Chemistry · Electrochemistry

Electrolytic Cells

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Multiple choice questions — 77 questions

1 Which of the following is correct for electrolytic cells?

(A) Electrons move in the electrolyte from the cathode to the anode.

(B) Electrons move in the external circuit towards the cathode.

(C) Negative ions move in the external circuit towards the anode.

(D) Positive ions move in the electrolyte towards the anode.

---

2 All of the following happen in electrolytic cells, except .....

(A) An oxidation process at the anode.

(B) Non-spontaneous reactions.

(C) Chemical decomposition of the electrolyte.

(D) Transfer of electrons from the cathode to the anode.

---

3 An electrolyte solution is electrically neutral, because .....

(A) the number of cations in it equals the number of anions.

(B) the sum of the positive charges of its cations equals the sum of the negative charges of its anions.

(C) the positive charge on each cation equals the negative charge on each anion.

(D) the solvent can separate the cations from the anions.

---

4 Which of the following describes the movement of cations in electrolyte solutions?

(A) In electrolytic cells: towards the cathode, in galvanic cells: towards the anode half-cell

(B) In electrolytic cells: towards the anode, in galvanic cells: towards the cathode half-cell

(C) In electrolytic cells: towards the cathode, in galvanic cells: towards the cathode half-cell

(D) In electrolytic cells: towards the anode, in galvanic cells: towards the anode half-cell

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5 Electrolysis of $CuCl_2$ solution between graphite electrodes causes all of the following, except .....

(A) Copper cations are reduced at the cathode.

(B) Chlorine is evolved at the anode.

(C) The blue colour gradually fades.

(D) Graphite atoms are oxidized at the anode.

---

6 Which of the following happens during the electrolysis of copper(II) chloride solution? [Cu = 63.5 , Cl = 35.5]

(A) Reduction of copper atoms.

(B) Reduction of chloride ions.

(C) The mass of copper deposited is less than the mass of chlorine evolved.

(D) The mass of copper deposited is greater than the mass of chlorine evolved.

---

7 What substance is formed at the anode during the electrolysis of molten NaCl?

(A) Chlorine. (B) Oxygen. (C) Sodium. (D) Hydrogen.

---

8 Which of the following processes takes place at the cathode of the cell for the electrolysis of molten lead(II) bromide?

(A) $Pb \rightarrow Pb^{2+} + 2e^-$

(B) $Br_2 + 2e^- \rightarrow 2Br^-$

(C) $2Br^- \rightarrow Br_2 + 2e^-$

(D) $Pb^{2+} + 2e^- \rightarrow Pb$

---

9 The reaction at the cathode of an electrolytic cell with graphite electrodes and dilute hydrochloric acid as the electrolyte is .....

(A) $H_2(g) \rightarrow 2H^+(aq) + 2e^-$

(B) $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$

(C) $2H^+(aq) + 2e^- \rightarrow H_2(g)$

(D) $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

---

10 The process $Ni^{2+}(aq) + 2e^- \rightarrow Ni^0(s)$ takes place in .....

(A) a galvanic cell at the cathode or an electrolytic cell at the anode.

(B) a galvanic cell or an electrolytic cell at the cathode.

(C) a galvanic cell or an electrolytic cell at the anode.

(D) a galvanic cell at the anode or an electrolytic cell at the cathode.

---

11 Comparing the galvanic cell with the electrolytic cell, we find that oxidation takes place at .....

(A) the anode in both cells.
(B) the cathode in both cells.
(C) the anode in the galvanic cell only.
(D) the cathode in the galvanic cell only.

12 Which of the following describes the electrolytic cell?

(A) Electrode: anode, process: oxidation, electrons: gain
(B) Electrode: anode, process: reduction, electrons: loss
(C) Electrode: cathode, process: oxidation, electrons: loss
(D) Electrode: cathode, process: reduction, electrons: gain

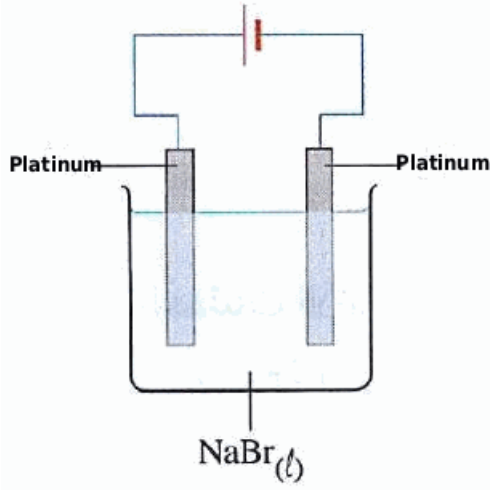
13 In the electrolysis of molten NaI, a metal formed at one electrode and a brown substance at the other electrode.

What is the cathode half-cell reaction?

(A) $2I^- \rightarrow I_2 + 2e^-$ (B) $Na^+ + e^- \rightarrow Na$ (C) $I_2 + 2e^- \rightarrow 2I^-$ (D) $Na \rightarrow Na^+ + e^-$

14 From the opposite cell,

which of the following choices is correct?



(A) The mass of the positive electrode decreases.
(B) The mass of the negative electrode decreases.
(C) The mass of the cathode increases and the electrolyte concentration stays constant.
(D) The mass of the cathode increases and no change occurs in the mass of the anode.

15 Magnesium can be extracted from molten magnesium chloride by electrolysis.
Which of the following gives the two reactions at the positive and negative electrodes of the cell used?

(A) Positive electrode: $Cl_2 + 2e^- \rightarrow 2Cl^-$, negative electrode: $Mg \rightarrow Mg^{2+} + 2e^-$

(B) Positive electrode: $Cl_2 + 2e^- \rightarrow 2Cl^-$, negative electrode: $Mg^{2+} + 2e^- \rightarrow Mg$

(C) Positive electrode: $2Cl^- \rightarrow Cl_2 + 2e^-$, negative electrode: $Mg \rightarrow Mg^{2+} + 2e^-$

(D) Positive electrode: $2Cl^- \rightarrow Cl_2 + 2e^-$, negative electrode: $Mg^{2+} + 2e^- \rightarrow Mg$

---

16 An electrolytic cell in which the two processes at its electrodes are given by the following equations:

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

$2Y^-(aq) \rightarrow Y_2(g) + 2e^-$

Which of the following describes this cell?

(A) Process: $X^{2+} + 2e^- \rightarrow X$ at the anode, electrode mass after 15 min: decreases

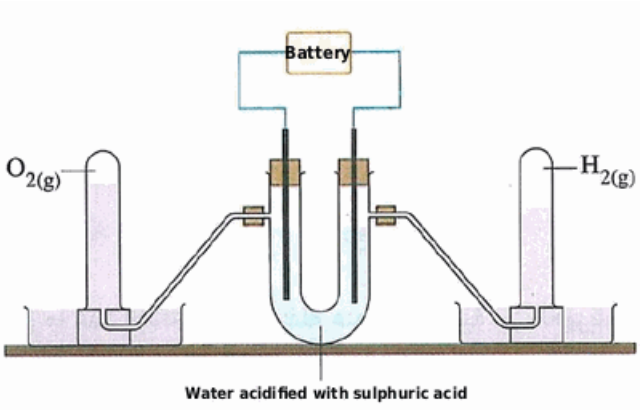
(B) Process: $X^{2+} + 2e^- \rightarrow X$ at the cathode, electrode mass after 15 min: increases

(C) Process: $2Y^- \rightarrow Y_2 + 2e^-$ at the anode, electrode mass after 15 min: increases

(D) Process: $2Y^- \rightarrow Y_2 + 2e^-$ at the cathode, electrode mass after 15 min: decreases

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17 The opposite figure shows the electrolysis of water acidified with sulphuric acid using two graphite electrodes. When X mol of electrons pass through the cell, 0.25X mol of ..... is evolved.



(A) O_2 at the anode.

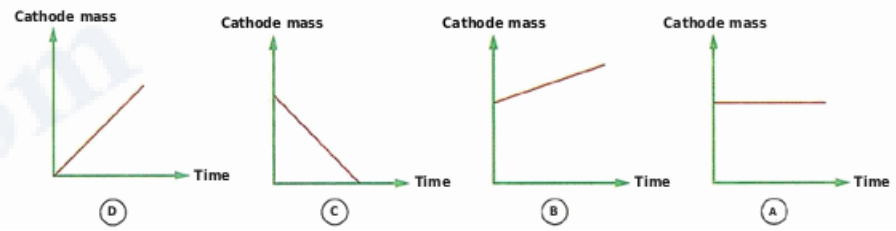
(B) H_2 at the anode.

(C) O_2 at the cathode.

(D) H_2 at the cathode.

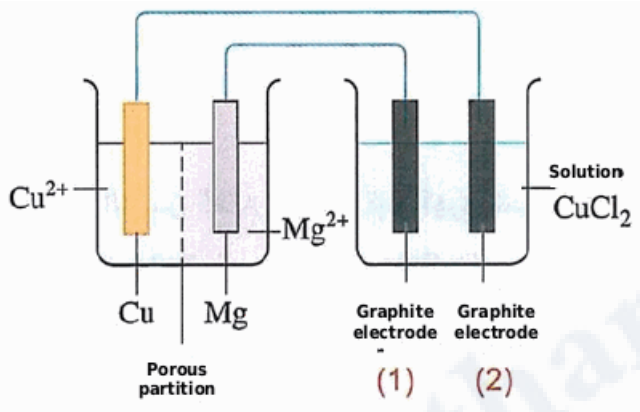
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18 Which graph shows the mass of the cathode with time when a constant current is passed through molten zinc sulphate using zinc electrodes?



19 From the opposite figure:

What substances are formed at the graphite electrodes (1) and (2)?



(A) (1): Cu , (2): H_2

(B) (1): Cl_2 , (2): Cu

(C) (1): Cu , (2): Cl_2

(D) (1): O_2 , (2): Cu

20 From the opposite potentials:

$A^+ + e^- \rightarrow A$, $E^\circ = +0.52$ V

$B^+ + e^- \rightarrow B$, $E^\circ = +0.25$ V

Which of the following describes the electrochemical cell in which the metal A is oxidized and the ion B^+ is reduced?

(A) A galvanic cell, its emf is +0.27 V

(B) A galvanic cell, its emf is −0.27 V

(C) An electrolytic cell, its emf is +0.27 V

(D) An electrolytic cell, its emf is −0.27 V

21 What quantity of electricity is consumed when a current of 0.1 A passes through an electrolyte solution for half an hour?

(A) 80 C

(B) 180 C

(C) 360 C

(D) 3600 C

22 What quantity of electricity results from passing a current of 2.5 A for 120 min through an electrolyte solution?

(A) 0.332 F (B) 0.187 F (C) 0.233 F (D) 1.076 F

23 What current is needed to pass 0.36 F through an electrolyte solution in one hour?

(A) 9.65 A (B) 579 A (C) 0.965 A (D) 5.79 A

24 What is the number of electrons transferred when a quantity of electricity of one coulomb passes through an electrolyte?

(A) $6.023 \times 10^{23}\ e^-$ (B) $1.64 \times 10^{-24}\ e^-$ (C) $6.24 \times 10^{18}\ e^-$ (D) $6.24 \times 10^{-24}\ e^-$

25 When a current of 5 A passes through molten calcium chloride, 30 g of calcium [Ca = 40] is deposited in a time of .....

