are used in our voltaic and electrolytic cells

- We need to see how the charge is determined before we can calculate moles of electrons

- Charge, *Q*, in coulombs, is determined from the electric current, *I*, in amperes (coulombs per second), and the time, *t*, in seconds, according to the following definition:

charge

time

$$Q = It$$

current

$300{,}000\ A = \frac{C}{s}$

Ex: Modern electrolytic cells may use up to 300 kA of current. What is the charge that passes through one of these cells in a 24 h period?

$$Q = It$$

$$= (300000\ \tfrac{C}{\cancel{s}})(86400\ \cancel{s})$$

$$= \boxed{2.6 \times 10^{10}\ C}$$

Faraday's Law

96500 C

- Micheal Faraday found that *9.65 x 10^4 C of charge is transferred for every mole of electrons that flows in the cell.*

6.02×10^{23}

- In modern terms, this value is the molar charge of electrons, also called the **Faraday constant, *F*.**

$$F = 9.65 \times 10^4 \frac{\text{C}}{\text{mol e}^-}$$

- We can use this number to convert electric charge to an amount in moles of electrons

- This is similar to the way that molar mass is used to convert mass to a chemical amount

current (A)

time (s)

$$n_{e^-} = \frac{Q}{F} \xrightarrow{\text{because } \boxed{Q = It}} \boxed{n_{e^-} = \frac{It}{F}}$$

n_{e^-} 0.150 C/s

Ex. What amount of electrons is transferred in a cell that operates for 1.25 h at a current of 0.150 A?

$I = 0.150\ A$

$t = 1.25 h \times \frac{3600 s}{h}$

$= 4500 s$

$n_{e^-} = \frac{It}{F} = \frac{(0.150 \frac{C}{s})(4500 s)}{96500\ C/mol}$

$\boxed{n_{e^-} = 0.006994 \ldots mol}$

$\rightarrow \boxed{= 0.00699 mol\ e^-}$

Half-Cell Calculations

- the mass of an element produced at an electrode depends on the amount of transferred electrons

- a half-reaction equation showing the number of electrons involved is necessary to do stoichiometric calculations

- Separate calculations are carried out for each electrode, although the *same charge and, therefore, the same amount of electrons passes through each electrode in a cell*

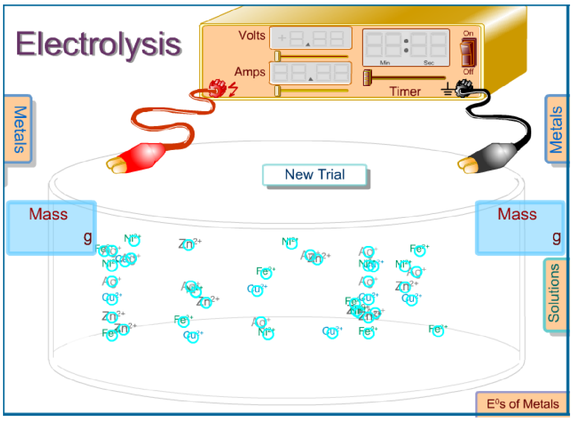
SUMMARY **Procedure for Half-Cell Stoichiometry**

Step 1: Write the balanced equation for the half-cell reaction of the substance produced or consumed. List the measurements and conversion factors for the given and required entities.

Step 2: Convert the given measurements to an amount in moles by using the appropriate conversion factor (*M*, *c*, *F*).

Step 3: Calculate the amount of the required substance by using the mole ratio from the half-reaction equation.

Step 4: Convert the calculated amount to the final quantity by using the appropriate conversion factor (*M*, *c*, *F*).



__Ex.__ What is the mass of copper deposited at the cathode of a copper electrorefining cell operated at 4.0 A for 20.0 min?

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

$I = 4.0A$ $\quad m = ??$

$t = 1200s$

$n_e = \frac{It}{F} = 0.0497...mol \xrightarrow{\times \frac{1}{2}} n = 0.02487...mol$

$M = 63.55 g/mol$

$\boxed{m = 1.6g}$

__Ex.__ Silver is deposited on objects in a silver electroplating cell. If 0.175 g of silver is to be deposited from a silver cyanide solution in a time of 10.0 min, predict the current required.

