Chemistry · Electrochemistry

# Galvanic Cells

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## Multiple choice questions — 72 questions

**1** From the opposite reactions:

(1) $BaCl_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$

(2) $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

(3) $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

(4) $Fe_2O_3(s) + 2Al(s) \rightarrow Al_2O_3(s) + 2Fe(s)$

Which of the following are two redox (oxidation-reduction) reactions?

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

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**2** Which of the following equations represents a redox reaction?

(A) $Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

(B) $CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$

(C) $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

(D) $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

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**3** Which of the following redox equations is balanced?

(A) $Cr(s) + 3Fe^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + Fe(s)$

(B) $Pb(s) + 2H^+(aq) \rightarrow Pb^{2+}(aq) + H_2(g)$

(C) $Pb(s) + Ag^+(aq) \rightarrow Pb^{2+}(aq) + Ag(s)$

(D) $Zn(s) + Br_2(aq) \rightarrow Zn^{2+}(aq) + Br^-(aq)$

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**4** In the reaction: $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$ the correct oxidation half-reaction is .....

(A) $Mg + 2e^- \rightarrow Mg^{2+}$ (B) $Cl_2 + 2e^- \rightarrow 2Cl^-$

(C) $Mg \rightarrow Mg^{2+} + 2e^-$ (D) $Cl_2 \rightarrow 2Cl^- + 2e^-$

---

**5** Oxidation and reduction take place when iron(II) sulphate solution is added to .....

(A) sodium hydroxide solution.

(B) acidified potassium permanganate solution.

(C) silver nitrate solution.

(D) ammonia solution.

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**6** In the redox reaction: $Al + MnO_4^- + 2H_2O \rightarrow Al(OH)_4^- + MnO_2$

oxidation takes place for .....

(A) Al (B) $MnO_4^-$ (C) $Al(OH)_4^-$ (D) $MnO_2$

**7** In the reaction: $Pb(s) + Cu^{2+}(aq) \rightarrow Pb^{2+}(aq) + Cu(s)$ the reducing agent is .....

(A) $Pb^{2+}(aq)$ (B) $Cu^{2+}(aq)$ (C) Pb(s) (D) Cu(s)

**8** In the reaction: $2Cs(s) + Cl_2(g) \rightarrow 2CsCl(s)$

Which of the following describes the role of $Cl_2$ in this reaction?

(A) Reducing agent. (B) Oxidizing agent. (C) Electron donor. (D) Catalyst.

**9** Which of the following equations represents a reduction that forms a stable ion?

(A) $Cr^{2+} \rightarrow Cr^{3+} + e^-$ (B) $V^{2+} \rightarrow V^{3+} + e^-$

(C) $Mn^{3+} + e^- \rightarrow Mn^{2+}$ (D) $Ti^{4+} + e^- \rightarrow Ti^{3+}$

**10** The oxidation number of a reducing agent can change from .....

(A) -1 → -3 (B) -2 → -1 (C) 3 → -1 (D) 4 → -3

**11** The following equation represents a redox reaction: $Mn^{2+} + IO_3^- + H_2O \rightarrow MnO_4^- + \frac{1}{2}I_2 + 2H^+$

The reducing agent in it is .....

(A) $I_2$ (B) $IO_3^-$ (C) $H_2O$ (D) $Mn^{2+}$

**12** From the ionic equation: $xFe^{2+}(aq) + Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow xFe^{3+}(aq) + 2Cr^{3+}(aq) + 7H_2O(l)$

What is the value of the coefficient (x)?

(A) 6 (B) 4 (C) 3 (D) 2

**13** How many moles of electrons are needed to decolourize 1 mol of permanganate ion in an acidic medium?

(A) 5 (B) 4 (C) 3 (D) 2

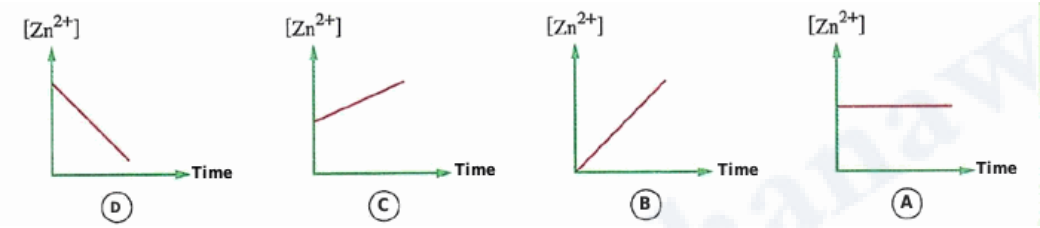
**14** Which of the following describes the change in the oxidation numbers of iodine and sulphur in the reaction of hydrogen iodide with concentrated sulphuric acid?

(A) Iodine: -1 → 0 , sulphur: +6 → +2
(B) Iodine: +1 → 0 , sulphur: -6 → -2
(C) Iodine: +1 → 0 , sulphur: -6 → +4
(D) Iodine: -1 → 0 , sulphur: +6 → +4

**15** Passing $CO_2$ gas over hot coal produces compound (X).
Which of the following reactions cannot take place?

(A) $X + A_2(SO_4)_3 \rightarrow ASO_4 + Y$
(B) $X + ZO_4^- \rightarrow ZO_2 + Y$
(C) $X + A(NO_3)_3 \rightarrow A(NO_3)_2 + Y$
(D) $X + ZSO_3 \rightarrow ZSO_4 + Y$

**16** A zinc rod is placed in copper(II) sulphate solution.
Which of the following graphs represents $[Zn^{2+}]$ with time?




**17** A scandium rod is placed in iron(III) chloride solution.
Which of the following is incorrect?

(A) Oxidation and reduction take place.
(B) $[Fe^{3+}]$ decreases and $[Cl^-]$ does not change.
(C) The mass of the consumed substance is greater than the mass of the deposited substance.
(D) The number of moles of the deposited substance = the number of moles of the consumed substance.

**18** All of the following happen when a piece of zinc is placed in copper(II) sulphate solution, except .....

(A) The piece of zinc becomes coated with a layer of copper.
(B) The mass of zinc decreases.
(C) Electrical energy can be obtained.
(D) The colour of the solution fades.

**19** The reaction in a galvanic cell is represented by the ionic equation: $Co^{2+}(aq) + Fe(s) \rightarrow Co(s) + Fe^{2+}(aq)$

Which of the following describes the reaction that takes place?

(A) Cobalt ions are oxidized.

(B) Cobalt ions are reduced by losing electrons.

(C) Iron atoms are oxidized.

(D) Iron atoms are reduced by gaining electrons.

---

**20** The following redox reaction takes place in a galvanic cell: $2Au^{3+} + 3Cu \rightarrow 3Cu^{2+} + 2Au$

What happens at the cathode of this cell?

(A) Gold atoms lose electrons.

(B) Gold ions gain electrons.

(C) Copper atoms lose electrons.

(D) Copper ions gain electrons.

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**21** Reduction in a galvanic cell happens to .....

(A) the anode rod.

(B) the cathode rod.

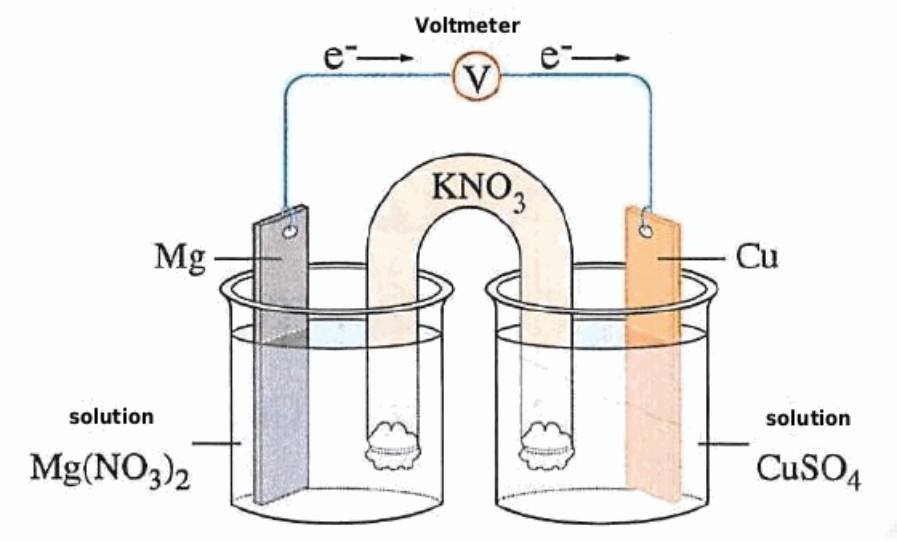
(C) the cations of the anode electrolyte.

(D) the cations of the cathode electrolyte.

---

**22** The opposite figure represents a galvanic cell.

Which of the following represents the anode reaction in this cell?




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

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

(C) $Mg \rightarrow Mg^{2+} + 2e^-$

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

---

**23** When a Daniell cell works, the ions move between the two half-cells through .....

(A) the anode.
(B) the cathode.
(C) the salt bridge.
(D) the external circuit.

---

**24** The opposite reaction takes place in a galvanic cell: $Zn + Fe^{2+} \rightarrow Zn^{2+} + Fe$

Which of the following is correct?

(A) Zinc is the cathode.
(B) Iron is the cathode.
(C) Reduction takes place at the zinc electrode.
(D) Oxidation takes place at the iron electrode.

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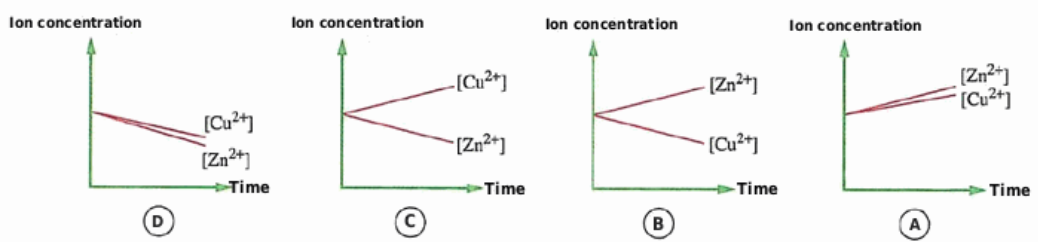
**25** Which of the following represents a half-cell reaction?

(A) $IO_3^- + 6H^+ + 5e^- \rightarrow I_2 + 3H_2O$
(B) $IO_3^- + 6H^+ + 4e^- \rightarrow \frac{1}{2}I_2 + 3H_2O$
(C) $IO_3^- + 6H^+ \rightarrow \frac{1}{2}I_2 + 3H_2O + 5e^-$
(D) $IO_3^- + 6H^+ + 5e^- \rightarrow \frac{1}{2}I_2 + 3H_2O$

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**26** A galvanic cell consists of a Zn electrode dipped in a 0.1 M solution of $Zn^{2+}(aq)$ ions, and a Cu electrode dipped in a 0.1 M solution of $Cu^{2+}(aq)$ ions.

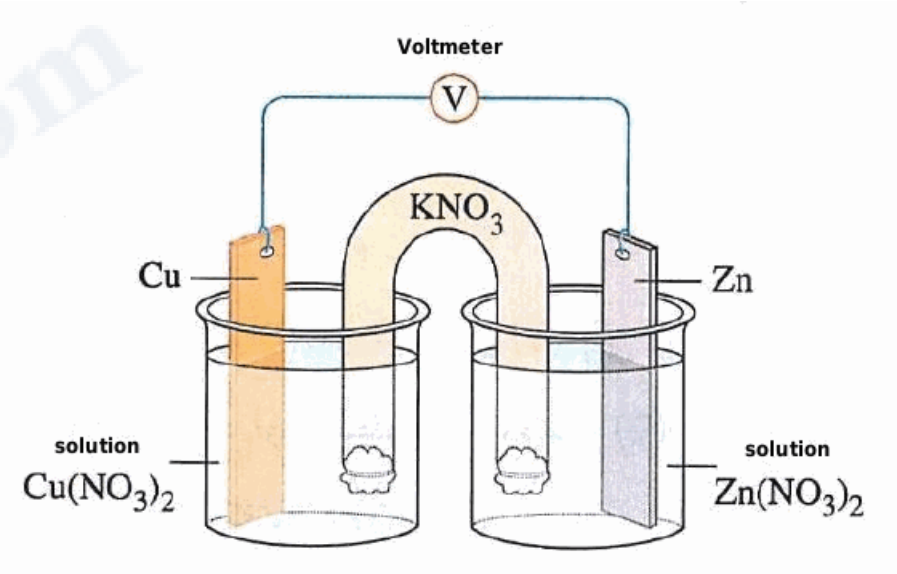
Which of the following graphs represents the concentration of $Zn^{2+}(aq)$ and $Cu^{2+}(aq)$ ions in the electrolyte of each half-cell with time?




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**27** The opposite figure represents a galvanic cell.

What happens when this cell works?




(A) Copper ions move towards the salt bridge.

(B) Cations move towards the zinc electrode.

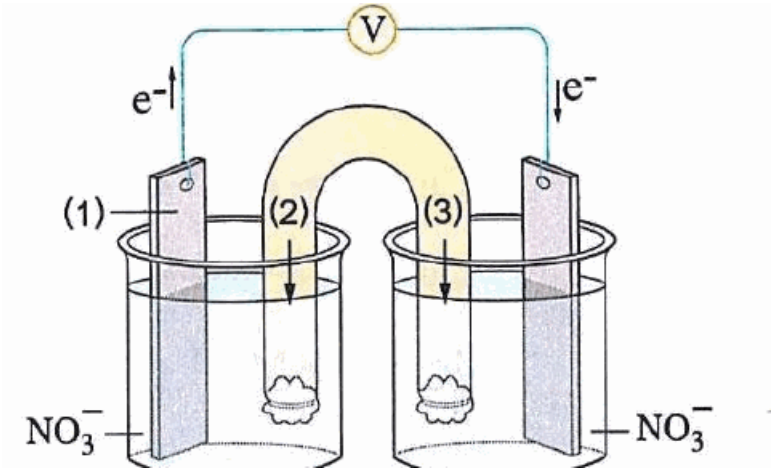
(C) The mass of the copper electrode increases.

(D) The mass of the zinc electrode increases.

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**28** The opposite figure represents an electric cell made of iron and copper electrodes.

Which of the following is correct?




(A) Fe : (1) , $K^+$ : (2)

(B) Cu : (1) , $NO_3^-$ : (2)

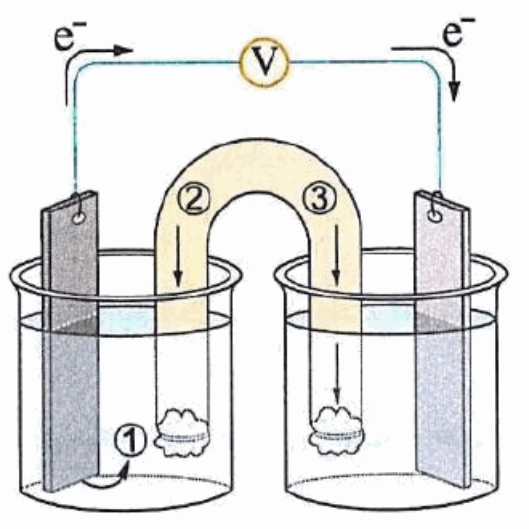
(C) Cu : (1) , $K^+$ : (3)

(D) Fe : (1) , $K^+$ : (3)

---

**29** The opposite figure represents a galvanic cell.

Which of the following is correct?




