

# SOLUTION DPP-7

## Electrolytic Cell (Quantitative)

Chapter: Electrochemistry

### TYPE-1 : Basic Faraday's Law Calculations

#### Solution 1

##### Explanation

$\text{Ba}^{2+} + 2\text{e}^- \longrightarrow \text{Ba}$ . So  $n$ -factor = 2.

Charge = 96500 C = 1 F.

Moles of Ba deposited =  $\frac{Q}{nF} = \frac{96500}{2 \times 96500} = 0.5$  mol.

##### Approach

Faraday's law is like a vending machine: you insert coins (charge) and get items (metal).  $\text{Ba}^{2+}$  needs 2 coins (electrons) per atom. With 1 F (= 1 mol of electrons), you get  $\frac{1}{2}$  mol of Ba.

##### Answer

(1) 0.5 mol

#### Solution 2

##### Explanation

$\text{Ag}^+ + \text{e}^- \longrightarrow \text{Ag}$ ,  $n = 1$ .

$Q = I \times t = 965 \times 1 = 965$  C.

Mass =  $\frac{M \times Q}{n \times F} = \frac{107.87 \times 965}{1 \times 96500} = 1.0787$  g.

##### Approach

965 C is exactly  $\frac{1}{100}$  of a Faraday. Silver needs 1 electron per atom, so you get  $\frac{1}{100}$  mol of Ag =  $\frac{107.87}{100} = 1.0787$  g.

##### Answer

(3) 1.0787 g

#### Solution 3

##### Explanation

$\text{Zn}^{2+} + 2\text{e}^- \longrightarrow \text{Zn}$ ,  $n = 2$ ,  $M = 65$  g/mol.

$Q = I \times t = 5 \times 40 \times 60 = 12000$  C.

Mass =  $\frac{65 \times 12000}{2 \times 96500} = \frac{780000}{193000} = 4.041$  g  $\approx 4.065$  g.

### Approach

Convert time to seconds first (40 min = 2400 s). Then plug into the “Faraday ATM”: deposit =  $\frac{M \times I \times t}{nF}$ . Zinc needs 2 electrons per atom (like paying double fare).

### Answer

(4) 4.065 g

### Solution 4

#### Explanation

$\text{Cu}^{2+} + 2e^- \longrightarrow \text{Cu}$ ,  $n = 2$ ,  $M = 63.5$  g/mol.

$Q = 10 \times 10 \times 60 = 6000$  C.

$$\text{Mass} = \frac{63.5 \times 6000}{2 \times 96500} = \frac{381000}{193000} = 1.974 \text{ g} \approx 1.97 \text{ g}.$$

### Approach

Same formula, same vending machine. Cu needs 2 electrons. 10 A for 10 min = 6000 C. Plug and chug!

### Answer

(1) 1.97 g

### Solution 5

#### Explanation

$\text{Co}^{2+} + 2e^- \longrightarrow \text{Co}$ ,  $n = 2$ ,  $M = 59$  g/mol.

$Q = 10 \times 109 \times 60 = 65400$  C.

$$\text{Mass} = \frac{59 \times 65400}{2 \times 96500} = \frac{3858600}{193000} = 19.99 \text{ g} \approx 20.0 \text{ g}.$$

### Approach

Cobalt(II) means  $\text{Co}^{2+}$  — needs 2 electrons.  $10 \times 109 \times 60 = 65400$  C, which is about  $\frac{2}{3}$  F. So moles of Co =  $\frac{65400}{2 \times 96500} = 0.3388$  mol  $\times 59 \approx 20$  g.

### Answer

(3) 20.0

### Solution 6

#### Explanation

$\text{Al}^{3+} + 3e^- \longrightarrow \text{Al}$ ,  $n = 3$ ,  $M = 27$  g/mol.

$Q = 4.0 \times 10^4 \times 6 \times 3600 = 8.64 \times 10^8$  C.

$$\text{Mass} = \frac{27 \times 8.64 \times 10^8}{3 \times 96500} = \frac{2.3328 \times 10^{10}}{289500} = 8.058 \times 10^4 \text{ g} \approx 8.1 \times 10^4 \text{ g}.$$

### Approach

Industrial scale problem! Huge current ( $4 \times 10^4$  A) for 6 hours. Al needs 3 electrons per atom (expensive metal to electrolyse — like buying a luxury item from the vending machine). Convert hours to seconds, then use the master formula.

### Answer

(3)  $8.1 \times 10^4$  g

### Solution 7

#### Explanation

$\text{Ag}^+ + \text{e}^- \longrightarrow \text{Ag}$ ,  $n = 1$ ,  $M = 108$  g/mol.

$Q = 9650$  C =  $\frac{9650}{96500} = 0.1$  F.

Mass =  $0.1 \times \frac{108}{1} = 10.8$  g.

### Approach

$9650$  C =  $\frac{1}{10}$  of a Faraday. Silver needs just 1 electron per atom. So you deposit  $\frac{1}{10}$  mol Ag = 10.8 g. Clean numbers = quick solve!

### Answer

(4) 10.8 g

### Solution 8

#### Explanation

At cathode:  $2\text{H}_2\text{O} + 2\text{e}^- \longrightarrow \text{H}_2 + 2\text{OH}^-$ . The  $\text{OH}^-$  combines with  $\text{Na}^+$  to form NaOH.

$Q = 9.65 \times 1000 = 9650$  C = 0.1 F.

Moles of  $\text{OH}^-$  = moles of electrons = 0.1 mol.

Mass of NaOH =  $0.1 \times 40 = 4.0$  g.

### Approach

In NaCl electrolysis,  $\text{Na}^+$  doesn't deposit — instead water breaks at cathode giving  $\text{OH}^-$ . Each electron produces one  $\text{OH}^-$ , which pairs with  $\text{Na}^+$  to give NaOH. So moles of NaOH = moles of electrons = 0.1.

