Electrochemistry Review

Knowledge

30–B1.1k Define oxidation and reduction operationally and theoretically.

Use the following information to answer the next question.

Statements

I	Sulfur forms stable ions by gaining electrons.
II	Magnesium forms stable ions by losing electrons.
III	The oxidation number of iron changes from +3 to +2.
IV	The oxidation number of oxygen changes from –2 to –1.

1. The statements numbered above that refer to oxidation are
   A. I and III **B.** I and IV **C.** II and III **D.** II and IV

Use the following information to answer the next question.

Leaching technology is used in the mining and refining of copper ore. In the first step of the leaching process, concentrated aqueous sulfuric acid flows through a copper ore deposit. Solid copper(II) oxide reacts with sulfuric acid as represented by the following net ionic equation.

$$CuO(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l)$$

The resulting solution that contains copper(II) ions is transferred to an electrolytic cell where pure copper is produced.

2. In the reaction represented by the equation above, copper undergoes
   A. reduction only **B.** oxidation only
   C. both oxidation and reduction **D.** neither oxidation nor reduction

3. Which of the following statements is an operational definition of the metal undergoing reduction?
   A. Iron metal undergoes a formation reaction with oxygen gas.
   B. Magnesium metal increases in mass when heated in air.
   C. Iron (III) hydroxide reacts with oxygen in the air to form ionic compounds.
   D. Zinc sulfide ore is roasted in the presence of oxygen gas to produce zincmetal.

Use the following information to answer the next question.

The following reaction will occur at high temperatures.

$$2Na_{(g)} + Cl_{2(g)} \rightarrow 2NaCl_{(g)} + \text{energy}$$

4. The half-reaction for the reduction that occurs in this reaction is

A. $Na(g) \rightarrow Na^{+}(g) + e^{-}$

B. $Na(g) + e^{-} \rightarrow Na^{+}(g)$

C. $Cl_2(g) + 2e^{-} \rightarrow 2Cl^{-}(g)$

D. $Cl_2(g) \rightarrow 2Cl^{-}(g) + 2e^{-}$

Knowledge
30–B1.2k Define oxidizing agent, reducing agent, oxidation number, half-reaction, disproportionation.

5. Define the following terms:
   a. Oxidizing agent

   b. Reducing agent

   c. Oxidation number

   d. disproportionation

Use the following equation to answer the next question.

$$2\,H_2S(g) + 3\,O_2(g) \rightarrow 2\,SO_2(g) + 2\,H_2O(g)$$

6. In the reaction represented by the equation above, oxygen acts as the **__i__** agent, and the oxidation number of the sulfur atom increases by__ **ii__** .
   The statement above is completed by the information in row

ROW	i	ii
A.	oxidizing	2
B.	oxidizing	6
C.	reducing	2
D.	reducing	6

7. Which of the following equations represent a disproportionation reaction?
   A. $2Na_{(s)} + I_{2(s)}$ ––> $2NaI_{(s)}$
   B. $2F_{2(g)} + O_{2(g)}$ ––> $2OF_{2(g)}$
   C. $Cl_{2(aq)} + H_2O_{(l)}$ ––> $HOCl_{(aq)} + H^+_{(aq)} + Cl^-_{(aq)}$
   D. $2NH_{3(aq)} + NaOCl_{(aq)}$ ––> $N_2H_{4(aq)} + NaCl_{(aq)} + H_2O_{(l)}$

8. In a reaction, $Sn^{2+}(aq)$
   A. will undergo oxidation when combined with $Pb(NO_3)_2(aq)$
   B. act as a reducing agent when combined with Ni(s)
   C. always act as an oxidizing agent
   D. act as an oxidizing agent when combined with Cd(s)

9. A redox reaction occurs when an iron nail is placed in a solution of copper(II) sulphate. Elemental copper begins to form, and the colour of the solution changes. In this reaction, the reducing agent is
   A. Fe(s) B. Cu(s) C. $Fe^{2+}(aq)$ D. $Cu^{2+}(aq)$

Knowledge

30–B1.3k Differentiate between redox reactions and other reactions, using half-reactions and/or oxidation numbers.

10. Which of the following equations represent a redox reaction?
   A. $NaOH_{(aq)} + HNO_{3(aq)}$ —–> $NaNO_{3(aq)} + H_2O_{(l)}$
   B. $2AgNO_{3(aq)} + Cu_{(s)}$ —–> $2Ag_{(s)} + Cu(NO_3)_{2(aq)}$
   C. $H_2SO_{4(aq)} + 2\ KOH_{(aq)}$ —–> $K_2SO_{4(aq)} + 2H_2O_{(l)}$
   D. $CaCl_{2(aq)} + Ba(OH)_{2(aq)}$ —–> $Ca(OH)_{2(aq)} + BaCl_{2(aq)}$

Use the following equations to answer the next question.