(A) $Zn^{2+}$ : (1) , $SO_4^{2-}$ : (2) , $Ba^{2+}$ : (3)

(B) $Ba^{2+}$ : (1) , $SO_4^{2-}$ : (2) , $Cu^{2+}$ : (3)

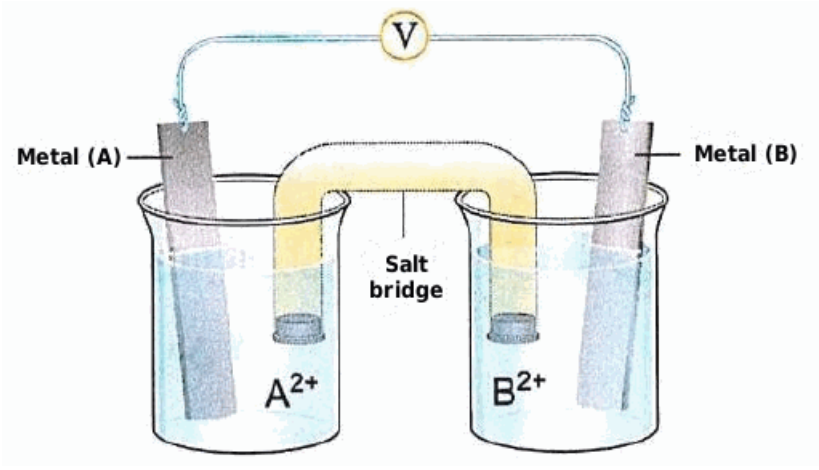
(C) $Cu^{2+}$ : (1) , $Cl^{-}$ : (2) , $Pb^{2+}$ : (3)

(D) $Zn^{2+}$ : (1) , $SO_4^{2-}$ : (2) , $K^{+}$ : (3)

**30** After the galvanic cell shown in the opposite figure worked for several hours, it was noticed that:

- the colour of the $A^{2+}$ solution fades.
- the colour of the $B^{2+}$ solution deepens.

Which of the following describes this cell?




(A) Electrons move from electrode (A) to electrode (B).

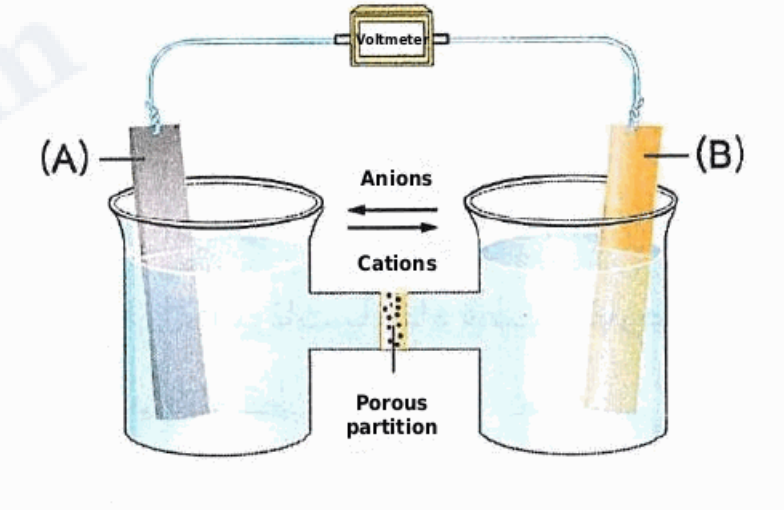
(B) The mass of electrode (B) increased after hours of operation.

(C) The salt-bridge anions move to half-cell (B).

(D) Electrode (A) is the reducing agent.

**31** In the opposite galvanic cell.

Which of the following is correct?




(A) The mass of (A) decreases.

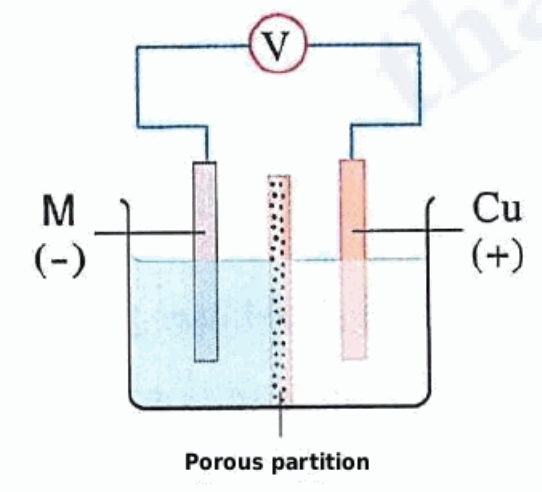
(B) Electrons move from (B) to (A).

(C) The colour of electrolyte (B) deepens.

(D) (A) is the cathode , (B) is the anode.

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**32** Which of the following describes the galvanic cell shown in the opposite figure?




(A) Electrons flow in the external wire from Cu to (M).

(B) The reduction potential of (M) is greater than that of Cu

(C) The cell reaction is: $2M + 3Cu^{2+} \rightarrow 2M^{3+} + 3Cu$

(D) The cell reaction is: $Cu + M^{2+} \rightarrow Cu^{2+} + M$

---

**33** A galvanic cell whose anode is a divalent metal (A) and whose cathode is a trivalent metal (B); the balanced overall reaction of this cell is .....

(A) $A + B^{3+} \rightarrow A^{2+} + B$
(B) $B + A^{2+} \rightarrow B^{3+} + A$
(C) $3A + 2B^{3+} \rightarrow 3A^{2+} + 2B$
(D) $2B + 3A^{2+} \rightarrow 2B^{3+} + 3A$

**34** The salt bridge in galvanic cells neutralizes .....

(A) anions only.
(B) electrons.
(C) cations only.
(D) ions.

**35** The following redox equation represents a galvanic cell: $Ni(s) + Pb^{2+}(aq) \rightarrow Ni^{2+}(aq) + Pb(s)$
The salt bridge in this cell is used to connect .....

(A) Pb(s) with Ni(s)
(B) $Ni^{2+}(aq)$ with $Pb^{2+}(aq)$
(C) $Ni^{2+}(aq)$ with Ni(s)
(D) $Pb^{2+}(aq)$ with Pb(s)

**36** All of the following are roles of the porous partition in a galvanic cell, except .....

(A) Neutralizing the ions produced by oxidation and reduction.
(B) Reacting with the ions of the two half-cell solutions.
(C) Connecting the two half-cell solutions indirectly.
(D) Keeping the oxidation and reduction reactions going.

**37** Which of the following solutions affects the emf of the cell when it is used in the salt bridge of a Daniell cell?

(A) $K_2SO_4$
(B) $NaNO_3$
(C) $KNO_3$
(D) $Na_2CO_3$

**38** A galvanic cell whose two half-cell electrolytes contain ($SO_4^{2-}$), and whose salt bridge contains a cation of the fifth analytical group.
What happens after this cell has worked for some time?

(A) $[SO_4^{2-}]$ decreases in the cathode half-cell.
(B) The bridge cations move to the negative half-cell.
(C) The emf of the cell increases.
(D) The colour of the positive half-cell electrolyte deepens.

**39** A single half-cell .....

(A) is an open circuit, as no electrons flow from or to it.

(B) has only oxidation taking place on the surface of the electrode dipped in it.

(C) has only reduction taking place on the surface of the electrode dipped in it.

(D) always has an electrode reduction potential equal to zero.

---

**40** All of the following are components of the standard hydrogen electrode, except .....

(A) $H^+$ ions (B) $H_2$ gas (C) H atoms (D) Pt plate

---

**41** From the opposite reduction potential of iron(II) ions: $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ $E^\circ = -0.44$ V

What is the emf of the reaction: $Fe(s) + 2H^+(aq) \rightarrow H_2(g) + Fe^{2+}(aq)$?

(A) −0.22 V (B) −0.44 V (C) +0.22 V (D) +0.44 V

---

**42** A galvanic cell is represented by the following equation: $2M^+(aq) + H_2(g) \rightarrow 2M(s) + 2H^+(aq)$

emf = +0.8 V

What is the standard oxidation potential of electrode (M)?

(A) −0.2 V (B) +0.4 V (C) −0.8 V (D) +0.8 V

---

**43** Which of the following solutions turns blue when copper turnings are added to it?

(A) $AgNO_3$ (B) $Zn(NO_3)_2$ (C) $Ba(NO_3)_2$ (D) $NaNO_3$

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**44** The opposite equation: $Fe(s) + CuCl_2(aq) \rightarrow Cu(s) + FeCl_2(aq)$ represents a reaction that is .....

(A) spontaneous, in which the colour of the solution used changes with time.

(B) non-spontaneous, in which the mass of Cu increases with time.

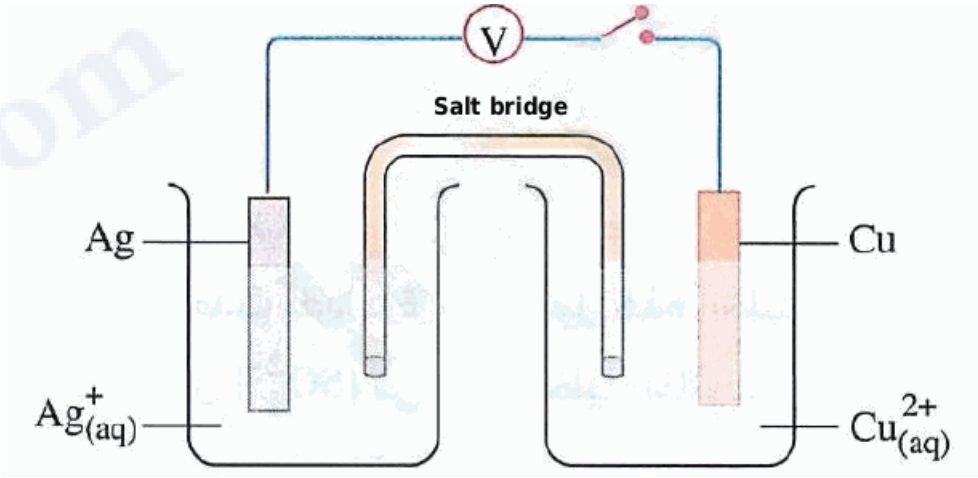
(C) spontaneous, in which the mass of Fe increases with time.

(D) non-spontaneous, in which the mass of Fe decreases with time.

---

**45** From the opposite figure:

What happens when the cell works?




(A) $Ag^+$ ions in the solution increase.

(B) The copper electrode loses electrons.

(C) Electrons move through the electrolyte.

(D) The mass of the copper electrode increases.

---

**46** The E°cell of a galvanic cell is calculated from the relation: E°cell = E°(1) – E°(2)

Which of the following represents (1) and (2)?

(A) (1): reduction potential of $H^+$ , (2): reduction potential of $Cu^{2+}$

(B) (1): oxidation potential of $H_2$ , (2): oxidation potential of Fe

(C) (1): oxidation potential of Fe , (2): oxidation potential of Au

(D) (1): reduction potential of $Cu^{2+}$ , (2): reduction potential of $Pt^{2+}$

---

**47** From the opposite table: the standard reduction potential E° of the half-cell:

$Cu^{2+}/Cu$ : +0.34 V , $Fe^{3+}/Fe^{2+}$ : +0.77 V , $Al^{3+}/Al$ : –1.66 V , $Zn^{2+}/Zn$ : –0.76 V , $Ag^+/Ag$ : +0.8 V

Which pair of the following half-cells forms a cell whose E°cell equals + 0.43 V?

(A) $Cu^{2+}/Cu$ , $Zn^{2+}/Zn$

(B) $Al^{3+}/Al$ , $Zn^{2+}/Zn$

(C) $Fe^{3+}/Fe^{2+}$ , $Cu^{2+}/Cu$

(D) $Ag^+/Ag$ , $Cu^{2+}/Cu$

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**48** The two opposite equations represent the reduction potentials of the two halves of a galvanic cell.

- $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$ $E^\circ = -0.727$ V
- $Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$ $E^\circ = +1.09$ V

What is the emf of this galvanic cell?

(A) 1.817 V (B) 0.363 V (C) 4.724 V (D) 1.316 V

**49** From the galvanic cells shown in the following table:

Cell (1): anode Ga , cathode Pd , $E^\circ_{cell} = +1.18$ V

Cell (2): anode Ga , cathode Cd , $E^\circ_{cell} = +0.16$ V

Cell (3): anode Cd , cathode Pd , $E^\circ_{cell} = ?$

What is the $E^\circ_{cell}$ of cell (3)?

(A) −1.34 V (B) −1.02 V (C) +1.02 V (D) +1.34 V

**50** Metal (X) displaces the hydrogen of HCl acid, while metal (Y) cannot. If the emf of the cell made of the two half-cells (X) and (Y) is 2.46 V and the reduction potential of electrode (Y) is 0.8 V,

then the reduction potential of electrode (X) is .....

(A) +1.66 V (B) −3.26 V (C) +3.26 V (D) −1.66 V

**51** The value of the standard hydrogen potential changes when each of the following is used, except .....

(A) (1 M) $H_2SO_4$ (B) (0.5 M) $H_2SO_4$ (C) (0.5 M) HCl (D) (1.5 M) HCl

**52** A galvanic cell is made of the standard zinc half-cell and the standard hydrogen half-cell. Which of the following describes this cell after it has worked for some time?

(A) The pH around the hydrogen electrode decreases.

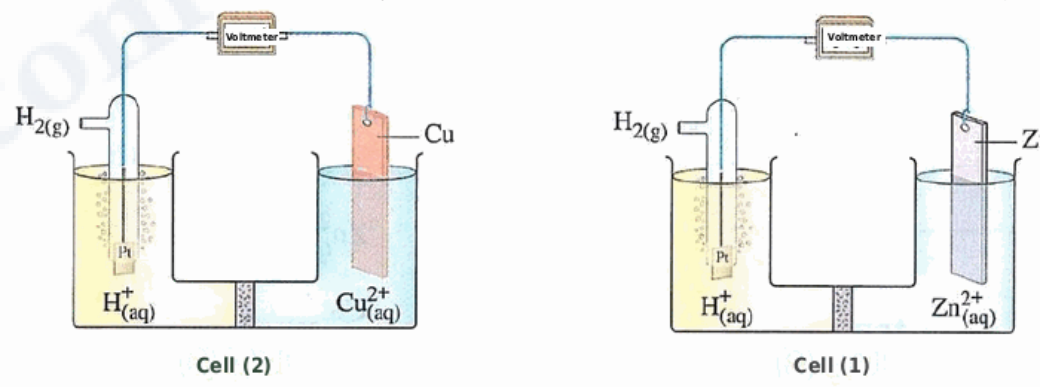
(B) The pOH around the hydrogen electrode decreases.

(C) $[OH^-]$ does not change in the hydrogen half-cell.

(D) $[H^+]$ does not change in the hydrogen half-cell.

**53** The following two figures represent two galvanic cells in which the standard hydrogen electrode is used:

What role does the standard hydrogen electrode play in each of cells (1) and (2)?




(A) (1): cathode half-cell , (2): cathode half-cell.

(B) (1): anode half-cell , (2): cathode half-cell.

(C) (1): cathode half-cell , (2): anode half-cell.

(D) (1): anode half-cell , (2): anode half-cell.

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**54** From the standard potentials of the opposite half-cells:

$2H^+(aq) + 2e^- \rightarrow H_2(g)$ $E^\circ = 0$

$K^+(aq) + e^- \rightarrow K(s)$ $E^\circ = -2.93$ V

$F_2(g) + 2e^- \rightarrow 2F^-(aq)$ $E^\circ = +2.87$ V

$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$ $E^\circ = -1.66$ V

Which of the following is the strongest reducing agent?

(A) $H^+(aq)$ (B) K(s) (C) $F^-(aq)$ (D) Al(s)

---

**55** In the half-cell: $X^{2+} + 2e^- \rightarrow X$ $E^\circ = -0.13$ V it can be concluded that .....

(A) (X): does not displace the hydrogen of acids.

(B) (X): is a reducing agent.

(C) ($X^{2+}$): is a reducing agent.

