

# **EXERCISE (0-I)**

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## **GALVANIC CELL**



In galvanic cell:

~ gallium anode,  
Cathode = (+)ve electrode, At cathode reduction occurs.

Anode = (+)ve electrode, At anode oxidation occurs.

2. A standard hydrogen electrode has zero electrode potential because
- (A) hydrogen is easier to oxidise                          ✓ (B) electrode potential is assumed to be zero  
(C) hydrogen atom has only one electron              (D) hydrogen is the lightest element.

$$\left. \begin{array}{l} E^\circ_{H^+/H_2} = 0 \text{ volt} \\ E^\circ_{H_2/H^+} = 0 \text{ volt} \end{array} \right\} \text{Arbitrarily assumed.}$$

3.

A standard reduction electrode potentials of four metals are

$$A = -0.250 \text{ V}, \quad B = -0.140 \text{ V} \quad C = -0.126 \text{ V}, \quad D = -0.402 \text{ V}$$

The metal that displaces A from its aqueous solution is :-

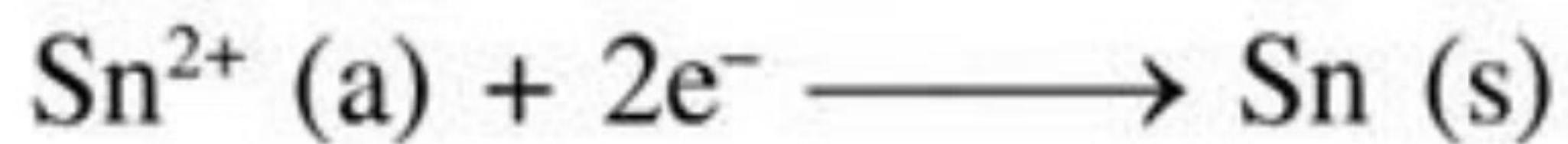
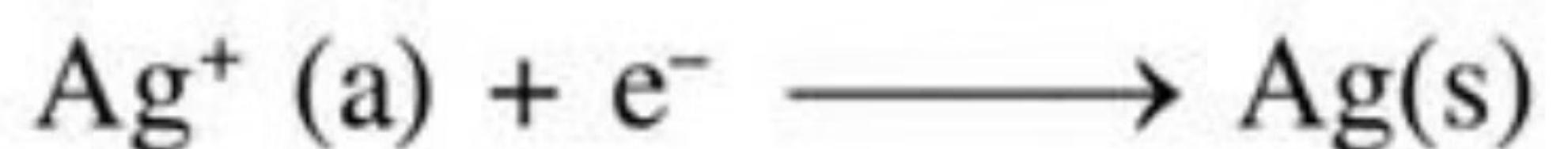
- (A) B                      (B) C                      ~~(C) D~~                      (D) None of the above

To displace 'A' from solution,  $E^{\circ}_{\text{Reduction}}$  of metal

must be less than  $-0.25 \text{ V}$ .

4.

The standard electrode potentials for the reactions



at 25 °C are 0.80 volt and -0.14 volt, respectively. The standard emf of the cell.

$\text{Sn}_{(\text{s})} | \text{Sn}^{2+}_{(\text{aq})} (1\text{M}) || \text{Ag}^+_{(\text{aq})} (1\text{M}) | \text{Ag}_{(\text{s})}$  is :

- (A) 0.66 volt      (B) 0.80 volt      (C) 1.08 volt       (D) 0.94 volt

$$\begin{aligned}
 E^\circ_{\text{cell}} &= (\text{SRP})_C - (\text{SRP})_A \\
 &= E^\circ_{\text{Ag}^+/\text{Ag}} - E^\circ_{\text{Sn}^{2+}/\text{Sn}} \\
 &= 0.8 - (-0.14) = 0.94\text{V}.
 \end{aligned}$$

5.

The thermodynamic efficiency of cell is given by-

(A)  $\frac{\Delta H}{\Delta G}$

(B)  $\frac{nFE_{cell}}{\Delta G}$

(C)  $-\frac{nFE_{cell}}{\Delta H}$

(D) Zero

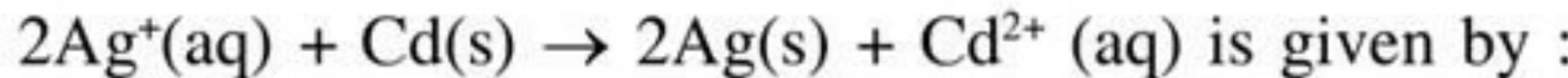
$$\text{efficiency} = \frac{\Delta G_f}{\Delta H} = -\frac{nFE_{cell}}{\Delta H}$$

6.

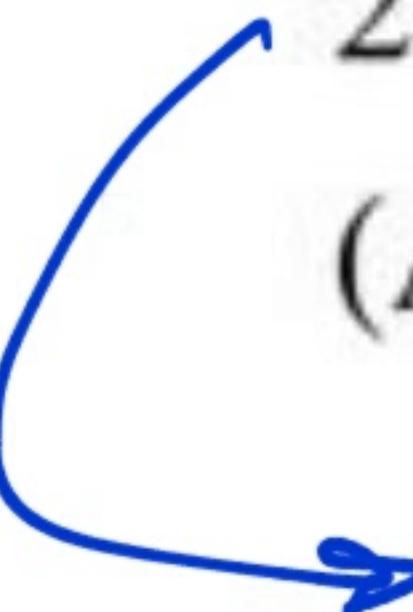
The standard reduction potentials for two half-cell reactions are given below,



The standard free energy change for the reaction



- (A) 115.8 KJ      (B) -115.8 KJ      ~~(C)~~ -231.6 KJ      (D) 231.6KJ

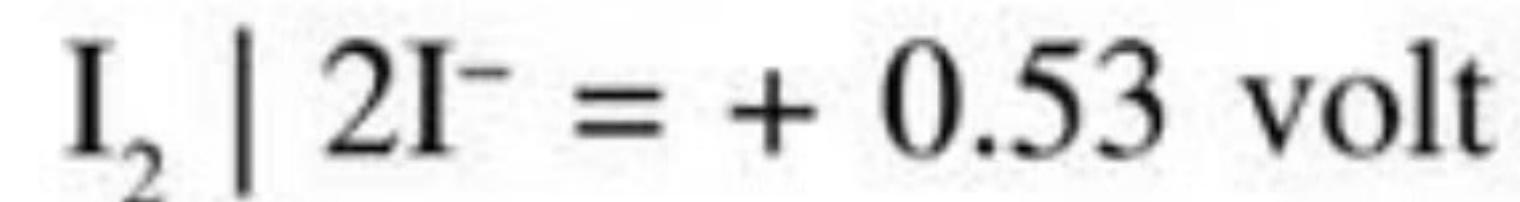
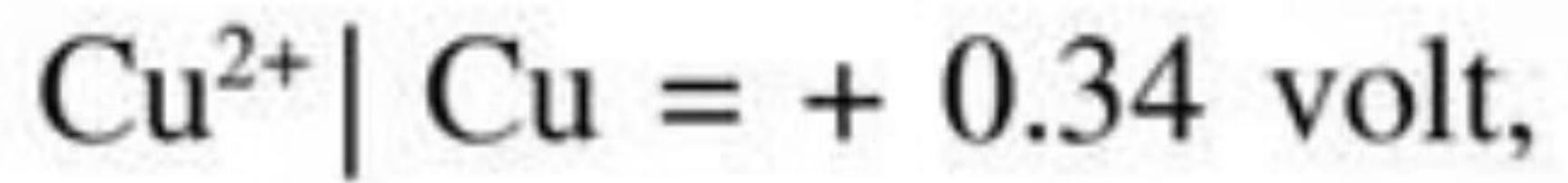
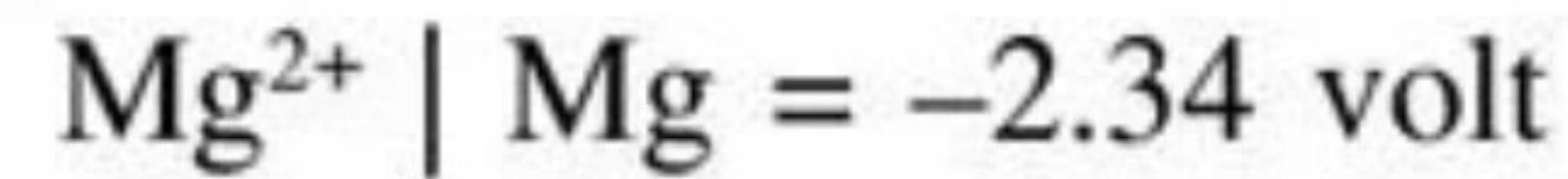
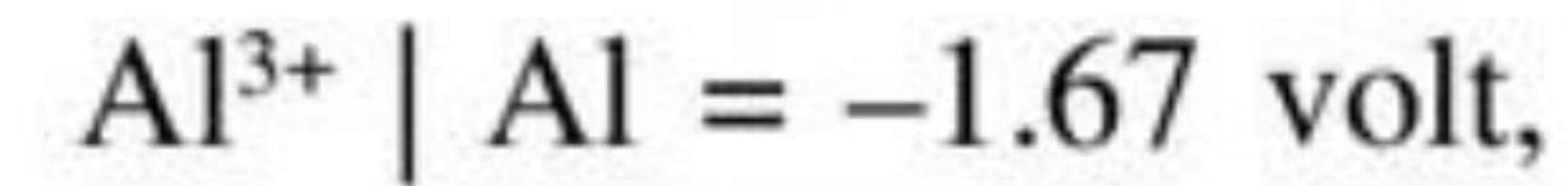


$$E_{\text{cell}}^\circ = E_{\text{Ag}^+/\text{Ag}}^\circ - E_{\text{Cd}^{2+}/\text{Cd}}^\circ = 0.8 - (-0.4) = 1.2 \text{ V}.$$

$$\Delta G^\circ = -nFE_{\text{cell}}^\circ = -2 \times 96500 \times 1.2 \text{ J} = -231.6 \text{ KJ}.$$

7.

