

## Questions

MathonGo

## Q1 - 24 June - Shift 1

The cell potential for the following cell



is 0.576 V at 298 K. The pH of the solution is \_\_\_\_.

(Nearest integer)

Space for your notes:

## Q2 - 24 June - Shift 2

The resistance of conductivity cell containing

0.01 M KCl solution at 298 K is 1750  $\Omega$ . If the

conductivity of 0.01 M KCl solution at 298 K is

$0.152 \times 10^{-3} \text{ S cm}^{-1}$ , then the cell constant of the

conductivity cell is \_\_\_\_\_  $\times 10^{-3} \text{ cm}^{-1}$ .

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## Q3 - 25 June - Shift 1

In a cell, the following reactions take place



The standard electrode potential for the

spontaneous reaction in the cell is  $x \times 10^{-2} \text{ V}$  298

K. The value of x is \_\_\_\_\_ (Nearest

Integer)

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## Q4 - 25 June - Shift 2

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The correct order of reduction potentials of the following pairs is

- A.  $\text{Cl}_2/\text{Cl}^-$
- B.  $\text{I}_2/\text{I}^-$
- C.  $\text{Ag}^+/\text{Ag}$
- D.  $\text{Na}^+/\text{Na}$
- E.  $\text{Li}^+/\text{Li}$

Choose the correct answer from the options given below.

- (A)  $A > C > B > D > E$
- (B)  $A > B > C > D > E$
- (C)  $A > C > B > E > D$
- (D)  $A > B > C > E > D$

**Q5 - 25 June - Shift 2**

A solution of  $\text{Fe}_2(\text{SO}_4)_3$  is electrolyzed for 'x' min with a current of 1.5 A to deposit 0.3482 g of Fe.

The value of x is \_\_\_\_\_. [nearest integer]

Given :  $1 \text{ F} = 96500 \text{ C mol}^{-1}$

Atomic mass of Fe =  $56 \text{ g mol}^{-1}$

**Q6 - 26 June - Shift 1**

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The  $\left(\frac{\partial E}{\partial T}\right)_P$  of different types of half cells are as follows :

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A B C D

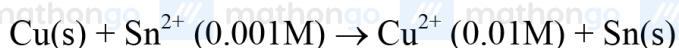
$1 \times 10^{-4}$   $2 \times 10^{-4}$   $0.1 \times 10^{-4}$   $0.2 \times 10^{-4}$

(Where E is the electromotive force)

Which of the above half cells would be preferred to be used as reference electrode ?

- (A) A (B) B  
(C) C (D) D

## Q7 - 26 June - Shift 2



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The Gibbs free energy change for the above reaction at 298 K is  $x \times 10^{-1}$  kJ mol<sup>-1</sup>;

The value of x is \_\_\_\_\_ . [nearest integer]

[Given :  $E_{\text{Cu}^{2+}/\text{Cu}}^{\ominus} = 0.34\text{V}$ ;  $E_{\text{Sn}^{2+}/\text{Sn}}^{\ominus} = -0.14\text{V}$ ;  $F = 96500\text{C mol}^{-1}$ ]

## Q8 - 27 June - Shift 1

The limiting molar conductivities of NaI, NaNO<sub>3</sub> and AgNO<sub>3</sub> are 12.7, 12.0 and 13.3 mS m<sup>2</sup> mol<sup>-1</sup>, respectively (all at 25°C). The limiting molar conductivity of AgI at this temperature is \_\_\_\_\_ mS m<sup>2</sup> mol<sup>-1</sup>

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## Q9 - 27 June - Shift 2

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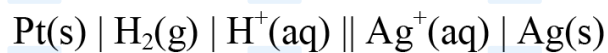
In 3d series, the metal having the highest  $M^{2+}/M$  standard electrode potential is

- (A) Cr (B) Fe  
(C) Cu (D) Zn

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**Q10 - 27 June - Shift 2**

For the reaction taking place in the cell:



$$E^\circ_{\text{Cell}} = +0.5332 \text{ V.}$$

The value of  $\Delta_r G^\circ$  is \_\_\_\_\_  $\text{kJ mol}^{-1}$ . (in nearest integer)

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**Q11 - 28 June - Shift 1**

The solubility product of a sparingly soluble salt  $\text{A}_2\text{X}_3$  is  $1.1 \times 10^{-23}$ . If specific conductance of the solution is  $3 \times 10^{-5} \text{ S m}^{-1}$ , the limiting molar conductivity of the solution is  $x \times 10^{-3} \text{ S m}^2 \text{ mol}^{-1}$ .

The value of x is \_\_\_\_\_.

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**Q12 - 28 June - Shift 1**

The quantity of electricity in Faraday needed to reduce 1 mol of  $\text{Cr}_2\text{O}_7^{2-}$  to  $\text{Cr}^{3+}$  is \_\_\_\_\_.