### Answer

(3) 4.0 g

### Solution 9

#### Explanation

Same as Q8.  $Q = 9.65 \times 1000 = 9650$  C = 0.1 F.

Moles of NaOH = 0.1 mol.

Mass =  $0.1 \times 40 = 4.0$  g.

### Approach

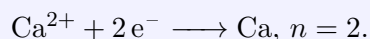
Identical to Q8 — same current, same time, same answer. Sometimes exams repeat concepts to check if you're consistent!

### Answer

(2) 4.0 g

## Solution 10

### Explanation



$$Q = 25 \times 10^{-3} \times 60 = 1.5 \text{ C.}$$

$$\text{Moles of Ca} = \frac{Q}{nF} = \frac{1.5}{2 \times 96500} = 7.772 \times 10^{-6} \text{ mol.}$$

$$\text{Atoms} = 7.772 \times 10^{-6} \times 6.022 \times 10^{23} = 4.68 \times 10^{18}.$$

### Approach

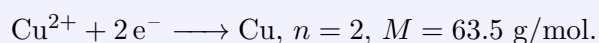
Tiny current (milliamperes) for short time = tiny charge. Convert to moles using Faraday's law, then multiply by Avogadro's number to get atoms. Like counting individual coins from a small piggy bank.

### Answer

(1)  $4.68 \times 10^{18}$

## Solution 11

### Explanation



$$Q = 20 \times 965 = 19300 \text{ C.}$$

$$\text{Theoretical mass} = \frac{63.5 \times 19300}{2 \times 96500} = \frac{1225550}{193000} = 6.35 \text{ g.}$$

$$\text{At 50\% efficiency: Actual mass} = 6.35 \times 0.5 = 3.175 \text{ g.}$$

### Approach

First calculate as if 100% efficient, then multiply by efficiency. Think of it like a leaky pipe — only 50% of the water (current) actually reaches the plant (deposits Cu). So halve the theoretical answer.

### Answer

(2) 3.175 g

## Solution 12

### Explanation

$\text{Cd}^{2+} + 2e^- \longrightarrow \text{Cd}$ ,  $n = 2$ ,  $M = 112.4 \text{ g/mol}$ .

Current from lamp:  $I = \frac{P}{V} = \frac{100}{110} = \frac{10}{11} \text{ A}$ .

$Q = \frac{10}{11} \times 10 \times 3600 = \frac{360000}{11} = 32727.3 \text{ C}$ .

Mass =  $\frac{112.4 \times 32727.3}{2 \times 96500} = \frac{3678630}{193000} = 19.06 \text{ g}$ .

### Approach

Series circuit trick! The lamp controls the current ( $I = P/V$ ). That same current flows through the electrolytic cell. It's like the lamp is the "gatekeeper" deciding how much current the cell gets. Once you find  $I$ , the rest is standard Faraday.

### Answer

(1) 19.06 g

## TYPE-2 : Time-Based Electrolysis Problems

### Solution 13

### Explanation

1 mol of electrons = 1 F = 96500 C.

$Q = I \times t \Rightarrow t = \frac{Q}{I} = \frac{96500}{10^{-16}} = 9.65 \times 10^{20} \text{ s}$ .

### Approach

Tiny current ( $10^{-16} \text{ A}$ ) delivering 1 F of charge? That's like filling an Olympic swimming pool with an eyedropper. Time = total charge  $\div$  current. The math is simple, the wait is astronomical!

### Answer

(3)  $9.65 \times 10^{20} \text{ s}$

### Solution 14

### Explanation

$2\text{H}_2\text{O} \longrightarrow 2\text{H}_2 + \text{O}_2$ . For 2 mol  $\text{H}_2\text{O}$ , 4 mol electrons needed (2 for  $\text{H}_2$  at cathode + 2 for  $\text{O}_2$  at anode per mol  $\text{H}_2\text{O}$  — actually total  $4e^-$  per 2  $\text{H}_2\text{O}$ ).

36 g  $\text{H}_2\text{O} = 2 \text{ mol}$ . Total electrons needed = 4 mol = 4 F.

$t = \frac{Q}{I} = \frac{4 \times 96500}{3} = 128667 \text{ s} = 35.74 \text{ hr}$ .

### Approach

Breaking water needs 2 electrons per  $\text{H}_2\text{O}$  molecule (considering the overall decomposition). 2 mol water needs 4 F. Then just divide total charge by current to get time. Like asking: "How long to empty a tank if the drain rate is 3 litres/second?"

**Answer****(3) 35.74 hour****Solution 15****Explanation**

$2 \text{Cl}^- \longrightarrow \text{Cl}_2 + 2 \text{e}^-$ . For 0.10 mol  $\text{Cl}_2$ , electrons needed =  $0.10 \times 2 = 0.2 \text{ mol} = 0.2 \text{ F}$ .

$Q = 0.2 \times 96500 = 19300 \text{ C}$ .

$t = \frac{19300}{3} = 6433 \text{ s} = 107.2 \text{ min} \approx 110 \text{ min}$ .

**Approach**

Each  $\text{Cl}_2$  molecule needs 2 electrons (two  $\text{Cl}^-$  each donate one). So 0.1 mol  $\text{Cl}_2$  needs 0.2 F of charge. Divide by current to get time.

**Answer****(4) 110 minutes****TYPE-3 : Electrons, Charge & Faraday Constants****Solution 16****Explanation**

1 mole of electrons =  $N_A \times e = 6.022 \times 10^{23} \times 1.6 \times 10^{-19} \text{ C} = 96500 \text{ C} = 1 \text{ F}$ .

**Approach**

Faraday's constant is just Avogadro's number  $\times$  charge on one electron. It's the "bulk price" of electrons — how much charge a whole mole of them carries.