The reduction potential values are given below:



Which one is the best reducing agent ?

(A) Al

~~(B) Mg~~

(C) Cu

(D) I<sub>2</sub>

*Smaller the value of reduction potential data  
better is the reducing agent.*

8.  $E^\circ(\text{Ni}^{2+}|\text{Ni}) = -0.25$  volt,  $E^\circ(\text{Au}^{3+}|\text{Au}) = 1.50$  volt. The standard emf of the voltaic cell.  $\text{Ni}_{(s)} \mid \text{Ni}^{2+}_{(\text{aq})}(1.0 \text{ M}) \parallel \text{Au}^{3+}_{(\text{aq})}(1.0 \text{ M}) \mid \text{Au}_{(s)}$  is :  
(A) 1.25 volt      (B) -1.75 volt      ~~(C) 1.75 volt~~      (D) 4.0 volt

$$\begin{aligned}E_{\text{cell}} &= E^\circ_{\text{Au}^{3+}/\text{Au}} - E^\circ_{\text{Ni}^{2+}/\text{Ni}} \\&= 1.5 - (-0.25) = \boxed{1.75 \text{ volt}}.\end{aligned}$$

Ans.

9.  $E^\circ$  for  $F_2 + 2e^- = 2F^-$  is 2.8 V,  $E^\circ$  for  $\frac{1}{2}F_2 + e^- = F^-$  is ?

- (A) 2.8 V      (B) 1.4 V      (C) -2.8 V      (D) -1.4 V

$E^\circ$  is independent of stoichiometric coefficients

10. From the following  $E^\circ$  values of half cells,
- (i)  $A + e \rightarrow A^-; E^\circ = -0.24 \text{ V}$       (ii)  $B^- + e \rightarrow B^{2-}; E^\circ = +1.25 \text{ V}$
- (iii)  $C^- + 2e \rightarrow C^{3-}; E^\circ = -1.25 \text{ V}$       (iv)  $D + 2e \rightarrow D^{2-}; E^\circ = +0.68 \text{ V}$

What combination of two half cells would result in a cell with the largest potential ?

- (A) (ii) and (iii)      (B) (ii) and (iv)      (C) (i) and (iii)      (D) (i) and (iv)

Combination of highest and lowest reduction potential.

11. If  $\Delta G^\circ$  of the cell reaction,

$\text{AgCl(s)} + \frac{1}{2}\text{H}_2\text{(g)} \rightarrow \text{Ag(s)} + \text{H}^+ + \text{Cl}^-$  is  $-21.52 \text{ KJ}$  then  $\Delta G^\circ$  of  $2\text{AgCl(s)} + \text{H}_2\text{(g)} \rightarrow 2\text{Ag(s)} + 2\text{H}^+ + 2\text{Cl}^-$  is :

- (A)  $-21.52 \text{ KJ}$       (B)  $-10.76 \text{ KJ}$       ~~(C)~~  $-43.04 \text{ KJ}$       (D)  $43.04 \text{ KJ}$

$$\Delta G^\circ = 2 \times (-21.52)$$

$$= -43.04 \text{ KJ.}$$

12. The reduction potential of hydrogen electrode ( $P_{H_2} = 1 \text{ atm}$ ;  $[H^+] = 0.1 \text{ M}$ ) at  $25^\circ\text{C}$  will be -
- (A) 0.00 V      (B)  $-0.059 \text{ V}$       (C)  $0.118 \text{ V}$       (D)  $0.059 \text{ V}$



$$E_{H^+/H_2} = E^{\circ}_{H^+/H_2} - \frac{0.059}{1} \log \frac{P_{H_2(g)}}{[H^+]}$$

or,  $E_{H^+/H_2} = 0 - \frac{0.059}{1} \log \frac{1}{0.1}$

$$= \boxed{-0.059 \text{ volt}}$$

Ans.

13. Which of the following represents the reduction potential of silver wire dipped into 0.1 M AgNO<sub>3</sub> solution at 25° C ?

- (A) E<sub>red</sub><sup>o</sup>      (B) (E<sub>red</sub><sup>o</sup> + 0.059)      (C) (E<sub>oxi</sub><sup>o</sup> - 0.059)      (D) (E<sub>red</sub><sup>o</sup> - 0.059)



$$E_{Ag^+/Ag} = E_{Ag^+/Ag}^o - \frac{0.059}{1} \log \frac{1}{[Ag^+]}$$

or,  $E_{Ag^+/Ag} = E_{Ag^+/Ag}^o - \frac{0.059}{1} \log \frac{1}{0.1}$

or,  $E_{Ag^+/Ag} = \boxed{E_{Ag^+/Ag}^o - 0.059}$  Ans.

14. Which of the following will increase the voltage of the cell with following cell reaction



- (A) Decrease in the concentration of  $\text{Ag}^+$  ions  
(B) Increase in the concentration of  $\text{Sn}^{+2}$  ions  
 (C) Increase in the concentration of  $\text{Ag}^+$  ions  
(D) (A) & (B) both

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{2} \log \frac{[\text{Sn}^{+2}]}{[\text{Ag}^{+}]^2}$$

$\Rightarrow$  (C) is correct.

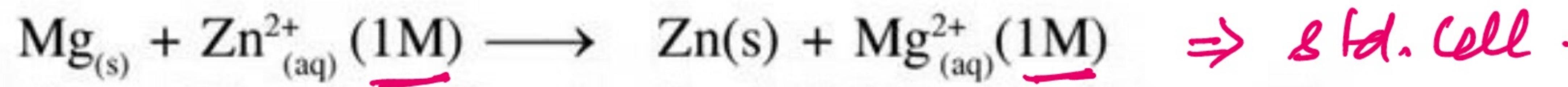
15. For a reaction -  $A(s) + 2B_{(aq)}^+ \rightarrow A_{(aq)}^{2+} + 2B_{(s)}$ ;  $K_C$  has been found to be  $10^{12}$ . The  $E^\circ$ cell is:
- (A) 0.354 V      (B) 0.708 V      (C) 0.0098 V      (D) 1.36 V

$$K_C = 10^{12} = 10 \frac{2 \times E^\circ_{cell}}{0.059}$$

16. At 25°C the standard emf of cell having reactions involving two electrons change is found to be 0.295V. The equilibrium constant of the reaction is -
- (A)  $29.5 \times 10^{-2}$       (B) 10      ✓(C)  $10^{10}$       (D)  $29.5 \times 10^{10}$

$$K = 10^{\frac{n E_{cell}^\circ}{0.059}} = 10^{\frac{2 \times 0.295}{0.059}} = \boxed{10} \quad \text{Ans.}$$

17. For the cell reaction



The emf has been found to be 1.60 V,  $E^\circ$  of the cell is :

- (A) -1.60 V      (B) 1.60 V      (C) 0.0 V      (D) 0.16 V

∴  $E_{\text{cell}} = E^\circ = \boxed{1.6 \text{ volt}}$  & ( $\because$  cell is at std. state  
 $\therefore E_{\text{cell}} = E^\circ_{\text{cell}}$ ) .

18. The emf of the cell in which the following reaction,  
 $\text{Zn(s)} + \text{Ni}^{2+}_{(\text{aq})} (\text{a} = 0.1) \rightarrow \text{Zn}^{2+}_{(\text{aq})} (\text{a} = 1.0) + \text{Ni(s)}$   
occurs, is found to be 0.5105 V at 298 K. The standard e.m.f. of the cell is :-  
(A) -0.5105 V      (B) 0.5400 V      (C) 0.4810 V      (D) 0.5696 V

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.059}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{Ni}^{2+}]}$$

$$0.5105 = E^{\circ}_{\text{cell}} - \frac{0.059}{2} \log \frac{1}{0.1}$$

$$\text{or, } 0.5105 = E^{\circ}_{\text{cell}} - \frac{0.059}{2}$$

$$\text{or, } E^{\circ}_{\text{cell}} = 0.5105 + \frac{0.059}{2} = \boxed{0.5400 \text{ Volt}}$$

19. The standard emf for the cell reaction,

$\text{Zn}_{(\text{s})} + \text{Cu}^{2+}_{(\text{aq})} \longrightarrow \text{Zn}^{2+}_{(\text{aq})} + \text{Cu}_{(\text{s})}$  is 1.10 volt at 25 °C. The emf for the cell reaction when 0.1 M Cu<sup>2+</sup> and 0.1 M Zn<sup>2+</sup> solution are used at 25°C is :

- (A) 1.10 volt      (B) 0.110 volt      (C) -1.10 volt      (D) -0.110 volt

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

or,  $E_{\text{cell}} = 1.1 - \frac{0.059}{2} \log \frac{0.1}{0.1}$

$$= \boxed{1.1 \text{ volt}} \quad \underline{\text{Ans}}$$

20. What is the potential of the cell containing two hydrogen electrodes as represented below



- (A) - 0.295 V      (B) - 0.0591 V      (C) 0.295 V      (D) 0.0591 V

$$E_{\text{Cell}} = 0 - \frac{0.059}{1} \log \frac{[\text{H}^+]_A}{[\text{H}^+]_C} \quad (\text{Assume } P_{\text{H}_2} = 1 \text{ bar})$$

so,  $E_{\text{Cell}} = -\frac{0.059}{1} \log \frac{10^{-8}}{(0.001)}$

$$= 0.059 \times 5 = \boxed{0.295 \text{ V.}}$$

21. Consider the cell,  $\text{Cu}|\text{Cu}^{+2}\|\text{Ag}^+|\text{Ag}$ . If the concentration of  $\text{Cu}^{+2}$  and  $\text{Ag}^+$  ions becomes ten times the emf of the cell :-