$Ag^+_{(aq)} + 1e^- \rightarrow 1 Ag_{(s)}$

$I = ??$ $\quad m = 0.175 g$

$t = 600 s$ $\quad M = 107.87 g/mol$

$n_e = 0.0016...mol \xleftarrow{\times \frac{1}{1}} n = 0.001622...mol$

$n_e = \frac{It}{F}$

$I = \frac{n_e F}{t} = \boxed{0.261 A}$

__Ex.__ A 25.72 g piece of copper metal is the anode in a cell in which a current of 0.876 A flows for 75.0 min. Determine the final mass of the copper electrode

$Cu_{(s)} \rightarrow Cu^{2+} + 2e^-$

$n = 0.02042...mol \xleftarrow{\times \frac{1}{2}} n_e = 0.0408...mol$

$M = 63.55 g/mol$ $\quad I = 0.876A$

$m_{used} = 1.2976...$ $\quad t = 4500s$

$\boxed{m_{final} = 24.4g}$

$1.298...g$ $\quad n_e = \frac{It}{F}$

$=$

$m_{left} = 25.72g - 1.298...g$

$= 24.4219...g$

$\boxed{= 24.4g}$

Practice Sheet 11

1. Calculate the charge transferred by a current of 1.5 A flowing for 30 s.

$Q = It = (1.5\frac{C}{s})(30s) = \boxed{45C}$

2. In an electrolytic cell, 87.6 C of charge is transferred in 22.5 s. Determine the electric current.

$I = \frac{Q}{t} = \frac{87.6C}{22.5s} = \boxed{3.89A}$

3. Calculate the charge transferred by a current of 250 mA in a time of 28.5 s.

$Q = It = (0.250A)(28.5s) = \boxed{7.13C}$

4. How long, in minutes, does it take a current of 1.60 A to transfer a charge of 375 C?

$t = \frac{Q}{I} = \frac{375C}{1.60A} = 234.375s = \boxed{3.91min}$

5. An electroplating cell operates for 35 min with a current of 1.9 A. Calculate the amount, in moles, of electrons transferred.

$n_e = \frac{It}{F} = \frac{(1.9A)(2100s)}{96500\ C/mole^-} = 0.0413...\ mol\ e^- \quad \boxed{0.041\ mol\ e^-}$

6. A cell transferred 0.146 mol of electrons with a constant current of 1.24 A. How long, in hours, did this take?

3.16 h

7. Calculate the current required to transfer 0.015 mol of electrons in 20 min.

1.2 A

8. A family wishes to plate an antique teapot with 10.00 g of silver. If the current to be used is 1.80 A, what length of time, in minutes, is required?

$Ag^+ + 1e^- \rightarrow 1Ag_{(s)}$

$n_e = 0.0927...$ $\quad m = 10.00g$

$I = 1.80A$ $\quad M = 107.87g/mol$

$t = ??$ $\quad n = 0.0927...\ mol$

$n_e = \frac{It}{F}$

$t = \frac{n_e F}{I} = 4969.97...s = \boxed{82.8min}$

9. A typical Hall–Héroult cell produces 425 kg of molten aluminium in 24.0 h. Calculate the current used.

$Al^{3+}_{(l)} + 3e^- \rightarrow Al_{(l)}$

$n_e = 47257.22... \xleftarrow{\times 3} n = 15752.409...\ mol$

$m = 425000g$

$M = 26.98g/mol$

$t = 86400s$

$F = 96500\frac{C}{mol}$

$n_e = \frac{It}{F}$

$I = \frac{n_e F}{t} = 52781.5...A$

$= 52.8\ kA$

10. Magnesium metal is produced in an electrolytic cell containing molten magnesium chloride. A current of 2.0×10^5 A is passed through the cell for 18.0 h.
(a) Determine the mass of magnesium produced.

1.6×10^6 g

(b) What mass of chlorine is produced at the same time?

4.8×10^6 g

11. Cobalt metal is plated from 250.0 mL of cobalt(II) sulphate solution. What is the minimum concentration of cobalt(II) sulfate required for this cell to operate for 2.05 h with a current of 1.14 A?

0.174 mol/L

12. A student reconstructs Volta's electric battery using sheets of copper and zinc, and a current of 0.500 A is produced for 10.0 min. Calculate the mass of zinc oxidized to aqueous zinc ions.

0.102 g

13. Electroplating is a common technological process for coating objects with a metal to enhance the appearance of the object or its resistance to corrosion.
(a) A car bumper is plated with chromium using chromium(III) ions in solution. If a current of 54 A flows in the cell for 45 min 30 s, determine the mass of chromium deposited on the bumper.

26 g

(b) For corrosion resistance, a steel bolt is plated with nickel from a solution of nickel(II) sulfate. If 0.250 g of nickel produces a plating of the required thickness and a current of 0.540 A is used, predict how long in minutes the process will take.

25.4 min