**Answer****(2) 96500 C****Solution 17****Explanation**

$Q = I \times t = 1 \times 60 = 60 \text{ C}$ .

Number of electrons =  $\frac{Q}{e} = \frac{60}{1.6 \times 10^{-19}} = 3.75 \times 10^{20}$ .

**Approach**

Total charge  $\div$  charge per electron = number of electrons. Simple division — like counting how many 1 coins make up 60.

**Answer****(1)  $3.75 \times 10^{20}$**

### Solution 18

#### Explanation

1 gram equivalent of any substance is deposited by exactly 1 Faraday of charge. This is Faraday's first law. Regardless of the element or its valency, 1 equivalent always needs 1 F.

#### Approach

This is the definition of a Faraday! 1 F deposits 1 gram equivalent — always. It doesn't matter if it's Al, Cu, or Ag. The "equivalent weight" already accounts for the number of electrons needed.

#### Answer

(1) 1

### Solution 19

#### Explanation

$\text{Cr}_2\text{O}_7^{2-} \longrightarrow 2\text{Cr}^{3+}$ . Cr goes from +6 to +3, i.e., gains  $3e^-$  per Cr atom.  
For 1 mol  $\text{Cr}_2\text{O}_7^{2-}$ : 2 Cr atoms  $\times 3e^-$  each = 6 mol electrons = 6 F.

#### Approach

Two Cr atoms in dichromate, each dropping from +6 to +3 (3 electron gain each). Total =  $2 \times 3 = 6$  electrons per formula unit. Like two people each climbing down 3 stairs.

#### Answer

(1) 6 F

### Solution 20

#### Explanation

$\text{MnO}_4^- \longrightarrow \text{Mn}^{2+}$ . Mn goes from +7 to +2, gaining 5 electrons.  
For 1 mol  $\text{MnO}_4^-$ : 5 mol electrons = 5 F.

#### Approach

Permanganate in acidic medium: Mn drops from +7 to +2. That's a 5-electron jump — like an elevator going down 5 floors. Each floor = 1 Faraday.

#### Answer

(1) 5

### Solution 21

#### Explanation

$\text{MnO}_4^{2-} \longrightarrow \text{MnO}_4^-$ . Mn goes from +6 to +7, losing 1 electron.  
For 0.1 mol: charge =  $0.1 \times 1 \times 96500 = 9650$  C.

### Approach

Manganate to permanganate: Mn goes from +6 to +7 — only 1 electron lost per ion. For 0.1 mol, that's 0.1 F = 9650 C. Small oxidation state change = small charge needed.

### Answer

**(3) 9650 C**

## Solution 22

### Explanation

$\text{MnO}_4^- \longrightarrow \text{MnO}_2$ . Mn goes from +7 to +4, gaining 3 electrons.  
For 1 mol: 3 F of charge required.

### Approach

Permanganate to  $\text{MnO}_2$ : Mn drops from +7 to +4. That's 3 floors down in the elevator = 3 F.

### Answer

**(2) 3F**

## Solution 23

### Explanation

$\text{KMnO}_4$ : Mn is +7.  $\text{Mn}^{2+}$ : Mn is +2.  
Change = +7  $\rightarrow$  +2 = 5 electrons gained per Mn.  
Charge = 5 F per mol.

### Approach

Same as Q20!  $\text{KMnO}_4/\text{Mn}^{2+}$  is the classic acidic medium permanganate reaction. 7 to 2 = 5 electron drop = 5F. This is one of the most frequently tested conversions.

### Answer

**(3) 5F**

## Solution 24

### Explanation

$\text{Ca}^{2+} + 2e^- \longrightarrow \text{Ca}$ . So 1 mol of Ca needs 2 mol of electrons.

### Approach

Calcium has +2 charge, so it needs 2 electrons to become neutral — like paying a 2 bill with two 1 coins.

### Answer

**(2) 2 mole of electrons**

## Solution 25

### Explanation

1 F liberates 1 mol of metal only when  $n = 1$  (monovalent cation).

$\text{AuCl}_3$ :  $\text{Au}^{3+}$ ,  $n = 3$  — needs 3 F.

$\text{AgNO}_3$ :  $\text{Ag}^+$ ,  $n = 1$  — needs 1 F.

$\text{CaCl}_2$ :  $\text{Ca}^{2+}$ ,  $n = 2$  — needs 2 F.

$\text{CuSO}_4$ :  $\text{Cu}^{2+}$ ,  $n = 2$  — needs 2 F.

### Approach

1 F = 1 mol electrons. Only a +1 ion needs exactly 1 electron per atom, so 1 F gives exactly 1 mol. Higher charges need more electrons per atom, so 1 F won't complete the job for a full mole. Silver nitrate is the only +1 option.

### Answer

(2) Silver nitrate

## Solution 26

### Explanation

$\text{Al}^{3+} + 3e^- \longrightarrow \text{Al}$ ,  $M = 27$  g/mol.

Moles of Al =  $\frac{13.5}{27} = 0.5$  mol.

Faradays =  $0.5 \times 3 = 1.5$  F.

### Approach

Work backwards: find moles of Al deposited, then multiply by  $n$ -factor (3 for  $\text{Al}^{3+}$ ). Each Al atom needs 3 electrons, so 0.5 mol Al needs  $0.5 \times 3 = 1.5$  mol electrons = 1.5 F.

### Answer

(3) 1.5

## Solution 27

### Explanation

$\text{H}_2\text{O}_2 \longrightarrow \text{O}_2 + 2\text{H}^+ + 2e^-$ . O goes from  $-1$  in  $\text{H}_2\text{O}_2$  to  $0$  in  $\text{O}_2$ , losing  $1e^-$  per O atom,  $2e^-$  per molecule.

$17$  g  $\text{H}_2\text{O}_2 = \frac{17}{34} = 0.5$  mol.