- (A) Becomes 10 times  
 ✓ (C) Increase by 0.0295 V

- (B) Remains same  
 (D) Decrease by 0.0295 V

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.059}{2} \log \frac{[\text{Cu}^{+2}]}{[\text{Ag}^+]^2}$$

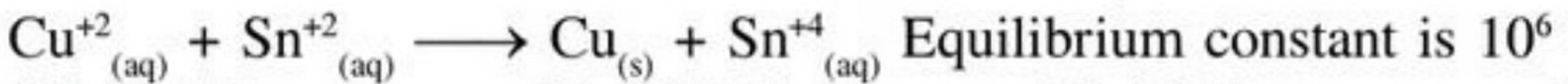
$$E_1 = E_{\text{cell}}^\circ - \frac{0.059}{2} \log \frac{[\text{Cu}^{+2}]}{[\text{Ag}^+]^2}$$

$$E_2 = E_{\text{cell}}^\circ - \frac{0.059}{2} \log \frac{10[\text{Cu}^{+2}]}{100[\text{Ag}^+]^2}$$

$$\therefore E_2 - E_1 = \frac{0.059}{2} \log 10 = 0.0295$$

$$\therefore E_2 = E_1 + 0.0295 \Rightarrow \boxed{\text{increases by } 0.0295 \text{ V}}$$

22. Determine the value of  $E^\circ$  cell for the following reaction -



- Equilibrium constant is  $10^6$
- (A) 0.177      (B) 0.0177      (C) 0.215      (D) 1.77

✓  $K_C = 10^{\frac{nE^\circ_{\text{cell}}}{0.059}}$

or,  $10^6 = 10^{\frac{2xE^\circ_{\text{cell}}}{0.059}}$

or,  $6 = \frac{2xE^\circ_{\text{cell}}}{0.059}$

or,  $E^\circ_{\text{cell}} = 3 \times 0.059 = \boxed{0.177 \text{ volt}}$

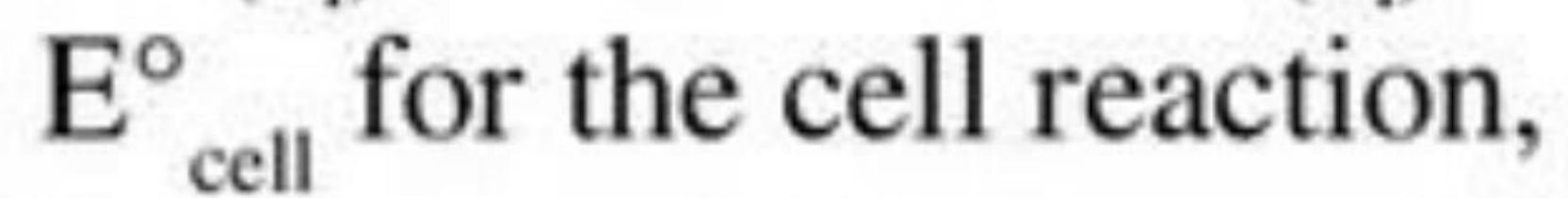
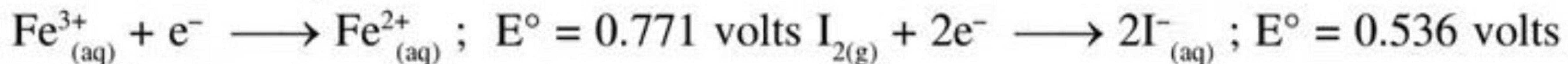
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23. The standard emf of a galvanic cell involving cell reaction with  $n = 4$  is found to be 0.295 V at 25°C. The equilibrium constant of the reaction would be,

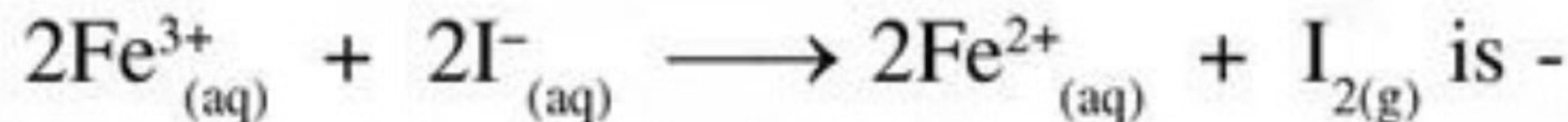
- (A)  $1.0 \times 10^{20}$       (B)  $2.0 \times 10^{11}$       (C)  $4.0 \times 10^{12}$       (D)  $1.0 \times 10^2$

$K = 10^{\frac{nE^\circ_{cell}}{0.059}} = 10^{\frac{4 \times 0.295}{0.059}} = \boxed{10^{20}}$  Ans.

24. Given electrode potentials :



$E^\circ_{\text{cell}}$  for the cell reaction,



(A)  $(2 \times 0.771 - 0.536) = 1.006 \text{ volts}$       (B)  $(0.771 - 0.5 \times 0.536) = 0.503 \text{ volts}$

(C)  $0.771 - 0.536 = 0.235 \text{ volts}$       (D)  $0.536 - 0.771 = -0.235 \text{ volts}$

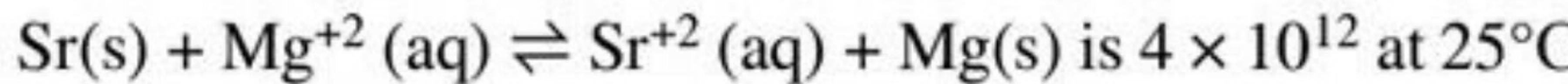
$$E^\circ_{\text{cell}} = E^\circ_{\text{Fe}^{3+}/\text{Fe}^{2+}} - E^\circ_{\text{I}_2/\text{I}^-}$$

$$= 0.771 - 0.536$$

$$= \boxed{0.235 \text{ volt.}} \quad \underline{\text{Ans.}}$$

25.

The equilibrium constant for the reaction



is  $4 \times 10^{12}$  at  $25^\circ\text{C}$

(log 2 = 0.3)

- (A) 0.3717 V      (B) 0.7434 V      (C) 0.1858 V      (D) 0.135 V

$$K = 10^{\frac{nE^\circ}{0.059}}$$

$$\text{or, } \log K = \frac{nE^\circ}{0.059}$$

$$\text{or, } \log (4 \times 10^{12}) = \frac{2 \times E^\circ}{0.059}$$

$$\text{or, } (0.6 + 12) = \frac{2 \times E^\circ}{0.059}$$

$$\text{or, } E^\circ = \frac{12.6 \times 0.059}{2}$$

$$\text{or, } E^\circ = \boxed{0.3717 \text{ volt}}$$

Ans.

26. By how much times will potential of half cell  $\text{Cu}^{2+}|\text{Cu}$  change if, the solution is diluted to 100 times at 298 K :-
- (A) Increases by 59 mV      ✓ (B) Decrease by 59 mV  
 (C) Increases by 29.5 mV      (D) Decreases by 29.5 mV



$$E_{\text{Cu}^{2+}|\text{Cu}} = E_{\text{Cu}^{2+}|\text{Cu}}^{\circ} - \frac{0.059}{2} \log \frac{1}{[\text{Cu}^{2+}]}$$

let  $[\text{Cu}^{2+}]_{\text{initial}} = x \quad \therefore [\text{Cu}^{2+}]_{\text{after dilution}} = \frac{x}{100}$

$$E_1 = E_{\text{Cu}^{2+}|\text{Cu}}^{\circ} - \frac{0.059}{2} \log \frac{1}{x} \quad \text{--- (1)}$$

$$E_2 = E_{\text{Cu}^{2+}|\text{Cu}}^{\circ} - \frac{0.059}{2} \log \left( \frac{1 \times 100}{x} \right) \quad \text{--- (2)}$$

$$E_1 - E_2 = \frac{0.059}{2} \log 100 = 0.059$$

$\therefore E_2 = E_1 - 0.059 \Rightarrow \text{decrease by } 0.059 \text{ volt} = 59 \text{ mV.}$

## ELECTROLYTIC CELL

27. When an electric current is passed through a cell containing an electrolyte, positive ions move towards the cathode and negative ions towards the anode. What will happen if the cathode is pulled out of the solution?
- (A) The positive ions will start moving towards the anode and negative ions will stop moving.
  - (B) The negative ions will continue to move towards the anode and the positive ions will stop moving
  - (C) Both positive and negative ions will move towards the anode.
  - (D) None of these movements will take place.

Circuit will break so no movement of ion  
towards electrodes:-

28. Which of the substances Na, Hg, S, Pt and graphite can be used as electrodes in electrolytic cells having aqueous solution ?

(A) Hg and Pt

(C) Na, S

~~(B) Hg, Pt and graphite~~

(D) Na, Hg, S

memory based

29. The products formed when an aqueous solution of NaBr is electrolyzed in a cell having inert electrodes are :

- (A) Na and Br<sub>2</sub>      (B) Na and O<sub>2</sub>      ~~(C) H<sub>2</sub>, Br<sub>2</sub> and NaOH~~ (D) H<sub>2</sub> and O<sub>2</sub>

At C



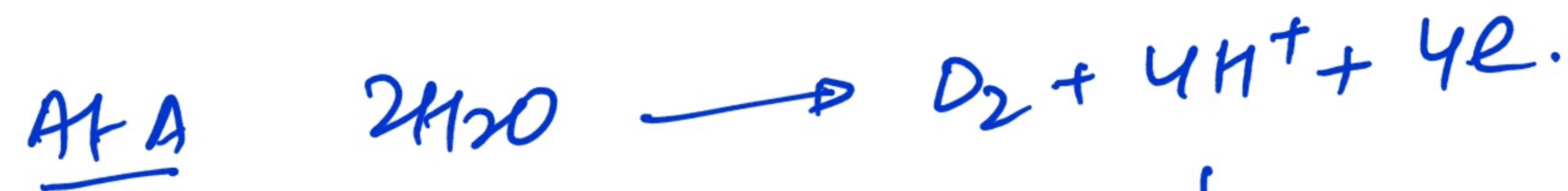
At A



30. Electrolysis of a  $\text{CuSO}_4$  produces :-

- (A) An increase in pH  
(C) Either decrease or increase

- (B) A decrease in pH  
(D) None



↓  
as  $\text{H}^+$  is produced  $\Rightarrow$  pH decreases.