1	$HSO_3{}^-{}_{(aq)} + HCO_3{}^-{}_{(aq)} \rightarrow H_2CO_{3(aq)} + SO_3{}^{2-}{}_{(aq)}$
2	$C_6H_{12}O_{6(aq)} + 6\,O_{2(g)} \rightarrow 6\,CO_{2(g)} + 6\,H_2O_{(l)}$
3	$Ni^{2+}{}_{(aq)} + Fe_{(s)} \rightarrow Fe^{2+}{}_{(aq)} + Ni_{(s)}$
4	$Co^{2+}{}_{(aq)} + 2\,Fe^{2+}{}_{(aq)} \rightarrow 2\,Fe^{3+}{}_{(aq)} + Co_{(s)}$
5	$6\,CO_{2(g)} + 6\,H_2O_{(l)} \rightarrow C_6H_{12}O_{6(aq)} + 6\,O_{2(g)}$

11. Match the equations, as numbered above, with the corresponding descriptionslisted below.

A biological redox reaction carried out in a plant cell but not in an animal cell ___________

A biological redox reaction carried out in both animal and plant cells ___________

A spontaneous, non-biological redox reaction ___________

A non-spontaneous, non-biological redox reaction ___________

Knowledge

30–B1.4k Identify electron transfer, oxidizing agents and reducing agents in redox reactions that occur in everyday life, in both living systems (*e.g., cellular respiration, photosynthesis*) and nonliving systems; i.e., corrosion.

Use the following equation to answer the next question.

Cellular Respiration

$$C_6H_{12}O_6(s) + 6\,O_2(g) \rightarrow 6\,CO_2(g) + 6\,H_2O(l)$$

12. During cellular respiration, the oxidizing agent is
    A. $O_2(g)$ **B.** $CO_2(g)$ **C.** $H_2O(l)$ **D.** $C_6H_{12}O_6(s)$

Use the following equations to answer the next question.

Metallurgical Processes

I	$2\,Al_2O_3(s) \rightarrow 4\,Al(s) + 3\,O_2(g)$
II	$2\,PbO(s) + C(s) \rightarrow 2\,Pb(s) + CO_2(g)$
III	$Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
IV	$Cu(s) + 4\,HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2\,NO_2(g) + 2\,H_2O(l)$

13. The metallurgical processes in which the metal loses electrons are
    A. I and II
    B. I and III
    C. II and IV
    D. III and IV

Knowledge

30–B1.5k Compare the relative strengths of oxidizing and reducing agents, using empirical data.

Use the following data to answer the next question.

	$Be^{2+}(aq)$	$Cd^{2+}(aq)$	$Ra^{2+}(aq)$	$V^{2+}(aq)$
Be(s)	×	✓	×	✓
Cd(s)	×	×	×	×
Ra(s)	✓	✓	×	✓
V(s)	×	✓	×	×

✓ evidence of a spontaneous reaction
× no spontaneous reaction

14. When listed in order from strongest to weakest, the oxidizing agents are
A. Ra(s), Be(s), V(s), Cd(s)
B. Cd(s), V(s), Be(s), Ra(s)
C. $Ra^{2+}(aq)$, $Be^{2+}(aq)$, $V^{2+}(aq)$, $Cd^{2+}(aq)$
D. $Cd^{2+}(aq)$, $V^{2+}(aq)$, $Be^{2+}(aq)$, $Ra^{2+}(aq)$

Use the following information to answer the next question.

A student collected the data below.

$$M^{2+}(aq) + 2e^- \rightarrow M(s) \qquad E° = +1.21\ V$$
$$Q^{+}(aq) + e^- \rightarrow Q(s) \qquad E° = +1.03\ V$$
$$Z^{3+}(aq) + 3e^- \rightarrow Z(s) \qquad E° = -0.21\ V$$
$$X_2(aq) + 2e^- \rightarrow 2X^-(aq) \qquad E° = -1.23\ V$$

15. From the student's data, the strongest reducing agent is
A. $M^{2+}(aq)$ **B.** $X^-(aq)$ **C.** $X_2(aq)$ **D.** M(s)

16. Four metals represented by the symbols R, S, T, and V and their ions combine with each other in the following manner:

$$S^{2+}_{(aq)} + 2\,T_{(s)} \rightarrow 2\,T^{+}_{(aq)} + S_{(s)}$$
$$R^{3+}_{(aq)} + T_{(s)} \rightarrow \text{No Reaction}$$
$$2\,R^{3+}_{(aq)} + 3\,V_{(s)} \rightarrow 3\,V^{2+}_{(aq)} + 2\,R_{(s)}$$

When the oxidizing agents are arranged from strongest to weakest, the order is

A. $S^{2+}_{(aq)},\ T^{+}_{(aq)},\ R^{3+}_{(aq)},\ V^{2+}_{(aq)}$

B. $V^{2+}_{(aq)},\ R^{3+}_{(aq)},\ T^{+}_{(aq)},\ S^{2+}_{(aq)}$

C. $V_{(s)},\ R_{(s)},\ T_{(s)},\ S_{(s)}$

D. $S_{(s)},\ T_{(s)},\ R_{(s)},\ V_{(s)}$

Knowledge

30–B1.6k Predict the spontaneity of a redox reaction, based on standard reduction potentials, and compare the predictions to experimental results.