Charge =  $0.5 \times 2 = 1$  F.

### Approach

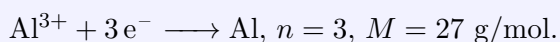
In  $\text{H}_2\text{O}_2$ , oxygen is in  $-1$  state (unusual!). Converting to  $\text{O}_2$  means oxygen goes to  $0$  — each O loses 1 electron. Two O atoms per molecule = 2 electrons per  $\text{H}_2\text{O}_2$ . For 0.5 mol:  $0.5 \times 2 = 1$  F.

### Answer

(1) 1F

### Solution 28

#### Explanation



$$\text{Moles of Al} = \frac{5120}{27} = 189.63 \text{ mol.}$$

$$Q = 189.63 \times 3 \times 96500 = 5.49 \times 10^7 \text{ C.}$$

#### Approach

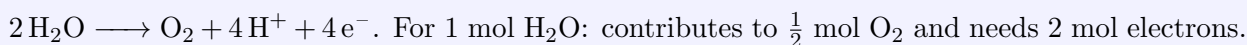
5.12 kg = 5120 g. Find moles of Al, multiply by 3 (each Al needs 3 electrons), then multiply by 96500 to convert to coulombs. Industrial-scale problem!

#### Answer

**(3)**  $5.49 \times 10^7 \text{ C}$  of electric charge

### Solution 29

#### Explanation



$$Q = 2 \times 96500 = 1.93 \times 10^5 \text{ C.}$$

#### Approach

Overall: 2 H<sub>2</sub>O gives 1 O<sub>2</sub> needing 4 electrons. So 1 mol H<sub>2</sub>O needs 2 electrons. Charge =  $2 \times 96500 = 1.93 \times 10^5 \text{ C}$ .

#### Answer

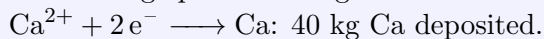
**(3)**  $1.93 \times 10^5 \text{ C}$

## TYPE-4 : Comparison & Ratio Problems (Series Cells)

### Solution 30

#### Explanation

Same charge passes through both cells.



$$\text{Moles of Ca} = \frac{40000}{40} = 1000 \text{ mol. Electrons used} = 2000 \text{ mol.}$$



$$\text{Mass of Al} = 666.7 \times 27 = 18000 \text{ g} = 18 \text{ kg.}$$

#### Approach

Same current, same time = same charge. Use the ratio:  $\frac{m_1}{m_2} = \frac{E_1}{E_2}$ , where  $E = \frac{M}{n}$  is the equivalent weight.  $\frac{40}{m_{\text{Al}}} = \frac{40/2}{27/3} = \frac{20}{9}$ . So  $m_{\text{Al}} = 40 \times \frac{9}{20} = 18 \text{ kg}$ .

#### Answer

**(2)** 18 kg

### Solution 31

#### Explanation

Same charge, series connection.

$$\frac{m_{Ni}}{m_{Cr}} = \frac{E_{Ni}}{E_{Cr}} = \frac{59/2}{52/3} = \frac{29.5}{17.33}$$

$$m_{Cr} = 0.3 \times \frac{17.33}{29.5} = 0.176 \text{ g} \approx 0.17 \text{ g}.$$

#### Approach

Series cells = same charge through both. Use the equivalent weight ratio shortcut:  $\frac{m_1}{m_2} = \frac{M_1/n_1}{M_2/n_2}$ . Ni is +2, Cr is +3. Plug in and solve for the unknown mass.

#### Answer

(2) 0.17 g

### Solution 32

#### Explanation

$$\frac{m_{Na}}{m_{Al}} = \frac{E_{Na}}{E_{Al}} = \frac{23/1}{27/3} = \frac{23}{9}$$

$$m_{Al} = 2.3 \times \frac{9}{23} = 0.9 \text{ g}.$$

#### Approach

Na is +1 (equivalent weight = 23), Al is +3 (equivalent weight = 9). Same charge means the ratio of masses = ratio of equivalent weights. Quick proportion gives 0.9 g.

#### Answer

(2) 0.9 gm

### Solution 33

#### Explanation

$$\frac{m_{Na}}{m_{Al}} = \frac{23/1}{27/3} = \frac{23}{9}$$

$$m_{Al} = 4.6 \times \frac{9}{23} = 1.8 \text{ g}.$$

#### Approach

Same as Q32 but with double the Na (4.6 g instead of 2.3 g). So Al deposited also doubles:  $0.9 \times 2 = 1.8 \text{ g}$ .

#### Answer

(2) 1.8 g

### Solution 34

#### Explanation

$$\frac{m_{Ag}}{m_{Al}} = \frac{E_{Ag}}{E_{Al}} = \frac{108/1}{27/3} = \frac{108}{9} = 12$$
$$m_{Al} = \frac{54}{12} = 4.5 \text{ g.}$$

#### Approach

Silver's equivalent weight is 108 (monovalent), aluminium's is 9 (trivalent). The ratio is 12:1. So if 54 g Ag deposits, Al = 54/12 = 4.5 g. Equivalent weight ratios are the fastest shortcut for series cell problems!

#### Answer

(1) 4.5 g

### Solution 35

#### Explanation

Cathode:  $\text{Al}^{3+} + 3\text{e}^- \longrightarrow \text{Al}$ . For 270 kg Al: moles =  $\frac{270000}{27} = 10000$  mol. Electrons = 30000 mol.

Anode:  $\text{C} + 2\text{O}^{2-} \longrightarrow \text{CO}_2 + 4\text{e}^-$ . For 30000 mol  $\text{e}^-$ : moles of C =  $\frac{30000}{4} = 7500$  mol.

Mass of C = 7500 × 12 = 90000 g = 90 kg.