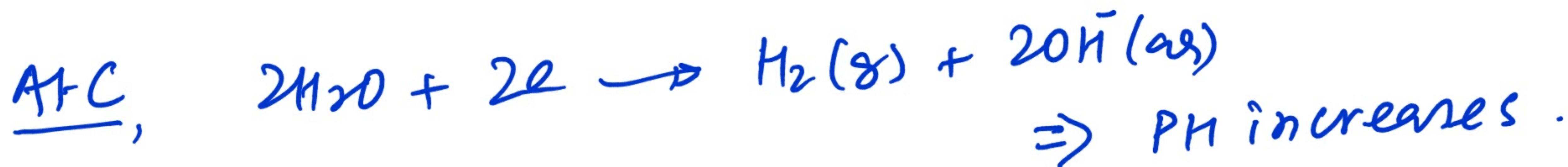
31. A solution of sodium sulphate in water is electrolysed using inert electrodes. The products at the cathode and anode are respectively.

- (A) H<sub>2</sub>, O<sub>2</sub>      (B) O<sub>2</sub>, H<sub>2</sub>      (C) O<sub>2</sub>, Na      (D) none



32. When an aqueous solution of lithium chloride is electrolysed using graphite electrodes
- (A)  $\text{Cl}_2$  is liberated at the anode.  
(B) Li is deposited at the cathode  
(C) as the current flows, pH of the solution remains constant  
(D) as the current flows, pH of the solution decreases.

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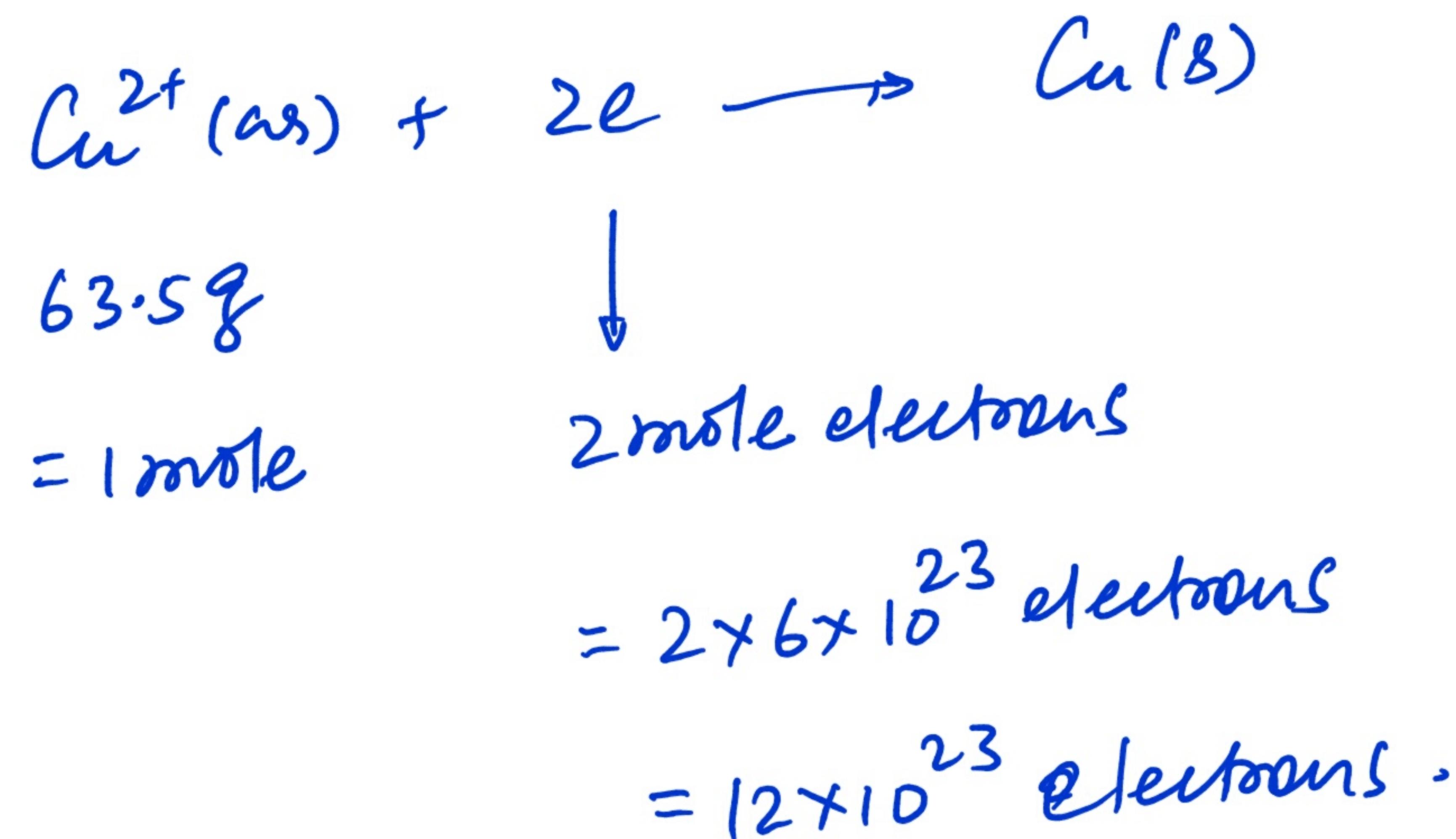


33. The amount of an ion discharged during electrolysis is not directly proportional to :

  - (A) resistance
  - (B) time ( $t$ )
  - (C) current strength ( $i$ )
  - (D) electrochemical equivalent of the element ( $z$ )

$w = z \dot{w}$

34. Number of electrons involved in the electrodeposition of 63.5 g of Cu from a solution of  $\text{CuSO}_4$  is : ( $N_A = 6 \times 10^{23}$ )
- (A)  $6 \times 10^{23}$       (B)  $3 \times 10^{23}$       ✓ (C)  $12 \times 10^{23}$       (D)  $6 \times 10^{22}$



$$w = z^2 t$$

$W = Z \cdot F$

where  $Z$  is the amount of substance deposited or liberated when 1 coulomb of electricity is supplied.

$Z$  is called electrochemical equivalence of the substance.

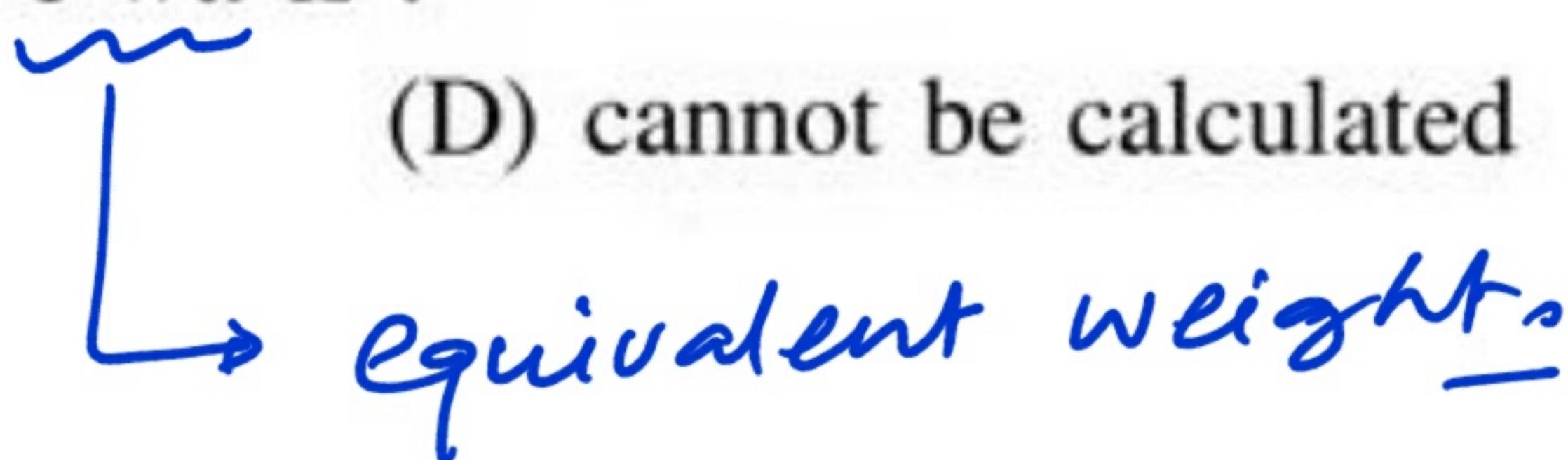
36. Electro chemical equivalent of a substance is 0.0006; its e wt. is :

(A) 57.9

(B) 28.95

(C) 115.8

(D) cannot be calculated

 → equivalent weight

$$E = FZ$$

$$\text{or, } E = 96500 \times 0.0006$$

=

$$57.98$$

37. W g of copper deposited in a copper voltameter when an electric current of 2 ampere is passed for 2 hours. If one ampere of electric current is passed for 4 hours in the same voltameter, copper deposited will be :

~~(A) W~~

(B)  $W/2$

(C)  $W/4$

(D)  $2W$

$$\text{initial charge} = i \times t = 2 \times 2 = 4 \text{ ampere hour.}$$

$$\text{final charge} = i \times t = 1 \times 4 = 4 \text{ ampere hour.}$$

$\because$  charge supplied is same  $\therefore$  amount of Cu deposited will be same.

38.

- When the same electric current is passed through the solution of different electrolytes in series the amounts of elements deposited on the electrodes are in the ratio of their:  
(A) atomic number (B) atomic masses (C) specific gravities (D) equivalent masses

$$W = Z i t$$

$$E = F Z$$

$E$  = equivalent mass.