(A) 8 h (B) 10 h (C) 16 h (D) 80 h

26 A current of 1.2 A was passed through a solution of copper(II) ions, and all the copper in it was deposited; its mass was 0.2 g

What is the time taken to deposit all the copper in the solution? [Cu = 63.5]

(A) 253 s (B) 507 s (C) 1012 s (D) 2024 s

27 What is the time needed to separate half the silver in 200 mL of 0.1 M silver nitrate solution using a current of 0.1 A?

(A) 40 min (B) 80.8 min (C) 120 min (D) 160.8 min

28 What is the time needed to generate a current of 1 A («assuming it can be obtained») when a 100 g zinc rod is placed in 1 L of 1 M $CuSO_4$ solution? [Zn = 65.5]

(A) 80.42 h (B) 53.6 h (C) 24.6 h (D) 1 h

29 What quantity of electricity is needed to deposit 6.57 g of gold [Au = 197] from an electrolyte containing Au^{3+} ions?

(A) 1 F (B) 0.1 F (C) 1 C (D) 0.1 C

30 How many gram-equivalent masses of copper can be deposited at the cathode of the cell for the electrolysis of $CuSO_4$ solution after a quantity of electricity of 241250 C passes through it?

(A) 1 (B) 2 (C) 2.5 (D) 1.25

31 When 1.5 F passes through a solution of the chloride of metal (M), 0.5 mol of the metal (M) is deposited.

What is the molecular formula of the chloride of metal (M)?

(A) MCl (B) MCl_2 (C) MCl_3 (D) M_2Cl

32 When a current of 1 A passes for 15 min through a solution of a metal salt, 0.173 g of the metal is deposited, so the equivalent mass of the metal is .....

(A) 18.55 g (B) 155.7 g (C) 9.27 g (D) 0.0016 g

33 When a current of 15 A passes for 50 min through a solution of a divalent metal, the mass of the cathode increased by 9.35 g

What is the gram atomic mass of this element?

(A) 20 g/atom (B) 30 g/atom (C) 40 g/atom (D) 50 g/atom

34 The masses of the following substances can gain (or lose) 1 mol of electrons during electrochemical processes, except .....

(A) 20 g of Ca^{2+} ions , [Ca = 40]
(B) 23 g of Na^{+} ions , [Na = 23]
(C) 16 g of O^{2-} ions , [O = 16]
(D) 35.5 g of Cl^{-} ions , [Cl = 35.5]

35 When 1 mol of electrons passes through an electrolytic cell, it can deposit .....

(A) 65 g of Zn , [Zn = 65]
(B) 24 g of Mg , [Mg = 24]
(C) 11.5 g of Na , [Na = 23]
(D) 9 g of Al , [Al = 27]

36 What is the mass of the substance formed at the positive electrode when 0.5 F passes through molten silver bromide? [Ag = 108 , Br = 80]

(A) 27 g (B) 40 g (C) 54 g (D) 80 g

37 A current of X A was passed for t s through an aqueous $CuSO_4$ solution, and an amount of copper [Cu = 63.5] was deposited that can be calculated from the relation .....

(A) X t / (63.5 × 96500)
(B) X t / (31.75 × 96500)
(C) X t / (31.75 × t)
(D) 31.75 × (X t) / 96500

38 What amount of copper [Cu = 63.5] can be deposited from $CuSO_4$ solution when a quantity of electricity of 0.5 F passes through it?

(A) 31.75 g
(B) 63.5 g
(C) 15.875 g
(D) 127 g

39 When a quantity of electricity of (X) faradays passes through molten aluminium oxide, 13.5 g of aluminium [Al = 27] is deposited.

What is the value of (X)?

(A) 2 F
(B) 1.5 F
(C) 1 F
(D) 0.5 F

40 What quantity of electricity is needed to deposit 39 g of potassium [K = 39] from molten potassium fluoride?

(A) 96500 C
(B) 48250 C
(C) 193000 C
(D) 10000 C

41 Separating one mol of magnesium from molten $MgCl_2$ needs a quantity of electricity of .....

(A) 1 F
(B) 2 F
(C) 3 F
(D) 4 F

42 The following chemical equation gives the products of the electrolysis of aqueous $NaClO_3$ solution:

$NaClO_3 + H_2O \rightarrow NaClO_4 + H_2$

What quantity of electricity is needed to produce 0.5 mol of $NaClO_4$?

(A) 1 F
(B) 1.5 F
(C) 2 F
(D) 3 F

43 What quantity of electricity is needed to evolve 355 g of Cl_2 gas [Cl = 35.5] by the electrolysis of molten NaCl?

(A) 9.25×10^4 C
(B) 9.65×10^5 C
(C) 9.65×10^4 C
(D) 4.83×10^5 C

44 What quantity of electricity is needed to obtain all the chlorine in 500 mL of 4 M sodium chloride solution by electrolysis?

(A) 24129 C (B) 48250 C (C) 96500 C (D) 193000 C

45 0.44 g of the metal (M), of atomic mass 88, is deposited when a current of 0.5 A passes for (32 min + 10 s) through its molten chloride.

What quantity of electricity is needed to deposit one mole of this metal?

(A) 0.2 F (B) 0.4 F (C) 2 F (D) 4 F

46 When a quantity of electricity of 4 F passes through $CuSO_4$ solution, the number of moles of copper deposited equals .....

(A) 3 mol (B) 2 mol (C) 1.5 mol (D) 4 mol

47 What is the number of moles of electrons produced by consuming 0.347 g of lithium [Li = 6.94] in a lithium-ion battery?

(A) 3.47 mol e^- (B) 1 mol e^- (C) 0.5 mol e^- (D) 0.05 mol e^-

48 What amount of chlorine gas [Cl = 35.5] is produced by the electrolysis of molten NaCl using a current of 1 A for 30 min?

(A) 0.66 mol (B) 0.33 mol (C) 0.66 g (D) 0.33 g

49 Chromium is deposited by the electrolysis of a solution containing Cr^{6+} ions.

What is the number of moles of chromium deposited using a current of 8 A for 2 h?

(A) 0.3 mol (B) 0.1 mol (C) 0.9 mol (D) 1.2 mol

50 In an electrolysis experiment, a quantity of electricity was passed through two electrolytes connected in series; 14 g of iron was deposited from a solution containing $Fe^{2+}(aq)$ ions, and 6.875 g of manganese.

What is the charge number of the manganese ion in its electrolyte solution? [Fe = 56 , Mn = 55]

(A) +1 (B) +2 (C) +3 (D) +4

51 The metal (M) used in galvanization can be obtained by the electrolysis of the molten chloride of this metal.

How many moles of the metal (M) can be obtained using a quantity of electricity of 9.65 C?

(A) 5×10^{-4} mol (B) 1×10^{-4} mol (C) 5×10^{-5} mol (D) 1×10^{-5} mol

52 What is the volume of H_2 gas produced (at STP) by the electrolysis of acidified water when a current of 4 A passes through it for 30 min?

(A) 0.0836 L (B) 0.0432 L (C) 0.1672 L (D) 0.836 L

53 What volume of oxygen gas can be liberated using a quantity of electricity of 96500 C (at STP)?

(A) 6.5 L (B) 5.6 L (C) 11.2 L (D) 22.4 L

54 What current must pass for one hour through KI solution to form 10 g of iodine? [I = 127]

(A) 1.05 A (B) 2.1 A (C) 13.9 A (D) 1.12 A

55 In the electrolysis of water acidified with sulphuric acid, 1.12 mL of hydrogen gas (at STP) is collected every second.

What is the current used in this process?

(A) 9.65 A (B) 19.3 A (C) 0.965 A (D) 1.93 A

56 Water acidified with dilute sulphuric acid is electrolysed, forming O_2 and H_2 gases.

What current must pass for 3 h through the acidified water to obtain 0.5 L of O_2 gas (at STP)?

(A) 0.5 A (B) 0.6 A (C) 0.8 A (D) 0.9 A

57 A current of 10 A was passed for 965 s through 1 L of 1 M HCl solution.

What is the pH of the solution at the end of the experiment?

(A) 0.8 (B) 0.2 (C) 0 (D) 0.0458

58 A quantity of electricity of 0.5 F was passed through 1 L of 2 M copper(II) chloride solution.

What is the concentration of the Cu^{2+} ions remaining in the electrolytic cell used?

(A) 0.25 M (B) 0.5 M (C) 4.5 M (D) 1.75 M

59 When the same quantity of electricity is used for electrolysis in each of the following two cases:

Case (1): the electrolyte used is 1 M H_2SO_4(aq)

Case (2): the electrolyte used is 1 M HCl(aq)

the amount of hydrogen evolved in case (1) compared with the amount evolved in case (2) is .....

(A) half the amount.
(B) a quarter of the amount.
(C) double the amount.
(D) the same amount.

---

60 When a quantity of electricity of 0.3 F passes through molten zinc chloride and molten scandium chloride, each separately, ... are deposited.

(A) 1.5 mol of Zn , 1 mol of Sc
(B) 1.5 mol of Zn , 0.1 mol of Sc
(C) 0.15 mol of Zn , 0.1 mol of Sc
(D) 0.15 mol of Zn , 1 mol of Sc

---

61 Which of the following is the maximum mass of oxygen gas that can be produced by a quantity of electricity equal to the quantity that can produce 0.504 g of hydrogen gas? [H = 1 , O = 16]

(A) 0.4 g
(B) 2 g
(C) 4 g
(D) 8 g

---

62 A galvanic cell made of a Cd electrode in $Cd(NO_3)_2$ solution and a Zn electrode in $Zn(NO_3)_2$ solution. At the start of the experiment each electrode had a mass of 5 g, and after some time of operation one electrode had a mass of 4.75 g

Which electrode lost mass, and by how much did the mass of the other electrode increase? [Cd = 112.4 , Zn = 65.37]

(A) 0.43 g / Zn
(B) 0.145 g / Cd
(C) 0.145 g / Zn
(D) 0.43 g / Cd

---

63 The quantity of electricity used to deposit 18.25 g of copper [Cu = 63.5] from $CuSO_4$ solution is used to deposit 80 g of cerium [Ce = 140] from a solution of .....

(A) $CeNO_3$
(B) $Ce(NO_3)_2$
(C) $CeSO_4$
(D) $CeCl_4$

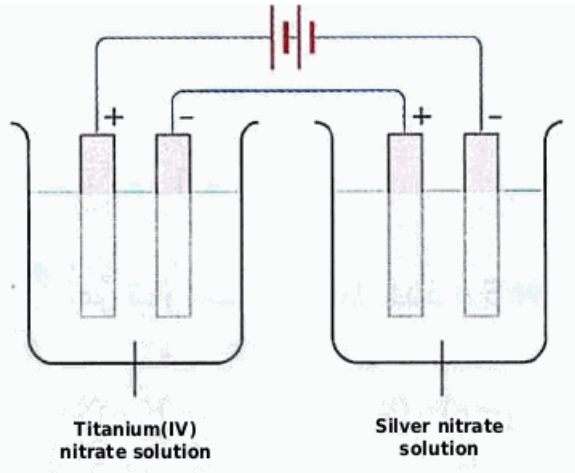
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64 A current of 1 A was passed for 1 h through three electrolyte solutions connected in series, containing Cu^{2+}(aq), Ag^{+}(aq) and Au^{3+}(aq) ions.

