Which of the following changes is **not** an example of oxidation?

A. Corrosion of metals
B. Plating of metals
C. Rusting of iron
D. Reaction at the anode of a battery

An equation that represents a redox reaction is

A. $NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$
B. $AgNO_{3(aq)} + KI_{(aq)} \rightarrow AgI_{(s)} + KNO_{3(aq)}$
C. $Mg(OH)_{2(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(aq)} + 2\,H_2O_{(l)}$
D. $Cu_{(s)} + 4\,HNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2\,NO_{2(g)} + 2\,H_2O_{(l)}$

Use the following equation to answer the next question.

$$4\,NH_{3(g)} + 7O_{2(g)} \rightarrow 4NO_{2(g)} + 6H_2O_{(g)}$$

(+3, 0, +4, −2; 4e⁻; 7e⁻)

Numerical Response

When balanced in terms of lowest whole numbers, the coefficients for this equation are, respectively, 4, 7, 4, and 6.
(Record all four digits on the answer sheet.)

A spontaneous reaction would occur between a 1.0 mol/L $Fe^{3+}_{(aq)}$ solution and

A. $I_{2(s)}$
B. $Zn_{(s)}$
C. $Hg_{(l)}$
D. 1.0 mol/L $Fe^{2+}_{(aq)}$

Numerical Response

8. The volume of 0.160 mol/L $K_2Cr_2O_{7(aq)}$ required to completely react with 10.0 mL of acidic 0.881 mol/L $H_2O_{2(aq)}$ is 18.4 mL.

(Record your answer to three digits on the answer sheet.)

$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow$
$3[H_2O_2 \rightarrow 2e^- + \ldots]$

$Cr_2O_{7\,(aq)}^{2-} + 3H_2O_{2\,(aq)}$
$c = 0.160$ mol/L; $n = 0.0029\ldots$ mol
$V = \frac{n}{c}$
$V = 0.010$ L; $c = 0.881$ mol/L; $n = 0.00881$ mol
$\times \frac{1}{3}$

Numerical Response

9. The oxidation numbers of carbon in $HCOOH_{(aq)}$, $C_6H_{12}O_{6(s)}$, $CO_{2(g)}$, and $CHCl_{3(g)}$, respectively, are 2042.

(Record all four digits on the answer sheet.)

The reaction $2\,H_2O_{(l)} \rightarrow 2\,H_{2(g)} + O_{2(g)}$ is an example of an

A. exothermic redox reaction
B. endothermic redox reaction
C. exothermic reaction that absorbs energy
D. endothermic reaction that releases energy

In which change are electrons gained?

A. $Ca^{2+}_{(aq)}$ to $Ca_{(s)}$
B. $2\,Cl^-_{(aq)}$ to $Cl_{2(g)}$
C. $Fe^{2+}_{(aq)}$ to $Fe^{3+}_{(aq)}$
D. $NaCl_{(s)}$ to $Na^+_{(aq)}$ and $Cl^-_{(aq)}$

Which ion could **not** act as both an oxidizing agent and a reducing agent?

A. $Cu^{2+}_{(aq)}$
B. $Sn^{2+}_{(aq)}$
C. $Fe^{2+}_{(aq)}$
D. $Cr^{2+}_{(aq)}$

Bacteria in our mouths and digestive systems convert sodium nitrate and other nitrate salts into nitrites, as indicated by the incomplete and unbalanced half-reaction $NaNO_{3(aq)} \rightarrow NaNO_{2(aq)}$. In this half-reaction,

A. the oxidation number for sodium changes from –9 to –7
B. the oxidation number of nitrogen increases
C. the oxidation number of oxygen increases
D. reduction occurs

Two reagents that will oxidize $Pb_{(s)}$ to $Pb^{2+}_{(aq)}$ but that will **not** oxidize $I^-_{(aq)}$ to $I_{2(s)}$ are

A. $F_{2(g)}$ and $Fe^{3+}_{(aq)}$
B. $Fe^{3+}_{(aq)}$ and $Br_{2(l)}$
C. $Cd^{2+}_{(aq)}$ and $Ag^+_{(aq)}$
D. $Cu^{2+}_{(aq)}$ and $Sn^{4+}_{(aq)}$

Metals $W_{(s)}$, $X_{(s)}$, $Y_{(s)}$, and $Z_{(s)}$ were placed in solutions of each of their respective ionic salts. The results are summarized in the data table. A check mark indicates that a reaction occurred.