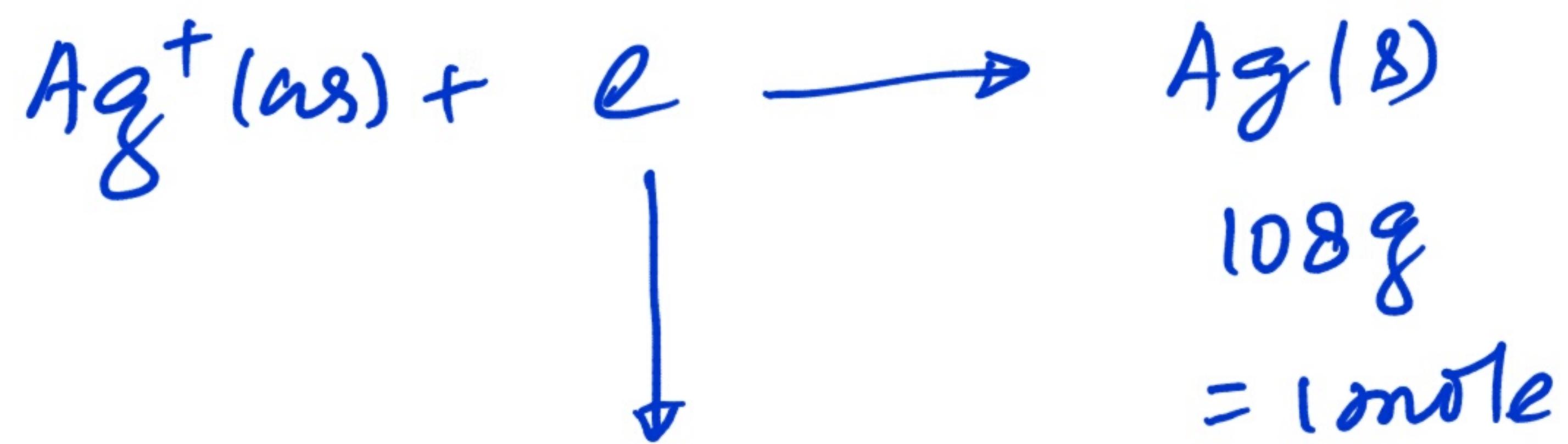
$$\therefore W = \frac{E}{F} \times i \times t$$

for same charge supplied

$$W \propto E$$

39. The amount of electricity that can deposit 108 g. of silver from silver nitrate solution is:

  - (A) 1 ampere
  - (B) 1 coulomb
  - (C) 1 Faraday
  - (D) 2 ampere



$$1 \text{ mole} = 1 \text{ Faraday}$$

40. The ratio of weights of hydrogen and magnesium deposited by the same amount of electricity from aqueous  $\text{H}_2\text{SO}_4$  and fused  $\text{MgSO}_4$  are :
- (A) 1 : 8      (B) 1 : 12      (C) 1 : 16      (D) None of these

$$\frac{W_{\text{H}_2}}{W_{\text{Mg}}} = \frac{E_{\text{H}_2}}{E_{\text{Mg}}} = \boxed{\frac{1}{12}}$$

Ans.

41. A current of 9.65 amp. flowing for 10 minute deposits 3.0 g of a metal. The equivalent wt. of the metal is :

(A) 10

(B) 30

(C) 50

(D) 96.5

$$\frac{W}{E} = \frac{ixt}{96500}$$

or,  $\frac{3}{E} = \frac{9.65 \times 10 \times 60}{96500}$

or,  $E = 508$

42. The same amount of electricity was passed through two separate electrolytic cells containing solutions of nickel nitrate  $[Ni(NO_3)_2]$  and chromium nitrate  $[Cr(NO_3)_3]$  respectively. If 0.3 g of nickel was deposited in the first cell, the amount of chromium deposited is :

(at. wt. of Ni = 59, at. wt. of Cr = 52)

- (A) 0.1 g      ~~(B) 0.17 g~~      (C) 0.3 g      (D) 0.6 g

$$\frac{w_{Ni}}{w_{Cr}} = \frac{E_{Ni}}{E_{Cr}}$$

$$\text{or, } \frac{0.3}{w_{Cr}} = \frac{(59/2)}{(52/3)} \Rightarrow w_{Cr} = \frac{0.3 \times 52 \times 2}{3 \times 59} = \boxed{0.17 \text{ g}}$$

Ans.

43. The electric charge for electro deposition of 1 equivalent of a substance is :

  - (A) one ampere per second
  - (B) 96500 coulomb per second
  - (C) one ampere for one hour
  - (D) ~~charge on one mole of electrons~~

no. of faraday = no. of equivalent

二 IF

$\equiv$  charge on 1 mole of electrons.

44. 3.17 g., of a substance was deposited by the flow of 0.1 mole of electrons. The equivalent weight of the substance is :  
(A) 3.17      (B) 0.317      (C) 317       $\checkmark$  (D) 31.7

$$\text{no. of equivalent} = \frac{\text{mass}}{E}$$

$E = \text{equivalent mass}$

$$\therefore 0.1 = \frac{3.17}{E}$$

$$\therefore E = 31.7 \text{ g}$$

$\therefore \text{no. of equivalent} = \text{no. of moles of e}^-$

45. A current of 9.65 amp. passing for 16 min. 40 sec. through a molten tin salt deposits 5.95 g. of tin. The oxidation state of the tin in the salt is : (at. wt of Sn = 119)

$$\frac{w}{E} = \frac{ixt}{96500}$$

$$t = 16 \times 60 + 40$$
$$= 960 + 40$$

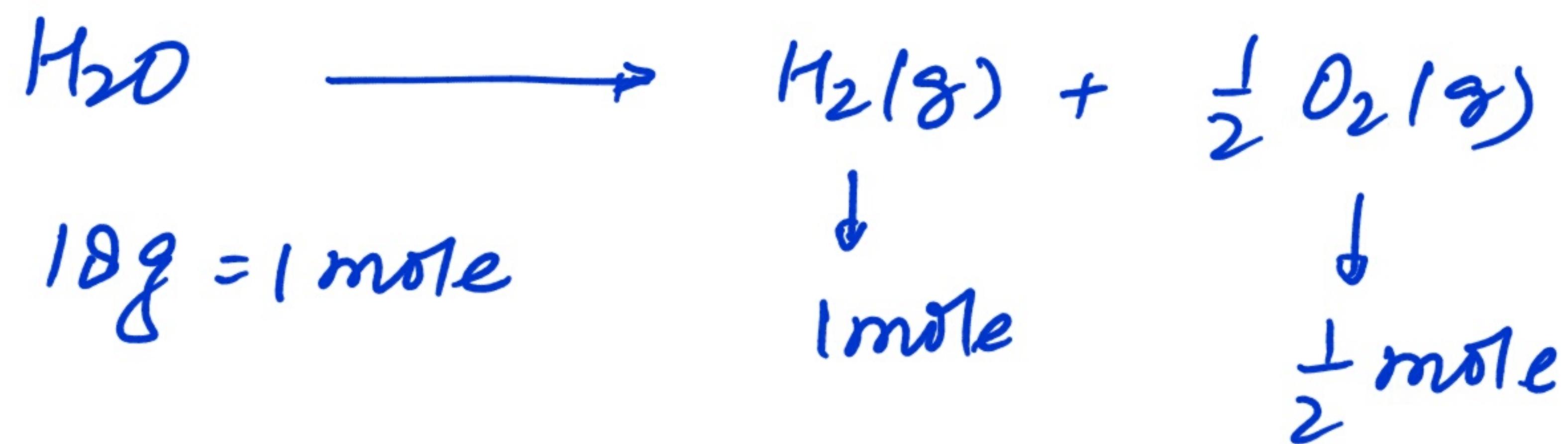
$$\text{or, } \frac{5.95}{\left(\frac{119}{n}\right)} = \frac{9.65 \times 10^9 P}{96599}$$

$$= 1000 \text{ sec}^-1$$

$$\sigma_1 = \frac{119}{n} = 59.5$$

$$\text{or, } \eta = \frac{119}{59.5} = \boxed{2} \quad \text{Ans}$$

46. The time required for a current of 3 amp. to decompose electrolytically 18 g of H<sub>2</sub>O is:  
 (A) 18 hour      (B) 36 hour      (C) 9 hour      (D) 18 seconds



$$\frac{W}{E} = \frac{ixt}{96500}$$

for H<sub>2</sub>,

$$\frac{2}{1} = \frac{3 \times t}{96500} \Rightarrow t = \frac{2 \times 96500}{3} = 64333 \text{ sec} \approx \boxed{18 \text{ hours}}$$