Which order gives the masses of the metals deposited on the cathode of each cell? [Cu = 63.5 , Ag = 108 , Au = 197]

(A) Cu < Au < Ag
(B) Ag < Cu < Au
(C) Au < Cu < Ag
(D) Ag < Au < Cu

65 From the opposite figure: after a certain time, 6 g of titanium [Ti = 48] is deposited.
What is the mass of silver [Ag = 108] deposited on the cathode of the other cell?



(A) 6.8 g (B) 13.5 g (C) 27 g (D) 54 g

---

66 The same quantity of electricity was used to deposit:
• silver from a solution containing Ag^+ ions
• copper from a solution containing Cu^{2+} ions
Which of the following describes the amount of silver deposited? [Ag = 108 , Cu = 63.5]

(A) Its mass equals half the mass of copper deposited.
(B) Its mass equals double the mass of copper deposited.
(C) Its number of moles equals half the number of moles of copper deposited.
(D) Its number of moles equals double the number of moles of copper deposited.

---

67 An electric current was passed through two cells connected in series; the first contains $AgNO_3$ solution and the second $ZnSO_4$ solution, and the electrodes in both are platinum.
What is the ratio of the mass of silver deposited to the mass of zinc deposited, respectively?
[Ag = 108 , Zn = 65]

(A) 108 / 65 (B) 65 / 108 (C) 108 / 32.5 (D) 32.5 / 108

---

68 What masses of aluminium and silver could be deposited when the same quantity of electricity passes through two solutions of their salts connected in series? [Al = 27 , Ag = 108]

(A) 1 g of Al , 12 g of Ag
(B) 2 g of Al , 6 g of Ag
(C) 12 g of Al , 1 g of Ag
(D) 3 g of Al , 8 g of Ag

---

69 The same quantity of electricity was used in the electrolysis of molten magnesium chloride and molten titanium chloride, and 4.8 g of each of Mg and Ti was deposited.
What is the formula of the titanium chloride used? [Mg = 24 , Ti = 48]

(A) $TiCl$ (B) $TiCl_2$ (C) $TiCl_3$ (D) $TiCl_4$

70 A current of a A was passed for 3 min through molten copper(II) chloride, and 10 mL of chlorine gas (at STP) was evolved at the anode.
What current is needed to produce the same volume of oxygen gas in the same time by the electrolysis of sulphuric acid (at STP)?

(A) 0.5a A (B) a A (C) 2a A (D) 4a A

71 When a quantity of electricity passes through a sulphuric acid solution, 2.24 L of hydrogen gas (at STP) can be collected. When the same quantity of electricity passes through copper(II) sulphate solution, a mass of copper [Cu = 63.5] of ..... can be deposited.

(A) 1.59 g (B) 3.18 g (C) 6.35 g (D) 12.7 g

72 Copper can be deposited from either of the two cases:
• (X): copper(II) sulphate solution.
• (Y): copper(I) chloride solution.
Using the same quantity of electricity, the amount of copper deposited in electrolytic cell .....

(A) (X) is double. (B) (Y) is double. (C) (Y) is a quarter. (D) (X) is a quarter.

73 In the electrolysis of silver nitrate solution, 1.08 g of silver [Ag = 108] was deposited on the cathode.
What is the volume of oxygen gas evolved at the anode (at STP)?

(A) 0.028 L (B) 0.056 L (C) 0.168 L (D) 0.224 L

74 From the opposite table:

The same quantity of electricity was passed through three electrolytic cells containing three different solutions of salts of the metals (X), (Y) and (Z).

What are the valencies of these metals, respectively?

Mass of the element deposited at the cathode	Gram atomic mass of the element	Metal
2.1 g	7 g/atom	(X)
2.7 g	27 g/atom	(Y)
9.6 g	64 g/atom	(Z)

(A) 3 , 1 , 2 (B) 1 , 3 , 2 (C) 3 , 1 , 3 (D) 1 , 3 , 3

75 The quantity of electricity needed to form 1 mol of O_2 gas equals the quantity of electricity needed to form .....

(A) 1 mol of Cl_2 gas (B) 2 mol of F_2 gas

(C) 1 mol of Ca metal (D) 1 mol of Al metal

76 An object was coated with a layer of the metal (X) by electrolysis, using the data shown in the opposite table:

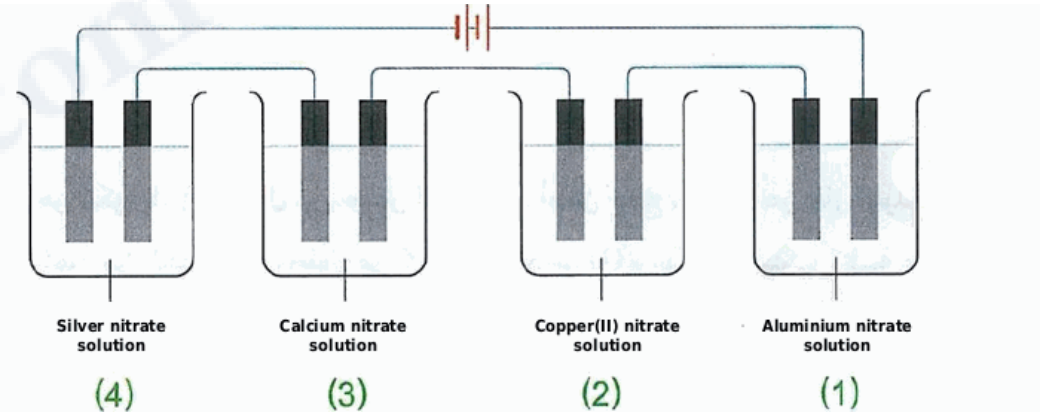
What is the oxidation number of (X) in the electrolyte used?

$0.15\ cm^3$	Volume of metal (X) deposited
$19.3\ g/cm^3$	Density of metal (X)
196.97 g/mol	Molar atomic mass of metal (X)
4 A	Current used
17.75 min	Cell operating time

(A) +1 (B) +2 (C) +3 (D) +4

77 The following figure shows four cells with identical graphite electrodes:

After some time, the increase in the mass of the cathode is greatest in cell ..... [Ag = 108 , Ca = 40 , Cu = 63.5 , Al = 27]



A (1). B (2). C (3). D (4).

Essay questions — 9 questions

1 An electrochemical cell whose overall reaction is given by the equation:

$Cu(s) + Mn^{2+}(aq) \rightarrow Cu^{2+}(aq) + Mn(s)$

Given that:

- $Mn(s) \rightarrow Mn^{2+}(aq) + 2e^{-}$, $E^{\circ} = +1.03$ V
- $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$, $E^{\circ} = +0.34$ V

What type of electrochemical cell is this? Explain.

2 From the opposite figure:

(1) What is the type of each of:

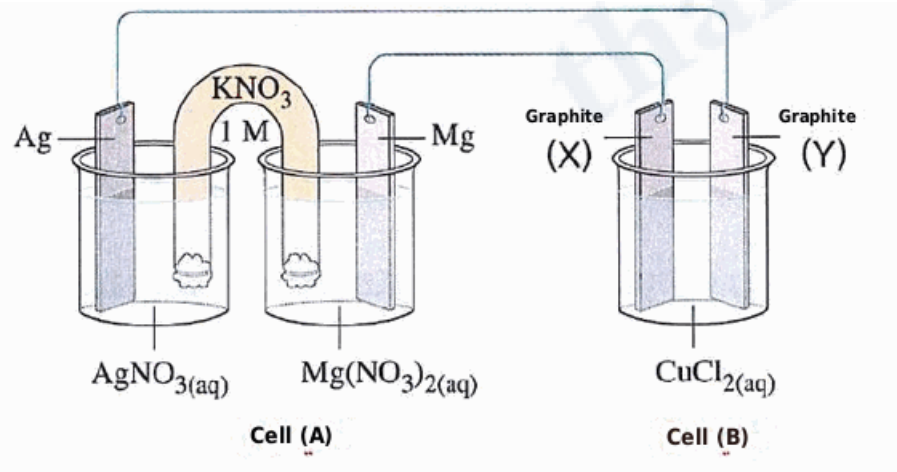
(a) cell (A).

(b) cell (B).

(2) Write the equation of the reaction at:

(a) the graphite electrode (X).

(b) the graphite electrode (Y).



3 The electrolysis of molten strontium chloride $SrCl_2$ produces strontium metal and chlorine gas.

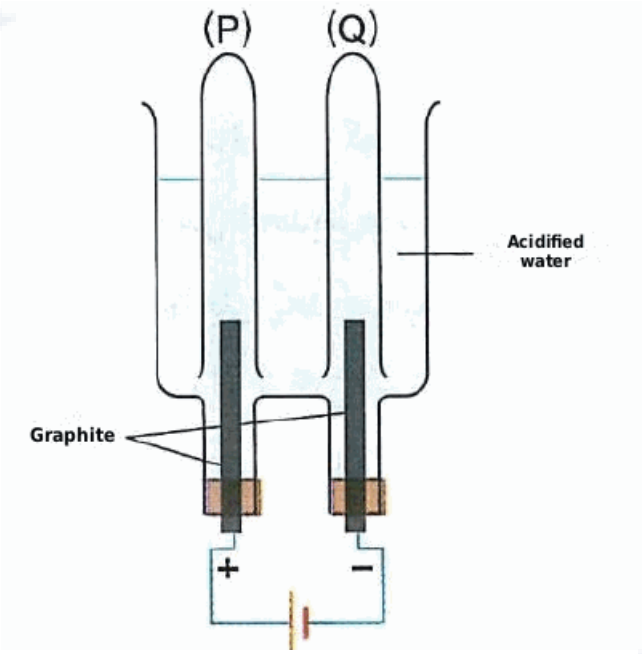
Write the ionic equation for the reaction at each of:

(1) the anode.

(2) the cathode.

4 From the experiment shown in the opposite figure:

What is the ratio between the volumes of the two gases evolved in the tubes (P) and (Q) from the electrolysis of acidified water? Write the ionic equations that support your answer.



5 Fe^{3+} ions can be converted into Fe^{2+} ions and vice versa:

(1) Which of these two processes happens more easily? Explain.

(2) Calculate the quantity of electricity in faradays needed to convert 1 mol of Fe^{3+} into Fe^{2+}

6 A silver nitrate solution was divided between two electrolytic cells. A current of 0.965 A was passed through the first cell for 1000 s, and a quantity of electricity of 9650 C through the second cell:

(1) Calculate the mass of silver [Ag = 108] deposited on the cathode of each cell, and explain.

(2) What scientific rule supports your answer?

7 Two electrolytic cells connected in series; the first contains silver nitrate solution $AgNO_3$ and the second copper sulphate solution $CuSO_4$. After the current passed through them for a certain time, the mass of the cathode in the first cell increased by 1.078 g

By how much did the mass of the cathode of the second cell increase? [Ag = 108 , Cu = 63.5]

8 Calculate the volume of oxygen gas (at STP) produced by passing a quantity of electricity of 5 F through an electrolyte solution according to the anode reaction:

$2O^{2-}(aq) \rightarrow O_2(g) + 4e^-$

9 Give the scientific reason for each of the following:

(1) The gram-equivalent mass of sodium [Na = 23] equals its molar mass, while the gram-equivalent mass of magnesium [Mg = 24] equals half its molar mass.