#### Approach

In the Hall process, carbon anode reacts with oxide ions to form  $\text{CO}_2$ . Each C atom gives up 4 electrons (going from 0 to +4 in  $\text{CO}_2$ ). Balance the electrons: Al side needs 30000 mol  $\text{e}^-$ , carbon side provides 4 per atom. So 30000/4 = 7500 mol C = 90 kg.

#### Answer

(3) 90 kg

### Solution 36

#### Explanation

Series connection: same charge through both cells.

$\text{Ag}^+ + \text{e}^- \longrightarrow \text{Ag}$ : moles of Ag =  $\frac{Q}{F}$ .

$\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}$ : moles of Cu =  $\frac{Q}{2F}$ .

Ratio =  $\frac{\text{mol Cu}}{\text{mol Ag}} = \frac{Q/2F}{Q/F} = \frac{1}{2}$ .

#### Approach

Same charge, but Cu needs 2 electrons while Ag needs only 1. So Cu deposits at half the molar rate of Ag. Like two vending machines — one charges 1 per item, other charges 2 per item. Same money = half the items from the expensive one.

#### Answer

(1)  $\frac{1}{2}$

### Solution 37

#### Explanation

$\text{Fe}^{2+} + 2\text{e}^- \longrightarrow \text{Fe}$  (from  $\text{FeCl}_2$ ): moles of Fe =  $\frac{Q}{2F}$ .  
 $\text{Fe}^{3+} + 3\text{e}^- \longrightarrow \text{Fe}$  (from  $\text{FeCl}_3$ ): moles of Fe =  $\frac{Q}{3F}$ .  
Mass ratio =  $\frac{m_1}{m_2} = \frac{Q \times 56 / (2F)}{Q \times 56 / (3F)} = \frac{3}{2}$ .

#### Approach

Same element (Fe) in two different oxidation states!  $\text{Fe}^{2+}$  needs fewer electrons per atom (2 vs 3), so more Fe deposits from the ferrous cell. Ratio =  $\frac{n_2}{n_1} = \frac{3}{2}$ .

#### Answer

(4) 3 : 2

### Solution 38

#### Explanation

For same charge  $Q$ :  
Moles of Ag =  $\frac{Q}{1 \times F}$ , Cu =  $\frac{Q}{2F}$ , Fe (from  $\text{Fe}^{3+}$ ) =  $\frac{Q}{3F}$ .  
Ratio =  $1 : \frac{1}{2} : \frac{1}{3}$ .

#### Approach

The mole ratio is simply the inverse of  $n$ -factors.  $\text{Ag}^+$  ( $n = 1$ ),  $\text{Cu}^{2+}$  ( $n = 2$ ),  $\text{Fe}^{3+}$  ( $n = 3$ ). Moles  $\propto \frac{1}{n}$ . So ratio =  $1 : \frac{1}{2} : \frac{1}{3}$ .

#### Answer

(2)  $1 : \frac{1}{2} : \frac{1}{3}$

### Solution 39

#### Explanation

4 F passed through each cell.  
 $\text{Ag}^+$  ( $n = 1$ ): mol =  $4/1 = 4$ .  
 $\text{Cd}^{2+}$  ( $n = 2$ ): mol =  $4/2 = 2$ .  
 $\text{Al}^{3+}$  ( $n = 3$ ): mol =  $4/3$ .  
 $\text{Pb}^{4+}$  ( $n = 4$ ): mol =  $4/4 = 1$ .  
Ratio =  $4 : 2 : \frac{4}{3} : 1 = 12 : 6 : 4 : 3$ .

#### Approach

Each metal grabs electrons at its own rate. Silver is cheapest ( $1\text{e}^-$ ), lead is most expensive ( $4\text{e}^-$ ). With 4 F available, divide by each  $n$ -factor. Multiply through by 3 to clear fractions: 12 : 6 : 4 : 3.

#### Answer

(3) 12 : 6 : 4 : 3

### Solution 40

#### Explanation

3 F passed through each cell.  
 $\text{Ag}^+$  ( $n = 1$ ): mol =  $3/1 = 3$ .  
 $\text{Ca}^{2+}$  ( $n = 2$ ): mol =  $3/2 = 1.5$ .  
 $\text{Al}^{3+}$  ( $n = 3$ ): mol =  $3/3 = 1$ .  
Ratio =  $3 : 1.5 : 1 = 6 : 3 : 2$ .

#### Approach

Same pattern: divide total Faradays by  $n$ -factor for each.  $3 : 1.5 : 1$ , multiply by 2 to get whole numbers:  $6 : 3 : 2$ .

#### Answer

(3)  $6 : 3 : 2$

### Solution 41

#### Explanation

Series cells: same charge through all.  
From cell P: mol Ag =  $\frac{0.216}{108} = 0.002$  mol. Charge =  $0.002 \times 1 \times F = 0.002$  F.  
Cell Q ( $\text{Hg}^{2+}$ ,  $n = 2$ ): mol Hg =  $\frac{0.002}{2} = 0.001$  mol. Mass =  $0.001 \times 200.59 = 0.2006$  g.  
Cell R ( $\text{Hg}_2^{2+}$  i.e.  $\text{Hg}^+$ ,  $n = 1$ ): mol Hg =  $\frac{0.002}{1} = 0.002$  mol. Mass =  $0.002 \times 200.59 = 0.4012$  g.

#### Approach

Key distinction: mercuric ( $\text{Hg}^{2+}$ ) needs 2 electrons, mercurous ( $\text{Hg}^+$  in  $\text{Hg}_2^{2+}$ ) needs 1 electron per Hg atom. Same charge but different  $n$ -factors = different amounts deposited. Mercurous gives double the mercury!