17. Which of the following equations represents a spontaneous redox reaction?

A. $Zn^{2+}(aq) + Pb(s) \longrightarrow Zn(s) + Pb^{2+}(aq)$

B. $Sn^{4+}(aq) + Fe(s) \longrightarrow Sn^{2+}(aq) + Fe^{2+}(aq)$

C. $Zn^{2+}(aq) + Co(s) \longrightarrow Zn(s) + Co^{2+}(aq)$

D. $O_2(g) + 2H_2O(l) + 4\ Br^-(aq) \longrightarrow 2\ Br_2(l) + 4\ OH^-(aq)$

18. The reducing agent that can convert 1.0 mol/L $Sn^{4+}(aq)$ ions to $Sn^{2+}(aq)$ but **not** 1.0 mol/L $Sn^{2+}(aq)$ to Sn(s) is

A. Cu(s)

B. Pb(s)

C. Ni(s)

D. Cr(s)

Knowledge

30–B1.7k Write and balance equations for redox reactions in acidic and neutral solutions by

- using half-reaction equations obtained from a standard reduction potential table
- developing simple half-reaction equations from information provided about redox changes
- assigning oxidation numbers, where appropriate, to the species undergoing chemical change

Use the following equation to answer the next question.

$$\_\_OCl^-(aq) + \_\_I^-(aq) + \_\_H^+(aq) \rightarrow \_I_2(aq) + \_Cl^-(aq) + \_H_2O(l)$$

19. When the equation above is balanced under acidic conditions, the whole number coefficient for H+(aq) is **__i__** and the amount of electrons transferred is **__ ii__** .

The statement above is completed by the information in row

ROW	i	ii
A.	1	1 mol
B.	1	2 mol
C.	2	1 mol
D.	2	2 mol

Use the following information to answer the next question.

Chlorine gas and aqueous sodium hyposulfite react as represented by the following **unbalanced** equation.

$$Cl_2(g) + S_2O_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + H^+(aq) + Cl^-(aq)$$

20. The balanced oxidation half-reaction equation is

A. $Cl_2(g) + 2\ e^- \longrightarrow 2\ Cl^-(aq)$

B. $S_2O_3^{2-}(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2H^+(aq) + 4\ e^-$

C. $S_2O_3^{2-}(aq) + 5H_2O(l) \longrightarrow 2\ SO_4^{2-}(aq) + 10H^+(aq) + 8\ e^-$

D. $S_2O_3^{2-}(aq) + 5H_2O(l) + 4\ e^- \longrightarrow 2\ SO_4^{2-}(aq) + 10H^+(aq)$

Use the following information to answer the next question.

The colour of brick depends upon the type of clay and additives used. The reaction that occurs when clays containing iron(II) persulphide, $FeS_{2(s)}$, are heated in the kiln is

$$4\,FeS_{2(s)} + 11\,O_{2(g)} \rightarrow 2\,Fe_2O_{3(s)} + 8\,SO_{2(g)}$$

21. In this reaction, the oxidation state of iron

A. changes from 0 to +3
B. changes from +2 to +3
C. changes from +2 to 0
D. does not change

22. In the balanced redox reaction equation

$$3\,Cu_{(s)} + 2\,NO_3^{-}{}_{(aq)} + 8\,H^{+}{}_{(aq)} \rightarrow 3\,Cu^{2+}{}_{(aq)} + 2\,NO_{(g)} + 4\,H_2O_{(l)},$$

the oxidation number of nitrogen

A. decreases by 3
B. increases by 3
C. increases by 2
D. decreases by 6

23. When the equation

$$V_2O_{5(s)} + Mn_{(s)} \rightarrow VO_{(s)} + MnO_{2(s)}$$

is balanced using the lowest whole number coefficients, the coefficient of

$V_2O_{5(s)}$ is ___________

$Mn_{(s)}$ is ___________

$VO_{(s)}$ is ___________

$MnO_{2(s)}$ is ___________

Knowledge
30–B1.8k Perform calculations to determine quantities of substances involved in redox titrations.

Use the following information to answer the next two questions.

A standardized 0.125 mol/L potassium dichromate solution was used to titrate 20.0 mL samples of acidified $Sn^{2+}(aq)$. The data is represented in the following table.

Titration Data

Trial	**I**	**II**	**III**
Final burette reading (mL)	27.2	44.5	30.1
Initial burette reading (mL)	10.1	27.2	12.9

24. The amount of potassium dichromate solution required to complete this titration is
A. 8.33×10^{-4} mol
B. 6.45×10^{-3} mol
C. 2.50×10^{-3} mol
D. 2.15×10^{-3} mol

Use your recorded answer from the above question to answer the next question

25. The concentration of $Sn^{2+}(aq)$ in the sample used in the titration, expressed in scientific notation, is **a.bc** $\times 10^{-\mathbf{d}}$ mol/L. The values of**a**, **b**, **c**, and **d** are ______, ______, ______, and ______.