(2) The quantity of electricity needed to produce 32 g of oxygen gas O_2 by electrolysis is double the quantity needed to produce 2 g of hydrogen gas H_2 [O = 16 , H = 1]

Past exam questions — 35 questions

1 An electric current is passed through copper(II) nitrate solution using a pure copper electrode connected to the positive terminal of the battery and an iron electrode connected to the negative terminal of the battery.

Which of the following is correct while the current passes?

(A) The concentration of nitrate ions in the solution increases and the concentration of copper ions stays constant.

(B) Copper ions are reduced at the cathode and the solution stays electrically neutral.

(C) The concentration of copper ions increases and the concentration of nitrate ions in the solution decreases.

(D) The copper of the anode is oxidized and the total charge of the solution becomes positive.

---

2 Three electrolytic cells connected in series, each containing inert electrodes and the following electrolytes:

- molten Al_2O_3
- molten Mg_3N_2
- molten NaCl

Which of the following gives the ratio between the volumes of the gases evolved from these cells?

(A) Oxygen gas: 2 , nitrogen gas: 3 , chlorine gas: 1

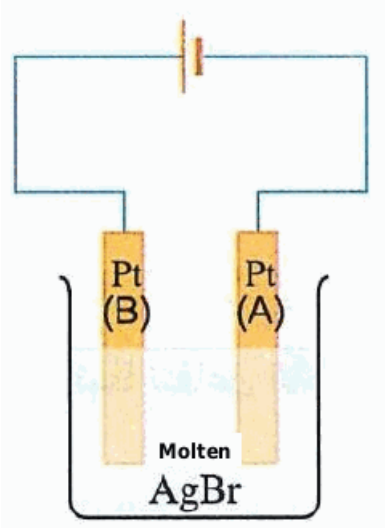
(B) Oxygen gas: 0.166 , nitrogen gas: 0.250 , chlorine gas: 0.330

(C) Oxygen gas: 1 , nitrogen gas: 2 , chlorine gas: 3

(D) Oxygen gas: 0.250 , nitrogen gas: 0.166 , chlorine gas: 0.500

---

3 Which of the following happens in the electrolytic cell in the opposite figure?



(A) Reduction of bromide ions $Br^-(aq)$ at electrode (A).

(B) Oxidation of bromide ions $Br^-(aq)$ at electrode (B).

(C) Reduction of silver ions $Ag^+(aq)$, which are deposited at electrode (B).

(D) Oxidation of silver ions $Ag^+(aq)$, which are deposited at electrode (A).

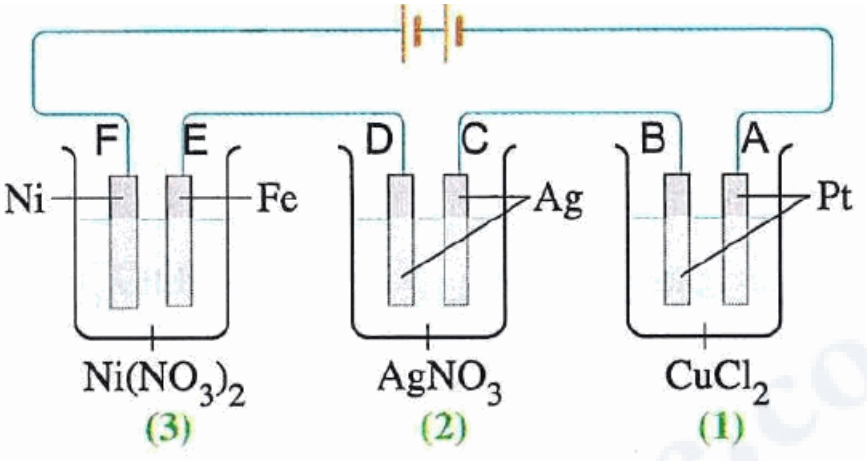
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4 A current of 10 amperes was passed for 10 hours through molten sodium chloride.
Which of the following gives the mass of sodium produced? [Na = 23]

(A) 42.1 g (B) 423.2 g (C) 85.8 g (D) 21.1 g

---

5 The following figure shows three cells connected in series:
When a current passes through the circuit for a quarter of an hour, which electrode increases in mass at the greatest rate? [Ni = 58.7 g/mol , Ag = 108 g/mol , Cu = 63.5 g/mol]



---

6 One faraday is passed through molten sodium chloride.
Which of the following gives the number of moles of gas evolved?

(A) 0.5 mol (B) 1 mol (C) 1.5 mol (D) 2 mol

7 Which of the following is the anode half-reaction during the electrolysis of molten sodium hydride NaH between graphite electrodes?

(A) $2Na^+ + 2e^- \rightarrow 2Na$
(B) $2Na \rightarrow 2Na^+ + 2e^-$
(C) $H_2^- \rightarrow 2H^+ + 2e^-$
(D) $2H^- \rightarrow H_2(g) + 2e^-$

8 Which of the following metals deposits 18 g when 1.5 F passes through one of its molten salts? [Na = 23 , Mg = 24 , Ca = 40 , K = 39]

(A) Na (B) Mg (C) Ca (D) K

9 In the Daniell cell, if the mass of zinc consumed is 6.5 g, the mass of copper deposited is ..... [Zn = 65 , Cu = 63.5]

(A) 6.5 g and it is deposited at the positive pole.
(B) 6.35 g and it is deposited in the anode half-cell.
(C) 6.35 g and it is deposited at the positive pole.
(D) 3.25 g and it is deposited in the cathode half-cell.

10 Which of the following is the reaction at the anode of the cell for the electrolysis of molten potassium bromide?

(A) $2K^+ + 2e^- \rightarrow 2K^0$
(B) $2Br^- \rightarrow Br_2 + 2e^-$
(C) $2K^+ \rightarrow 2K^0 - 2e^-$
(D) $Br_2 + 2e^- \rightarrow 2Br^-$

11 The quantity of electricity needed to deposit 1 mol of Au^{3+} can deposit 1.5 mol of .....

(A) Na^+ (B) Mg^{2+} (C) Cu^+ (D) Al^{3+}

12 Permanganate ions MnO_4^- can be prepared by the electrolysis of a solution containing manganese ions Mn^{2+} according to the opposite reaction:
$Mn^{2+} + 4H_2O \rightarrow MnO_4^- + 8H^+ + 5e^-$
The quantity of electricity needed to prepare 0.2 mol of permanganate ions equals .....

(A) 48250 C (B) 193000 C (C) 19300 C (D) 96500 C

13 Which of the following gives the reactions at the electrodes of a cell for the electrolysis of copper chloride solution using platinum electrodes?

(A) Reaction at the anode: $Cu \rightarrow Cu^{2+} + 2e^-$, reaction at the cathode: $Cl_2 + 2e^- \rightarrow 2Cl^-$

(B) Reaction at the anode: $2Cl^- \rightarrow Cl_2 + 2e^-$, reaction at the cathode: $Cu^{2+} + 2e^- \rightarrow Cu$

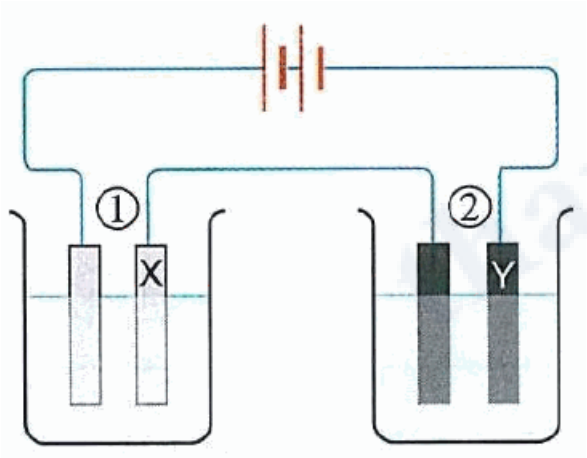
(C) Reaction at the anode: $Cu^{2+} + 2e^- \rightarrow Cu$, reaction at the cathode: $Cl_2 + 2e^- \rightarrow 2Cl^-$

(D) Reaction at the anode: $Cl_2 + 2e^- \rightarrow 2Cl^-$, reaction at the cathode: $Cu \rightarrow Cu^{2+} + 2e^-$

---

14 The opposite figure shows two cells containing different molten salts. After the current passed through the two cells for some time, the mass of the substance formed at the cathode of cell (1) was double that formed at the cathode of cell (2).

Which of the following gives the molten salts in the two cells and the substance formed at the electrodes (X) and (Y)? [Cl = 35.5 , Al = 27 , Au = 195 , Mg = 24 , Zn = 65 , Cr = 52]



(A) Cell (1): $AuCl_3$, gold is deposited at (X) / Cell (2): $ZnCl_2$, chlorine gas is evolved at (Y)

(B) Cell (1): $AuCl_3$, gold is deposited at (X) / Cell (2): $ZnCl_2$, zinc is deposited at (Y)

(C) Cell (1): $CrCl_3$, chlorine gas is evolved at (X) / Cell (2): $AlCl_3$, chlorine gas is evolved at (Y)

(D) Cell (1): $CrCl_3$, chromium is deposited at (X) / Cell (2): $MgCl_2$, chlorine gas is evolved at (Y)

---

15 When a spoon is plated with a layer of gold using a cell containing an electrolyte solution of $AuCl_3$, 0.5 mol of electrons was passed through the solution.

Which of the following describes the changes that happen during the electrolysis? [Au = 197 g/mol]

(A) Increase in the mass of the spoon: 10.94 g , concentration of gold ions in the solution: decreases

(B) Increase in the mass of the spoon: 32.8 g , concentration of gold ions in the solution: increases

(C) Increase in the mass of the spoon: 32.8 g , concentration of gold ions in the solution: stays constant

(D) Increase in the mass of the spoon: 10.94 g , concentration of gold ions in the solution: stays constant

---

16 Which of the following describes what happens in the electrolysis of copper sulphate solution between copper electrodes?

(A) The mass of the anode increases and the concentration of copper ions in the solution does not change.

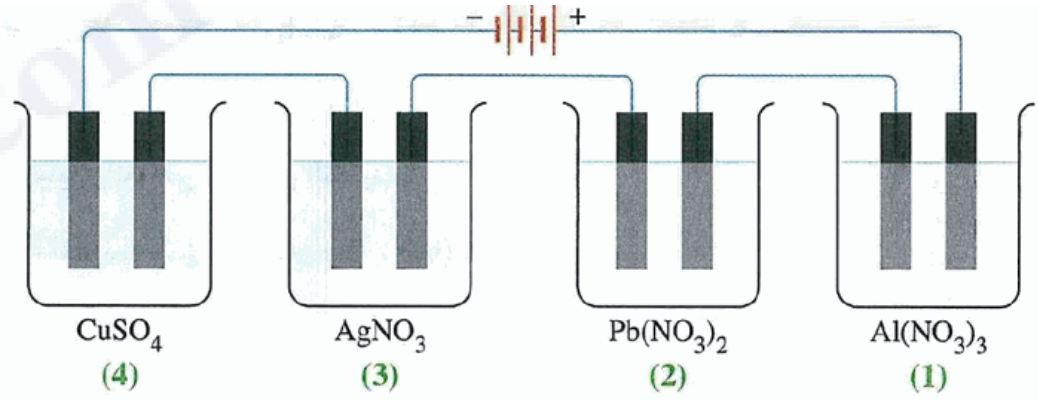
(B) The mass of the cathode decreases and the concentration of copper ions in the solution increases.