#### Answer

(3)  $0.2006$  g and  $0.4012$  g

### Solution 42

#### Explanation

Electrolysis of dilute  $\text{H}_2\text{SO}_4$ :  
Cathode:  $2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$ . Anode:  $2\text{H}_2\text{O} \longrightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$ .  
For 4 F of charge:  $\text{H}_2$  produced = 2 mol = 4 g.  $\text{O}_2$  produced = 1 mol = 32 g.  
Mass ratio (cathode : anode) =  $4 : 32 = 1 : 8$ .

#### Approach

$\text{H}_2$  is super light ( $M = 2$ ),  $\text{O}_2$  is heavy ( $M = 32$ ). Even though 2 mol  $\text{H}_2$  forms vs 1 mol  $\text{O}_2$ , the mass ratio still favours oxygen.  $\frac{2 \times 2}{1 \times 32} = \frac{4}{32} = 1 : 8$ .

#### Answer

(1)  $1 : 8$

## TYPE-5 : Mass / Volume Deposited or Liberated

### Solution 43

#### Explanation

112 mL H<sub>2</sub> at STP = 0.112 L =  $\frac{0.112}{22.4} = 0.005$  mol.  
 $2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$ : electrons =  $2 \times 0.005 = 0.01$  mol.  
 $Q = 0.01 \times 96500 = 965$  C.  
 $I = \frac{Q}{t} = \frac{965}{965} = 1.0$  A.

#### Approach

Convert volume to moles (divide by 22400 mL), find electrons needed (2 per H<sub>2</sub>), convert to charge, then divide by time to get current. Work backwards through Faraday's law!

#### Answer

(1) 1.0

### Solution 44

#### Explanation

$2\text{H}_2\text{O} \longrightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$ . For 1 C of charge:  
Moles of  $\text{e}^- = \frac{1}{96500}$ . Moles of  $\text{O}_2 = \frac{1}{4 \times 96500}$ .  
Volume at STP =  $\frac{1}{4 \times 96500} \times 22.4 \text{ L} = \frac{22.4}{4 \times 96500} = \frac{5.6}{96500} \text{ L}$ .

#### Approach

1 coulomb is a tiny amount of charge. Find what fraction of a Faraday it is, calculate moles of O<sub>2</sub> (need 4e<sup>-</sup> per molecule), then convert to volume at STP using 22.4 L/mol.

#### Answer

(2)  $\frac{5.6}{96500} \text{ L}$

### Solution 45

#### Explanation

$0.5 \text{ g H}_2 = \frac{0.5}{2} = 0.25$  mol. Electrons =  $0.25 \times 2 = 0.5$  mol = 0.5 F.  
Same charge deposits Cu:  $\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}$ .  
Moles of Cu =  $\frac{0.5}{2} = 0.25$  mol. Mass =  $0.25 \times 63.5 = 15.875 \text{ g} \approx 15.9 \text{ g}$ .

#### Approach

First find the charge from the hydrogen cell, then use it in the copper cell. Both H<sub>2</sub> and Cu need 2 electrons, so moles of Cu = moles of H<sub>2</sub> = 0.25 mol. But masses differ because atomic weights differ!

#### Answer

(2) 15.9 g

### Solution 46

#### Explanation

5600 mL O<sub>2</sub> at STP = 5.6 L = 0.25 mol.

O<sub>2</sub> production:  $2\text{H}_2\text{O} \longrightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$ . Electrons =  $0.25 \times 4 = 1 \text{ mol} = 1 \text{ F}$ .

$\text{Ag}^+ + \text{e}^- \longrightarrow \text{Ag}$ ; 1 F deposits 1 mol Ag = 108 g.

#### Approach

Convert O<sub>2</sub> volume to moles, find electrons needed (4 per O<sub>2</sub>), then use same electrons for Ag (1 per Ag).  $0.25 \times 4 = 1 \text{ F} \rightarrow 108 \text{ g Ag}$ .

#### Answer

(4) 108.0 g

### Solution 47

#### Explanation

1.6 g O<sub>2</sub> at anode =  $\frac{1.6}{32} = 0.05 \text{ mol}$ .

Electrons from anode:  $0.05 \times 4 = 0.2 \text{ mol}$ .

At cathode:  $\text{Ag}^+ + \text{e}^- \longrightarrow \text{Ag}$ . Moles of Ag = 0.2 mol.

Mass =  $0.2 \times 107.88 = 21.576 \text{ g} \approx 21.6 \text{ g}$ .

#### Approach

O<sub>2</sub> at anode tells you the total charge (4 electrons per O<sub>2</sub>). Same charge deposits Ag at cathode (1 electron per Ag). So moles of Ag = 4 × moles of O<sub>2</sub>.

#### Answer

(4) 21.6 g

### Solution 48

#### Explanation

1.6 g O<sub>2</sub> = 0.05 mol. Electrons =  $0.05 \times 4 = 0.2 \text{ mol}$ .

$\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}$ : moles of Cu =  $\frac{0.2}{2} = 0.1 \text{ mol}$ .

Mass =  $0.1 \times 63.5 = 6.35 \text{ g}$ .

#### Approach

Same idea as Q47 but with Cu instead of Ag. Cu needs 2 electrons per atom (vs 1 for Ag), so half the moles deposit.  $0.2/2 = 0.1 \text{ mol Cu} = 6.35 \text{ g}$ .

#### Answer

(1) 6.35 g

### Solution 49

#### Explanation

2.24 L H<sub>2</sub> at STP = 0.1 mol. Electrons = 0.1 × 2 = 0.2 mol.

Cu<sup>2+</sup> + 2e<sup>-</sup> → Cu: moles of Cu =  $\frac{0.2}{2}$  = 0.1 mol.

Mass = 0.1 × 63.5 = 6.35 g.

#### Approach

Find electrons from H<sub>2</sub> liberation (2 per molecule), then use for Cu deposition (also 2 per atom). Since both need 2 electrons, moles of Cu = moles of H<sub>2</sub> = 0.1 mol.