(D) (X): is an oxidizing agent.

---

**56** From the opposite standard potentials:

$Ag^+ + e^- \rightarrow Ag \quad E^\circ = +0.8\ V$

$Ni^{2+} + 2e^- \rightarrow Ni \quad E^\circ = -0.25\ V$

When the two opposite half-cells are used to make a galvanic cell, .....

(A) $Ag^+$ is an oxidizing agent, while $Ni^{2+}$ is a reducing agent.

(B) $Ni^{2+}$ can be reduced by silver metal.

(C) $Ni^{2+}$ is a stronger oxidizing agent than $Ag^+$

(D) Ni is a stronger reducing agent than Ag

---

**57** From the opposite spontaneous redox reactions:

(1) $X + Y \rightarrow X^- + Y^+$

(2) $Y^+ + Z \rightarrow Y + Z^+$

(3) $Z + X \rightarrow Z^+ + X^-$

The correct order of the strength of the oxidizing agents is .....

(A) $X^- < Y < Z$ (B) $Z < Y < X^-$ (C) $Z < Y^+ < X$ (D) $X < Y^+ < Z$

---

**58** Four metals (A), (B), (C) and (D) have the following properties:

- Only metals (A) and (C) can react with HCl acid.
- Metal (C) reacts with the salt solutions of all the other metals.
- Metal (D) reacts with the cations of metal (B).

Which of the following represents the strength of these metals as reducing agents?

(A) D < C < B < A (B) D < B < C < A (C) D < B < A < C (D) B < D < A < C

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**59** From the opposite table:

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$ $E^\circ = +0.535$ V

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

$2H^+(aq) + 2e^- \rightarrow H_2(g)$ $E^\circ = 0$ V

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

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $E^\circ = -0.762$ V

Which of the following is correct?

(A) $I_2(s)$ is the strongest oxidizing agent and Zn(s) the strongest reducing agent

(B) Zn(s) can reduce $I_2(s)$ and $Fe^{2+}(aq)$

(C) Cu(s) can oxidize Zn(s) and $Fe^{2+}(aq)$

(D) Fe(s) can oxidize Cu(s) and $Zn^{2+}(aq)$

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**60** From the opposite table: $A^{2+}/A$ : $E^\circ = -1.029$ V , $B^+/B$ : $E^\circ = +0.8$ V

What happens when a rod of metal (A) is dipped in a 1 M aqueous solution of $(B^+)$ ions?

(A) (B) gradually dissolves.

(B) (B) is deposited on (A).

(C) No reaction takes place.

(D) Water decomposes into $H_2$ and $O_2$

---

**61** From the opposite table:

$Cr^{3+}/Cr^{2+}$ : $E^\circ = -0.41$ V , $Mn^{3+}/Mn^{2+}$ : $E^\circ = +1.57$ V

$Fe^{3+}/Fe^{2+}$ : $E^\circ = +0.77$ V , $Co^{3+}/Co^{2+}$ : $E^\circ = +1.97$ V

Which of the following shows the correct order of the ease with which the ion changes from oxidation state $(+2)$ to oxidation state $(+3)$?

(A) $Co^{2+} < Mn^{2+} < Fe^{2+} < Cr^{2+}$

(B) $Cr^{2+} < Fe^{2+} < Mn^{2+} < Co^{2+}$

(C) $Mn^{2+} < Co^{2+} < Fe^{2+} < Cr^{2+}$

(D) $Cr^{2+} < Co^{2+} < Mn^{2+} < Fe^{2+}$

---

**62** From the opposite table and the following equation: $Mn^{3+} + e^- \rightarrow Mn^{2+}$ $E^\circ = -1.029$ V

$Pb^{2+} + 2e^- \rightarrow Pb$ $E^\circ = -0.126$ V , $Fe^{2+} + 2e^- \rightarrow Fe$ $E^\circ = -0.409$ V

$Mg^{2+} + 2e^- \rightarrow Mg$ $E^\circ = -2.375$ V , $Zn^{2+} + 2e^- \rightarrow Zn$ $E^\circ = -0.762$ V

Which of the following can reduce the $Mn^{3+}$ ion to the $Mn^{2+}$ ion?

(A) Mg only.

(B) Zn only.

(C) Fe and Pb only.

(D) Zn , Fe and Pb

---

**63** From the opposite table:

$Mn^{7+}(aq) + 5e^- \rightarrow Mn^{2+}(aq)$ $E^\circ = +1.51$ V

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $E^\circ = -0.76$ V

$Ag^{+}(aq) + e^- \rightarrow Ag(s)$ $E^\circ = +0.8$ V

$Au^{+}(aq) + e^- \rightarrow Au(s)$ $E^\circ = +1.69$ V

Which of the following metals can be oxidized using the $MnO_4^-$ ion?

(A) Zn and Ag only (B) Ag and Au only (C) Zn and Au only (D) Zn , Ag and Au

---

**64** The following equations represent three electrochemical processes:

(1) $Fe(s) + Ni^{2+}(aq) \rightarrow Fe^{2+}(aq) + Ni(s)$

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

(3) $Pb(s) + Fe^{2+}(aq) \rightarrow Pb^{2+}(aq) + Fe(s)$

and the standard reduction potentials: $Pb^{2+}/Pb$ : −0.13 V , $Ni^{2+}/Ni$ : −0.25 V , $Fe^{2+}/Fe$ : −0.44 V

Which of the following represents the spontaneous processes?

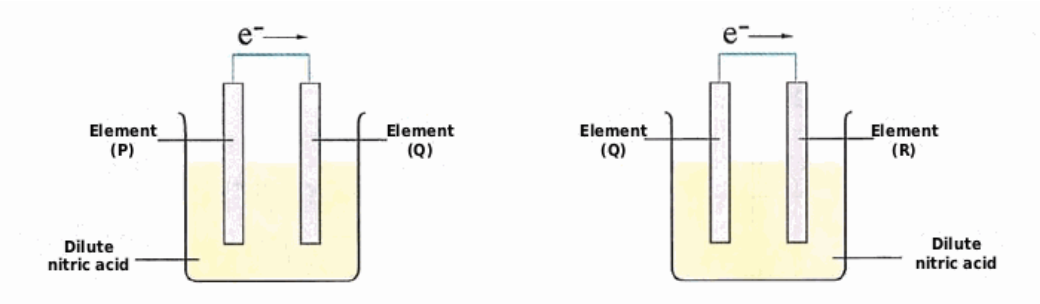
(A) Processes (1) , (2) and (3). (B) Processes (1) and (2) only.

(C) Processes (1) and (3) only. (D) Processes (2) and (3) only.

---

**65** From the following two cells:

Which of the following is correct?




(A) Cu : (R) , Ag : (Q) , Zn : (P) (B) Cu : (R) , Zn : (Q) , Ag : (P)

(C) Ag : (R) , Zn : (Q) , Cu : (P) (D) Ag : (R) , Cu : (Q) , Zn : (P)

---

**66** The circuit shown in the opposite figure is used to compare the activity of four unknown metals (X), (Q), (Y) and (Z), and the following table shows the voltmeter readings in the four cases:

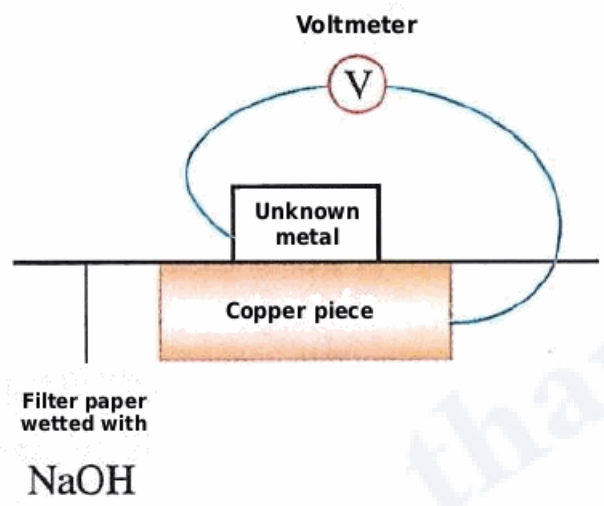
Metal (X): direction of electron flow in the external circuit X → Cu , voltmeter reading +0.87 V

Metal (Q): direction of electron flow Cu → Q , voltmeter reading +1.2 V

Metal (Y): direction of electron flow Y → Cu , voltmeter reading +1.58 V

Metal (Z): direction of electron flow Z → Cu , voltmeter reading +0.36 V

Which of the following shows the order of activity of these metals?




(A) Z < Y < Q < X (B) Q < Y < X < Z (C) Q < Z < X < Y (D) X < Q < Z < Y

**67** An excess of zinc powder was added to a small volume of a mixture of magnesium nitrate and copper(II) chloride solutions.

Which cations are present in the reaction mixture after the reaction ends?

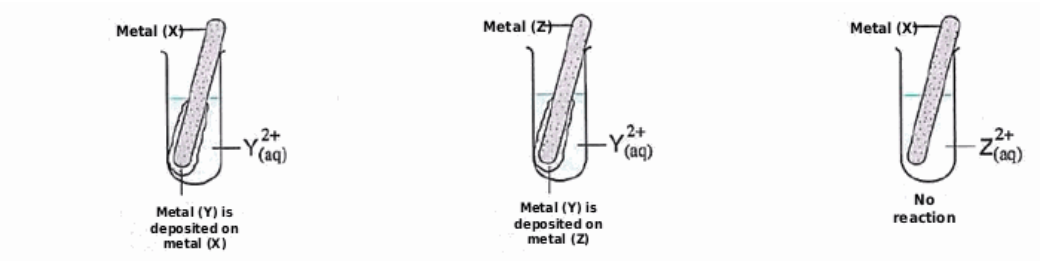
(A) $Mg^{2+}$ , $Cu^{2+}$ (B) $Zn^{2+}$ , $Cu^{2+}$

(C) $Zn^{2+}$ , $Mg^{2+}$ (D) $Cu^{2+}$ , $Zn^{2+}$ , $Mg^{2+}$

**68** The following figure represents three experiments done to find the chemical activity of three metals (X), (Y) and (Z):

- Metal (X) in $Z^{2+}$ solution: no reaction.
- Metal (X) in $Y^{2+}$ solution: metal (Y) is deposited on metal (X).
- Metal (Z) in $Y^{2+}$ solution: metal (Y) is deposited on metal (Z).

Which of the following represents the order of activity of these metals?




(A) Y < Z < X (B) Z < X < Y (C) X < Y < Z (D) Y < X < Z

**69** Which of the following is correct as a result of placing an iron rod in $CuSO_4$ solution for some time?

(A) Mass of iron: decreases , $[Cu^{2+}]$: decreases , number of moles of ions: decreases

(B) Mass of iron: increases , $[Cu^{2+}]$: increases , number of moles of ions: decreases

(C) Mass of iron: decreases , $[Cu^{2+}]$: decreases , number of moles of ions: does not change

(D) Mass of iron: increases , $[Cu^{2+}]$: increases , number of moles of ions: does not change

**70** When a rod of metal (A) was dipped in a concentrated aqueous solution (B), the solution turned blue, and when NaCl(aq) was added to solution (B), precipitate (C) formed, which dissolves when concentrated ammonia solution is added to it.

Which of the following is correct?

(A) Fe : (A) , $Zn(NO_3)_2$ : (B) , $ZnCl_2$ : (C)

(B) Cu : (A) , $AgNO_3$ : (B) , AgCl : (C)

(C) Ni : (A) , $Al(NO_3)_3$ : (B) , $AlCl_3$ : (C)

(D) Co : (A) , $Pb(NO_3)_2$ : (B) , $PbCl_2$ : (C)

**71** The following table shows whether 5 different metals (A), (B), (C), (D) and (E) react with solutions of their nitrates (✓ reaction takes place, ✗ no reaction):

$A(NO_3)_2$ solution: metals (B) ✓ , (C) ✓ , (D) ✗ , (E) ✓

$B(NO_3)_2$ solution: metals (A) ✗ , (C) ✗ , (D) ✗ , (E) ✗

$C(NO_3)_2$ solution: metals (A) ✗ , (B) ✓ , (D) ✗ , (E) ✗

$DNO_3$ solution: metals (A) ✓ , (B) ✓ , (C) ✓ , (E) ✓

$E(NO_3)_2$ solution: metals (A) ✗ , (B) ✓ , (C) ✓ , (D) ✗

What is the correct order of the chemical activity of these elements?

(A) D < E < A < C < B

(B) C < B < E < A < D

(C) D < A < E < C < B

(D) C < A < B < E < D

---

**72** The following table shows the results of adding 6 g of nickel to four solutions of the salts of metals (W), (X), (Y) and (Z), of the same volume and concentration, for a fixed time:

Salt solution of metal (W): mass of nickel after 15 min = 5 g

Salt solution of metal (X): mass of nickel after 15 min = 5.5 g

Salt solution of metal (Y): mass of nickel after 15 min = 4.5 g

Salt solution of metal (Z): mass of nickel after 15 min = 0

Which of the following is correct?

(A) Metal (W) comes before metal (X) in the electrochemical series.

(B) Metal (Y) can displace metal (Z) from its salt solution.

(C) Metal (Y) comes before nickel in the electrochemical series.

(D) Metal (Z) is nickel(II) chloride solution.

---

## Essay questions — 13 questions

**1** The reaction of zinc sulphite with hydrochloric acid is represented by the following ionic equation:

$ZnSO_3(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + SO_2(g) + H_2O(l)$

(1) Is this reaction a redox reaction? Explain.

(2) What is the effect of passing the gas produced on the colour of a $(H^+ + Cr_2O_7^{2-})$ solution?

**2** The two experiments shown in the following figures were done using two zinc plates:

(1) Is the reaction in experiment (1) complete or reversible? Explain.

(2) State two changes in experiment (2) that show that oxidation and reduction took place.




**3** A galvanic cell in which the electrolyte of the first half-cell contains $Ag^+(aq)$ ions and the second half contains $Ni^{2+}(aq)$ ions:

(1) Deduce the ionic equation of the overall reaction in this cell, with explanation.

(2) Calculate the emf of this cell, given the following two oxidation potentials:

$Ni \rightarrow Ni^{2+} + 2e^-$ $E^\circ = +0.23$ V

$Ag \rightarrow Ag^+ + e^-$ $E^\circ = -0.8$ V

**4** The following ionic equation represents the overall reaction in a galvanic cell:

$H_2(g) + Cu^{2+}(aq) \rightarrow 2H^+(aq) + Cu(s)$

(1) What is the oxidizing agent in this reaction? Explain.

(2) Calculate the emf of this cell.

**5** • Metal (A): monovalent, its oxidation potential is +0.7 V

• Metal (B): divalent, its oxidation potential is −0.4 V

(1) Calculate the emf of the galvanic cell made of them.

(2) Write the equation of the overall reaction in this cell.

**6** From the opposite electrochemical potentials:

$A^{4+}(aq) + e^- \rightarrow A^{3+}(aq)$ $E^\circ = 0.01$ V

$B^{3+}(aq) + e^- \rightarrow B^{2+}(aq)$ $E^\circ = 0.77$ V

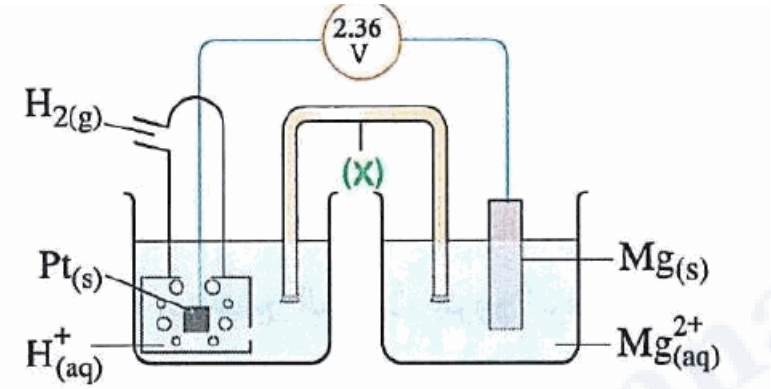
Show whether the following reaction is spontaneous, with explanation:

$A^{4+}(aq) + B^{2+}(aq) \rightarrow A^{3+}(aq) + B^{3+}(aq)$

**7** From the opposite figure:

(1) Does magnesium act as the cathode or the anode in this cell? Explain your answer.