(C) The mass of the anode decreases and the concentration of copper ions in the solution decreases.

(D) The mass of the cathode increases and the concentration of copper ions in the solution does not change.

---

17 The following figure shows several electrolytic cells connected in series:
In which of the cells shown does the mass of the cathode increase the most? [Al = 27 , Pb = 208 , Ag = 108 , Cu = 63.5]



(A) Cell (1). (B) Cell (2). (C) Cell (3). (D) Cell (4).

18 When 0.1 F is passed through 5.85 g of pure molten sodium chloride, which of the following results? [Na = 23 , Cl = 35.5]

(A) The melt decomposes partly and 4.6 g of sodium is deposited.
(B) The melt decomposes completely and 2.3 g of sodium is deposited.
(C) The melt decomposes completely and 7.1 g of chlorine gas is evolved.
(D) The melt decomposes partly and 3.55 g of chlorine gas is evolved.

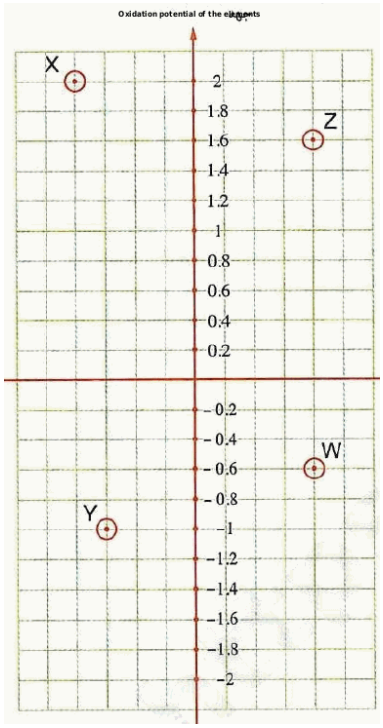
19 Which of the following represents the reaction at the cathode of an electrolytic cell containing molten potassium bromide?

(A) $Br_2 + 2e^- \rightarrow 2Br^-$
(B) $2K \rightarrow 2K^+ + 2e^-$
(C) $2Br^- \rightarrow Br_2 + 2e^-$
(D) $2K^+ + 2e^- \rightarrow 2K$

20 When a current of 5 A is passed for 15 minutes through 5 g of pure molten potassium chloride KCl, which of the following is correct? [K = 39 , Cl = 35.5]

(A) The melt decomposed completely and 1.82 g of potassium was deposited.
(B) The melt decomposed completely and 1.65 g of chlorine was evolved.
(C) 3.47 g of the melt decomposed and 1.65 g of potassium was deposited.
(D) The melt did not decompose completely and 1.53 g remained undecomposed.

21 Four elements (X), (Y), (Z) and (W) whose electrode potentials are shown in the opposite graph. Which of the following choices is correct?



(A) The cell made of the electrodes (W) and (Z) is electrolytic and the element (W) is the cathode.

(B) The cell made of the electrodes (Y) and (Z) is galvanic, gives (emf = 0.6 V), and the element (Z) is the anode.

(C) The cell made of the electrodes (W) and (Y) is electrolytic and the element (Y) is the cathode.

(D) The cell made of the electrodes (W) and (X) is galvanic, gives (emf = 2.6 V), and the element (X) is the anode.

---

22 Given that:

- $X^{2+} + 2e^- \rightarrow X$, $E^\circ = -0.23$ V
- $Y - 2e^- \rightarrow Y^{2+}$, $E^\circ = -0.4$ V

When an electric current is passed between graphite electrodes through a solution containing the chlorides of X^{2+} and Y^{2+} at equal concentrations, which of the following choices is correct?

(A) The mass of the cathode increases because the metal (Y) is deposited.

(B) The mass of the anode increases because the metal (X) is deposited.

(C) Chlorine gas is evolved at the cathode.

(D) The metal (X) is deposited at the anode.

23 When an electric current is passed through molten XCl_4, 33.6 L of chlorine gas at STP is evolved at the anode, so the number of moles of the element (X) deposited at the cathode equals .....

(A) 1.5 mol (B) 0.5 mol (C) 0.75 mol (D) 0.375 mol

24 If the quantity of electricity needed to deposit the equivalent mass of a metal equals the quantity needed to deposit 1 mol of it, which of the following correctly describes this process?

(A) One mole of the metal ion gains one mole of electrons.

(B) One mole of the metal loses one mole of electrons.

(C) One mole of the metal ion gains 2 moles of electrons.

(D) One mole of the metal loses 2 moles of electrons.

25 When 10 g of the element (A) is deposited according to the equation: $A^{2+} + 2e^- \rightarrow A$, the quantity of electricity equals ..... [A = 63.5]

(A) 0.315 F (B) 0.675 C (C) 15196 C (D) 30393 F

26 The quantity of electricity needed to evolve 1.204×10^{23} molecules of oxygen gas in the electrolysis of acidified water is .....

(A) 0.8 F (B) 0.4 F (C) 9650 C (D) 19300 C

27 The quantity of electricity in faradays needed to deposit 0.5 g of gold on a metal medal by electrolysis, according to the equation $Au^{3+} + 3e^- \rightarrow Au^0$, given that [Au = 196.98], equals .....

(A) 2.53×10^{-3} F (B) 7.61 F (C) 7.61×10^{-3} F (D) 2.53 F

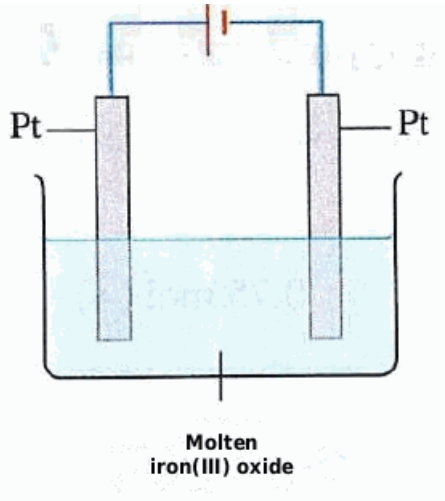
28 When a quantity of electricity of 5000 C passes through an aqueous solution of the chloride of the element (X), 3.4 g of the element (X) is deposited, so its equivalent mass equals .....

(A) 32.8 g (B) 65.6 g (C) 98.4 g (D) 196.9 g

29 When a quantity of electricity passes through molten magnesium nitride, 48 g of magnesium is deposited at the cathode, so the volume of nitrogen gas evolved at the anode (STP) is ..... [given that Mg = 24 , N = 14]

(A) 14.93 L (B) 22.4 L (C) 44.8 L (D) 33.6 L

30 The opposite figure shows an electrolytic cell for molten iron(III) oxide. When a current of 10 A passes for two hours through molten iron(III) oxide, the volume of the gas evolved at the anode (STP) is .....



(A) 8.34 L (B) 16.68 L (C) 12.51 L (D) 4.18 L

---

31 In the opposite reaction taking place in an electric cell:

$Ni^0(s) + 2Ag^+(aq) \rightarrow Ni^{2+}(aq) + 2Ag^0(s)$

Given that:

- $Ni^0(s) \rightarrow Ni^{2+}(aq) + 2e^-$, $E^\circ = +0.23$ V
- $2Ag^+(aq) + 2e^- \rightarrow 2Ag^0(s)$, $E^\circ = +0.8$ V

which of the following is correct?

(A) The cell is electrolytic, emf = −1.03 V (B) The cell is galvanic, emf = 1.03 V

(C) The cell is galvanic, emf = 0.564 V (D) The cell is electrolytic, emf = −0.564 V

---

32 In the cell whose electrodes are iron and tin, given that:

- $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$, $E^\circ = 0.409$ V
- $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$, $E^\circ = 0.150$ V

which of the following is correct?

(A) Iron is the anode, and the emf of the cell is positive.

(B) Iron is the cathode, and the emf of the cell is negative.

(C) Tin is the anode, and the emf of the cell is positive.

(D) Tin is the cathode, and the emf of the cell is negative.

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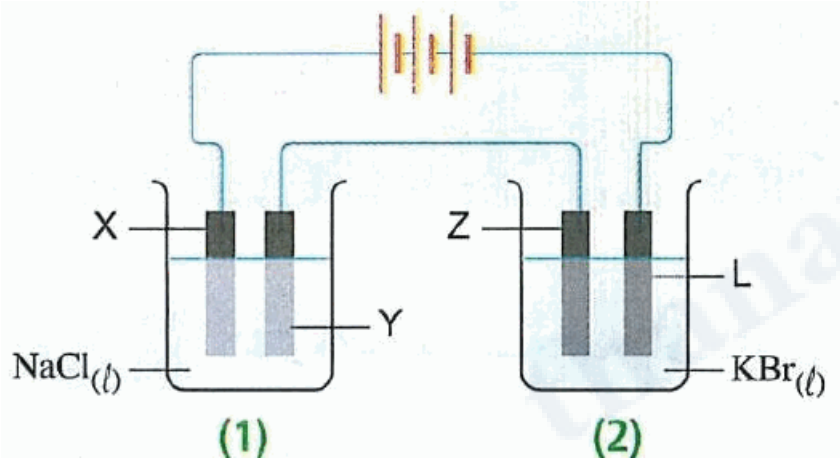
33 The following equations give the two half-cell reactions of an electric cell:

- $2Ni^{3+} + 2e^- \rightarrow 2Ni^{2+}$, $E^\circ = +0.898$ V
- $Cd^{2+} + 2e^- \rightarrow Cd^0$, $E^\circ = -0.402$ V

The non-spontaneous oxidation reaction in the cell is .....

(A) $Cd^0 \rightarrow Cd^{2+} + 2e^-$, $E^\circ = +0.402$ V
(B) $2Ni^{2+} \rightarrow 2Ni^{3+} + 2e^-$, $E^\circ = -0.898$ V
(C) $Cd^{2+} + 2e^- \rightarrow Cd^0$, $E^\circ = -0.402$ V
(D) $2Ni^{3+} + 2e^- \rightarrow 2Ni^{2+}$, $E^\circ = +0.898$ V

34 In the opposite figure: cell (1) contains molten sodium chloride and cell (2) contains molten potassium bromide. When each of them is electrolysed, the substances formed at the electrodes (X , Y , Z , L) are .....



(A) (X): H_2 , (Y): Cl_2 , (Z): Br_2 , (L): K
(B) (X): Cl_2 , (Y): Na , (Z): K , (L): Br_2
(C) (X): Cl_2 , (Y): Na , (Z): Br_2 , (L): K
(D) (X): H_2 , (Y): Cl_2 , (Z): K , (L): Br_2

35 When a quantity of electricity passes through two cells connected in series, the first containing $Pb(NO_3)_2$ solution, 8.28 g of lead is deposited, while in the second cell the reaction $X^{4+} + 3e^- \rightarrow X^+$ takes place.