#### Answer

(3) 6.35 g

### Solution 50

#### Explanation

0.108 g Ag deposited. Moles of Ag =  $\frac{0.108}{108}$  = 0.001 mol. Electrons = 0.001 mol.

At water voltameter anode: 2H<sub>2</sub>O → O<sub>2</sub> + 4H<sup>+</sup> + 4e<sup>-</sup>.

Moles of O<sub>2</sub> =  $\frac{0.001}{4}$  = 0.00025 mol.

Volume at STP = 0.00025 × 22400 = 5.6 cm<sup>3</sup>.

#### Approach

Ag tells you the charge (1:1 with electrons). Same charge goes to water voltameter. O<sub>2</sub> needs 4 electrons per molecule, so divide by 4, then convert to volume. 0.001/4 = 0.00025 mol × 22400 = 5.6 cm<sup>3</sup>.

#### Answer

(3) 5.6 cm<sup>3</sup>

### Solution 51

#### Explanation

4.5 g Al =  $\frac{4.5}{27}$  =  $\frac{1}{6}$  mol. Electrons =  $\frac{1}{6}$  × 3 = 0.5 mol.

2H<sup>+</sup> + 2e<sup>-</sup> → H<sub>2</sub>: moles of H<sub>2</sub> =  $\frac{0.5}{2}$  = 0.25 mol.

Volume at STP = 0.25 × 22.4 = 5.6 L.

#### Approach

Al gives you the electron count (×3), then use those electrons to make H<sub>2</sub> (÷2).  $\frac{1}{6}$  × 3 ÷ 2 = 0.25 mol H<sub>2</sub> = 5.6 L.

#### Answer

(4) 5.6 L

### Solution 52

#### Explanation

$2\text{H}_2\text{O} \longrightarrow 2\text{H}_2 + \text{O}_2$ . Mole ratio:  $\text{H}_2 : \text{O}_2 = 2 : 1$ .  
At same conditions, volume ratio = mole ratio =  $2 : 1$ .  
If  $0.224 \text{ L H}_2$ , then  $\text{O}_2 = \frac{0.224}{2} = 0.112 \text{ L}$ .

#### Approach

From Avogadro's law: equal moles = equal volumes at same conditions. Water electrolysis always gives  $\text{H}_2 : \text{O}_2$  in  $2:1$  volume ratio. Half the hydrogen volume = oxygen volume.

#### Answer

**(3) 0.112 L**

### Solution 53

#### Explanation

$1 \text{ L of } 0.01 \text{ M H}_2\text{SO}_4 = 0.01 \text{ mol H}_2\text{SO}_4$ .  
 $\text{H}^+ = 0.02 \text{ mol}$ . Complete electrolysis:  
 $\text{H}_2 = \frac{0.02}{2} = 0.01 \text{ mol}$ .  $\text{O}_2 = \frac{0.01}{2} = 0.005 \text{ mol}$ .  
Total gas =  $0.015 \text{ mol}$ . Volume at STP =  $0.015 \times 22400 = 336 \text{ mL}$ .

#### Approach

All the water associated with  $0.01 \text{ mol H}_2\text{SO}_4$  electrolyses. Total  $\text{H}^+$  available =  $0.02 \text{ mol}$ , giving  $0.01 \text{ mol H}_2$ . The corresponding  $\text{O}_2$  is half that in moles. Add volumes together.

#### Answer

**(2) 336 ml**

### Solution 54

#### Explanation

$4 \text{ mol electrons transferred from anode to cathode}$ .  
At cathode:  $2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$ . Moles of  $\text{H}_2 = \frac{4}{2} = 2 \text{ mol}$ .  
At anode:  $2\text{H}_2\text{O} \longrightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$ . Moles of  $\text{O}_2 = \frac{4}{4} = 1 \text{ mol}$ .  
Total gas =  $2 + 1 = 3 \text{ mol}$ . Volume at STP =  $3 \times 22.4 = 67.2 \text{ L}$ .

#### Approach

Same electrons work at both electrodes simultaneously! Cathode uses them to make  $\text{H}_2$  (2 per molecule), anode releases them while making  $\text{O}_2$  (4 per molecule). Total gas =  $\text{H}_2 + \text{O}_2 = 2+1 = 3 \text{ mol}$ .

#### Answer

**(1) 67.2 L**

## TYPE-6 : Electrochemical Equivalent & Reactions

### Solution 55

#### Explanation

$$\text{Electrochemical equivalent (ECE)} = \frac{M}{nF} = \frac{63.5}{2 \times 96500}.$$

#### Approach

ECE tells you “grams deposited per coulomb.” It’s the atomic mass divided by ( $n$ -factor  $\times$  Faraday constant). Think of it as the “price per coulomb” — how many grams you get for each coulomb you spend.

#### Answer

$$(2) \frac{63.5}{96500 \times 2}$$

### Solution 56

#### Explanation

$$m = Z \times I \times t = 0.001118 \times 0.5 \times 200 = 0.1118 \text{ g.}$$

#### Approach

ECE ( $Z$ ) is the direct multiplier:  $\text{mass} = Z \times I \times t$ . No need for  $n$ -factor or Faraday constant — ECE already has them baked in. Just multiply three numbers!

#### Answer

$$(2) 0.11180 \text{ g}$$

### Solution 57

#### Explanation

$$Q = I \times t = 9.65 \times 600 = 5790 \text{ C.}$$
$$m = \frac{E \times Q}{F}, \text{ where } E = \text{equivalent weight.}$$
$$3.0 = \frac{E \times 5790}{96500}$$
$$E = \frac{3.0 \times 96500}{5790} = 50.$$

#### Approach

Rearrange Faraday’s law to find equivalent weight. It’s like reverse-engineering the vending machine: you know what came out (3 g) and how many coins you put in (5790 C), now find the “price tag” (equivalent weight).