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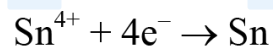
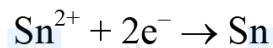
**Q13 - 28 June - Shift 2**

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For the given reactions



The electrode potentials are;  $E_{\text{Sn}^{2+}/\text{Sn}}^{\circ} = -0.140 \text{ V}$

and  $E_{\text{Sn}^{4+}/\text{Sn}}^{\circ} = 0.010 \text{ V}$ . The magnitude of

standard electrode potential for  $\text{Sn}^{4+}/\text{Sn}^{2+}$  i.e.

$E_{\text{Sn}^{4+}/\text{Sn}^{2+}}^{\circ}$  is \_\_\_\_\_  $\times 10^{-2} \text{ V}$ . (Nearest integer)

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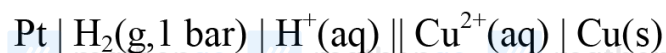
**Q14 - 29 June - Shift 1**

A dilute solution of sulphuric acid is electrolysed using a current of 0.10 A for 2 hours to produce hydrogen and oxygen gas. The total volume of gases produced at STP is \_\_\_\_\_  $\text{cm}^3$ . (Nearest integer) [Given : Faraday constant  $F = 96500 \text{ C mol}^{-1}$  at STP, molar volume of an ideal gas is  $22.7 \text{ L mol}^{-1}$ ]

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**Q15 - 29 June - Shift 2**

The cell potential for the given cell at 298 K



is 0.31V. The pH of the acidic solution is found to be 3, whereas the concentration of  $\text{Cu}^{2+}$  is  $10^{-x} \text{ M}$ .

The value of x is \_\_\_\_\_.

(Given:  $E_{\text{Cu}^{2+}/\text{Cu}}^{\circ} = 0.34 \text{ V}$  and  $\frac{2.303RT}{F} = 0.06\text{V}$ )

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**Answer Key**

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**Q1 (5)**                      **Q2 (266)**                      **Q3 (23)**                      **Q4 (A)**  
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**Q5 (20)**                      **Q6 (C)**                      **Q7 (983)**                      **Q8 (14)**  
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**Q9 (C)**                      **Q10 (51)**                      **Q11 (3)**                      **Q12 (6)**  
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**Q13 (16)**                      **Q14 (127)**                      **Q15 (7)**  
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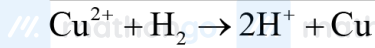
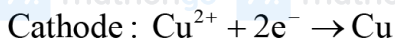
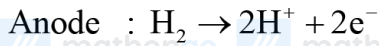
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Q1 (5)



$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{0.06}{2} \log \frac{[\text{H}^+]^2}{[\text{Cu}^{2+}]}$$

$$0.576 = 0.34 - \frac{0.06}{2} \log \left\{ \frac{[\text{H}^+]^2}{(0.01)} \right\}$$

$$+ 3.93 - \log(\text{H}^+) + \log 0.1 \Rightarrow \text{pH} = 4.93 \simeq 5$$

Q2 (266)

$$K = \frac{1}{R} \times \text{cell constant}$$

$$0.152 \times 10^{-3} = \frac{1}{1750} \text{ cell constant}$$

$$\text{cell constant} = 266 \times 10^{-3}$$

Q3 (23)



$$E_{\text{Cell}}^0 = E_{\text{cathode}}^0 - E_{\text{anode}}^0$$

$$= 0.77 - 0.54$$

$$= 0.23$$

$$= 23 \times 10^{-2} \text{ V}$$

Q4 (A)

$$E_{\text{Cl}_2/\text{Cl}^{-}}^{\circ} = +1.36 \text{ V}$$

$$E_{\text{I}_2/\text{I}^{-}}^{\circ} = +0.54 \text{ V}$$

$$E_{\text{Ag}^{+}/\text{Ag}}^{\circ} = +0.80 \text{ V}$$

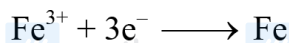
$$E_{\text{Na}^{+}/\text{Na}}^{\circ} = -2.71 \text{ V}$$

$$E_{\text{Li}^{+}/\text{Li}}^{\circ} = -3.05 \text{ V}$$

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## Hints and Solutions

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**Q5 (20)**

3F  $\longrightarrow$  1 mole Fe is deposited

For 56 g  $\longrightarrow$   $3 \times 96500$  (required charge)

For 1g  $\longrightarrow$   $\frac{3 \times 96500}{56}$  (required charge)

For 0.3482 g  $\longrightarrow$   $\frac{3 \times 96500}{56} \times 0.3482$   
 $= 1800.06$

Q = it

$$1800.06 = 1.5 t$$

t = 20 min

**Q6 (C)**

A cell with less variation in EMF with temperature is preferred as reference electrode because it can be used for wider range of temperature without much derivation from standard value so a cell with less

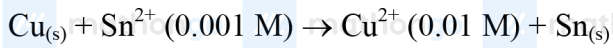
$\left(\frac{\partial E}{\partial T}\right)_p$  is preferred.