(2) State two of the changes that happen when the Mg rod is replaced by a Ag rod in this cell.



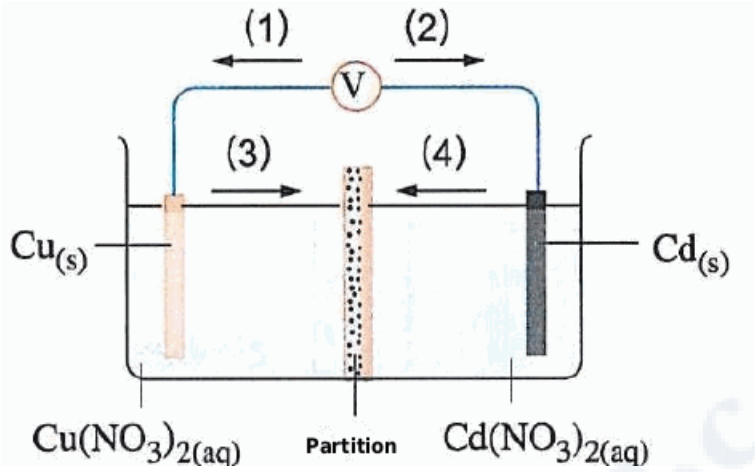

**8** The opposite figure represents a galvanic cell. Given that the oxidation potential of cadmium is 0.4 V and that of copper is −0.34 V:

(1) State the number that shows each of:

(a) the direction of electron movement.

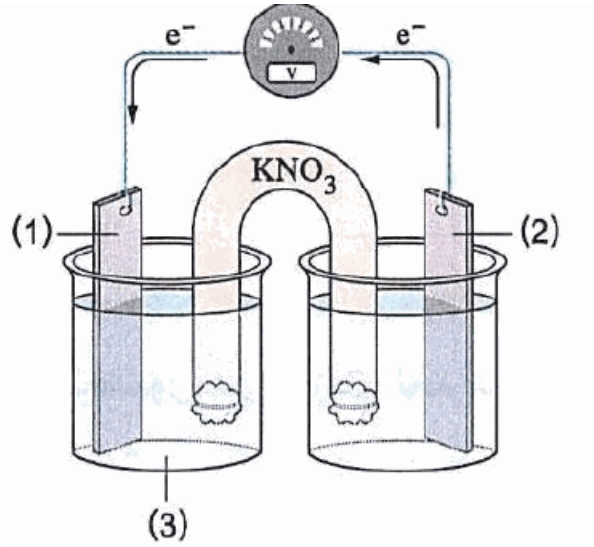
(b) the direction of anion movement.

(2) Write the equation of the overall reaction.



**9** From the following substances: nickel, aluminium, silver, $Ni(NO_3)_2(aq)$, $Al(NO_3)_3(aq)$, $AgNO_3(aq)$

Replace the numbers shown on the opposite figure with the suitable substances from the list above to give the largest voltmeter reading, and calculate this reading using the electrochemical potentials.




**10** From the following table:

Metal (A): reduction potential −1.66 V

Metal (B): reduction potential +0.34 V

Metal (C): reduction potential −0.44 V

Metal (D): reduction potential −2.73 V

Metal (E): reduction potential −0.76 V

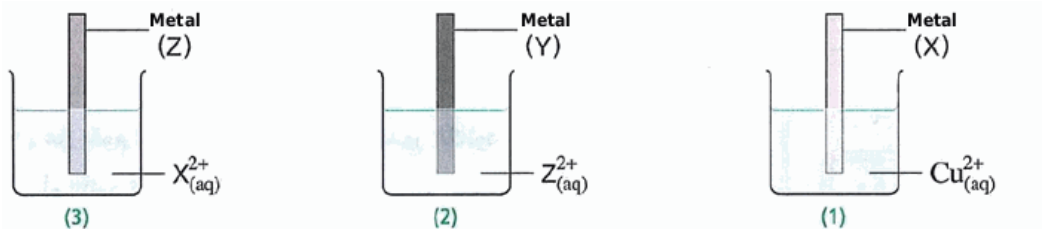
Arrange these metals according to their ability to displace one another from solutions of their salts, with explanation.

**11** The experiments shown in the following figures were done to compare the chemical activity of four metals; all of them were accompanied by a chemical change:

(1) What change happens with time in experiment (1) to each of the following, giving the reason:

(a) rod (X). (b) the solution of $Cu^{2+}$ ions

(2) Arrange the four metals in these experiments in ascending order of chemical activity.

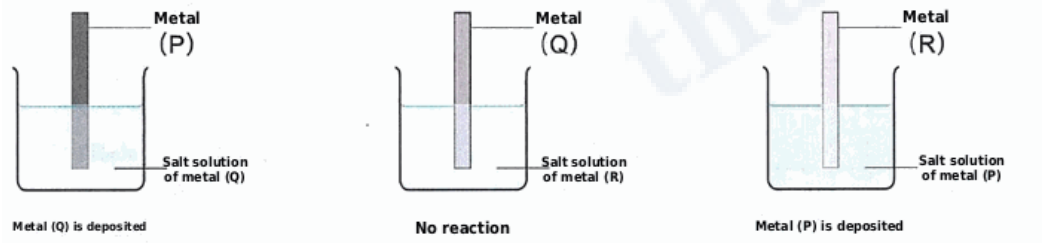



**12** The following figures represent three different experiments:

Two galvanic cells can be made by connecting:

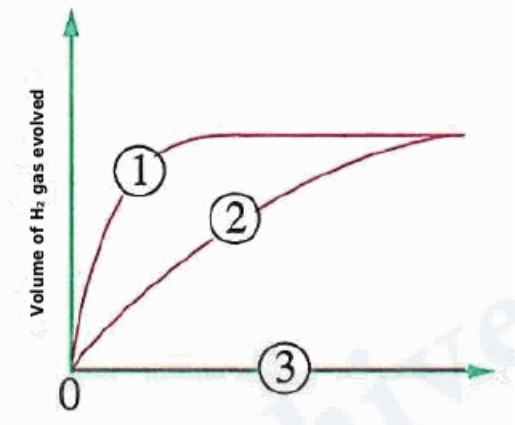
- metal (P) with metal (R).
- metal (P) with metal (Q).

State the role of metal (P) in each cell, with explanation.




**13** Dilute hydrochloric acid was added to three tubes containing iron, copper and magnesium metals, each separately, and the amounts of hydrogen given off with each metal were plotted as the curves shown in the opposite figure.

Match each metal with the curve that represents it, explaining your answer.



## Past exam questions — 64 questions

**1** A few drops of $Na_2S$ solution were added to the cathode half-cell of a Daniell cell. Which of the following is correct?

(A) The electromotive force increases.

(B) The life time of the cell decreases.

(C) The concentration of $Cu^{2+}$ cations increases.

(D) The current reverses and the emf decreases.

---

**2** In the following reaction:

$5Cl^- + 8H^+ + MnO_4^- \rightarrow (5/2)\ Cl_2 + 4H_2O + Mn^{2+}$

($Cl_2$ / $2Cl^-$ = +1.36 V , $Mn^{2+}$ / $Mn^{7+}$ = −1.52 V)

Which of the following gives the type and the potential of the reaction, respectively?

(A) Spontaneous, and the potential = +0.16 V

(B) Non-spontaneous, and the potential = −2.88 V

(C) Spontaneous, and the potential = −0.16 V

(D) Non-spontaneous, and the potential = +2.88 V

---

**3** The opposite table shows the components of two galvanic cells and the standard potential of each:

Cell 1: anode (X) , cathode Ag , emf = 0.80 V

Cell 2: anode (Y) , cathode Ag , emf = 1.56 V

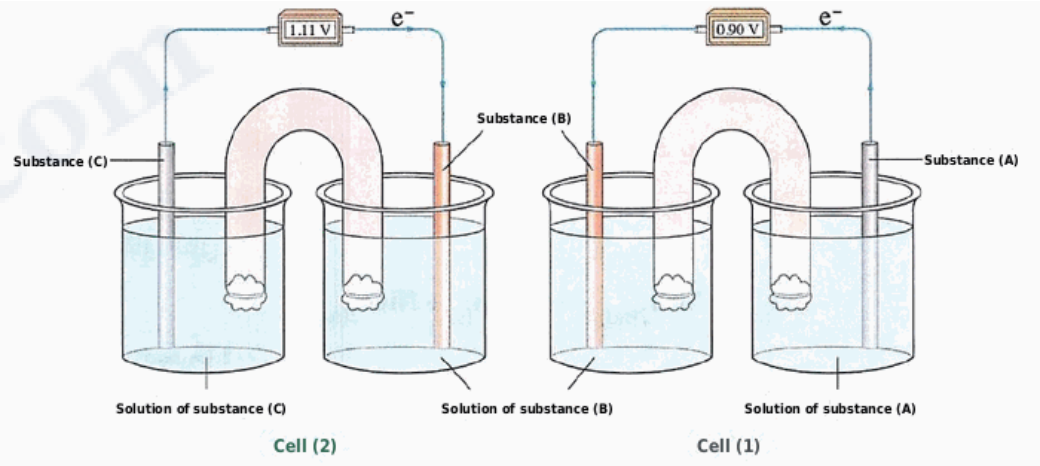
If a galvanic cell is made with electrodes (X , Y), its emf equals .....

(A) +2.36 V (B) −2.36 V (C) +0.76 V (D) −0.76 V

---

**4** The following figure shows two galvanic cells:

Which of the following is correct for electrodes (A , B , C)?




(A) The strongest reducing agent is (A).

(B) The lowest oxidation potential is (C).

(C) The least active electrode is (B).

(D) The mass of electrode (C) increases.

---

**5** In a galvanic cell made of electrodes (X) and (Y), electrons were observed to flow towards electrode (Y) while the cell works. Which of the following is correct?

(A) The reduction potential of (X) is smaller than that of (Y).

(B) The oxidation potential of (X) is smaller than that of (Y).

(C) Electrode (X) is the positive electrode of the cell.

(D) Anions flow through the bridge towards electrode (Y).

---

**6** The standard potentials of electrodes (X) and (Y) are as follows:

$X(s) + 2H^+(1\ M) \rightarrow X^{2+}(aq) + H_2(g)\ E^\circ = +0.76\ V$

$Y(s) + 3H^+(1\ M) \rightarrow Y^{3+}(aq) + (3/2)\ H_2(g)\ E^\circ = +1.66\ V$

Which of the following is correct when an electrochemical cell is made from electrodes (X) and (Y)?

(A) The anode is (X) and the emf equals +0.9 V

(B) The anode is (Y) and the emf equals +0.9 V

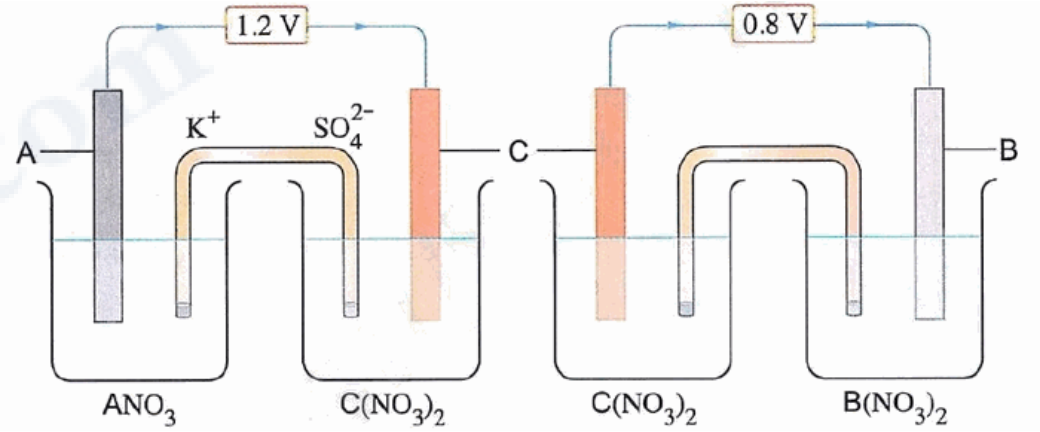
(C) The anode is (X) and the emf equals +2.42 V

(D) The anode is (Y) and the emf equals +2.42 V

---

**7** From the electric cells shown:

Which of the following describes the cell made of electrode (B) as anode and (A) as cathode?




(A) emf = +2 V , a spontaneous reaction takes place in it

(B) emf = −1.2 V , a non-spontaneous reaction takes place in it

(C) emf = −2 V , a non-spontaneous reaction takes place in it

(D) emf = +1.2 V , a spontaneous reaction takes place in it

---

**8** In the following spontaneous reaction:

$Ni(s) + CuCl_2(aq) \rightarrow Cu(s) + NiCl_2(aq)$

Which of the following is correct?

(A) The copper ion is a strong reducing agent.

(B) Solutions of copper salts can be kept in nickel containers.

(C) Solutions of nickel salts can be kept in a copper container.

(D) The chloride ion is a strong oxidizing agent.

---

**9** Which of the following reactions can represent a spontaneous reaction inside an electric cell?

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

(B) $Mg(s) + Zn^{2+}(aq) \rightarrow Mg^{2+}(aq) + Zn(s)$

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

(D) $3Cu(s) + 2Al^{3+}(aq) \rightarrow 3Cu^{2+}(aq) + 2Al(s)$

---

**10** When a galvanic cell was made from electrodes (X) and (Y), the concentration of $Y^{2+}$ ions was observed to increase. Which of the following statements is correct?

(A) The oxidation potential of (X) is greater than that of (Y).

(B) $Y^{2+}$ is a reducing agent while the cell works.

(C) Electrons flow in the wire from (X) to (Y).

(D) $X^{2+}$ is an oxidizing agent while the cell works.

---

**11** Using the opposite table:

$Cd + Zn^{2+} \rightarrow Cd^{2+} + Zn$ non-spontaneous

$Cd + Cu^{2+} \rightarrow Cd^{2+} + Cu$ spontaneous

Which of the following gives the correct order of the opposite metal ions by their strength as oxidizing agents?

(A) $Cd^{2+} > Cu^{2+} > Zn^{2+}$

(B) $Zn^{2+} > Cu^{2+} > Cd^{2+}$

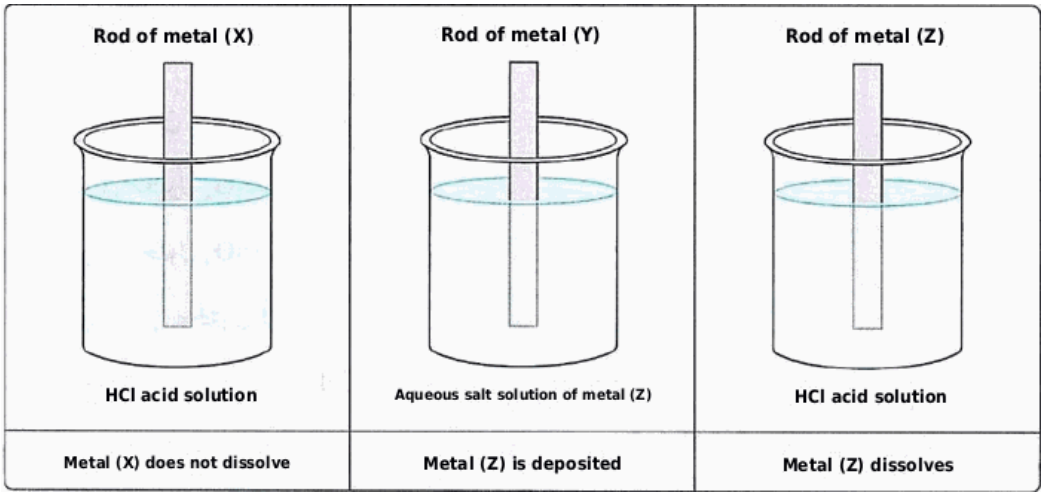
(C) $Cd^{2+} > Zn^{2+} > Cu^{2+}$

(D) $Cu^{2+} > Cd^{2+} > Zn^{2+}$

---

**12** From the following figures:

Which of the following gives the electrodes that form a galvanic cell with the greatest electromotive force?