26. In an experiment, a student used 11.33 mL of $H_2O_2(aq)$ to titrate a 17.00 mL sample of acidified 8.0×10^{-3} mol/L $KMnO_4(aq)$. If $Mn^{2+}(aq)$ is one of the products, then the concentration of the $H_2O_2(aq)$ is
A. 1.2×10^{-2} mol/L
B. 1.5×10^{-2} mol/L
C. 3.0×10^{-2} mol/L
D. 6.0×10^{-2} mol/L

27. Iron metal is easily oxidized to $Fe^{2+}(aq)$ by an acidified potassium dichromate solution during a redox titration.

a. Write the net ionic equation for this process and **determine** the mass of iron metal oxidized by 50.0 mL of a 0.250 mol/L acidified $K_2Cr_2O_7(aq)$ solution. **(3 marks)**

b. Oxidation of iron is often an undesirable reaction in the environment. There are several methods to prevent corrosion of iron. **Describe** one of these methods and **explain** how this method prevents iron from corroding. **(2 marks)**

Knowledge

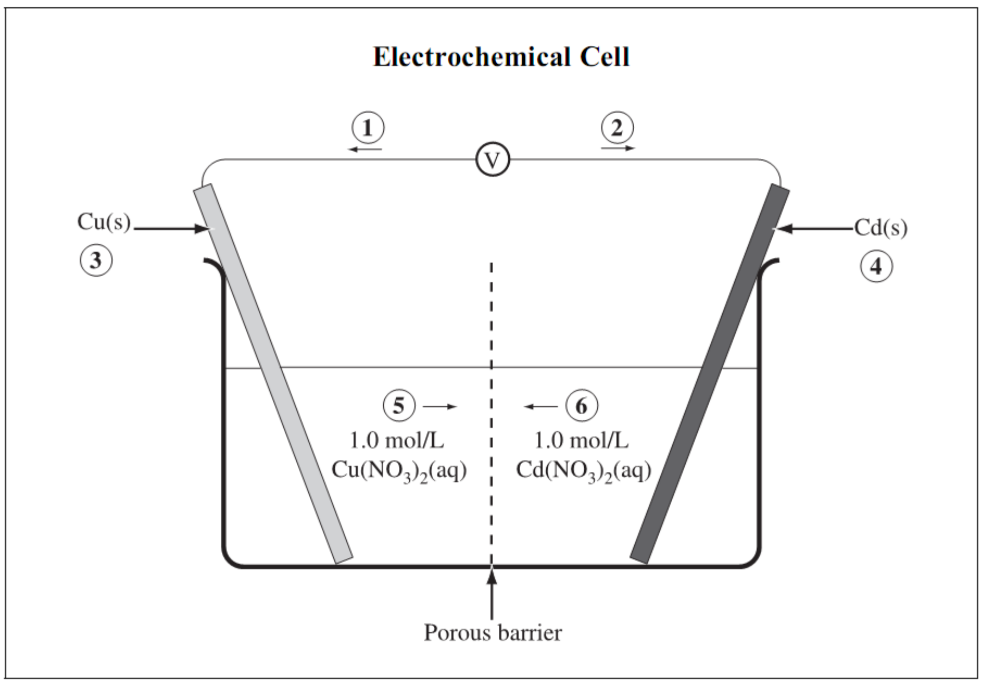
30–B2.1k Define anode, cathode, anion, cation, salt bridge/porous cup, electrolyte, external circuit, power supply, voltaic cell and electrolytic cell.

28. The anode of an electrochemical cell is the site at which
  A. oxidation occurs
  B. cations gain electrons
  C. cations are attracted to the electrode
  D. electrons are attracted to the electrode

29. Illustrate and describe a working voltaic cell that incorporates a standard nickel half-cell and has a net cell potential greater than 1.00 V.

Your response should include
- relevant balanced half-reaction equations and an E°net calculation
- a labelled cell diagram
- evidence that a reaction has occurred in each half-cell

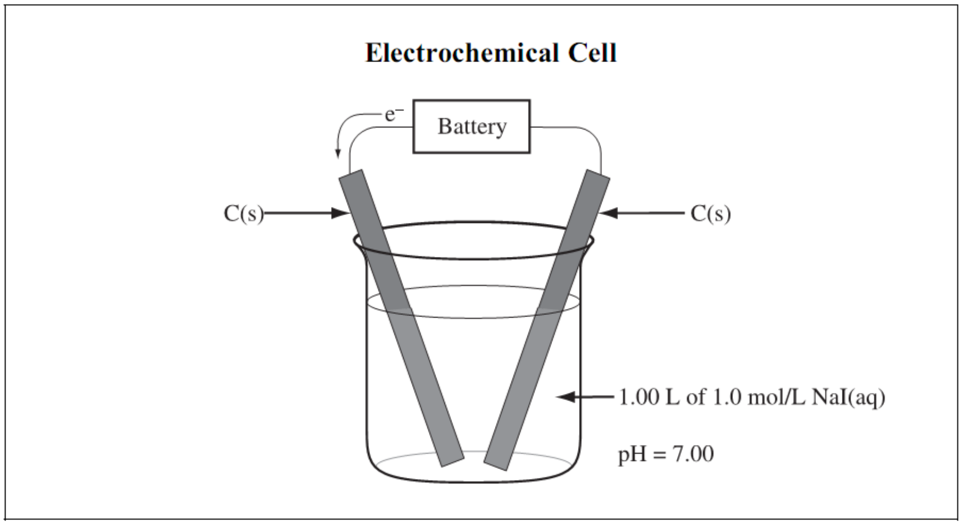
Use the following diagram to answer the next question.