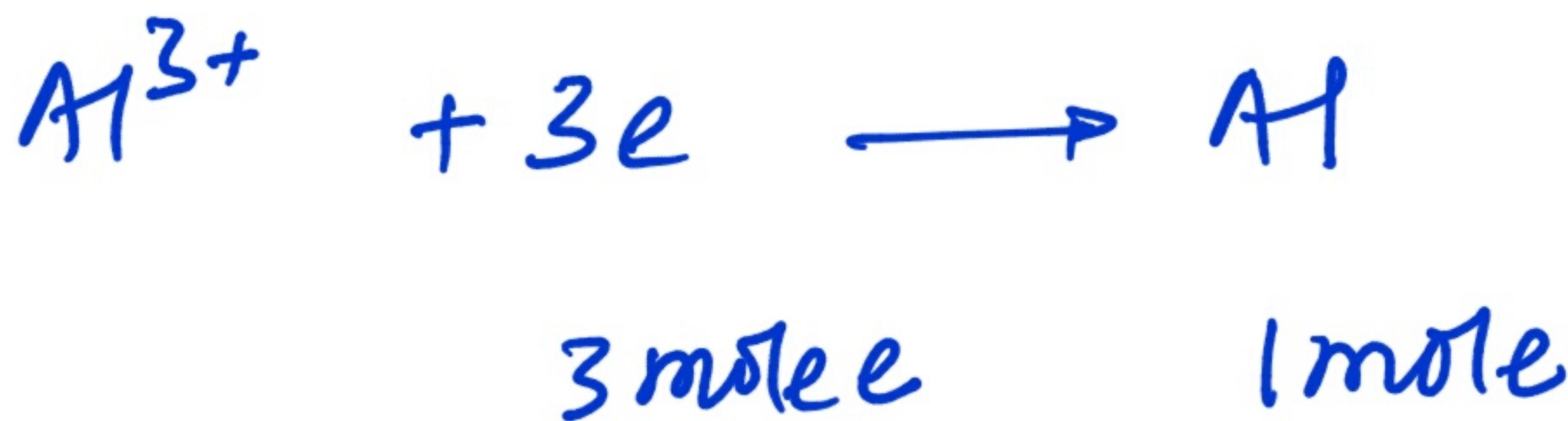
47. 1 mole of Al is deposited by X coulomb of electricity passing through aluminium nitrate solution. The number of moles of silver deposited by X coulomb of electricity from silver nitrate solution is :

~~(A) 3~~

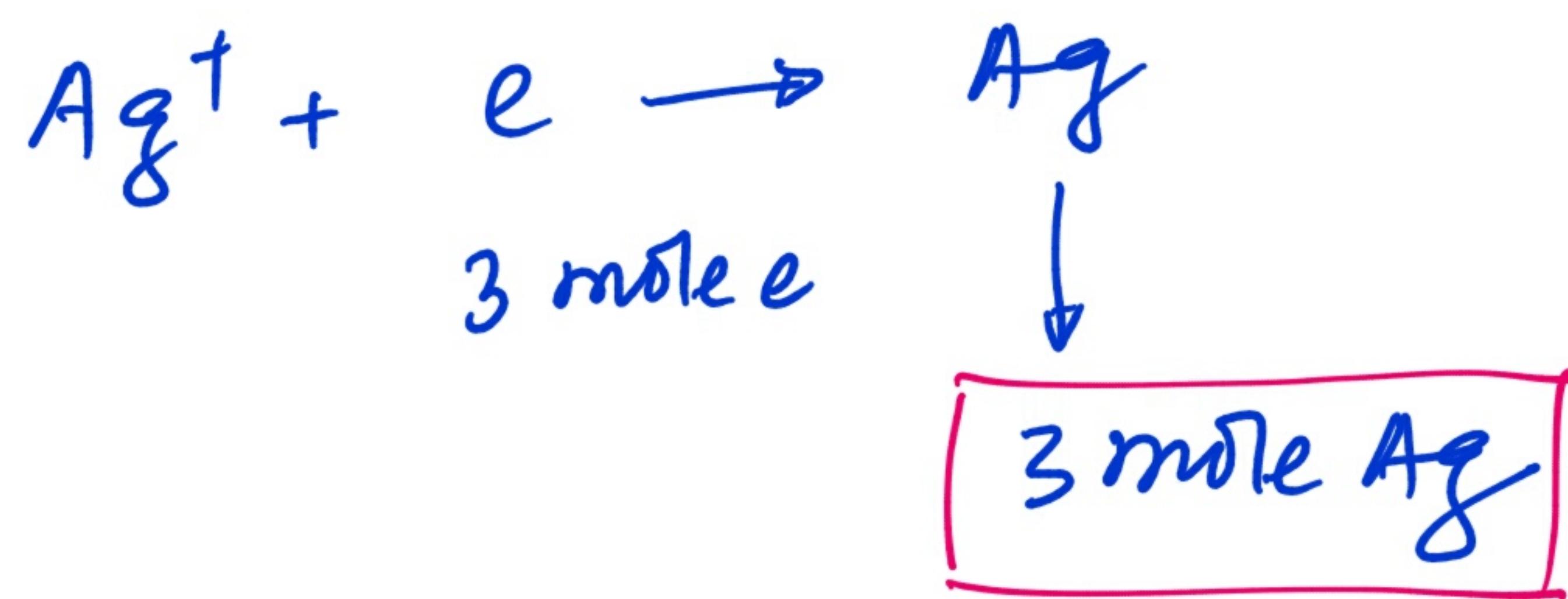
(B) 4

(C) 2

(D) 1



$$X = 3 \times 96500 \text{ Coulomb.}$$



Ans.

48. An ion is reduced to the element when it absorbs  $6 \times 10^{20}$  electrons. The number of gm equivalents of the ion is :

(A) 0.10

(B) 0.01

~~(C) 0.001~~

(D) 0.0001

$$\text{no. of moles of } e = \frac{6 \times 10^{20}}{6 \times 10^{23}} = 0.001 \text{ mole of } e.$$

$$\text{no. of equivalent} = \text{no. of moles of } e = \boxed{0.001} \text{ Ans.}$$

49. When an electric current is passed through acid diluted water, 112 ml. of hydrogen gas at STP collects at the cathode in 965 second. The current passed, in ampere is :
- (A) 1.0      (B) 0.5      (C) 0.1      (D) 2.0



112 ml at STP

$$= \frac{112}{22400} \text{ mole} = \frac{1}{200} \text{ mole} = \frac{1}{200} \times 2 \text{ g} \\ = \frac{1}{100} \text{ g}$$

$$\frac{W}{E} = \frac{i \times t}{96500}$$

$$E_{H_2} = \frac{2}{2} = 1 \text{ g}$$

$$\therefore \frac{1}{199 \times 1} = \frac{i \times 965}{96500}$$

$$\boxed{i = 1 \text{ A}} \quad \underline{\text{Ans.}}$$

50. How many coulombs of electric charge are required for the oxidation of 1 mole of  $\text{H}_2\text{O}$  to  $\text{O}_2$ ?  
(A)  $9.65 \times 10^4 \text{ C}$    (B)  $4.825 \times 10^5 \text{ C}$    (C)  $1.93 \times 10^5 \text{ C}$    (D)  $1.93 \times 10^4 \text{ C}$



$$\text{no. of moles of e transferred} = n = 2$$

$$= 2 \times 96500 \text{ C}$$

$$= 193000 \text{ C}$$

$$= \boxed{1.93 \times 10^5 \text{ C}}$$

51. A factory produces 40 kg. of calcium in two hours by electrolysis. How much aluminium can be produced by the same current in two hours :-  
(At wt. of Ca = 40, Al = 27)
- (A) 22 kg.      (B) 18 kg.      (C) 9 kg.      (D) 27 kg.

for same charge

$$\frac{W_{Ca}}{W_{Al}} = \frac{E_{Ca}}{E_{Al}}$$

or,  $\frac{40}{W_{Al}} = \frac{(40/2)}{(27/3)}$

or,  $\frac{40}{W_{Al}} = \frac{20}{9}$

$$\Rightarrow W_{Al} = 18 \text{ kg}$$

52. Calculate the volume of hydrogen at STP obtained by passing a current of 0.536 ampere through acidified water for 30 minutes.

- (A) 0.112 litre      (B) 0.224 litre      (C) 0.056 litre      (D) 0.448 litre



$$E_{H_2} = 1$$

$$i = 0.536 A, t = 30 \text{ min},$$

$$\frac{w_{H_2}}{E} = \frac{i \times t}{96500}$$

$$\text{or, } \frac{w_{H_2}}{1} = \frac{0.536 \times 30 \times 60}{96500} = 0.018$$

$$\therefore n_{H_2} = \frac{0.01}{2} \text{ mole} \quad \therefore V_{H_2 \text{ at NTP}} = \frac{0.01}{2} \times 22.4$$

$$= \boxed{0.112 L} \quad \text{Ans}$$

53. An electric current is passed through silver voltameter connected to a water voltameter in series. The cathode of the silver voltameter weighed 0.108g more at the end of the electrolysis. The volume of oxygen evolved at STP is :

- (A) 56cm<sup>3</sup>      (B) 550 cm<sup>3</sup>      ~~(C) 5.6 cm<sup>3</sup>~~      (D) 11.2 cm<sup>3</sup>

$$\text{wt of Ag deposited} = 0.108 \text{ g.}, E_{\text{Ag}} = \frac{108}{1} = 108 \text{ g}$$

$$\frac{W_{\text{Ag}}}{W_{\text{O}_2}} = \frac{E_{\text{Ag}}}{E_{\text{O}_2}}$$

$$\text{or, } \frac{0.108}{W_{\text{O}_2}} = \frac{108}{8} \Rightarrow W_{\text{O}_2} = 0.008 \text{ g}$$

$$\therefore V_{\text{O}_2 \text{ at STP}} = \frac{0.008}{32} \times 22400 \text{ ml} \\ = \boxed{5.6 \text{ ml.}}$$

54. The time required to coat a meter surface of  $80 \text{ cm}^2$  with  $5 \times 10^{-3} \text{ cm}$  thick layer of silver (density  $1.08 \text{ g cm}^{-3}$ ) with the passage of  $9.65\text{A}$  current through a silver nitrate solution is :  
 (A) 10 sec.      ~~(B) 40 sec.~~      (C) 30 sec.      (D) 20 sec.



$$\frac{W}{E} = \frac{ixt}{96500}$$

$$\text{or, } \frac{d \times A \times x}{E} = \frac{ixt}{96500}$$

$d$  = density in  $\text{g/ml}$   
 $A$  = surface area in  $\text{cm}^2$   
 $x$  = thickness in  $\text{cm.}$

$$\text{or, } \frac{1.08 \times 80 \times 5 \times 10^{-3}}{108} = \frac{9.65 \times t}{96500} \Rightarrow t = 40 \text{ sec.}$$

55. One gm metal  $M^{+2}$  was discharged by the passage of  $1.2 \times 10^{22}$  electrons. What is the atomic weight of metal?