(A) (Z) , (X).

(B) (Y) , $H_2$.

(C) (Y) , (X).

(D) (Z) , (Y).

---

**13** The following reaction takes place inside an electrochemical cell:

$3Cu^{2+} + 2Cr \rightarrow 3Cu + 2Cr^{3+}$

Given that: ($Cu^{2+} + 2e^- \rightarrow Cu$ , $E^\circ = +0.340$ V , $Cr \rightarrow Cr^{3+} + 3e^-$ , $E^\circ = +0.740$ V)

Which of the following gives the type of the cell and its emf?

(A) Galvanic , emf = +1.08 V

(B) Electrolytic , emf = −1.08 V

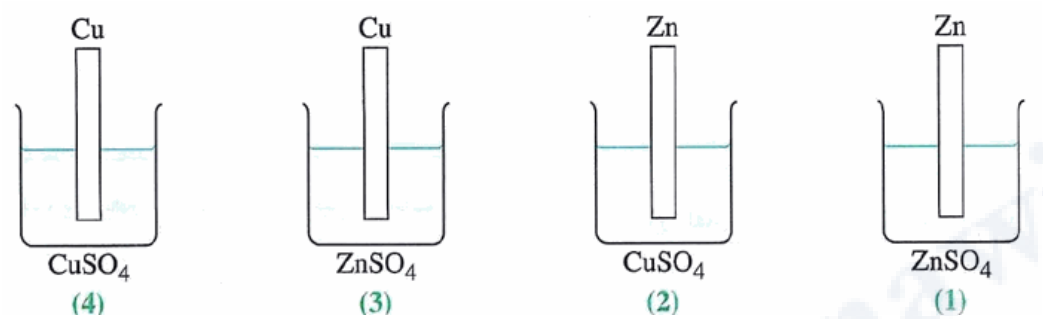
(C) Galvanic , emf = +0.4 V

(D) Electrolytic , emf = −0.4 V

---

**14** From the following figures:

Which of the following is correct?




(A) When (1) is connected to (4) by a salt bridge and an external wire, (4) becomes the anode and (1) the cathode.

(B) A spontaneous reaction takes place in (2) and the concentration of copper ions increases.

(C) When (1) is connected to (4) by a salt bridge and a wire, electrons move spontaneously from (1) to (4).

(D) A spontaneous reaction takes place in vessel (3) and the concentration of zinc ions decreases.

---

**15** A galvanic cell consists of:

• electrode (X) dipped in $XSO_4$ solution

• electrode (Y) dipped in $YSO_4$ solution

If metal (X) is a stronger reducing agent than (Y), which of the following describes what happens in this cell?

(A) Electrons move from (Y) to (X) through the external wire.

(B) The mass of electrode (X) increases with time.

(C) Negative ions move from the bridge to half-cell (X).

(D) ($X^{2+}$) ions are reduced at the anode.

**16** An electric cell in which the following reaction takes place: $2Y(s) + 3X^{2+}(aq) \rightarrow 2Y^{3+}(aq) + 3X(s)$ , emf = +1.6 V

Which of the following is correct for both the type of electrode (Y) and the type of this cell?

(A) Electrode (Y): cathode / cell type: galvanic

(B) Electrode (Y): anode / cell type: electrolytic

(C) Electrode (Y): anode / cell type: galvanic

(D) Electrode (Y): cathode / cell type: electrolytic

---

**17** Four electrodes (X), (Y), (Z) and (L), where:

$X^{2+}/X$ : −0.44 V , $Y^{2+}/Y$ : +1.5 V , $Z/Z^{2+}$ : −0.34 V , $L/L^{2+}$ : +1.18 V

Which of the following reactions represents a cell that produces the greatest electromotive force (emf)?

(A) $Y + L^{2+} \rightarrow Y^{2+} + L$

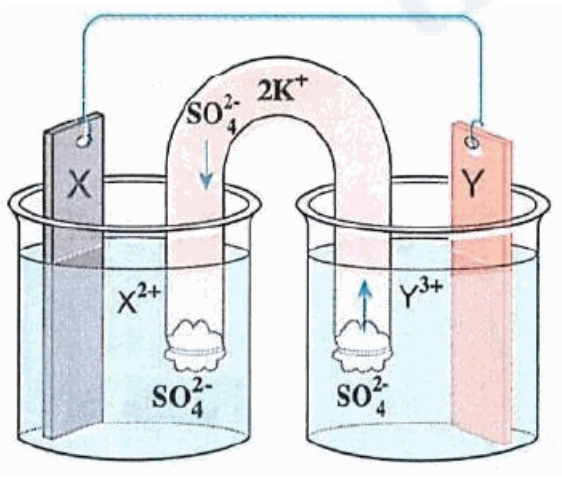
(B) $X + L^{2+} \rightarrow X^{2+} + L$

(C) $L + Y^{2+} \rightarrow L^{2+} + Y$

(D) $X + Y^{2+} \rightarrow X^{2+} + Y$

---

**18** The opposite figure shows a galvanic cell made of electrodes (X , Y). Which of the following represents the overall cell reaction?




(A) $X(s) + Y^{3+}(aq) \rightarrow X^{2+}(aq) + Y(s)$

(B) $Y(s) + X^{2+}(aq) \rightarrow Y^{3+}(aq) + X(s)$

(C) $2X(s) + 3Y^{3+}(aq) \rightarrow 2X^{2+}(aq) + 3Y(s)$

(D) $3X(s) + 2Y^{3+}(aq) \rightarrow 3X^{2+}(aq) + 2Y(s)$

---

**19** Which of the following reactions represents a non-spontaneous reaction at the anode of an electrolytic cell?

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

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

(C) $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$

(D) $2H_2O(l) + 2e^- \rightarrow H_2(g) + 4OH^-(aq)$

**20** Galvanic cell (X) is made of electrodes (A , B), and electrons move in the external wire from (B) to (A). Electric cell (Y), made of the same electrodes, has the following reaction: $A(s) + B^{2+}(aq) \rightarrow A^{2+}(aq) + B(s)$

Which of the following choices describes cell (Y)?

(A) Galvanic with positive emf, electrode (A) is negative.

(B) Galvanic with positive emf, electrode (B) is positive.

(C) Electrolytic with negative emf, electrode (A) is negative.

(D) Electrolytic with negative emf, electrode (B) is negative.

---

**21** Two galvanic cells:

• First cell: electrode (X) is the anode and the hydrogen electrode is the cathode, and it gives an emf of +0.23 V

• Second cell: electrode (Y) is the cathode and the hydrogen electrode is the anode, and it gives an emf of +0.8 V

Which of the following describes the cell made of electrodes (X) and (Y)?

(A) (X) anode and (Y) cathode , emf = 1.03 V

(B) (X) anode and (Y) cathode , emf = 0.57 V

(C) (Y) anode and (X) cathode , emf = 1.03 V

(D) (Y) anode and (X) cathode , emf = 0.57 V

---

**22** A galvanic cell made of iron and tin electrodes:

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

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

Which of the following gives the anode or cathode and the emf of the cell?

(A) The iron electrode is the anode and the emf of the cell is +0.559 V

(B) The iron electrode is the cathode and the emf of the cell is −0.559 V

(C) The tin electrode is the anode and the emf of the cell is +0.259 V

(D) The tin electrode is the cathode and the emf of the cell is −0.259 V

---

**23** The following reaction represents an electrochemical cell:

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

Given that:

$X(s) \rightarrow X^{2+}(aq) + 2e^{-}$ $E^{\circ} = 0.23$ V

$Y^{+} + e^{-} \rightarrow Y(s)$ $E^{\circ} = 0.8$ V

Which of the following gives the type of the cell and its emf?

(A) Electrolytic and emf = −1.03 V
(B) Galvanic and emf = −0.564 V
(C) Electrolytic and emf = +0.564 V
(D) Galvanic and emf = +1.03 V

---

**24** The following reaction represents a galvanic cell: $Fe(s) + Pb^{2+}(aq) \rightarrow Fe^{2+}(aq) + Pb(s)$

Which of the following can be used as the electrolyte solution in the salt bridge of this cell?

(A) Sodium chloride.
(B) Sodium carbonate.
(C) Sodium nitrate.
(D) Ethyl alcohol.

---

**25** The following equation represents the reaction in an electric cell:

$2X + 3CuSO_4 \rightarrow X_2(SO_4)_3 + 3Cu$ , E_cell = +1.8 V

Which of the following gives the reduction potential of (X)? ($Cu \rightarrow Cu^{2+} + 2e^{-}$ , $E^{\circ} = -0.34$ V)

(A) −2.14 V
(B) +1.46 V
(C) +2.14 V
(D) −1.46 V

---

**26** The opposite table shows the oxidation potentials of some electrodes:

X : 0.76 , Y : −0.34 , Z : 1.67 , W : 2.37 , E : −0.8

An electric cell was made of electrodes (X , Y). Which of the following changes to this cell increases the electromotive force and changes the direction of the current?

(A) Replacing the anode with electrode W
(B) Replacing the anode with electrode Z
(C) Replacing the cathode with electrode W
(D) Replacing the cathode with electrode E

---

**27** The following table shows the standard electrode potentials of four elements:

$A^{2+} + 2e^- \rightarrow A$ $E^\circ = -2.71$ V

$B \rightarrow B^{2+} + 2e^-$ $E^\circ = +0.40$ V

$C^{2+} + 2e^- \rightarrow C$ $E^\circ = +0.80$ V

$D \rightarrow D^{2+} + 2e^-$ $E^\circ = -1.20$ V

Which of the following gives the number of galvanic cells that can be made from these elements, and the highest emf that can be obtained?

(A) 4 cells , emf = 3.91 V

(B) 6 cells , emf = 3.91 V

(C) 6 cells , emf = 3.11 V

(D) 12 cells , emf = 3.51 V

---

**28** The following reaction takes place in an electric cell:

$2M(s) + 3Zn^{2+}(aq) \rightarrow 2M^{3+}(aq) + 3Zn(s)$ , $E^\circ_cell = +1$ V

Given that: $Zn^{2+}(aq) \rightarrow Zn(s)$ $E^\circ = -0.76$ V

Which of the following gives the type of electrode (M), its oxidation potential and the type of the cell?

(A) Electrode (M) is the anode, its oxidation potential = +1.76 and the cell is galvanic.

(B) Electrode (M) is the anode, its oxidation potential = +0.24 and the cell is electrolytic.

(C) Electrode (M) is the cathode, its oxidation potential = −0.24 and the cell is electrolytic.

(D) Electrode (M) is the cathode, its oxidation potential = −1.76 and the cell is galvanic.

---

**29** From the opposite electrode potentials:

$X^{2+} + 2e^- \rightarrow X$ $E^\circ = -0.41$ V

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

Which of the following choices describes the reaction: $X + Y^{2+} \rightarrow X^{2+} + Y$?

(A) Non-spontaneous, and the mass of electrode (X) decreases.

(B) Non-spontaneous, and the mass of electrode (Y) decreases.

(C) Spontaneous, and the mass of electrode (X) decreases.

(D) Spontaneous, and the mass of electrode (Y) decreases.

---

**30** The following table shows the potentials of some elements:

$X \rightarrow X^{2+}$ $E^\circ = 0.76$

$W \rightarrow W^{2+}$ $E^\circ = 2.37$

$Z^{2+} \rightarrow Z$ $E^\circ = 0.23$

When two of the electrodes above are used to make a cell with an emf of 1.61 V, which of the following describes the change when the cathode of this cell is replaced with the third electrode?

(A) The emf increases and the direction of the current stays the same.

(B) The emf decreases and the direction of the current stays the same.

(C) The emf increases and the direction of the current reverses.

(D) The emf decreases and the direction of the current reverses.

---

**31** The following reaction represents an electrochemical cell: $X^{2+}(aq) + 2Y(s) \rightarrow X(s) + 2Y^{+}(aq)$

If:

$X^{2+}(aq) \rightarrow X(s)$ $E^\circ = -0.56$ V

$Y^{+}(aq) \rightarrow Y(s)$ $E^\circ = +0.23$ V

which of the following gives the type of the cell and its electromotive force (emf)?

(A) Galvanic , emf = +0.23 V

(B) Galvanic , emf = +0.79 V

(C) Electrolytic , emf = −0.23 V

(D) Electrolytic , emf = −0.79 V

---

**32** From the opposite standard potentials:

$X^{2+} + 2e^{-} \rightarrow X$ $E^\circ = -0.402$ V

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

Which of the following choices describes the following reaction? $Y + X^{2+} \rightarrow Y^{2+} + X$

(A) Non-spontaneous, and the mass of (X) decreases.

(B) Non-spontaneous, and electrode (Y) is the anode.

(C) Spontaneous, and electrode (Y) is the anode.

(D) Spontaneous, and the mass of (X) decreases.

---

**33** The following table shows the potentials of some metals:

Al : −1.67 V , Ag : 0.8 V , Zn : −0.76 V , Cu : 0.34 V , Fe : −0.409 V

When equal masses of iron pieces are placed in several solutions of equal concentration of copper sulphate, aluminium sulphate, zinc sulphate and silver nitrate, which of the following choices is correct about the corrosion of the iron pieces in these solutions?

(A) In zinc sulphate it is slower than in aluminium sulphate.

(B) In copper sulphate it is faster than in silver nitrate.

(C) In silver nitrate it is faster than in copper sulphate.

(D) In aluminium sulphate it is slower than in zinc sulphate.

---

**34** When a magnesium ribbon is placed in silver nitrate solution, the following reaction takes place:

$Mg(s) + 2AgNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + 2Ag(s)$

Which of the following choices correctly describes what happened?

(A) Oxidation of magnesium and reduction of silver ions.

(B) Oxidation of magnesium and oxidation of silver.

(C) Reduction of magnesium and oxidation of silver.

(D) Reduction of magnesium and reduction of silver ions.

---

**35** When potassium dichromate solution acidified with sulphuric acid is added to iron(II) sulphate solution, the correct equation for the redox reaction that takes place is .....