#### Answer

$$(3) 50$$

### Solution 58

#### Explanation

$$m = \frac{M \times Q}{n \times F}$$
$$2.977 = \frac{106.4 \times 10800}{n \times 96500}$$
$$n = \frac{106.4 \times 10800}{2.977 \times 96500} = \frac{1149120}{287278.5} = 4.0$$

Charge on cation = +4.

#### Approach

Rearrange for  $n$  (charge). You know mass deposited, atomic mass, and charge passed. Solve for  $n$  to find the oxidation state. It's like figuring out how many tickets each ride costs when you know total tickets spent and rides taken.

#### Answer

(1) +4

### Solution 59

#### Explanation

$$Q = 2.0 \times 5 \times 3600 = 36000 \text{ C.}$$
$$m = \frac{M \times Q}{n \times F}$$
$$22.2 = \frac{177 \times 36000}{n \times 96500}$$
$$n = \frac{177 \times 36000}{22.2 \times 96500} = \frac{6372000}{2142300} = 2.97 \approx 3.$$

#### Approach

Same reverse-engineering as Q58. Find  $n$  by plugging in known values.  $n \approx 3$  means the metal exists as  $M^{3+}$  in the salt.

#### Answer

(3) +3

## TYPE-7 : Advanced & Miscellaneous Problems

### Solution 60

#### Explanation

$$Q = 0.965 \times 10 \times 60 = 579 \text{ C.}$$
$$\text{Moles of Zn deposited} = \frac{579}{2 \times 96500} = 0.003 \text{ mol.}$$
$$\text{Initial moles of Zn}^{2+} = 0.2 \times 0.5 = 0.1 \text{ mol.}$$
$$\text{Remaining Zn}^{2+} = 0.1 - 0.003 = 0.097 \text{ mol.}$$
$$\text{Molarity} = \frac{0.097}{0.5} = 0.194 \text{ M.}$$

### Approach

First find how much  $\text{Zn}^{2+}$  is removed from solution (deposited as metal). Subtract from initial moles, then recalculate molarity. The volume stays the same (500 mL). It's like removing items from a jar — the jar size doesn't change, just the count inside.

### Answer

(4) 0.194 M

## Solution 61

### Explanation

Initial: 1 L of 1 M  $\text{CuSO}_4 = 1 \text{ mol Cu}^{2+}$ .  
 $\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}$ : 2 F deposits 1 mol Cu.  
After passing 2 F: all 1 mol of  $\text{Cu}^{2+}$  is deposited.  
Remaining  $\text{CuSO}_4 = 0 \text{ mol}$ . Molarity = 0.

### Approach

2 F = 2 mol electrons. Cu needs 2 electrons per atom. So 2 F deposits exactly 1 mol Cu — which is all the  $\text{Cu}^{2+}$  available. The solution is completely depleted!

### Answer

(4) zero

## Solution 62

### Explanation

$\text{C}_6\text{H}_5\text{NO}_2 + 6\text{H}^+ + 6\text{e}^- \longrightarrow \text{C}_6\text{H}_5\text{NH}_2 + 2\text{H}_2\text{O}$ .  
N goes from +3 to -3: 6 electrons per molecule.  
12.3 g nitrobenzene =  $\frac{12.3}{123} = 0.1 \text{ mol}$ .  
Theoretical charge =  $0.1 \times 6 \times 96500 = 57900 \text{ C}$ .  
At 50% efficiency: actual charge =  $\frac{57900}{0.5} = 115800 \text{ C}$ .

### Approach

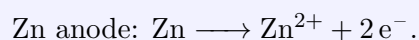
Nitrobenzene to aniline: the  $\text{NO}_2$  group converts to  $\text{NH}_2$ . That's N going from +3 to -3 = 6 electrons needed. But 50% efficiency means you need to pass *double* the theoretical charge — like a leaky bucket that loses half the water.

### Answer

(1) 115800 C

### Solution 63

#### Explanation



100 g Zn =  $\frac{100}{65} = 1.538$  mol. Electrons = 3.077 mol.

But also limited by  $\text{Cu}^{2+}$ : 1.0 mol  $\text{Cu}^{2+}$  needs 2 mol  $\text{e}^{-}$ .

Limiting reagent:  $\text{Cu}^{2+}$  (needs only 2 mol  $\text{e}^{-}$  vs 3.077 available from Zn).

$$Q = 2 \times 96500 = 193000 \text{ C.}$$

$$t = \frac{193000}{1.0} = 193000 \text{ s} = 53.6 \text{ hr.}$$

#### Approach

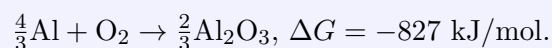
This is a galvanic cell, so it runs until one reactant is exhausted. Check both: Zn can provide 3.077 mol  $\text{e}^{-}$ , but  $\text{Cu}^{2+}$  can only accept 2 mol  $\text{e}^{-}$ .  $\text{Cu}^{2+}$  runs out first (limiting reagent). Time =  $\frac{\text{total charge}}{\text{current}}$ .

#### Answer

(3) 53.6 hr.

### Solution 64

#### Explanation



For the reverse (electrolysis):  $\Delta G = +827 \text{ kJ/mol.}$

$n$ -factor: each Al loses  $3\text{e}^{-}$ , and  $\frac{4}{3}$  Al atoms participate  $\Rightarrow n = \frac{4}{3} \times 3 = 4 \text{ mol e}^{-}$ .

$$E = \frac{-\Delta G}{nF} = \frac{827 \times 10^3}{4 \times 96500} = 2.14 \text{ V.}$$

#### Approach

$\Delta G = -nFE$ . For the given reaction as written, count total electrons transferred ( $n = 4$ ). The minimum EMF for electrolysis equals the cell potential of the forward (spontaneous) reaction. Plug into the formula:  $E = \frac{|\Delta G|}{nF}$ .

#### Answer

(1) 2.14 V