30. Match the numbers in the diagram above with their appropriate labels given below.

Anode	___________	(Record in the first box)
Cathode	___________	(Record in the second box)
Anion movement	___________	(Record in the third box)
Electron movement	___________	(Record in the fourth box)

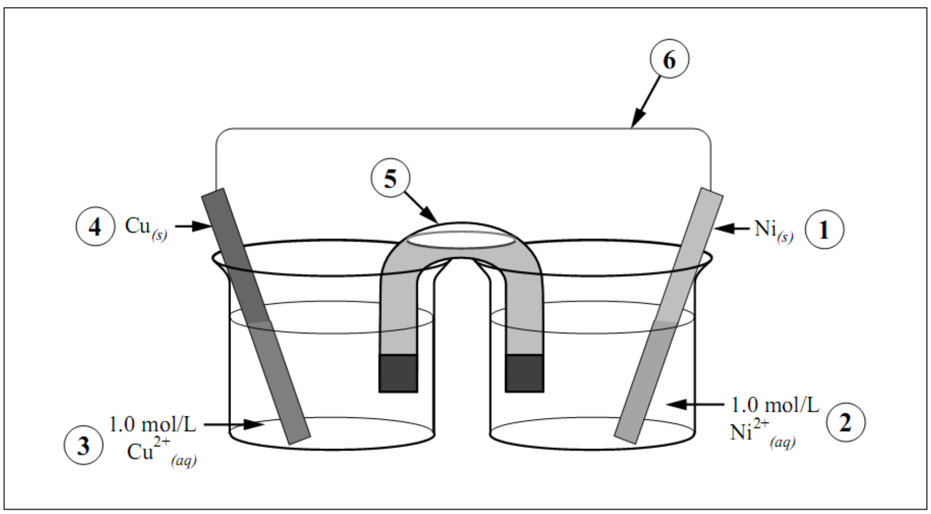
Use the following diagram to answer the next question.



31. Which of the following rows identifies the type of electrochemical cell in the diagram above and describes what happens during its operation?

ROW	Type of Cell	What Happens
A.	Voltaic	Electrons move toward the cathode
B.	Voltaic	$I^-_{(aq)}$ moves toward the cathode
C.	Electrolytic	Electrons move toward the cathode
D.	Electrolytic	$I^-_{(aq)}$ moves toward the cathode

Use the following diagram to answer the next question.



32. Identify the part of the electrochemical cell, as numbered above, that corresponds to the terms listed below.

Cathode ____________ (Record in the first box)
External electron circuit ____________ (Record in the second box)
Oxidizing agent ____________ (Record in the third box)
Anode ____________ (Record in the fourth box)

Knowledge

30–B2.2k Identify the similarities and differences between the operation of a voltaic cell and that of an electrolytic cell.

33. An electrolytic cell differs from a voltaic cell in that the electrolytic cell

A. is spontaneous
B. consumes electricity
C. has a positive E°net value
D. has an anode and a cathode

Use the following information to answer the next question.

Statements About Electrochemical Cells

1 Oxidation occurs at the anode.

2 The oxidizing agent reacts at the cathode.

3 Cations move through the wire to the cathode.

4 Cations move through the electrolyte to the cathode.

5 Electrons move through the wire to the cathode.

6 Electrical energy is converted to chemical energy.

7 Chemical energy is converted to electrical energy.

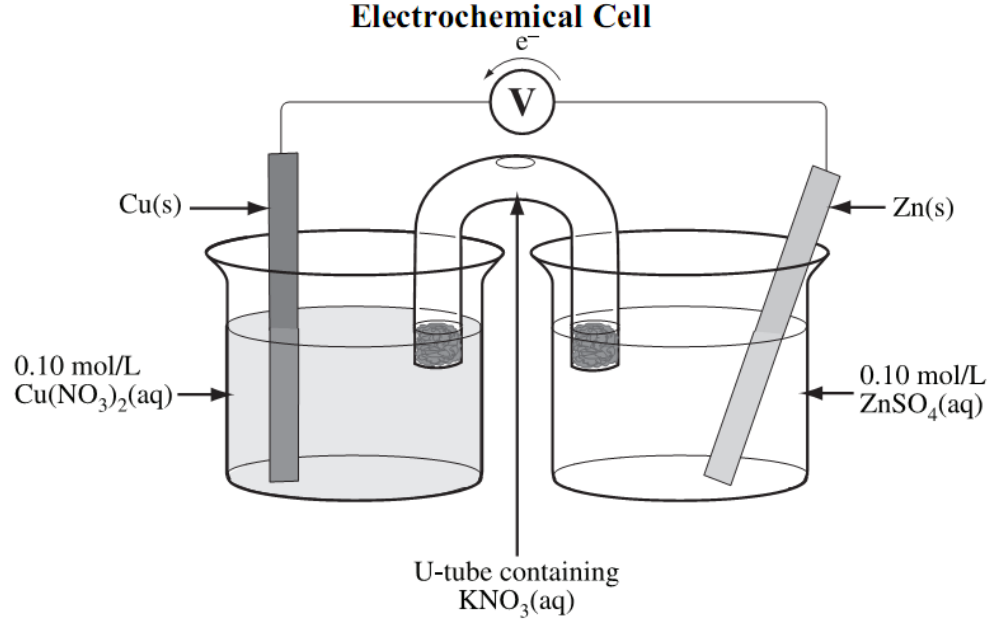
34. The statements numbered above that apply to both electrolytic cells and voltaic cells are _____ , _____ , _____ , and _____ .