(A) $Fe^{2+}(aq) + Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow Fe^{3+}(aq) + 2Cr^{3+}(aq) + 7H_2O(l)$

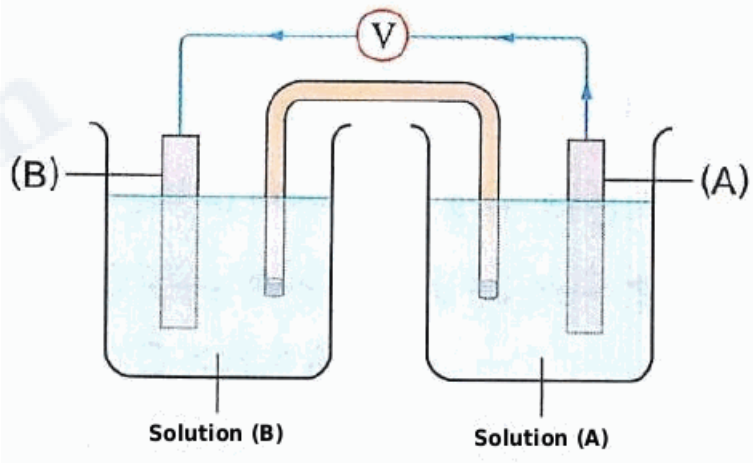
(B) $6Fe^{2+}(aq) + Cr_2O_7^{-}(aq) + 14H^+(aq) \rightarrow 6Fe^{3+}(aq) + 2Cr^{3+}(aq) + 7H_2O(l)$

(C) $Fe^{2+}(aq) + Cr_2O_7^{-}(aq) + 14H^+(aq) \rightarrow Fe^{3+}(aq) + 2Cr^{3+}(aq) + 7H_2O(l)$

(D) $6Fe^{2+}(aq) + Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow 6Fe^{3+}(aq) + 2Cr^{3+}(aq) + 7H_2O(l)$

---

**36** From the cell shown, which of the following is correct?




(A) The cell is galvanic and the concentration of solution (A) increases.

(B) The cell is galvanic and the concentration of solution (B) increases.

(C) The cell is electrolytic and the concentration of solution (A) decreases.

(D) The cell is electrolytic and the concentration of solution (B) decreases.

---

**37** The opposite reaction takes place in an electrochemical cell: $Sn(s) + 2Ag^{+}(aq) \rightarrow Sn^{2+}(aq) + 2Ag(s)$

The reaction represents .....

(A) a galvanic cell; electrons move from Ag to Sn

(B) an electrolytic cell; electrons move from $Sn^{2+}$ to Ag

(C) an electrolytic cell; electrons move from $Ag^{+}$ to Sn

(D) a galvanic cell; electrons move from Sn to $Ag^{+}$

---

**38** Which of the following changes increases the working time of the galvanic cell made of silver and nickel half-cells?

(A) Increasing the concentration of silver ions in the cathode half-cell.

(B) Decreasing the concentration of nickel ions in the anode half-cell.

(C) Decreasing the mass of the anode.

(D) Increasing the mass of the cathode.

---

**39** In the galvanic cell whose reaction is represented by the ionic equation: $Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$

When drops of HCl(aq) are added to each half-cell, which of the following is correct?

(A) The concentration of $Pb^{2+}(aq)$ ions increases.

(B) The emf of the cell increases.

(C) The life time of the battery decreases.

(D) The concentration of $Zn^{2+}(aq)$ ions decreases.

---

**40** When metal (X) was placed in a solution of the salt $YCl_2$, the concentration of $Y^{2+}$ cations changed from 0.1 M to 0.01 M. Which of the following is present in the solution?

(A) $Cl^-$ , $Y^{2+}$ and $X^{2+}$ ions

(B) $Cl^-$ and $X^{2+}$ ions only.

(C) $Cl^-$ and $Y^{2+}$ ions, and (X) settles at the bottom of the vessel.

(D) $Cl^-$ ions, and (X) and (Y) settle at the bottom of the vessel.

---

**41** Three rods of different elements (A , B , C) were placed in dilute HCl acid. (A , B) reacted and element (C) did not react. When element (A) was placed in a solution containing ions of element (B), it corroded. The order of these elements by their oxidation potentials is .....

(A) A > B > C (B) B > A > C (C) C > B > A (D) A > C > B

---

**42** A rod of element (A) is placed in a solution of ions of element (B). Given that element (A) is divalent and element (B) is monovalent, which of the following is correct?

(A) The number of moles of (A) dissolved is twice the number of moles of (B) deposited.

(B) The number of moles of (A) dissolved is half the number of moles of (B) deposited.

(C) The number of moles of (A) dissolved equals the number of moles of (B) deposited.

(D) The number of moles of (A) dissolved is three times the number of moles of (B) deposited.

---

**43** A piece of zinc was added to dilute sulphuric acid, then the gas produced was passed into four different solutions under the necessary conditions. Which of the following processes can take place?

(A) $YSO_4 \rightarrow Y_2(SO_4)_3$ (B) $WCl \rightarrow WCl_2$

(C) $X_2(SO_4)_3 \rightarrow XSO_4$ (D) $ZCl_2 \rightarrow ZCl_3$

**44** In the opposite reaction: $MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$
the changes that take place are .....

(A) $Mn^{4+} \rightarrow Mn^{2+}$ , $Cl_2 \rightarrow 2Cl^-$
(B) $Mn^{4+} \rightarrow Mn^{2+}$ , $2Cl^- \rightarrow Cl_2$
(C) $Mn^{2+} \rightarrow Mn^{4+}$ , $2Cl^- \rightarrow Cl_2$
(D) $Mn^{2+} \rightarrow Mn^{4+}$ , $Cl_2 \rightarrow 2Cl^-$

**45** Given that the divalent element (X) comes before the monovalent element (Y) in the electrochemical series, the overall reaction in the galvanic cell made of them is .....

(A) $X^{2+} + 2Y^+ \rightarrow X + 2Y$
(B) $X + 2Y^+ \rightarrow X^{2+} + 2Y$
(C) $2Y + X^{2+} \rightarrow 2Y^+ + X$
(D) $2Y + X \rightarrow 2Y^+ + X^{2+}$

**46** A galvanic cell is made of chromium and platinum electrodes. If the standard reduction potential of each is:
$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$ $E^\circ = -0.727$ V
$Pt^{2+}(aq) + 2e^- \rightarrow Pt(s)$ $E^\circ = +1.2$ V
then the overall reaction in the cell is .....

(A) $3Pt^{2+}(aq) + 2Cr^{3+}(aq) \rightarrow 3Pt(s) + 2Cr(s)$
(B) $Pt^{2+}(aq) + 2Cr(s) \rightarrow Pt(s) + 2Cr^{3+}(aq)$
(C) $Cr(s) + Pt^{2+}(aq) \rightarrow Cr^{3+}(aq) + Pt(s)$
(D) $2Cr(s) + 3Pt^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3Pt(s)$

**47** Given the potentials of the elements:
$K^+ \rightarrow K$ $E^\circ = -2.92$ V
$Cu^{2+} \rightarrow Cu$ $E^\circ = 0.34$ V
the overall reaction of the galvanic cell made of the two electrodes is .....

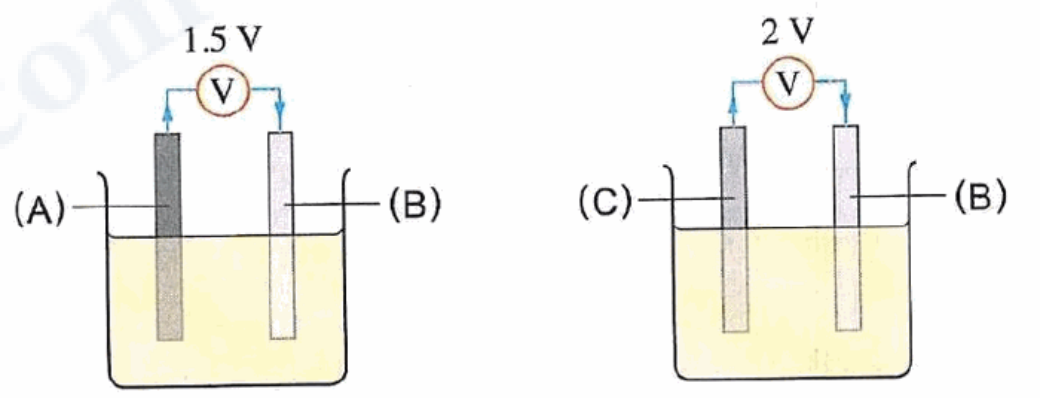
(A) $2K(s) + Cu^{2+}(aq) \rightarrow 2K^+(aq) + Cu(s)$
(B) $Cu(s) + 2K(s) \rightarrow Cu^{2+}(aq) + 2K^+(aq)$
(C) $Cu^{2+}(aq) + 2K^+(aq) \rightarrow Cu(s) + 2K(s)$
(D) $K^+(aq) + Cu(s) \rightarrow K(s) + Cu^{2+}(aq)$

**48** From the standard potentials of the two opposite half-cells:
$Al \rightarrow Al^{3+} + 3e^-$ $E^\circ = 1.67$ V
$Cu^{2+} + 2e^- \rightarrow Cu$ $E^\circ = 0.34$ V
What is the reaction in the galvanic cell made of them?

(A) $Al + Cu^{2+} \rightarrow Al^{3+} + Cu$
(B) $3Cu + 2Al^{3+} \rightarrow 3Cu^{2+} + 2Al$
(C) $2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$
(D) $Cu^{2+} + Al^{3+} \rightarrow Cu + Al$

**49** The following two figures represent two galvanic cells:

Given that (A) and (B) are divalent and (C) is trivalent, the overall reaction of the galvanic cell made of elements (A) and (C) is .....




(A) $2C(s) + 3A^{2+}(aq) \rightarrow 2C^{3+}(aq) + 3A(s)$

(B) $3A(s) + 2C^{3+}(aq) \rightarrow 3A^{2+}(aq) + 2C(s)$

(C) $2A(s) + 3C^{2+}(aq) \rightarrow 2A^{3+}(aq) + 3C(s)$

(D) $3C(s) + 2A^{3+}(aq) \rightarrow 3C^{2+}(aq) + 2A(s)$

---

**50** • The potential of a cell made of element (X) and the standard hydrogen electrode = 0.280 V

• The potential of a cell made of element (X) and element (Y) = 2.095 V

When element (Y) is placed in a solution of element (X), no reaction takes place.

Then the potential of the cell made of element (Y) and the standard hydrogen electrode equals .....

(A) −2.375 V (B) 2.375 V (C) 1.815 V (D) −1.815 V

---

**51** Which of the following represents the reaction of the galvanic cell made of silver and manganese electrodes, and the reducing agent in it?

(A) $Mn(s) + 2Ag^{+}(aq) \rightarrow Mn^{2+}(aq) + 2Ag(s)$ , the reducing agent is Mn

(B) $Ag(s) + Mn^{2+}(aq) \rightarrow Ag^{+}(aq) + Mn(s)$ , the reducing agent is Ag

(C) $Mn^{2+}(aq) + Ag^{+}(aq) \rightarrow Mn(s) + Ag(s)$ , the reducing agent is Mn

(D) $2Ag^{+}(aq) + Mn(s) \rightarrow 2Ag(s) + Mn^{2+}(aq)$ , the reducing agent is Ag

---

**52** The standard potentials of the two opposite half-cells:

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

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

What is the emf of the galvanic cell made of them?

(A) 1.639 V (B) 0.936 V (C) 0.396 V (D) 0.179 V

**53** In the galvanic cell whose electrodes are nickel and cadmium, given that:

$Cd(s) \rightarrow Cd^{2+}(aq) + 2e^{-}$ $E^{\circ} = +0.402$ V

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

the emf of the cell is .....

(A) 0.172 V (B) −0.632 V (C) 0.632 V (D) −0.172 V

**54** A cell made of elements (X) and (Y) has an emf of 0.94 V. Given that the standard oxidation potential of element (X) is 0.136 V and electrons move from (X) to (Y) through the wire, the oxidation potential of element (Y) equals .....

(A) +0.8 V (B) +1.076 V (C) −0.8 V (D) −1.076 V

**55** A galvanic cell whose electrodes are tin and silver. Given that the standard reduction potential of tin = −0.136 V and of silver = +0.8 V, which of the following represents the spontaneous reduction reaction in the cell?

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

(B) $2Ag(s) \rightarrow 2Ag^{+}(aq) + 2e^{-}$ $E^{\circ} = -1.6$ V

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

(D) $2Ag(s) \rightarrow 2Ag^{+}(aq) + 2e^{-}$ $E^{\circ} = +1.6$ V

**56** Given that:

$A \rightarrow A^{2+} + 2e^{-}$ $E^{\circ} = 0.409$ V

$B \rightarrow B^{+} + e^{-}$ $E^{\circ} = -0.800$ V

If a galvanic cell is made of elements (A) and (B), which of the following is correct?

(A) Anode reaction: $A \rightarrow A^{2+}$ , cathode reaction: $2B^{+} \rightarrow 2B$ , $E_cell = 1.209$ V

(B) Anode reaction: $2B^{+} \rightarrow 2B$ , cathode reaction: $A \rightarrow A^{2+}$ , $E_cell = 1.4$ V

(C) Anode reaction: $B^{+} \rightarrow B$ , cathode reaction: $2A \rightarrow 2A^{2+}$ , $E_cell = 0.896$ V

(D) Anode reaction: $2A \rightarrow 2A^{2+}$ , cathode reaction: $B^{+} \rightarrow B$ , $E_cell = 0.879$ V

**57** Standard oxidation potentials (volts) of four elements: (A): +2.711 , (B): +0.28 , (C): −1.2 , (D): −2.87

The table above gives the standard oxidation potentials of four elements (A), (B), (C) and (D).

The highest emf of a galvanic cell can be obtained from .....

(A) (B) anode , (D) cathode.

(B) (A) anode , (D) cathode.

(C) (D) anode , (C) cathode.

(D) (A) cathode , (D) anode.

---

**58** Which of the following represents a galvanic cell whose electrodes are chromium (Cr) and gold (Au), if the oxidation potential of chromium is (+0.74 V) and the reduction potential of gold is (+1.42 V)?

(A) Anode: $Cr(s) \rightarrow Cr^{3+}(aq) + 3e^-$ , cathode: $Au^{3+}(aq) + 3e^- \rightarrow Au(s)$ , E_cell = 2.16 V

(B) Anode: $Au^{3+}(aq) + 3e^- \rightarrow Au(s)$ , cathode: $Cr(s) \rightarrow Cr^{3+}(aq) + 3e^-$ , E_cell = 0.68 V

(C) Anode: $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$ , cathode: $Au(s) \rightarrow Au^{3+}(aq) + 3e^-$ , E_cell = 2.16 V

(D) Anode: $Au(s) \rightarrow Au^{3+}(aq) + 3e^-$ , cathode: $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$ , E_cell = 0.68 V

---

**59** $PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2Hg(l) + 2Cl^-(aq) \rightarrow Hg_2Cl_2(s) + PbSO_4(s) + 2H_2O(l)$

Given that: ($Pb^{2+}$ / $Pb^{4+}$ = −1.69 V , Hg / $Hg^+$ = −0.59 V)

The reaction above is .....

(A) non-spontaneous , emf = −1.1 V

(B) spontaneous , emf = 1.1 V

(C) non-spontaneous , emf = −2.28 V

(D) spontaneous , emf = 2.28 V

---

**60** From the opposite standard potentials:

$X^{2+} + 2e^- \rightarrow X$ E° = 1.5 V

$Y \rightarrow Y^+ + e^-$ E° = 0.75 V

$Z \rightarrow Z^{2+} + 2e^-$ E° = 2.32 V

$W^{3+} + 3e^- \rightarrow W$ E° = 1.4 V

Which of the following choices is correct?

(A) The reaction ($X^{2+} + 2Y \rightarrow X + 2Y^+$) represents a galvanic cell and emf = (+0.75 V).

(B) The reaction ($3Z + 2W^{3+} \rightarrow 3Z^{2+} + 2W$) represents an electrolytic cell and emf = (−3.44 V).

(C) The reaction ($Z + X^{2+} \rightarrow Z^{2+} + X$) represents a galvanic cell and emf = (+3.82 V).

(D) The reaction ($3Y + W^{3+} \rightarrow 3Y^+ + W$) represents an electrolytic cell and emf = (−2.15 V).

---

**61** The following reaction takes place in an electrochemical cell: $Ni(s) + Fe^{2+}(aq) \rightarrow Ni^{2+}(aq) + Fe(s)$

Given that (oxidation potential of Ni = +0.23 V , oxidation potential of Fe = +0.4 V), which of the following is correct?