(A) 25

(B) 50

~~(C) 100~~

(D) 75

$$no. \text{ of eq} = no. \text{ of Faraday}$$

$$\text{or, } \frac{W}{E} = \frac{1.2 \times 10^{22}}{6 \times 10^{23}}$$

$$\text{or, } \frac{I}{E} = 0.02 \Rightarrow E = 50 \text{ g.}$$

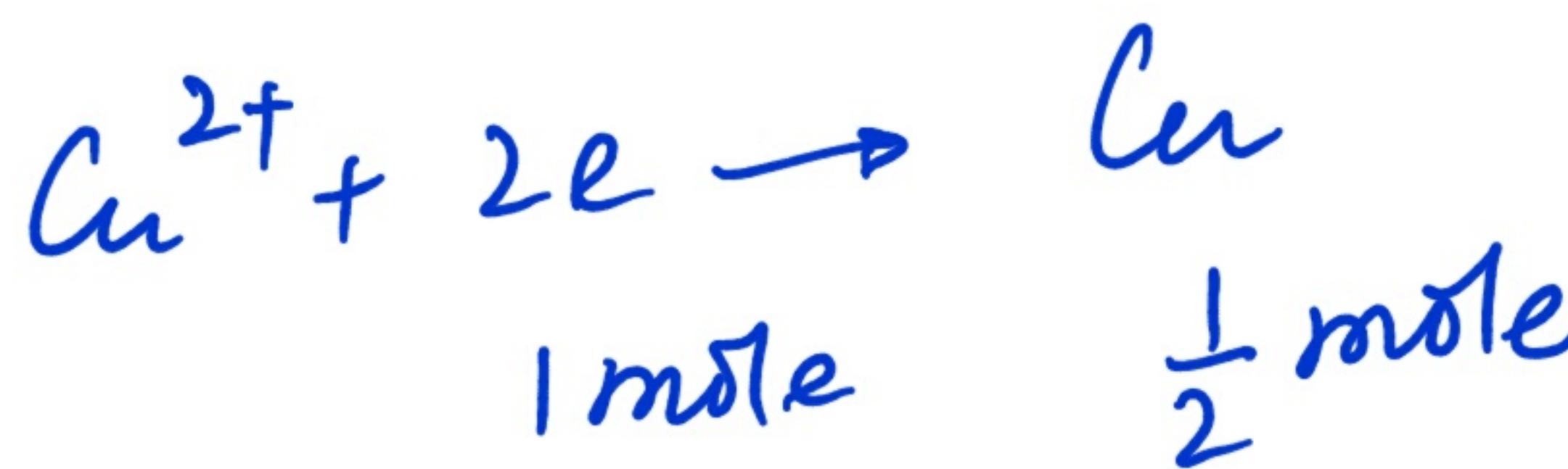
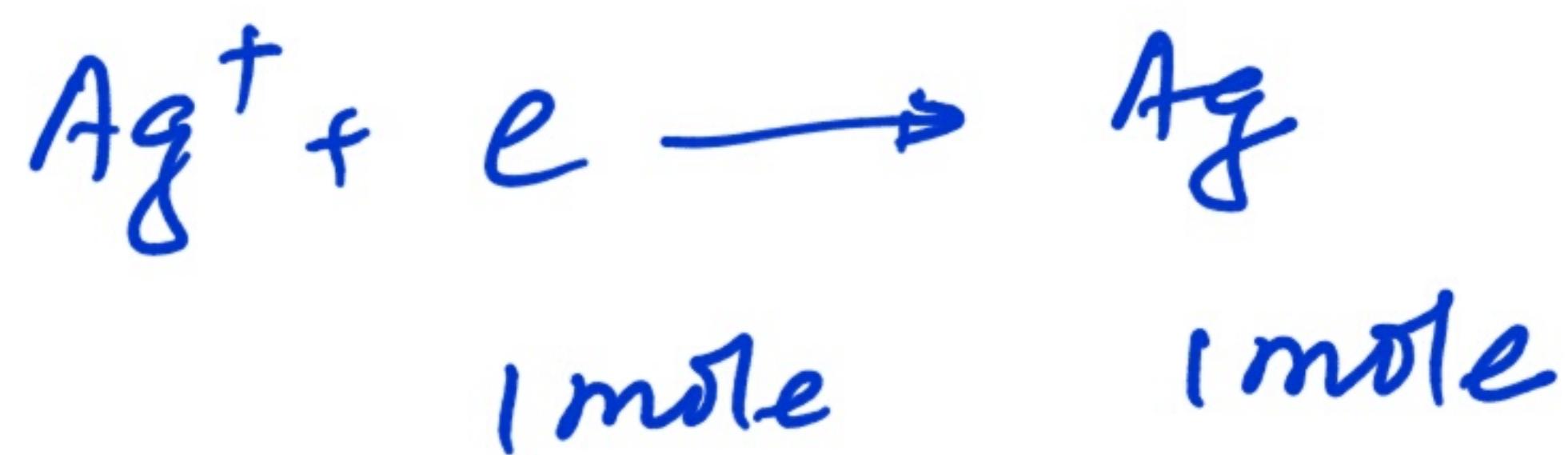
$$E = \frac{\text{Atomic wt}}{\text{valency}}$$

$$\text{or, } 50 = \frac{M}{2} \Rightarrow$$

$$M = 100$$

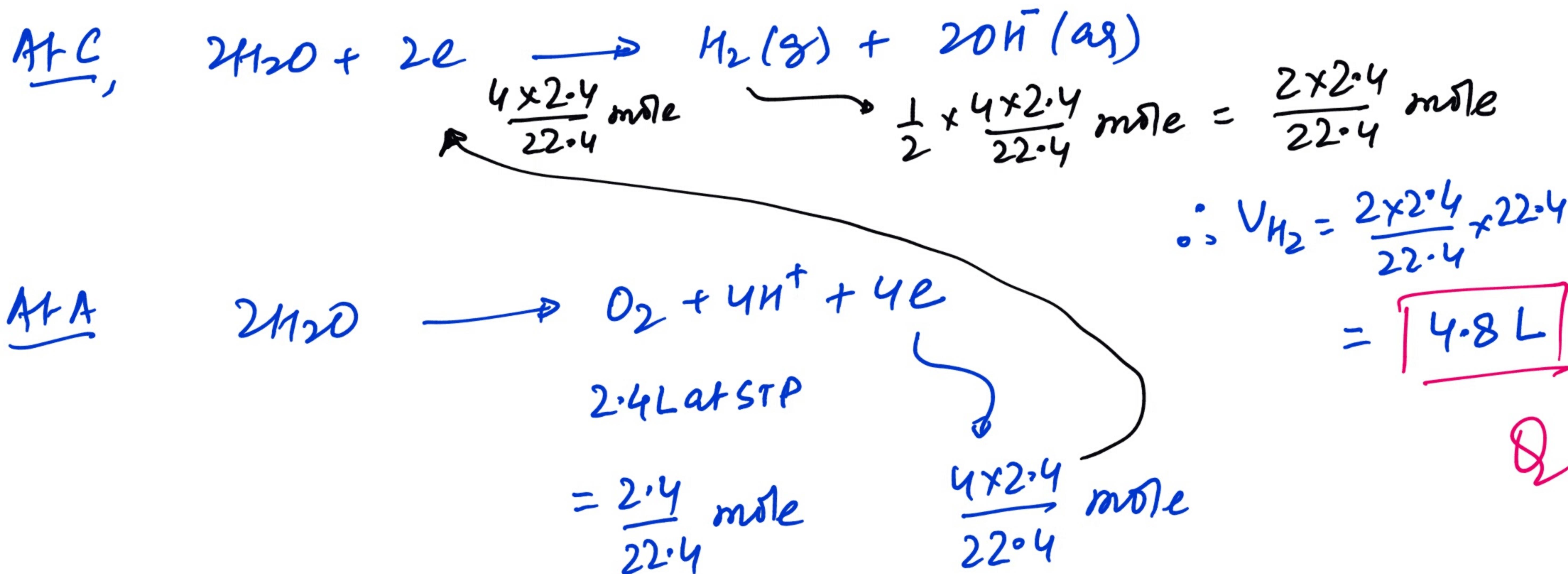
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56. One mole of electron passes through each of the solution of  $\text{AgNO}_3$ ,  $\text{CuSO}_4$  and  $\text{AlCl}_3$  when Ag, Cu and Al are deposited at cathode. The molar ratio of Ag, Cu and Al deposited are  
(A) 1 : 1 : 1      (B) 6 : 3 : 2      (C) 6 : 3 : 1      (D) 1 : 3 : 6

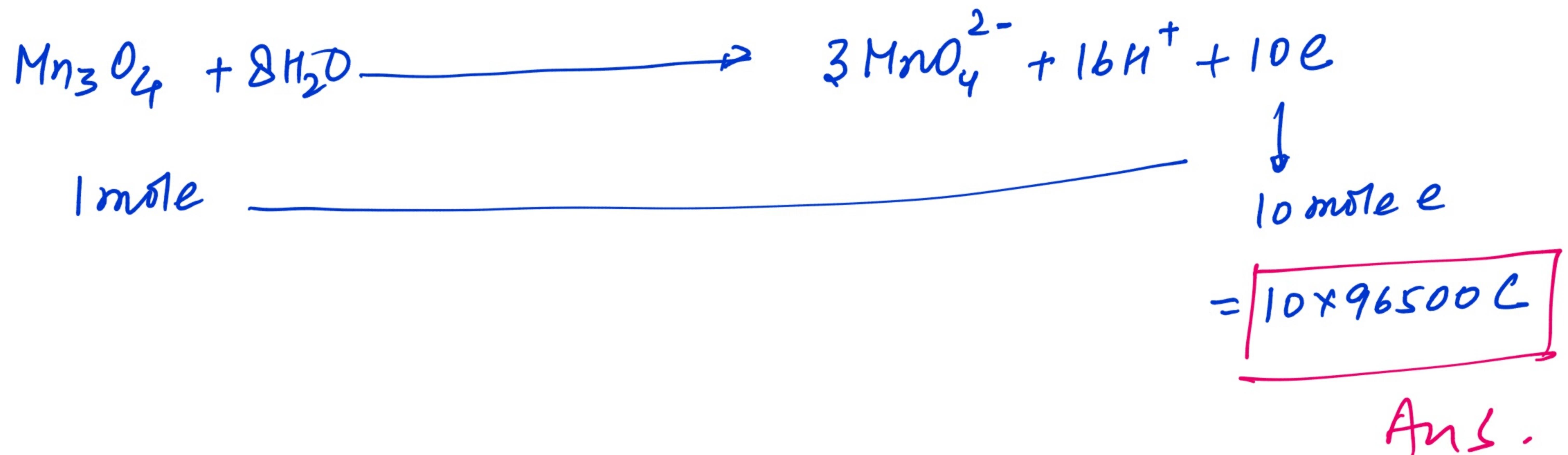


molar ratio of Ag, Cu and Al =  $1 : \frac{1}{2} : \frac{1}{3} = \boxed{6 : 3 : 2}$  Q