Knowledge

30–B2.3k Predict and write the half-reaction equation that occurs at each electrode in an electrochemical cell.

Use the following diagram to answer the next question.

Electrochemical Cell



35. The reduction half-reaction that occurs during the operation of the electrochemical cell represented in the diagram above is__i__, and this reaction occurs at the__ii__.

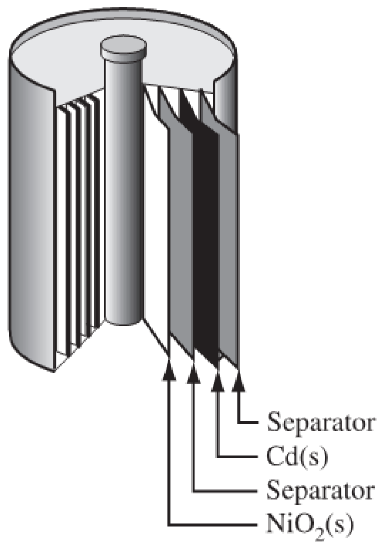
The statement above is completed by the information in row

ROW	i	ii
A.	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	anode
B.	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	cathode
C.	$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	anode
D.	$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	cathode

Use the following information to answer the next question.

Voltaic cells are used as portable sources of electrical energy. One common cell is the rechargeable nickel–cadmium cell.

Nickel–Cadmium Cell



The net equation representing the discharge of the nickel–cadmium cell is

$$NiO_2(s) + Cd(s) + 2\,H_2O(l) \rightarrow Cd(OH)_2(s) + Ni(OH)_2(s)$$

36. The half-reaction that occurs at the anode during the discharge of the nickel-cadmium cell is
    A. $Cd(s) + 2OH^-(aq) \rightarrow Cd(OH)_2(s) + 2\,e^-$
    B. $Cd(s) + 2OH^-(aq) + 2\,e^- \rightarrow Cd(OH)_2(s)$
    C. $NiO_2(s) + 2H_2O(l) + 2\,e^- \rightarrow Ni(OH)_2(s) + 2OH^-(aq)$
    D. $NiO_2(s) + 2H_2O(l) \rightarrow Ni(OH)_2(s) + 2OH^-(aq) + 2\,e^-$

Knowledge
30–B2.4k Recognize that predicted reactions do not always occur; *e.g., the production of chlorine gas from the electrolysis of brine.*

37. What is the exception that deals with the strength of water and the chloride ion as reducing agents?

Knowledge
30–B2.5k Explain that the values of standard reduction potential are all relative to 0 volts, as set for the hydrogen electrode at standard conditions.

38. For the standard reference half-cell, the reduction half-reaction and electrical potential are

A. $H_2(g) \rightarrow 2H^+(aq) + 2\,e^-$ E° = 0.00 V
B. $2H^+(aq) + 2\,e^- \rightarrow H_2(g)$ E° = 0.00 V
C. $2H2O(l) + 2\,e^- \rightarrow H_2(g) + 2OH^-(aq)$ E° = – 0.83 V
D. $H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2\,e^-$ E° = + 0.83 V

39. If the standard iodine half-cell is chosen as the reference half-cell instead of the hydrogen half-cell, then the cell potential for a silver–nickel cell is +/– ___________ V.

40. If the $Cu^{2+}(aq)$ / Cu(s) reduction half-reaction was assigned a reduction potential value of 0.00 V for an electrode potential table, then the $Ni^{2+}(aq)$ / Ni(s) half-reaction on that table would have a reduction potential value of
A. +0.26 V B. +0.08 V C. –0.26 V D. –0.60 V

Knowledge

30–B2.6k Calculate the standard cell potential for electrochemical cells.

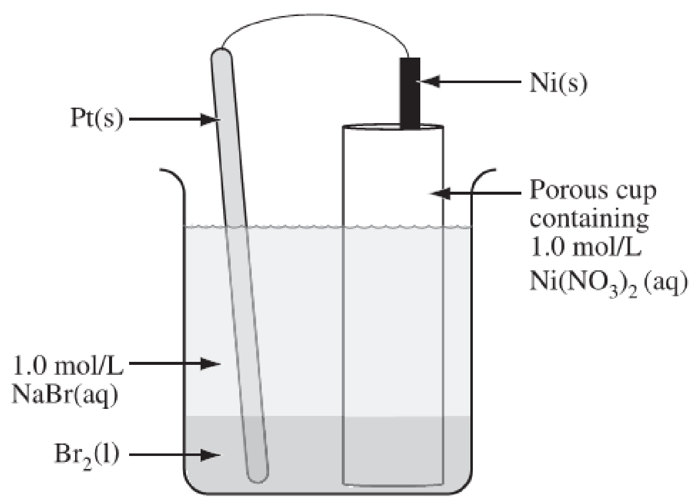
Use the following equation to answer the next question.

$$2\,Ag^{+}(aq) + Zn(s) \rightarrow 2\,Ag(s) + Zn^{2+}(aq)$$

41. The cell potential for the redox reaction represented by the equation above is
A. + 0.04 V B. + 0.84 V C. + 1.56 V D. + 2.36 V

Use the following diagram to answer the next question.