(A) The reaction is spontaneous , emf = −0.17 V

(B) The reaction is spontaneous , emf = +0.17 V

(C) The reaction is non-spontaneous , emf = −0.17 V

(D) The reaction is non-spontaneous , emf = +0.17 V

---

**62** The following reactions take place in galvanic cells under standard conditions:

$X + Y^{2+} \rightarrow X^{2+} + Y$ emf = 0.351 V

$Y + Z^{2+} \rightarrow Y^{2+} + Z$ emf = 0.749 V

From the reactions above, the emf of the cell $Z + X^{2+} \rightarrow Z^{2+} + X$ is .....

(A) −1.1 V (B) 1.1 V (C) 0.398 V (D) −0.398 V

---

**63** In a Daniell cell, when the zinc half-cell is replaced with a silver half-cell, which of the following is correct? Given the oxidation potentials of Ag and Zn: [E°(Ag) = −0.8 V , E°(Zn) = 0.76 V]

(A) The emf decreases and the direction of the current does not change.

(B) The emf increases and the direction of the current changes.

(C) The emf decreases and the direction of the current changes.

(D) The emf increases and the direction of the current does not change.

---

**64** Given the opposite standard potentials:

$X^{2+}(aq) + 2e^- \rightarrow X(s)$ E° = +0.34 V

$Y^{2+}(aq) + 2e^- \rightarrow Y(s)$ E° = −0.76 V

$Z^{2+}(aq) + 2e^- \rightarrow Z(s)$ E° = −2.375 V

If you have a primary galvanic cell made of two electrodes (X) and (Y), and half-cell (X) is replaced with half-cell (Z) under suitable conditions, which of the following choices is correct?

(A) The direction of the electric current changes and the emf decreases

(B) The direction of the electric current does not change and the emf decreases

(C) The direction of the electric current changes and the emf increases

(D) The direction of the electric current does not change and the emf increases

# Answer key and solutions

## Multiple choice questions

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

**1** **Answer (D)**

In a redox reaction the oxidation numbers of two elements change.

(2) Oxygen in $H_2O_2$ changes from -1 to -2 and 0.

(4) Al is oxidized and Fe is reduced.

(1) and (3) are precipitation and neutralization reactions with no change in oxidation numbers.

**2** **Answer (C)**

Mg is oxidized from 0 to +2 and $H^+$ is reduced to $H_2$.

The other equations are precipitation, combination or neutralization with no change in oxidation numbers.

**3** **Answer (B)**

In a balanced equation the atoms and the charges are equal on both sides.

$Pb + 2H^+ \rightarrow Pb^{2+} + H_2$ : the charge is +2 on both sides.

**4** **Answer (C)**

Oxidation is loss of electrons.

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

**5** **Answer (B)**

$Fe^{2+}$ is oxidized to $Fe^{3+}$ by acidified $KMnO_4$ solution.

$Mn^{7+}$ is reduced to $Mn^{2+}$, so oxidation and reduction take place.

(Note: $AgNO_3$ solution can also oxidize $Fe^{2+}$; the key is (B).)

**6** **Answer (A)**

Al changes to $Al(OH)_4^-$, so its oxidation number increases from 0 to +3.

Al is the one that is oxidized.

**7** **Answer (C)**

Pb loses electrons and is oxidized.

The substance that is oxidized is the reducing agent.

**8** **Answer (B)**

$Cl_2$ gains electrons and changes to $Cl^-$.

The substance that is reduced is the oxidizing agent.

**9** **Answer (C)**

Reduction is gain of electrons.

$Mn^{3+} + e^- \rightarrow Mn^{2+}$ gives $Mn^{2+}$, which is a stable ion because its configuration ($3d^5$) is half-filled.

$Ti^{4+}$ is more stable than $Ti^{3+}$, so this reduction does not form a stable ion.

**10** **Answer (B)**

The reducing agent is oxidized, so its oxidation number increases.

The only increase is -2 → -1

**11** **Answer (D)**

$Mn^{2+}$ changes to $MnO_4^-$, so the oxidation number of Mn increases from +2 to +7.

$Mn^{2+}$ is oxidized, so it is the reducing agent.

**12** **Answer (A)**

$Cr_2O_7^{2-}$ gains 6 electrons (Cr from +6 to +3 for each atom).

Each $Fe^{2+}$ loses one electron.

x = 6

**13** **Answer (A)**

$MnO_4^-$ is reduced in an acidic medium to $Mn^{2+}$.

$Mn^{7+} + 5e^- \rightarrow Mn^{2+}$

**14** **Answer (D)**

In HI: the oxidation number of iodine is -1 and it becomes 0 in $I_2$.

In $H_2SO_4$: the oxidation number of S is +6 and it becomes +4 in $SO_2$.

**15** **Answer (D)**

$CO_2 + C \rightarrow 2CO$ , so compound (X) is CO.

CO is a reducing agent, so it cannot oxidize $SO_3^{2-}$ to $SO_4^{2-}$.

(D) is an oxidation reaction that does not take place with CO.

**16** **Answer (B)**

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

$[Zn^{2+}]$ starts from zero and increases with time.

It becomes constant when the $Cu^{2+}$ ions are used up.

**17** **Answer (C)**

$Sc + Fe^{3+} \rightarrow Sc^{3+} + Fe$ (moles of consumed substance = moles of deposited substance).

The mass of Sc consumed $(45\ g/mol)$ is less than the mass of Fe deposited $(56\ g/mol)$.

So statement (C) is incorrect.

**18** **Answer (C)**

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

The zinc is coated with copper, its mass decreases and the blue colour of the solution fades.

Electrical energy is produced only in a galvanic cell with an external circuit.

**19** **Answer (C)**

Fe loses electrons (is oxidized) and $Co^{2+}$ gains them (is reduced).

**20** **Answer (B)**

The cathode is the electrode at which reduction takes place.

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

**21** **Answer (D)**

Reduction takes place at the cathode.

The cations of the cathode electrolyte gain the electrons.

**22** **Answer (C)**

The anode is the more active electrode, which is oxidized.

Mg is more active than Cu, so it is the anode.

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

**23** **Answer (C)**

The ions move between the two half-cells through the salt bridge.

The electrons move in the external circuit.

**24** **Answer (B)**

$Fe^{2+}$ is reduced to Fe, so the iron electrode is the cathode.

Zn is oxidized, so it is the anode.

**25** **Answer (D)**

A half-reaction is balanced by atoms and charges.

$IO_3^- + 6H^+ + 5e^- \rightarrow \frac{1}{2}I_2 + 3H_2O$

The charge is zero on both sides.

**26** **Answer (B)**

Zn is the anode; it is oxidized, so $[Zn^{2+}]$ increases.

$Cu^{2+}$ is reduced at the cathode, so $[Cu^{2+}]$ decreases.

**27** **Answer (C)**

Zn is more active than Cu, so it is the anode, and Cu is the cathode.

$Cu^{2+} + 2e^- \rightarrow Cu$ , so the mass of the copper electrode increases.

Cations move towards the cathode (copper).

**28** **Answer (D)**

Electrons leave electrode $(1)$, so it is the anode (Fe).

Anions move to the anode $(2)$ and $K^+$ cations to the cathode $(3)$.

**29** **Answer (D)**

Electrons leave the left electrode, so it is the anode.

$(1)$ is the anode cation $(Zn^{2+})$ and $(2)$ is the $SO_4^{2-}$ anion moving to the anode.

$(3)$ is the bridge cation $K^+$ moving to the cathode.

**30** **Answer (C)**

The colour of $A^{2+}$ fades: $A^{2+}$ is reduced, so electrode (A) is the cathode.

The colour of $B^{2+}$ deepens: B is oxidized, so it is the anode.

Anions move to the anode half-cell (B).

**31** **Answer (A)**

Anions move to the anode and cations to the cathode.

(A) is the anode, so its atoms are oxidized and its mass decreases.

**32** **Answer (C)**

M is the anode (negative electrode) and Cu is the cathode.

Electrons flow from M to Cu.

$2M + 3Cu^{2+} \rightarrow 2M^{3+} + 3Cu$

**33** **Answer (C)**

A loses two electrons and $B^{3+}$ gains 3 electrons.

The common multiple is 6 electrons: 3A and $2B^{3+}$

$3A + 2B^{3+} \rightarrow 3A^{2+} + 2B$

**34** **Answer (D)**

The salt bridge balances the charges by moving both anions and cations.

**35** **Answer (B)**

The salt bridge connects the solutions of the two half-cells.

$Ni^{2+}(aq)$ with $Pb^{2+}(aq)$

**36** **Answer (B)**

The porous partition lets ions pass and does not react with the ions of the two solutions.

(B) is not one of its roles.

**37** **Answer (D)**

$Na_2CO_3$ gives the $CO_3^{2-}$ ion, which precipitates $Zn^{2+}$ and $Cu^{2+}$.

The ion concentrations change, so the emf changes.

The other salts do not react with the ions of the cell.

**38** **Answer (A)**

Fifth-group cations ($Ba^{2+}$ , $Sr^{2+}$ , $Ca^{2+}$) precipitate their sulphates.

The cations move to the cathode (positive) half-cell, where some sulphate precipitates, so $[SO_4^{2-}]$ decreases.

**39** **Answer (A)**

The potential of a single half-cell cannot be measured alone.

No electrons flow from or to it, so it is an open circuit.

**40** **Answer (C)**

Standard hydrogen electrode: $H^+$ ions (1 M), $H_2$ gas (1 atm) and a platinum plate.

There are no separate H atoms in it.

**41** **Answer (D)**

E° of the cell = E° reduction of $H^+$ - E° reduction of $Fe^{2+}$

0 - (-0.44) = +0.44 V

**42** **Answer (C)**

emf = E° reduction of $M^+$ - E° reduction of $H^+$ = 0.8 - 0 = 0.8 V

Reduction potential of $M^+$ = +0.8 V

Oxidation potential = -0.8 V

**43** **Answer (A)**

Copper displaces only the ions that are less active than it.

$Ag^+$ has a reduction potential of +0.8 V, greater than $Cu^{2+}$ (+0.34 V).

$Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$ , so the blue colour of $Cu^{2+}$ appears.

**44** **Answer (A)**

Fe is more active than Cu, so it replaces it.

The reaction is spontaneous; the mass of Fe decreases and the colour of the solution changes.

**45** **Answer (B)**

Cu is more active than Ag, so it is the anode.

The copper electrode loses electrons and is oxidized.

**46** **Answer (C)**

emf = oxidation potential of the anode - oxidation potential of the cathode.

The more active metal (Fe) has a higher oxidation potential than Au.

**47** **Answer (C)**

emf = reduction potential of the cathode - reduction potential of the anode.

$Fe^{3+}/Fe^{2+}$ is the cathode $(0.77)$ and $Cu^{2+}/Cu$ the anode $(0.34)$

$0.77 - 0.34 = 0.43$ V

**48** **Answer (A)**

emf = $1.09 - (-0.727) = 1.817$ V

**49** **Answer (C)**

emf $(3)$ = emf $(1)$ - emf $(2)$

$1.18 - 0.16 = 1.02$ V

**50** **Answer (D)**

X displaces the hydrogen of the acid, so its reduction potential is negative.

emf = reduction potential of Y - reduction potential of X

$2.46 = 0.8$ - reduction potential of X

Reduction potential of X = -1.66 V

**51** **Answer (B)**

The standard hydrogen potential is measured at $[H^+] = 1$ M.

0.5 M $H_2SO_4$ gives $[H^+] = 1$ M, so the potential does not change.

The other concentrations give $[H^+]$ different from 1 M.

**52** **Answer (B)**

$H_2$ is evolved at the hydrogen electrode (the cathode): $2H^+ + 2e^- \rightarrow H_2$

$[H^+]$ decreases and $[OH^-]$ increases, so the pH increases and the pOH decreases.

**53** **Answer (C)**

Cell $(1)$: Zn is more active than hydrogen, so hydrogen is the cathode.

Cell $(2)$: Cu is less active than hydrogen, so hydrogen is the anode.

**54** **Answer (B)**

The strongest reducing agent has the lowest reduction potential.

$K^+$ has a potential of -2.93 V, the lowest.

Its reduced form K is the strongest reducing agent.

**55** **Answer (B)**

The reduction potential of $X^{2+}$ is negative $(-0.13\ V)$.

X is more active than hydrogen, so it displaces it and loses electrons.

X is a reducing agent.

**56** **Answer (D)**

**57** **Answer (C)**

$(1)$: X is a stronger oxidizing agent than $Y^{+}$.

$(2)$: $Y^{+}$ is a stronger oxidizing agent than $Z^{+}$.

$(3)$: X is stronger than $Z^{+}$.

The order: $Z < Y^{+} < X$

**58** **Answer (D)**

Only (A) and (C) react with HCl, so they are more active than hydrogen.

(C) reacts with the salts of all the others, so it is the most active.

(D) displaces (B), so it is more active than it.

The more active metal is the stronger reducing agent: B < D < A < C

**59** **Answer (B)**

Zn has the lowest reduction potential, so it is the strongest reducing agent.

Zn can reduce $I_2$ and $Fe^{2+}$ (their reduction potentials are higher).

**60** **Answer (B)**

The reduction potential of $B^{+}$ $(+0.8)$ is greater than that of $A^{2+}$ $(-1.029)$.

A is more active, so it displaces B and B is deposited on A.

**61** **Answer (A)**

The ease of oxidation from +2 to +3 increases as the reduction potential decreases.

The reduction potentials: Cr < Fe < Mn < Co.

The hardest to oxidize is $Co^{2+}$, then $Mn^{2+}$, then $Fe^{2+}$, then $Cr^{2+}$.

**62** **Answer (A)**

The reducing agent for $Mn^{3+}$ must have a reduction potential lower than -1.029 V.

Only Mg has a potential $(-2.375\ V)$ lower than -1.029 V.

**63** **Answer (A)**

$Mn^{7+}$ has a reduction potential of +1.51 V.

It oxidizes the metals with a lower reduction potential: Zn (-0.76) and Ag (+0.8).

$Au^{+}$ has a potential of +1.69 V, which is greater, so it is not oxidized.

**64** **Answer (B)**

In a spontaneous process the more active metal is the reducing agent.

(1) Fe is more active than Ni: spontaneous.

(2) Ni is more active than Pb: spontaneous.

(3) Pb is less active than Fe: non-spontaneous.

**65** **Answer (D)**

**66** **Answer (C)**

A metal from which electrons move to Cu is more active than Cu.

Q is less active than Cu.

Potential difference: Y is the largest, then X, then Z.

Q < Z < X < Y

**67** **Answer (C)**

Zn is more active than Cu, so it displaces it and Cu is deposited.

Zn is less active than Mg, so it does not displace it.

After the reaction: $Zn^{2+}$ and $Mg^{2+}$ only.

**68** **Answer (D)**

X does not displace Z: Z is more active than X.

Y is deposited on X and on Z: X and Z are more active than Y.

Y < X < Z

**69** **Answer (C)**

**70** **Answer (B)**

Metal (A) gives a blue colour in solution: $Cu^{2+}$ , so the metal is copper.

Solution (B) contains $Ag^{+}$ : $Cu + 2Ag^{+} \rightarrow Cu^{2+} + 2Ag$

$Ag^{+} + Cl^{-} \rightarrow AgCl$ , which dissolves in concentrated ammonia.

**71** **Answer (C)**

(B) is not displaced by any metal: it is the most active metal.

(A) does not displace (E), so A is less active than E.

The order in (C): D < A < E < C < B is the one that agrees with the table.

**72** **Answer (B)**

The salt of metal (Z) used up all the nickel, so (Z) is more active than it and the others.

Y comes before nickel in the potentials, so nickel deposits on it, and Y displacing Z from its salt solution is possible.