57. During electrolysis of an aqueous solution of sodium sulphate, 2.4 L of oxygen at STP was liberated at anode. The volume of hydrogen at STP, liberated at cathode would be
- (A) 1.2 L      (B) 2.4 L      (C) 2.6 L      ✓ (D) 4.8 L



58. The charge required for the oxidation of one mole  $\text{Mn}_3\text{O}_4$  into  $\text{MnO}_4^{2-}$  in presence of alkaline medium is



59. Electrolytic conduction differs from metallic conduction from the fact that in the former

- (A) The resistance increases with increasing temperature
- (B) The resistance decreases with increasing temperature
- (C) The resistance remains constant with increasing temperature
- (D) The resistance is independent of the length of the conductor

Resistance of an electrolytic sol<sup>n</sup> decreases with  
increase in temp while resistance of a metallic  
wire increases with increase in temp -

60. Which of the following solution of KCl has the lowest value of specific conductance :

- (A) 1 M                    (B) 0.1 M                    (C) 0.01 M                     (D) 0.001 M

specific conductance means conductivity -

conductivity always increases with increase in concentration and decreases with decrease in concentration

61. Which of the following solutions of KCl has the lowest value of equivalent conductance ?

- (A) 1 M      (B) 0.1 M      (C) .01 M      (D) .001 M

As conc. decreases  $\lambda_e$  increases.

62. The molar conductance at infinite dilution of  $\text{AgNO}_3$ ,  $\text{AgCl}$  and  $\text{NaCl}$  are 115, 120 and 110 respectively. The molar conductance of  $\text{NaNO}_3$  is :-

(A) 110

~~(B) 105~~

(C) 130

(D) 150

$$\Lambda_m^{\circ, \text{NaNO}_3} = \Lambda_m^{\circ, \text{NaCl}} + \Lambda_m^{\circ, \text{AgNO}_3} - \Lambda_m^{\circ, \text{AgCl}}$$

$$\text{or, } \Lambda_m^{\circ, \text{NaNO}_3} = (110 + 115 - 120) = \boxed{105} \text{ Ans.}$$

63. The equivalent conductivity of 0.1 N CH<sub>3</sub>COOH at 25 °C is 80 and at infinite dilution 400. The degree of dissociation of CH<sub>3</sub>COOH is :

- (A) 1                  ~~(B) 0.2~~                  (C) 0.1                  (D) 0.5

$$\alpha = \frac{\Lambda_e}{\Lambda_e^0} = \frac{80}{400} = \frac{1}{5} = 0.2$$

64. The specific conductance of a 0.01 M solution of KCl is  $0.0014 \text{ ohm}^{-1} \text{ cm}^{-1}$  at  $25^\circ \text{ C}$ . Its equivalent conductance ( $\text{cm}^2 \text{ ohm}^{-1} \text{ equiv}^{-1}$ ) is :-
- (A) 140      (B) 14      (C) 1.4      (D) 0.14

~~0.01M~~,  $K = 0.0014 \text{ S cm}^{-1}$ .

For KCl, Normality = Molarity.

$$\Lambda_e = \frac{K \times 1000}{N} = \frac{0.0014 \times 1000}{0.01} = \boxed{140 \text{ S cm}^2 \text{ eq}^{-1}}$$

65. The resistance of 0.01 N solution of an electrolyte was found to be 200 ohm at 298 K using a conductivity cell of cell constant  $1.5 \text{ cm}^{-1}$ . The equivalent conductance of solution is :-

(A)  $750 \text{ mho cm}^2 \text{ eq}^{-1}$       (B)  $75 \text{ mho cm}^2 \text{ eq}^{-1}$   
~~(C)  $750 \text{ mho}^{-1} \text{ cm}^2 \text{ eq}^{-1}$~~       (D)  $75 \text{ mho}^{-1} \text{ cm}^2 \text{ eq}^{-1}$

0.01 N, R = 200  $\Omega$

$$G^* = 1.5 \text{ cm}^{-1}. \quad \lambda_e = ?$$

$$\therefore G_1 G_1^* = k$$

$$\text{or } \frac{G^*}{R} = K \Rightarrow K = \frac{1.5}{200} \text{ s cm}^{-1}.$$

$$N_e = \frac{R \times 1000}{N} = \frac{1.5 \times 1000}{200 \times 0.01} = \frac{1500}{2} = 750 \text{ cm}^2 \text{ eq}^{-1}$$

66. The resistance of 0.1 N solution of acetic acid is 250 ohm. When measured in a cell of cell constant  $1.15 \text{ cm}^{-1}$ . The equivalent conductance (in  $\text{ohm}^{-1} \text{ cm}^2 \text{ equiv.}^{-1}$ ) of 0.1 N acetic acid is

$$0.1N, \quad R = 250\Omega, \quad g_1^* = 1.15 \text{ cm}^{-1}, \quad N_e = ?$$

$$G_1 G_1^* = K$$

$$r, \frac{G^*}{R} = K \Rightarrow K = \frac{1.15}{250} \text{ cm}^{-1}$$

$$N_e = \frac{K \times 1000}{N} = \frac{1.15 \times 1000}{250 \times 0.1} = 46 \text{ cm}^2 \text{ eq}^{-1}$$

67. If 0.01 M solution of an electrolyte has a resistance of 40 ohms in a cell having a cell constant of  $0.4 \text{ cm}^{-1}$  then its molar conductance in  $\text{ohm}^{-1} \text{ cm}^2 \text{ mol}^{-1}$  is :

(A) 10

(B)  $10^2$

~~(C)  $10^3$~~

(D)  $10^4$

$$0.01 \text{ M}, R = 40 \Omega, G_i^* = 0.4 \text{ cm}^{-1}. \Lambda_m = ?$$

$$G_i G_i^* = K$$

$$\text{or}, \frac{G_i^*}{R} = K \Rightarrow \frac{0.4}{40} = K \Rightarrow K = \frac{1}{100} \text{ S cm}^{-1}.$$

$$\Lambda_m = \frac{K \times 1000}{c} = \frac{1 \times 1000}{100 \times 0.01} = \boxed{10^3 \text{ S cm}^2 \text{ mol}^{-1}}.$$

Aus.

68. The conductivity of a saturated solution of  $\text{BaSO}_4$  is  $3.06 \times 10^{-6} \text{ ohm}^{-1} \text{ cm}^{-1}$  and its molar conductance is  $1.53 \text{ ohm}^{-1} \text{ cm}^2 \text{ mol}^{-1}$ . The  $K_{sp}$  of  $\text{BaSO}_4$  will be :
- (A) 379      (B) 139.5      (C) 203      (D) 330

$$\Lambda_m^0 = \frac{K \times 1000}{\gamma}$$

$$\text{or, } 1.53 = \frac{3.06 \times 10^{-6} \times 1000}{\gamma}$$

$$\text{or, } \gamma = \frac{3.06 \times 10^{-3}}{1.53} = 2 \times 10^{-3} \text{ mole/L}$$

$$K_{sp \text{ BaSO}_4} = \gamma^2 = (2 \times 10^{-3})^2 = \boxed{4 \times 10^{-6}} \quad \underline{\text{Ans.}}$$

→ All the given options are incorrect. Actual answer is  $4 \times 10^{-6}$ .

69. Equivalent conductances of  $\text{Ba}^{2+}$  and  $\text{Cl}^-$  ions are  $127 \text{ ohm}^{-1} \text{cm}^2 \text{eq}^{-1}$  respectively. Equivalent conductance of  $\text{BaCl}_2$  at infinite dilution is -
- (A) 379      (B) 139.5      ~~(C) 203~~      (D) 330

$$\begin{aligned}\lambda_e^{\circ}_{\text{BaCl}_2} &= \lambda_e^{\circ}_{\text{Ba}^{2+}} + \lambda_e^{\circ}_{\text{Cl}^-} \\ &= 127 + 76 \\ &= \boxed{203} \text{ S cm}^2 \text{eq}^{-1}.\end{aligned}$$

70. If  $x$  is specific resistance of the electrolyte solution and  $y$  is the molarity of the solution, then  $\Lambda_m$  is given by  $\hookrightarrow$  means resistivity

$$(A) \frac{1000x}{y}$$

$$(B) 1000 \frac{y}{x}$$

$$\checkmark (C) \frac{1000}{xy}$$

$$(D) \frac{xy}{1000}$$

$$\Lambda_m = \frac{K \times 1000}{M}$$

$$K = \frac{1}{\rho} = \frac{1}{x}$$

$$M = y$$

so,  $\Lambda_m = \frac{1 \times 1000}{x \times y}$

so,

$$\Lambda_m = \frac{1000}{xy}$$

Ans.