**Q7 (983)**

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## Hints and Solutions

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$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$

$$= -0.14 - (0.34)$$

$$= -0.48 \text{ V}$$

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

$$= -0.48 - \frac{0.059}{2} \log \frac{0.01}{0.001}$$

$$= -0.509$$

$$\Delta G = -nF E_{\text{cell}}$$

$$= -2 \times 96500 \times (-0.5095)$$

$$= 98333.5 \text{ J/mol}$$

$$= 98.335 \text{ kJ/mol}$$

$$= 98.35 \times 10^{-1} \text{ kJ/mol}$$

Nearest Integer : 983

**Q8 (14)**

Given

$$(1) \lambda_m^{\infty} (\text{NaI}) = 12.7 \text{ mS m}^2 \text{ mol}^{-1}$$

$$(2) \lambda_m^{\infty} (\text{NaNO}_3) = 12.0 \text{ mS m}^2 \text{ mol}^{-1}$$

$$(3) \lambda_m^{\infty} (\text{AgNO}_3) = 13.3 \text{ mS m}^2 \text{ mol}^{-1}$$

$$\lambda_m^{\infty} (\text{Ag I}) = (1) + (3) - (2)$$

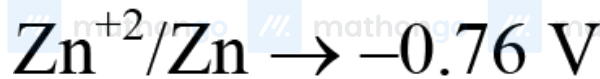
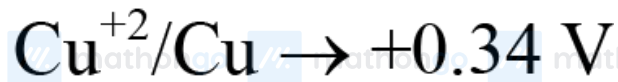
$$= 12.7 + 13.3 - 12.0$$

$$= 26.0 - 12.0$$

$$\lambda_m^{\infty} (\text{Ag I}) = 14.0$$

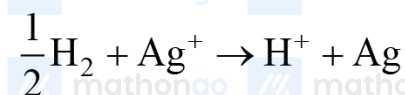
**Q9 (C)**

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So Ans.  $\text{Cu}^{+2}/\text{Cu}$

**Q10 (51)**



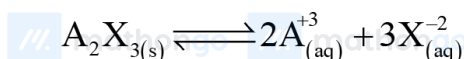
$$\Delta G^\circ = -nE^\circ F$$

$$= -1 \times 0.5332 \times 96500 \text{ J}$$

$$= -51.35 \text{ kJ}$$



**Q11 (3)**



$$\text{solubility} = s \text{ M} \quad 2s \quad 3s$$

$$(2s)^2(3s)^3 = 1.1 \times 10^{-23}$$

$$108 s^5 = 1.1 \times 10^{-23}$$

$$s \approx 10^{-5} \text{ M} = 10^{-5} \frac{\text{mol}}{\text{L}} = 0.01 \frac{\text{mol}}{\text{m}^3}$$

$$\text{Now } \wedge_m \approx \wedge_m^\infty = \frac{k}{m} = \frac{k}{s}$$

$$\Rightarrow \wedge_m^\infty = \frac{3 \times 10^{-5}}{0.01} = 3 \times 10^{-3} \text{ S-m}^2/\text{mol}$$

Ans. 3

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## Hints and Solutions

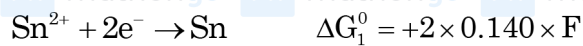
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Q12 (6)



1mol      6mol  
 $\Rightarrow$  number of faradays = moles of electrons  
 = 6

Q13 (16)



$$\Delta G_3^0 = \Delta G_2^0 - \Delta G_1^0$$

$$-2 \times E^0 \times F = -(0.04 + 0.28) \times F$$

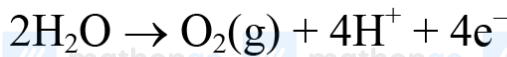
$$E^0 = 0.16 \text{ volt} = 16 \times 10^{-2} \text{ V}$$

Ans 16

Q14 (127)

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At anode



At cathode



$$\text{Now number of gm eq.} = \frac{i \times t}{96500}$$

$$= \frac{0.1 \times 2 \times 60 \times 60}{96500}$$

$$= 0.00746$$

$$V_{\text{O}_2} = \frac{0.00746}{4} \times 22.7 = 0.0423$$

$$V_{\text{H}_2} = \frac{0.00746}{2} \times 22.7 = 0.0846$$

$$V_{\text{Total}} \approx 127 \text{ ml or cc}$$

**Q15 (7)**



$$0.31 = 0.34 - \frac{0.06}{2} \log \frac{[\text{H}^+]^2}{[\text{Cu}^{2+}]}$$

$$[\text{Cu}^{2+}] = 10^{-7} \text{ M}$$

$$x = 7$$