Metal Y displaces Z from its salt solution.

## Essay questions

**1** **Model answer**

(1) It is not a redox reaction, because the oxidation numbers of Zn, S, O and H do not change between the reactants and the products.

(2) The colour of the solution changes from orange to green.

**2** **Model answer**

(1) A complete reaction, because the $H_2$ gas produced escapes from the open reaction vessel.

(2) • The colour of the copper(II) sulphate solution changes from blue to colourless.

• A layer of copper is deposited on the surface of the zinc plate.

**3** **Model answer**

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

Because Ni comes before Ag in the electrochemical series (the oxidation potential of Ni is greater than that of Ag), so the Ni electrode is the anode and the Ag electrode is the cathode.

(2) emf = 0.23 − (−0.8) = 1.03 V

**4** **Model answer**

(1) $Cu^{2+}$ ions, because the oxidizing agent is the one that is reduced.

(2) emf = 0 + 0.34 = 0.34 V

**5** **Model answer**

(1) emf = 0.7 − (−0.4) = 1.1 V

(2) $2A(s) + B^{2+}(aq) \rightarrow 2A^+(aq) + B(s)$

**6** **Model answer**

The equation shows that $A^{4+}$ ions are reduced and $B^{2+}$ ions are oxidized.

emf = 0.01 − 0.77 = −0.76 V

∴ The reaction is non-spontaneous, because the emf is negative.

**7** **Model answer**

(1) Magnesium acts as the anode in this cell, because the oxidation potential of the magnesium electrode is greater than that of the standard hydrogen electrode.

(2) • The direction of electron flow in the external wire changes.

• The voltmeter reading changes (the emf decreases).

**8** **Model answer**

(1) (a) Direction (1). (b) Direction (3).

Because the oxidation potential of cadmium is greater than that of copper, cadmium acts as the anode and copper as the cathode.

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

**9** **Model answer**

• Electrode (1): silver. • Electrode (2): aluminium. • Solution (3): $AgNO_3$

• Voltmeter reading: E_cell = 1.67 − (−0.8) = 2.47 V

**10** **Model answer**

Because metals with higher oxidation potentials displace metals with lower oxidation potentials from solutions of their salts.

Order of the metals: (B) < (C) < (E) < (A) < (D)

**11** **Model answer**

(1) (a) Rod (X) corrodes because its atoms are oxidized.

(b) The blue colour of the $Cu^{2+}$ solution gradually fades until it becomes colourless, because the ions of metal (X) replace its ions.

(2) Copper < metal (X) < metal (Z) < metal (Y)

**12** **Model answer**

Order of the metals by chemical activity: R > P > Q

• In the cell made of metals (P) and (R), metal (P) acts as the cathode, because the oxidation potential of metal (P) is lower than that of metal (R).

• In the cell made of metals (P) and (Q), metal (P) acts as the anode, because the oxidation potential of metal (P) is greater than that of metal (Q).

**13** **Model answer**

• Curve (3) represents copper, because copper does not react with dilute hydrochloric acid.

• Curve (1) represents magnesium and curve (2) represents iron, because the oxidation potential of magnesium is greater than that of iron, so the rate of reaction of magnesium with the acid is greater than that of iron.

## Past exam questions

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

**1** **Answer (B)**

$Na_2S$ precipitates $Cu^{2+}$ as CuS.

$[Cu^{2+}]$ at the cathode decreases, so the emf decreases and the life time of the cell decreases.

**2** **Answer (A)**

emf = reduction potential of $MnO_4^-$ + oxidation potential of $Cl^-$ = 1.52 - 1.36 = +0.16 V

The emf is positive, so the reaction is spontaneous.

**3** **Answer (C)**

Oxidation potential of X = 0.80 - 0.8 = 0 V

Oxidation potential of Y = 1.56 - 0.8 = 0.76 V

Cell (X , Y): Y is the anode and emf = 0.76 - 0 = +0.76 V

**4** **Answer (C)**

**5** **Answer (A)**

Electrons go to the electrode at which reduction takes place.

(Y) is the cathode, so the reduction potential of (X) is smaller than that of (Y).

**6** **Answer (B)**

Electrode Y has the higher oxidation potential, so it is the anode.

emf = 1.66 - 0.76 = +0.9 V

**7** **Answer (C)**

From the two cells: A has the highest oxidation potential, then C, then B.

With B as anode and A as cathode: emf = oxidation potential of B - oxidation potential of A

emf = −2 V , so the reaction is non-spontaneous.

**8** **Answer (C)**

Cu does not react with $Ni^{2+}$ ions because it is less active than Ni.

Nickel salt solutions can be kept in a copper container.

**9** **Answer (B)**

In a spontaneous reaction the more active metal is the reducing agent.

Mg is more active than Zn, so it displaces it.

**10** **Answer (D)**

An increase in $[Y^{2+}]$ means metal (Y) is oxidized, so it is the anode.

Electrode (X) is the cathode and $X^{2+}$ is reduced, so it is an oxidizing agent.

**11** **Answer (D)**

Cd does not displace $Zn^{2+}$: $Cd^{2+}$ is a stronger oxidizing agent than $Zn^{2+}$.

Cd displaces $Cu^{2+}$: $Cu^{2+}$ is stronger than $Cd^{2+}$.

$Cu^{2+} > Cd^{2+} > Zn^{2+}$

**12** **Answer (C)**

(X) does not dissolve in HCl, so it is less active than hydrogen.

(Z) dissolves in HCl and (Y) displaces (Z) from its salt solution, so it is more active than it.

Y > Z > H > X

The greatest emf comes from the largest difference in activity: (X) , (Y)

**13** **Answer (A)**

Oxidation potential of Cr = +0.74 V and reduction potential of $Cu^{2+}$ = +0.340 V

emf = 0.74 + 0.34 = +1.08 V

The emf is positive, so the cell is galvanic.

**14** **Answer (C)**

Zn is more active than Cu, so zinc is the anode (electrode 1) and copper is the cathode (electrode 4).

Electrons move spontaneously from (1) to (4).

**15** **Answer (C)**

(X) is the stronger reducing agent, so it is the anode.

Negative anions move from the bridge to the anode half-cell (X).

**16** **Answer (C)**

Y is oxidized from 0 to +3, so it is the anode.

The emf is positive, so the cell is galvanic.

**17** **Answer (C)**

The greatest emf comes from the reaction of the strongest reducing agent with the strongest oxidizing agent.

L (oxidation potential +1.18) with $Y^{2+}$ (reduction potential +1.5)

emf = 1.18 + 1.5 = 2.68 V

**18** **Answer (D)**

$SO_4^{2-}$ anions move to the anode (X).

$3X + 2Y^{3+} \rightarrow 3X^{2+} + 2Y$

To balance the number of electrons (6).

**19** **Answer (C)**

At the anode of an electrolytic cell a non-spontaneous oxidation takes place.

$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$

**20** **Answer (D)**

Cell (X): electrons move from (B) to (A), so (B) is the anode.

The reaction in cell (Y) is the reverse of the spontaneous one: electrolytic with negative emf.

(B) is the cathode in the electrolytic cell, so it is the negative electrode.

**21** **Answer (A)**

Cell 1: X is the anode, so its oxidation potential is +0.23 V

Cell 2: Y is the cathode, so its reduction potential is +0.8 V

emf = 0.23 + 0.8 = 1.03 V , X is the anode and Y the cathode.

**22** **Answer (A)**

Fe is the anode (oxidation potential 0.409 V) and Sn is the cathode (reduction potential 0.150 V).

emf = 0.409 + 0.150 = +0.559 V

**23** **Answer (D)**

emf = 0.23 + 0.8 = +1.03 V

The emf is positive, so the cell is galvanic.

**24** **Answer (C)**

The bridge needs a salt that does not react with the ions of the cell.

NaCl precipitates $Pb^{2+}$ as $PbCl_2$ and $Na_2CO_3$ precipitates it as $PbCO_3$.

Sodium nitrate does not react.

**25** **Answer (D)**

emf = oxidation potential of X + reduction potential of $Cu^{2+}$

1.8 = oxidation potential of X + 0.34

Oxidation potential of X = 1.46 V

Reduction potential of X = -1.46 V

**26** **Answer (C)**

The anode (X) has an oxidation potential of 0.76 V and the cathode (Y) -0.34 V.

W has an oxidation potential of 2.37 V, greater than X.

When the cathode is replaced with electrode W, W becomes the anode and X the cathode, so the current reverses and the emf increases.

**27** **Answer (B)**

Any two elements form a cell: number of cells = 4×3÷2 = 6 cells.

The highest emf is between the highest oxidation potential (A = +2.71 V) and the lowest oxidation potential (D = -1.20 V).

emf = 2.71 + 1.20 = 3.91 V

**28** **Answer (A)**

emf = oxidation potential of M + reduction potential of $Zn^{2+}$

1 = oxidation potential of M + (-0.76)

Oxidation potential of M = +1.76 V

M is the anode and the cell is galvanic.

**29** **Answer (C)**

Oxidation potential of X = +0.41 V and reduction potential of $Y^{2+}$ = +0.23 V

emf = 0.41 + 0.23 = +0.64 V , spontaneous.

X is oxidized, so the mass of electrode (X) decreases.

**30** **Answer (A)**

emf = oxidation potential of W - oxidation potential of X = 2.37 - 0.76 = 1.61 V

When the cathode X is replaced with electrode Z (reduction potential +0.23 V):

emf = 2.37 + 0.23 = 2.6 V , it increases and the direction of the current does not change.

**31** **Answer (D)**

emf = reduction potential of $X^{2+}$ - reduction potential of $Y^{+}$ = -0.56 - 0.23 = -0.79 V

The emf is negative, so the cell is electrolytic.

**32** **Answer (B)**

emf = -0.23 - 0.402 = -0.632 V

The emf is negative, so the reaction is non-spontaneous.

Y is oxidized, so it is the anode.

**33** **Answer (C)**

$Ag^{+}$ ions have a reduction potential of 0.8 V, greater than $Cu^{2+}$ (0.34 V).

Fe corrodes in silver nitrate faster than in copper sulphate.

$Zn^{2+}$ and $Al^{3+}$ ions do not corrode iron.

**34** **Answer (A)**

Mg is more active than Ag, so Mg is oxidized and $Ag^{+}$ ions are reduced.

**35** **Answer (D)**

$Cr_2O_7^{2-}$ carries a charge of 2-; each $Fe^{2+}$ loses one electron while $Cr_2O_7^{2-}$ gains 6 electrons.

$6Fe^{2+} + Cr_2O_7^{2-} + 14H^{+} \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$

**36** **Answer (A)**

Electrons leave electrode (A), so it is the anode and its atoms are oxidized.

The cell is galvanic and the concentration of solution (A) increases.

**37** **Answer (D)**

Sn loses electrons, which move to $Ag^+$ spontaneously.

The cell is galvanic.

**38** **Answer (A)**

The cell stops when the $Ag^+$ ions at the cathode are used up.

Increasing their concentration increases the working time of the cell.

**39** **Answer (C)**

HCl adds $Cl^-$ ions, which precipitate $Pb^{2+}$ as $PbCl_2$.

$[Pb^{2+}]$ decreases, so the emf decreases and the life time of the battery decreases.

**40** **Answer (A)**

(X) displaces $Y^{2+}$, so $[Y^{2+}]$ decreases from 0.1 to 0.01 M without becoming zero.

The solution contains $Cl^-$ , $Y^{2+}$ and $X^{2+}$.

**41** **Answer (A)**

A reacts with HCl and B reacts with HCl , C does not react: C is the least active.

A corrodes in B ions: A is more active than B.

A > B > C

**42** **Answer (B)**

$A^{2+}$ and $2B^+$ : $A + 2B^+ \rightarrow A^{2+} + 2B$

Each mole of A gives 2 moles of B.

The moles of A dissolved are half the moles of B deposited.

**43** **Answer (C)**

Hydrogen gas is a reducing agent: its oxidation number increases.

The only reaction in which reduction takes place: $X_2(SO_4)_3 \rightarrow XSO_4$

**44** **Answer (B)**

Mn changes from +4 to +2 (reduction).

$Cl^-$ is oxidized to $Cl_2$.

**45** **Answer (B)**

X comes before Y in the series, so it is the anode and is oxidized to $X^{2+}$.

$X + 2Y^{+} \rightarrow X^{2+} + 2Y$

**46** **Answer (D)**

The reduction potential of $Cr^{3+}$ is lower than that of $Pt^{2+}$, so chromium is the anode.

$2Cr + 3Pt^{2+} \rightarrow 2Cr^{3+} + 3Pt$

**47** **Answer (A)**

K is much more active than Cu, so K is oxidized and $Cu^{2+}$ is reduced.

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

**48** **Answer (C)**

Al has an oxidation potential of +1.67 V, so it is the anode.

$2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$

**49** **Answer (A)**

**50** **Answer (C)**

X displaces the hydrogen of the acid: oxidation potential of X = +0.280 V

Y does not displace X, so the oxidation potential of Y = 0.28 - 2.095 = -1.815 V

emf $(Y, H_2) = 0 - (-1.815) = 1.815$ V

**51** **Answer (A)**

Mn is more active than Ag, so it is the anode.

$Mn + 2Ag^{+} \rightarrow Mn^{2+} + 2Ag$

The reducing agent is Mn.

**52** **Answer (D)**

emf $= 0.409 + (-0.23) = 0.179$ V

**53** **Answer (A)**

emf $= 0.402 + (-0.23) = 0.172$ V

**54** **Answer (C)**

Electrons move from X to Y: X is the anode.

emf = oxidation potential of X - oxidation potential of Y

0.94 = 0.136 - oxidation potential of Y

Oxidation potential of Y = -0.8 V

**55** **Answer (A)**

The reduction potential of silver (+0.8 V) is greater than that of tin (-0.136 V).

Silver is the cathode, so $Ag^+$ is reduced with a potential of +0.8 V.

**56** **Answer (A)**

A: oxidation potential 0.409 V and B: oxidation potential -0.8 V, so the anode is A.

emf = 0.409 + 0.8 = 1.209 V

**57** **Answer (B)**

The highest emf comes from the highest oxidation potential (A = +2.711 V) as the anode and the lowest oxidation potential (D = -2.87 V) as the cathode.

**58** **Answer (A)**

Cr is the anode (oxidation potential +0.74 V) and Au is the cathode (reduction potential +1.42 V).

emf = 0.74 + 1.42 = 2.16 V

**59** **Answer (B)**

The reduction potential of $PbO_2$ equals +1.69 V and the oxidation potential of Hg = -0.59 V.

emf = 1.69 + (-0.59) = +1.1 V , spontaneous.

**60** **Answer (C)**

$Z + X^{2+} \rightarrow Z^{2+} + X$

emf = 2.32 + 1.5 = +3.82 V

The cell is galvanic.

**61** **Answer (C)**

emf = oxidation potential of Ni + reduction potential of $Fe^{2+}$

Reduction potential of $Fe^{2+}$ = -0.4 V

emf = 0.23 - 0.4 = -0.17 V

The reaction is non-spontaneous.

**62** **Answer (A)**

The emf (X , Z) is found by adding the two reactions above and reversing the result.

$X + Z^{2+} \rightarrow X^{2+} + Z$ : 0.351 + 0.749 = 1.1 V

Reversing the reaction: -1.1 V

**63** **Answer (C)**

In a Daniell cell Zn is the anode.

Ag (oxidation potential -0.8 V) is lower than Cu (oxidation potential -0.34 V), so Cu becomes the anode.

The direction of the current reverses and the emf decreases to 0.46 V.

**64** **Answer (C)**

X: +0.34 V and Y: -0.76 V in the original cell.

Z is more active than Y, so Z becomes the anode and the direction of the current reverses.

emf = 2.375 - 0.76 = 1.615 V , greater than 1.1 V