Electrochemical Cell



42. The cell potential for the electrochemical cell in the diagram above is __________ V.

43. The voltage of an electrochemical cell is +0.20 V. If one of the half-reactions is the reduction of Cu^{2+}(aq), then the other half-reaction that occurs could be
A. $2\,I^{-}_{(aq)} \rightarrow I_{2(s)} + 2\,e^{-}$
B. $S_{(s)} + 2\,H^{+}_{(aq)} + 2\,e^{-} \rightarrow H_2S_{(aq)}$
C. $H_2S_{(aq)} \rightarrow S_{(s)} + 2\,H^{+}_{(aq)} + 2\,e^{-}$
D. $I_{2(s)} + 2\,e^{-} \rightarrow 2\,I^{-}_{(aq)}$

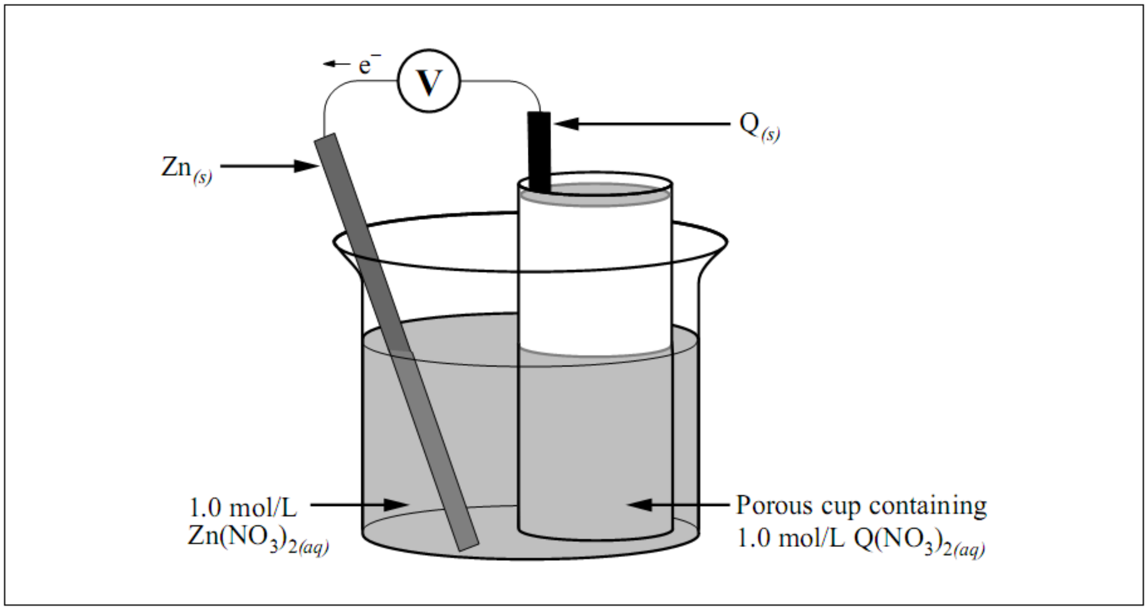
Use the following information to answer the next question.

During the operation of a NiCad battery, the two half-reactions that occur are

I	$Cd_{(s)} + 2OH^-_{(aq)} \rightarrow Cd(OH)_{2(s)} + 2e^-$	$E° = ?$ V
II	$NiO_{2(s)} + 2H_2O_{(l)} + 2e^- \rightarrow Ni(OH)_{2(s)} + 2OH^-_{(aq)}$	$E° = -0.49$ V

44. On discharging, the electrical potential of a NiCad battery is +1.40 V. The reduction potential for half-reaction I is –__________ V.

Use the following diagram to answer the next question.



45. Given that the reading on the voltmeter for this cell is +1.74 V, which of the following statements is correct?
A. The reduction potential of Q^{2+}(aq) is +2.50 V.
B. Zn(s) is a weaker reducing agent than Q(s).
C. Q^{2+}(aq) would react spontaneously with Cu(s).
D. Q^{2+}(aq) is a stronger oxidizing agent than Zn2+(aq).

46. A net cell potential value that would represent a spontaneous reaction is
A. –1.05 V
B. –0.08 V
C. 0.00 V
D. +0.15 V

Knowledge

30–B2.7k Predict the spontaneity or nonspontaneity of redox reactions, based on standard cell potential, and the relative positions of half-reaction equations on a standard reduction potential table.

Use the following information to answer the next question.

A student constructed two standard electrochemical cells using $Pb^{2+}(aq)$ and $Ni^{2+}(aq)$. In both cells a Pb(s) electrode was placed in the $Pb^{2+}(aq)$ solution. In the first cell a Ni(s) electrode was placed in the $Ni^{2+}(aq)$ solution. In the second cell an inert C(s) electrode was placed in the $Ni^{2+}(aq)$ solution instead of the Ni(s) electrode.

47. Which of the following statements describes what occurs in each cell?
A. In both cells a power source is needed.
B. In both cells a spontaneous reaction occurs and Pb(s) is produced.
C. In the first cell Ni(s) is produced, and in the second cell a power source is needed.
D. In the first cell the reaction is spontaneous, and in the second cell the reaction is nonspontaneous.

Use the following information to answer the next question.