Calculate the number of moles of X^+ formed [Pb = 207]

Answer key and solutions

Multiple choice questions

1 B	2 D	3 B	4 C	5 D	6 C	7 A	8 D	9 C	10 B
11 A	12 D	13 B	14 D	15 D	16 B	17 A	18 B	19 C	20 D
21 B	22 B	23 A	24 C	25 A	26 B	27 D	28 B	29 B	30 C
31 C	32 A	33 C	34 C	35 D	36 B	37 D	38 C	39 B	40 A
41 B	42 A	43 B	44 D	45 C	46 B	47 D	48 C	49 B	50 D
51 C	52 D	53 B	54 B	55 A	56 C	57 D	58 D	59 D	60 C
61 C	62 A	63 A	64 A	65 D	66 D	67 C	68 A	69 D	70 C
71 C	72 B	73 B	74 B	75 B	76 C	77 D			

1 **Answer (B)**

Electrons move in the external circuit from the anode to the cathode.

Ions move inside the electrolyte: cations towards the cathode and anions towards the anode.

2 **Answer (D)**

Electrons move in the external circuit from the anode to the cathode, not the opposite.

3 **Answer (B)**

The solution is neutral because the sum of positive charges equals the sum of negative charges.

The number of cations may differ from the number of anions (e.g. $Al_2(SO_4)_3$), so the first choice is not correct.

4 **Answer (C)**

Cations (positive ions) always move towards the cathode.

In the galvanic cell the cathode is the positive pole, so cations move to the cathode half-cell.

5 **Answer (D)**

Graphite is an inert electrode, so it is not oxidized.

What is oxidized at the anode is the chloride ions: $2Cl^- \rightarrow Cl_2 + 2e^-$

6 **Answer (C)**

$Cu^{2+} + 2e^- \rightarrow Cu$

$2Cl^- \rightarrow Cl_2 + 2e^-$

For the same quantity of electricity: 63.5 g of copper against 71 g of chlorine.

The mass of copper deposited is less than the mass of chlorine evolved.

7 **Answer (A)**

At the anode chloride ions are oxidized: $2Cl^- \rightarrow Cl_2 + 2e^-$

So chlorine gas is formed.

8 **Answer (D)**

At the cathode the cations are reduced.

$Pb^{2+} + 2e^- \rightarrow Pb$

9 **Answer (C)**

At the cathode hydrogen cations are reduced: $2H^+ + 2e^- \rightarrow H_2$

10 **Answer (B)**

$Ni^{2+} + 2e^- \rightarrow Ni$ is a reduction process.

Reduction takes place at the cathode in both the galvanic cell and the electrolytic cell.

11 **Answer (A)**

Oxidation takes place at the anode in both cells.

12 **Answer (D)**

In the electrolytic cell reduction takes place at the cathode.

The cathode gains electrons from the external source.

13 **Answer (B)**

At the cathode the cation is reduced: $Na^+ + e^- \rightarrow Na$ (the metal).

At the anode: $2I^- \rightarrow I_2 + 2e^-$ (the brown substance).

14 **Answer (D)**

At the cathode: $Na^+ + e^- \rightarrow Na$, so the mass of the cathode increases.

At the anode: $2Br^- \rightarrow Br_2 + 2e^-$ and bromine gas is evolved; platinum is inert, so its mass does not change.

15 **Answer (D)**

At the positive electrode (anode): $2Cl^- \rightarrow Cl_2 + 2e^-$

At the negative electrode (cathode): $Mg^{2+} + 2e^- \rightarrow Mg$

16 **Answer (B)**

The reduction $X^{2+} + 2e^- \rightarrow X$ takes place at the cathode.

The metal X is deposited on the cathode, so its mass increases.

17 **Answer (A)**

At the anode: $2O^{2-} \rightarrow O_2 + 4e^-$ (4 mol of electrons for each 1 mol of O_2 gas)

At the cathode: $4H^+ + 4e^- \rightarrow 2H_2$ (2 mol of electrons for each 1 mol of H_2 gas)

Number of moles of O_2 = X ÷ 4 = 0.25X mol at the anode.

18 **Answer (B)**

At the cathode: $Zn^{2+} + 2e^- \rightarrow Zn$

A layer of zinc forms on the cathode, so its mass increases at a constant rate (constant current).

The mass at the start is not zero because the cathode is a zinc electrode.

19 **Answer (C)**

In the galvanic cell, Mg is the anode (negative pole) and Cu is the cathode (positive pole).

Graphite (2) is connected to the positive pole, so it is the anode: $2Cl^- \rightarrow Cl_2 + 2e^-$

Graphite (1) is connected to the negative pole, so it is the cathode: $Cu^{2+} + 2e^- \rightarrow Cu$

20 **Answer (D)**

The metal A is oxidized and the ion B^+ is reduced.

emf = E°(cathode) − E°(anode) = 0.25 − 0.52 = −0.27 V

Negative emf: the reaction is non-spontaneous, so the cell is electrolytic.

21 **Answer (B)**

Quantity of electricity = current × time

= 0.1 × (30 × 60) = 180 C

22 **Answer (B)**

Quantity of electricity = $2.5 \times (120 \times 60) = 18000$ C

In faradays = $18000 \div 96500 = 0.187$ F

23 **Answer (A)**

Quantity of electricity = $0.36 \times 96500 = 34740$ C

Current = $34740 \div (60 \times 60) = 9.65$ A

24 **Answer (C)**

1 C = $1 \div (1.6 \times 10^{-19}) = 6.24 \times 10^{18}$ electrons

25 **Answer (A)**

$Ca^{2+} + 2e^- \rightarrow Ca$

Number of moles of Ca = $30 \div 40 = 0.75$ mol

Quantity of electricity = $0.75 \times 2 \times 96500 = 144750$ C

Time = $144750 \div 5 = 28950$ s $\approx$ 8 h

26 **Answer (B)**

$Cu^{2+} + 2e^- \rightarrow Cu$

Number of moles = $0.2 \div 63.5 = 0.00315$ mol

Quantity of electricity = $0.00315 \times 2 \times 96500 = 608$ C

Time = $608 \div 1.2 = 507$ s

27 **Answer (D)**

Number of moles of $Ag^+ = 0.1 \times 0.2 = 0.02$ mol

Half the amount = 0.01 mol

$Ag^+ + e^- \rightarrow Ag$: quantity of electricity = $0.01 \times 96500 = 965$ C

Time = $965 \div 0.1 = 9650$ s = 160.8 min

28 **Answer (B)**

Number of moles of $Cu^{2+} = 1 \times 1 = 1$ mol

$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

Only 1 mol of Zn reacts (the rest is in excess), producing 2 mol of electrons.

Quantity of electricity = $2 \times 96500 = 193000$ C

Time = $193000 \div 1 = 193000$ s = 53.6 h

29 **Answer (B)**

Number of moles of Au = 6.57 ÷ 197 = 0.0333 mol

$Au^{3+} + 3e^- \rightarrow Au$

Quantity of electricity = 3 × 0.0333 = 0.1 F

30 **Answer (C)**

96500 C is needed to deposit one gram-equivalent mass of copper.

Number of equivalent masses = 241250 ÷ 96500 = 2.5

31 **Answer (C)**

Quantity of electricity needed to deposit 1 mol = 1.5 ÷ 0.5 = 3 F

So the ion's charge is +3 and the formula is MCl_3.

32 **Answer (A)**

Quantity of electricity = $1 \times (15 \times 60) = 900$ C

Equivalent mass = $(0.173 \times 96500) \div 900 = 18.55$ g

33 **Answer (C)**

Quantity of electricity = $15 \times (50 \times 60) = 45000$ C

Equivalent mass = $(9.35 \times 96500) \div 45000 = 20$ g

Atomic mass = 20 × 2 (divalent) = 40 g/atom

34 **Answer (C)**

The amount that gains or loses 1 mol of electrons is one equivalent mass.

20 g of Ca^{2+} = 0.5 mol, which gains 1 mol of electrons.

23 g of Na^+ = 1 mol, which gains 1 mol of electrons.

35.5 g of Cl^- = 1 mol, which loses 1 mol of electrons.

16 g of O^{2-} = 1 mol, which loses 2 mol of electrons: this is the exception.

35 **Answer (D)**

1 mol of electrons deposits one equivalent mass.

Zn : 65 ÷ 2 = 32.5 g

Mg : 24 ÷ 2 = 12 g

Na : 23 g

Al : 27 ÷ 3 = 9 g

36 **Answer (B)**

The positive electrode is the anode: $2Br^- \rightarrow Br_2 + 2e^-$

0.5 F gives 0.5 mol of bromine atoms.

Mass = $0.5 \times 80 = 40$ g

37 **Answer (D)**

Equivalent mass of copper = $63.5 \div 2 = 31.75$ g

Quantity of electricity = X × t in coulombs.

Mass of copper = $(31.75 \times X \times t) \div 96500$

38 **Answer (C)**

$Cu^{2+} + 2e^- \rightarrow Cu$

Equivalent mass = $63.5 \div 2 = 31.75$ g for each 1 F

Mass = $0.5 \times 31.75 = 15.875$ g

39 **Answer (B)**

Number of moles of Al = $13.5 \div 27 = 0.5$ mol

$Al^{3+} + 3e^- \rightarrow Al$

Quantity of electricity = $0.5 \times 3 = 1.5$ F

40 **Answer (A)**

Number of moles of K = $39 \div 39 = 1$ mol

$K^+ + e^- \rightarrow K$

Quantity of electricity = 1 F = 96500 C

41 **Answer (B)**

$Mg^{2+} + 2e^- \rightarrow Mg$

2 F are needed to separate 1 mol of magnesium.

42 **Answer (A)**

$ClO_3^- \rightarrow ClO_4^-$: the oxidation number of Cl increases from +5 to +7.

2 mol of electrons are lost for each 1 mol of $NaClO_4$.

0.5 mol needs $0.5 \times 2 = 1$ F

43 **Answer (B)**

Number of moles of Cl_2 = 355 ÷ 71 = 5 mol

$2Cl^- \rightarrow Cl_2 + 2e^-$: quantity of electricity = 5 × 2 = 10 F

= 10 × 96500 = 9.65×10^5 C

44 **Answer (D)**

Number of moles of NaCl = 4 × 0.5 = 2 mol $(2 \text{ mol of } Cl^-)$

Each 1 mol of Cl^- loses 1 mol of electrons.

Quantity of electricity = 2 F = 193000 C

45 **Answer (C)**

Time = 32 × 60 + 10 = 1930 s

Quantity of electricity = 0.5 × 1930 = 965 C, which gives 0.44 g

To deposit 88 g: $(965 \times 88) \div 0.44$ = 193000 C = 2 F

46 **Answer (B)**

$Cu^{2+} + 2e^- \rightarrow Cu$

4 F deposit 4 ÷ 2 = 2 mol of copper.

47 **Answer (D)**

$Li \rightarrow Li^+ + e^-$

Number of moles of Li = 0.347 ÷ 6.94 = 0.05 mol

Number of moles of electrons = 0.05 mol

48 **Answer (C)**

Quantity of electricity = $1 \times (30 \times 60)$ = 1800 C = 0.01865 F

$2Cl^- \rightarrow Cl_2 + 2e^-$: number of moles of Cl_2 = 0.01865 ÷ 2 = 0.00933 mol

Mass = 0.00933 × 71 = 0.66 g

49 **Answer (B)**

Quantity of electricity = $8 \times (2 \times 3600)$ = 57600 C = 0.597 F

$Cr^{6+} + 6e^- \rightarrow Cr$

Number of moles of Cr = 0.597 ÷ 6 = 0.1 mol

50 **Answer (D)**

The two cells are in series: the quantity of electricity is the same.