In a laboratory, a student obtained the following results when testing, under standard conditions, reactions between various metals and their corresponding ions.

	$Ga_{(s)}$	$Fe_{(s)}$	$Zn_{(s)}$	$Mg_{(s)}$
$Ga^{3+}_{(aq)}$	–	×	✓	✓
$Fe^{2+}_{(aq)}$	✓	–	✓	✓
$Zn^{2+}_{(aq)}$	×	×	–	✓
$Mg^{2+}_{(aq)}$	×	×	×	–

Key

✓ denotes reaction
× denotes no reaction
– denotes no test performed

48. The reduction potential of the $Ga^{3+}(aq)$ could be
A. –0.53 V B. –1.41 V C. +1.21 V D. +1.92 V

Knowledge

30–B2.8k Calculate mass, amounts, current and time in single voltaic and electrolytic cells by applying Faraday's law and stoichiometry.

Use the following information to answer the next question.

The reduction half-reaction for a Hall-Héroult electrolytic cell is represented by the following equation.

$$Al^{3+}(l) + 3\,e^- \rightarrow Al(l)$$

49. If a current of 10.0 A is applied for 5.00 h to the Hall-Héroult electrolytic cell, then the amount of electrons transferred is

A. 5.60 mol B. 1.87 mol C. 6.22×10^{-1} mol D. 5.18×10^{-4} mol

Use the following information to answer the next question.

A Hall-Héroult electrolytic cell is used to produce molten aluminium from molten aluminium oxide, as represented by the following simplified equation.

$$2\,Al_2O_3(l) \rightarrow 4\,Al(l) + 3\,O_2(g)$$

50. In the Hall-Héroult electrolytic cell, the time required for the cell to operate at 5.55×10^3A to produce 20.0 kg of aluminum is ___________h.

51. If the electrochemical cell Cd(s) / Cd^{2+}(aq) //Ag^+(aq) / Ag(s) produces a 6.00A current for 2.00 h, the mass change of the anode will be a

A. 25.2 g decrease B. 2.25 g increase

C. 48.3 g decrease D. 48.3 g increase

Science, Technology, and Society

30–B2.2sts describe science and technology applications that have developed in response to human and environmental needs

- *investigate the use of technology, such as galvanism, metallurgy, magnesium coupling, painting, cathodic protection, to solve practical problems related to corrosion*

Use the following information to answer the next question.

A particular company manufactures plastic tape containing small pieces of magnesium. The tape is completely wrapped around iron pipes that will be buried underground.

52. Explain in chemical terms the purpose(s) of each component of the tape.

Your response should include

- an explanation of the corrosion of iron
- an explanation of how the plastic tape and magnesium pieces prevent the corrosion of iron
- relevant half-reaction equations

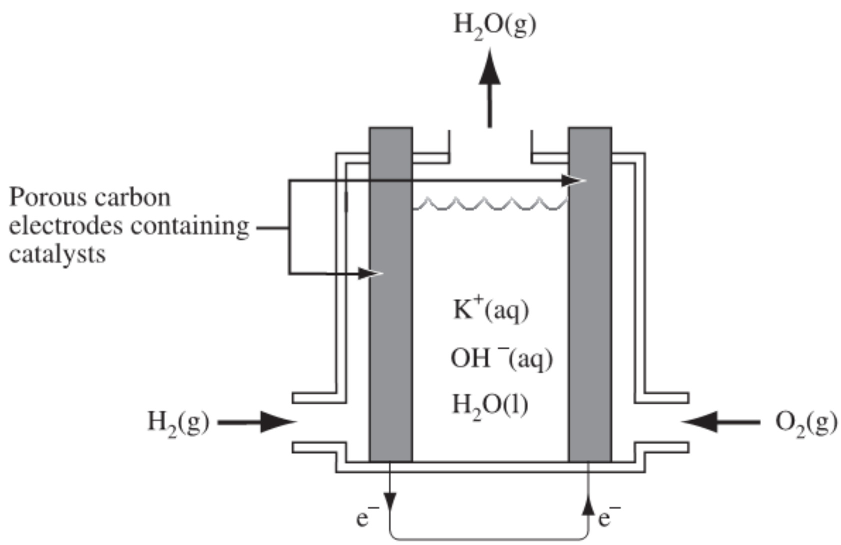
53. Sacrificial metals may be used to protect pipelines, septic tanks, and ship propellers. A metal that could be used as a sacrificial anode to protect iron is

A. magnesium
B. tin
C. lead
D. silver

Use the following information to answer the next question.

Hydrogen–oxygen fuel cells have been used for years in spacecraft and more recently in small-scale power plants to generate electricity. Now, some governments and companies are working together to perfect this type of fuel cell for automobile use, and experiments are currently being conducted with operational prototypes. A diagram of a hydrogen–oxygen fuel cell is shown below.

Hydrogen-Oxygen Fuel Cell



54. From an ecological perspective, a reason why hydrogen–oxygen fuel cells should not be used to power automobiles is that
A. hydrogen fuel can be produced through the electrolysis of seawater by using the energy produced from burning fossil fuels
B. cars powered by a hydrogen–oxygen fuel cell would be up to 30% more efficient than cars powered by gasoline
C. water vapour is the primary by-product of the cell
D. oxygen is readily available from the atmosphere