$Fe^{2+} + 2e^- \rightarrow Fe$: $14 \div 56 = 0.25$ mol, so the electricity is 0.5 F

Number of moles of Mn = $6.875 \div 55 = 0.125$ mol

Number of electrons per ion = $0.5 \div 0.125 = 4$

The charge of the manganese ion is +4.

51 **Answer (C)**

The metal used in galvanization is zinc: $Zn^{2+} + 2e^- \rightarrow Zn$

Number of moles of electrons = $9.65 \div 96500 = 1 \times 10^{-4}$ mol

Number of moles of Zn = $\left(1 \times 10^{-4}\right) \div 2 = 5 \times 10^{-5}$ mol

52 **Answer (D)**

Quantity of electricity = $4 \times \left(30 \times 60\right) = 7200$ C

$2H^+ + 2e^- \rightarrow H_2$: number of moles of $H_2 = 7200 \div \left(2 \times 96500\right) = 0.0373$ mol

Volume = $0.0373 \times 22.4 = 0.836$ L

53 **Answer (B)**

96500 C = 1 F

$2O^{2-} \rightarrow O_2 + 4e^-$: 4 F give 1 mol

1 F gives 0.25 mol : volume = $0.25 \times 22.4 = 5.6$ L

54 **Answer (B)**

$2I^- \rightarrow I_2 + 2e^-$

Number of moles of $I_2 = 10 \div 254 = 0.0394$ mol

Quantity of electricity = $0.0394 \times 2 \times 96500 = 7598$ C

Current = $7598 \div 3600 = 2.1$ A

55 **Answer (A)**

Number of moles of H_2 per second = $1.12 \div 22400 = 5 \times 10^{-5}$ mol

$2H^+ + 2e^- \rightarrow H_2$: quantity of electricity = $5 \times 10^{-5} \times 2 \times 96500 = 9.65$ C

Current = $9.65 \div 1 = 9.65$ A

56 **Answer (C)**

Number of moles of $O_2 = 0.5 \div 22.4 = 0.0223$ mol

$2O^{2-} \rightarrow O_2 + 4e^-$: quantity of electricity = $0.0223 \times 4 \times 96500 = 8616$ C

Current = $8616 \div \left(3 \times 3600\right) = 0.8$ A

57 **Answer (D)**

Quantity of electricity = $10 \times 965 = 9650$ C = 0.1 F

$2H^+ + 2e^- \rightarrow H_2$: 0.1 mol of H^+ is consumed.

Remaining H^+ concentration = $1 - 0.1 = 0.9$ M

pH = $-\log 0.9 = 0.0458$

58 **Answer (D)**

0.5 F reduces 0.25 mol of Cu^{2+} ($Cu^{2+} + 2e^- \rightarrow Cu$)

Remaining $Cu^{2+} = 2 - 0.25 = 1.75$ mol in 1 L

Concentration = 1.75 M

59 **Answer (D)**

In both cases: $2H^+ + 2e^- \rightarrow H_2$

The same quantity of electricity gives the same amount of hydrogen.

60 **Answer (C)**

$Zn^{2+} + 2e^- \rightarrow Zn$: $0.3 \div 2 = 0.15$ mol

$Sc^{3+} + 3e^- \rightarrow Sc$: $0.3 \div 3 = 0.1$ mol

61 **Answer (C)**

Number of moles of $H_2 = 0.504 \div 2 = 0.252$ mol, which needs 0.504 F

$2O^{2-} \rightarrow O_2 + 4e^-$: number of moles of $O_2 = 0.504 \div 4 = 0.126$ mol

Mass = $0.126 \times 32 = 4.03$ g, i.e. about 4 g

62 **Answer (A)**

Zn is the anode and loses mass; its new mass is 4.75 g (a loss of 0.25 g).

Number of moles of Zn oxidized = $0.25 \div 65.37 = 0.00382$ mol

$Cd^{2+} + 2e^- \rightarrow Cd$: the mass of Cd increases by $0.00382 \times 112.4 = 0.43$ g

63 **Answer (A)**

Copper: $18.25 \div 63.5 = 0.2874$ mol $\times 2 = 0.575$ F

Cerium: $80 \div 140 = 0.571$ mol

Each ion gains one electron $(0.575 \div 0.571 \approx 1)$, so it is Ce^+ and the salt is $CeNO_3$.

64 **Answer (A)**

The cells are in series: the quantity of electricity is the same.

The mass is proportional to the equivalent mass:

Cu : $63.5 \div 2 = 31.75$, Au : $197 \div 3 = 65.7$, Ag : $108 \div 1 = 108$

The order: Cu < Au < Ag

65 **Answer (D)**

$Ti^{4+} + 4e^- \rightarrow Ti$: number of moles of Ti = $6 \div 48 = 0.125$ mol

Quantity of electricity = $0.125 \times 4 = 0.5$ F

$Ag^+ + e^- \rightarrow Ag$: mass = $0.5 \times 108 = 54$ g

66 **Answer (D)**

$Ag^+ + e^- \rightarrow Ag$: 1 F deposits 1 mol

$Cu^{2+} + 2e^- \rightarrow Cu$: 1 F deposits 0.5 mol

The number of moles of silver is double the number of moles of copper.

67 **Answer (C)**

The two cells are in series: the quantity of electricity is the same.

Ratio = equivalent mass of silver ÷ equivalent mass of zinc

$= 108 \div (65 \div 2) = 108 \div 32.5$

68 **Answer (A)**

Equivalent mass: Al = $27 \div 3 = 9$ g , Ag = 108 g

Ratio of the masses of Al and Ag = 9 : 108 = 1 : 12

1 g of Al , 12 g of Ag

69 **Answer (D)**

Mg : $4.8 \div 24 = 0.2$ mol $\times 2 = 0.4$ F

Ti : $4.8 \div 48 = 0.1$ mol, which gains 0.4 F

Each Ti atom gains 4 electrons, so it is Ti^{4+} and the formula is $TiCl_4$.

70 **Answer (C)**

$2Cl^- \rightarrow Cl_2 + 2e^-$: 2 mol of electrons for each 1 mol of gas.

$2O^{2-} \rightarrow O_2 + 4e^-$: 4 mol of electrons for each 1 mol of gas.

The same volume means the same number of moles, so double the quantity of electricity is needed: 2a A

71 **Answer (C)**

Number of moles of H_2 = 2.24 ÷ 22.4 = 0.1 mol, which needs 0.2 F

$Cu^{2+} + 2e^- \rightarrow Cu$: number of moles of Cu = 0.2 ÷ 2 = 0.1 mol

Mass = 0.1 × 63.5 = 6.35 g

72 **Answer (B)**

(X): $Cu^{2+} + 2e^- \rightarrow Cu$: 1 F deposits 0.5 mol

(Y): $Cu^{+} + e^- \rightarrow Cu$: 1 F deposits 1 mol

The amount of copper deposited in cell (Y) is double.

73 **Answer (B)**

$Ag^{+} + e^- \rightarrow Ag$: 1.08 ÷ 108 = 0.01 mol, so the quantity of electricity is 0.01 F

$2O^{2-} \rightarrow O_2 + 4e^-$: number of moles of O_2 = 0.01 ÷ 4 = 0.0025 mol

Volume = 0.0025 × 22.4 = 0.056 L

74 **Answer (B)**

The quantity of electricity is the same in the three cells.

Number of moles of X = 2.1 ÷ 7 = 0.3 mol , Y = 2.7 ÷ 27 = 0.1 mol , Z = 9.6 ÷ 64 = 0.15 mol

The valency is inversely proportional to the number of moles: 1 ÷ 0.3 : 1 ÷ 0.1 : 1 ÷ 0.15 = 3.33 : 10 : 6.67

So the valencies are 1 , 3 , 2

75 **Answer (B)**

$2O^{2-} \rightarrow O_2 + 4e^-$: 4 F are needed to form 1 mol of O_2.

$2F^- \rightarrow F_2 + 2e^-$: 2 mol of F_2 need 2 × 2 = 4 F

76 **Answer (C)**

Mass of the metal = 19.3 × 0.15 = 2.895 g

Quantity of electricity = $4 \times (17.75 \times 60) = 4260$ C

Equivalent mass = $(2.895 \times 96500) \div 4260 = 65.58$ g

Oxidation number = 196.97 ÷ 65.58 = 3

77 **Answer (D)**

The cells are in series: the quantity of electricity is the same.

(1) $Al(NO_3)_3$ solution: hydrogen is evolved at the cathode and aluminium is not deposited.

(3) Calcium nitrate solution: hydrogen is also evolved.

(2) Cu^{2+} : equivalent mass 31.75 g , (4) Ag^{+} : equivalent mass 108 g

The greatest increase in the mass of the cathode is in cell (4).

Essay questions

1 **Model answer**

An electrolytic cell, because copper is oxidized (anode) and the manganese(II) ion is reduced (cathode).

emf = E°(cathode) − E°(anode) = (−1.03) − (+0.34) = −1.37 V

The emf is negative, so the reaction is non-spontaneous and needs an external current source.

2 **Model answer**

(1) (a) Cell (A): a galvanic cell.

(b) Cell (B): an electrolytic cell.

(2) (a) Graphite electrode (X): $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$

(b) Graphite electrode (Y): $2Cl^{-}(aq) \rightarrow Cl_2(g) + 2e^{-}$

3 **Model answer**

(1) Anode: $2Cl^{-}(l) \rightarrow Cl_2(g) + 2e^{-}$

(2) Cathode: $Sr^{2+}(l) + 2e^{-} \rightarrow Sr(s)$

4 **Model answer**

Tube (P) at the positive electrode (anode): $2O^{2-} \rightarrow O_2 + 4e^{-}$, so oxygen gas is evolved.

Tube (Q) at the negative electrode (cathode): $4H^{+} + 4e^{-} \rightarrow 2H_2$, so hydrogen gas is evolved.

Volume of O_2 gas : volume of H_2 gas = 1 : 2 (i.e. volume of oxygen gas (P) : volume of hydrogen gas (Q) = 1 : 2)

5 **Model answer**

Electron configuration: Fe : [Ar] $3d^6\ 4s^2$, Fe^{2+} : [Ar] $3d^6$, Fe^{3+} : [Ar] $3d^5$

(1) The conversion of Fe^{2+} into Fe^{3+} happens more easily, because the 3d sublevel in the Fe^{3+} ion is half-filled ($3d^5$), so it is more stable.

(2) $Fe^{3+} + e^{-} \rightarrow Fe^{2+}$

Quantity of electricity needed to convert 1 mol = 1 F

6 **Model answer**

(1) First cell: quantity of electricity = 0.965 × 1000 = 965 C , mass of silver deposited = 1.08 g

Second cell: quantity of electricity = 9650 C , mass of silver deposited = 10.8 g

Explanation: the greater the quantity of electricity passing through the cell, the greater the mass of the substance deposited at the electrode.

(2) Faraday's first law: the mass of the substance deposited or evolved at an electrode is directly proportional to the quantity of electricity passing through the electrolyte.

7 **Model answer**

Equivalent mass of silver = 108 g , equivalent mass of copper = 63.5 ÷ 2 = 31.75 g

The two cells are in series, so the quantity of electricity is the same:

Mass of silver ÷ mass of copper = equivalent mass of silver ÷ equivalent mass of copper

1.078 ÷ mass of copper = 108 ÷ 31.75

Mass of copper = (1.078 × 31.75) ÷ 108 = 0.317 g

Increase in the mass of the cathode of the second cell = 0.317 g

8 **Model answer**

$2O^{2-} \rightarrow O_2 + 4e^-$

Every 4 F produce 1 mol of O_2, so 5 F produce 5 ÷ 4 = 1.25 mol

Volume = 1.25 × 22.4 = 28 L

9 **Model answer**

(1) Because the oxidation number of sodium is +1 in all its compounds and that of magnesium is +2, and equivalent mass = atomic mass ÷ oxidation number.

Sodium: 23 ÷ 1 = 23 g , which equals its molar mass.

Magnesium: 24 ÷ 2 = 12 g , which equals half its molar mass.

(2) 2 g of H_2 = 1 mol, and producing it needs 2 F ($2H^+ + 2e^- \rightarrow H_2$)

32 g of O_2 = 1 mol, and producing it needs 4 F ($2O^{2-} \rightarrow O_2 + 4e^-$)

So the quantity of electricity needed to produce oxygen is double that needed to produce hydrogen.

Past exam questions

1	2	3	4	5	6	7	8	9	10
B	D	B	C	B	A	D	B	C	B
11	12	13	14	15	16	17	18	19	20
B	D	B	B	C	D	C	B	D	D

21	22	23	24	25	26	27	28	29	30
D	A	C	A	A	A	C	B	A	D

31	32	33	34	35
B	A	B	C	—

1 **Answer (B)**

At the anode (copper): $Cu \rightarrow Cu^{2+} + 2e^-$

At the cathode (iron): $Cu^{2+} + 2e^- \rightarrow Cu$

Copper ions are reduced at the cathode and the solution stays neutral, and the concentration of copper ions stays constant.

2 **Answer (D)**

The cells are in series: the quantity of electricity is the same $(1\ F)$.

$Al_2O_3 : 2O^{2-} \rightarrow O_2 + 4e^-$, producing 0.25 mol of O_2 gas

$Mg_3N_2 : 2N^{3-} \rightarrow N_2 + 6e^-$, producing 0.166 mol of N_2 gas

$NaCl : 2Cl^- \rightarrow Cl_2 + 2e^-$, producing 0.5 mol of Cl_2 gas

Oxygen : nitrogen : chlorine = 0.250 : 0.166 : 0.500

3 **Answer (B)**

Electrode (B) is connected to the positive terminal of the battery (the long line), so it is the anode.

At the anode the bromide ions are oxidized: $2Br^- \rightarrow Br_2 + 2e^-$

4 **Answer (C)**

Quantity of electricity = $10 \times (10 \times 3600) = 360000$ C

Number of faradays = $360000 \div 96500 = 3.73$ F

$Na^+ + e^- \rightarrow Na$: mass = $3.73 \times 23 = 85.8$ g

5 **Answer (B)**

The cells are in series: the quantity of electricity is the same.

The mass of the cathode in each cell increases in proportion to the equivalent mass of the metal deposited:

Cu : $63.5 \div 2 = 31.75$, Ag : 108 , Ni : $58.7 \div 2 = 29.35$

The greatest increase is in cell (2) at its cathode (C).

6 **Answer (A)**

$Na^+ + e^- \rightarrow Na$ (a solid metal, not a gas)

$2Cl^- \rightarrow Cl_2 + 2e^-$: 1 F gives 0.5 mol of Cl_2 gas.

7 **Answer (D)**

Oxidation takes place at the anode.

The hydride ion is oxidized: $2H^- \rightarrow H_2(g) + 2e^-$

8 **Answer (B)**

Equivalent mass = $(18 \times 1) \div 1.5 = 12$ g

Mg : $24 \div 2 = 12$ g

9 **Answer (C)**

Number of moles of Zn = $6.5 \div 65 = 0.1$ mol, giving 0.2 mol of electrons.

$Cu^{2+} + 2e^- \rightarrow Cu$: 0.1 mol = $0.1 \times 63.5 = 6.35$ g

Copper is deposited at the cathode, which is the positive pole in the Daniell cell.

10 **Answer (B)**

Oxidation takes place at the anode of the electrolytic cell.

Bromide ions are oxidized: $2Br^- \rightarrow Br_2 + 2e^-$

11 **Answer (B)**

Au^{3+} : 1 mol needs 3 F.

1.5 mol of the other ion also needs 3 F, so each 1 mol needs 2 F.

So the ion's charge is +2, and it is Mg^{2+}.

12 **Answer (D)**

$Mn^{2+} + 4H_2O \rightarrow MnO_4^- + 8H^+ + 5e^-$

0.2 mol needs $0.2 \times 5 = 1$ mol of electrons = 1 F = 96500 C

13 **Answer (B)**

At the anode chloride ions are oxidized: $2Cl^- \rightarrow Cl_2 + 2e^-$

At the cathode copper ions are reduced: $Cu^{2+} + 2e^- \rightarrow Cu$

Platinum electrodes are inert, so they do not take part in the reaction.

14 **Answer (B)**

The two cells are in series: the quantity of electricity is the same.

Mass of gold ÷ mass of zinc = $(195 \div 3) \div (65 \div 2) = 65 \div 32.5 = 2$

The two metals are deposited at the cathodes (X) and (Y), and the mass in the first cell is double that in the second.

15 **Answer (C)**

$Au^{3+} + 3e^- \rightarrow Au$

Number of moles of Au = $0.5 \div 3 = 0.1667$ mol

Increase in the mass of the spoon = $0.1667 \times 197 = 32.8$ g

In gold plating the anode is made of gold, so it dissolves by the same amount and the ion concentration stays constant.

16 **Answer (D)**

At the cathode: $Cu^{2+} + 2e^- \rightarrow Cu$, so its mass increases.

At the anode: $Cu \rightarrow Cu^{2+} + 2e^-$, so its mass decreases.

The Cu^{2+} consumed at the cathode is replaced by that formed at the anode, so the concentration stays constant.

17 **Answer (C)**

The cells are in series: the quantity of electricity is the same.

$Al(NO_3)_3$ in water: hydrogen is evolved and aluminium is not deposited.

Equivalent mass: Pb = $208 \div 2 = 104$, Ag = 108 , Cu = $63.5 \div 2 = 31.75$

The greatest increase in the mass of the cathode is in cell (3) (silver).

18 **Answer (B)**

Number of moles of NaCl = $5.85 \div 58.5 = 0.1$ mol

0.1 F is enough to decompose all 0.1 mol ($Na^+ + e^- \rightarrow Na$).

Mass of Na deposited = $0.1 \times 23 = 2.3$ g

19 **Answer (D)**

Reduction takes place at the cathode: $K^+ + e^- \rightarrow K$

i.e. $2K^+ + 2e^- \rightarrow 2K$

20 **Answer (D)**

Quantity of electricity = $5 \times (15 \times 60) = 4500$ C = 0.0466 F

$K^+ + e^- \rightarrow K$: mass of K = $0.0466 \times 39 = 1.82$ g

Mass of KCl decomposed = $0.0466 \times 74.5 = 3.47$ g

Remaining melt = $5 - 3.47 = 1.53$ g : it did not decompose completely.

21 **Answer (D)**

The vertical axis shows the oxidation potential.

Cell (W) and (X): emf = $2.0 - (-0.6) = 2.6$ V (positive, so it is galvanic).

The element (X) has the greatest oxidation potential, so it is the anode.

22 **Answer (A)**

Reduction potential of $X^{2+} = -0.23$ V

Reduction potential of $Y^{2+} = +0.4$ V (the reverse of the oxidation potential -0.4 V)

Y^{2+} is easier to reduce, so it is reduced first at the cathode and the metal (Y) is deposited, so the mass of the cathode increases.

23 **Answer (C)**

Number of moles of $Cl_2 = 33.6 \div 22.4 = 1.5$ mol

$2Cl^- \rightarrow Cl_2 + 2e^-$: quantity of electricity $= 1.5 \times 2 = 3$ F

$X^{4+} + 4e^- \rightarrow X$: number of moles of X $= 3 \div 4 = 0.75$ mol

24 **Answer (A)**

The equivalent mass needs 1 F to be deposited.

Depositing 1 mol also needs 1 F, so n = 1:

1 mol of the metal ion gains 1 mol of electrons ($M^+ + e^- \rightarrow M$).

25 **Answer (A)**

Number of moles of A $= 10 \div 63.5 = 0.1575$ mol

$A^{2+} + 2e^- \rightarrow A$: quantity of electricity $= 0.1575 \times 2 = 0.315$ F

26 **Answer (A)**

Number of moles of $O_2 = (1.204 \times 10^{23}) \div (6.02 \times 10^{23}) = 0.2$ mol

$2O^{2-} \rightarrow O_2 + 4e^-$: quantity of electricity $= 0.2 \times 4 = 0.8$ F

27 **Answer (C)**

Number of moles of Au $= 0.5 \div 196.98 = 0.002538$ mol

$Au^{3+} + 3e^- \rightarrow Au$: quantity of electricity $= 0.002538 \times 3 = 7.61 \times 10^{-3}$ F

28 **Answer (B)**

Equivalent mass $= (3.4 \times 96500) \div 5000 = 65.6$ g

29 **Answer (A)**

Number of moles of Mg = 48 ÷ 24 = 2 mol, which needs 4 F

$2N^{3-} \rightarrow N_2 + 6e^-$: number of moles of N_2 = 4 ÷ 6 = 0.667 mol

Volume = 0.667 × 22.4 = 14.93 L

30 **Answer (D)**

Quantity of electricity = $10 \times (2 \times 3600)$ = 72000 C = 0.746 F

$2O^{2-} \rightarrow O_2 + 4e^-$: number of moles of O_2 = 0.746 ÷ 4 = 0.1865 mol

Volume = 0.1865 × 22.4 = 4.18 L

31 **Answer (B)**

$Ni \rightarrow Ni^{2+} + 2e^-$ (E° = +0.23 V) , $2Ag^+ + 2e^- \rightarrow 2Ag$ (E° = +0.8 V)

emf = 0.23 + 0.8 = 1.03 V

The value is positive, so the reaction is spontaneous and the cell is galvanic.

32 **Answer (A)**

$Fe \rightarrow Fe^{2+} + 2e^-$ (oxidation potential 0.409 V) , $Sn^{2+} + 2e^- \rightarrow Sn$ (reduction potential 0.150 V)

emf = 0.409 + 0.150 = 0.559 V , positive.

Iron is oxidized, so it is the anode.

33 **Answer (B)**

The reduction potential of Ni^{3+} (+0.898 V) is greater than that of Cd^{2+} (−0.402 V).

Spontaneous: Ni^{3+} is reduced and Cd is oxidized.

The non-spontaneous oxidation is the reverse one: $2Ni^{2+} \rightarrow 2Ni^{3+} + 2e^-$, E° = −0.898 V

34 **Answer (C)**

Anode: the anions are oxidized, so Cl_2 forms at (X) and Br_2 at (Z).

Cathode: the cations are reduced, so Na forms at (Y) and K at (L).

35 **Model answer**

$Pb^{2+} + 2e^- \rightarrow Pb$: 2 F are needed to deposit 207 g

Quantity of electricity = $(2 \times 8.28) \div 207$ = 0.08 F

$X^{4+} + 3e^- \rightarrow X^+$: 3 F are needed to form 1 mol

Number of moles of X^+ = $(1 \times 0.08) \div 3$ = 0.0267 mol