	$X^{+}_{(aq)}$	$Y^{2+}_{(aq)}$	$Z^{3+}_{(aq)}$	$W^{+}_{(aq)}$
$X_{(s)}$	—	no reaction	√	√
$Y_{(s)}$	√	—	√	√
$Z_{(s)}$	no reaction	no reaction	—	√
$W_{(s)}$	no reaction	no reaction	no reaction	—

According to the results, the strongest reducing agent is

A. $Y^{2+}_{(aq)}$
B. $W_{(s)}$
C. $Y_{(s)}$
D. $W^{+}_{(aq)}$

For the standard reference half-cell, the oxidation half-reaction and $E°$ are

A. $H_{2(g)} \rightarrow 2H^{+}_{(aq)} + 2e^{-}$ $\quad E° = 0.00\ V$
B. $2H^{+}_{(aq)} + 2e^{-} \rightarrow H_{2(g)}$ $\quad E° = 0.00\ V$
C. $2H_2O_{(l)} + 2e^{-} \rightarrow H_{2(g)} + 2OH^{-}_{(aq)}$ $\quad E° = -0.83\ V$
D. $H_{2(g)} + 2OH^{-}_{(aq)} \rightarrow 2e^{-} + 2H_2O_{(l)}$ $\quad E° = +0.83\ V$

Electrolytic cells are used commercially in

A. cameras
B. fuel cells
C. flashlights
D. metal plating

Which of the following aqueous ions can either gain or lose electrons in a redox reaction?

A. $Sn^{2+}_{(aq)}$
B. $Cl^{-}_{(aq)}$
C. $Ca^{2+}_{(aq)}$
D. $S^{2-}_{(aq)}$

If the lithium reduction half-reaction, $Li^{+}_{(aq)} + e^{-} \rightarrow Li_{(s)}$, had been assigned an $E°$ value of 0.00 V, the predicted $E°_{rxn}$ value for the reaction $Cu_{(s)} + Zn^{2+}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + Zn_{(s)}$ would be

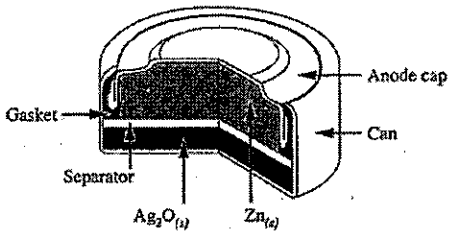
A. +3.38 V
B. −2.28 V
C. −0.42 V
D. −1.10 V

The equation representing a spontaneous reaction at standard conditions is

A. $Co^{2+}_{(aq)} + 2Fe^{2+}_{(aq)} \rightarrow Co_{(s)} + 2Fe^{3+}_{(aq)}$
B. $Sn^{4+}_{(aq)} + 2Br^{-}_{(aq)} \rightarrow Sn^{2+}_{(aq)} + Br_{2(l)}$
C. $2I^{-}_{(aq)} + Cl_{2(g)} \rightarrow I_{2(s)} + 2Cl^{-}_{(aq)}$
D. $Pb_{(s)} + Fe^{2+}_{(aq)} \rightarrow Pb^{2+}_{(aq)} + Fe_{(s)}$

Use the following information to answer the next 5 questions

Silver oxide cells are efficient but expensive because they contain silver. The diagram illustrates the construction of a silver oxide cell.



The half reactions are

$Ag_2O_{(s)} + H_2O_{(l)} + 2e^- \rightarrow 2\,Ag_{(s)} + 2\,OH^-_{(aq)}$ $\quad E^\circ = +0.34$ V

$Zn(OH)_{2(s)} + 2e^- \rightarrow Zn_{(s)} + 2OH^-_{(aq)}$ $\quad Eo = -1.25$ V

The anode of the cell is

A. $Ag_{(s)}$
B. $Zn_{(s)}$
C. $Ag_2O_{(s)}$
D. $Zn(OH)_{2(s)}$

Using lowest whole number coefficients, the coefficient for $H_2O_{(l)}$ in the balanced oxidation-reduction reaction that occurs during discharging of the cell is

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

As the cell operates, the species oxidized is

A. $Ag_{(s)}$
B. $Zn_{(s)}$
C. $Ag_2O_{(s)}$
D. $Zn(OH)_{2(s)}$

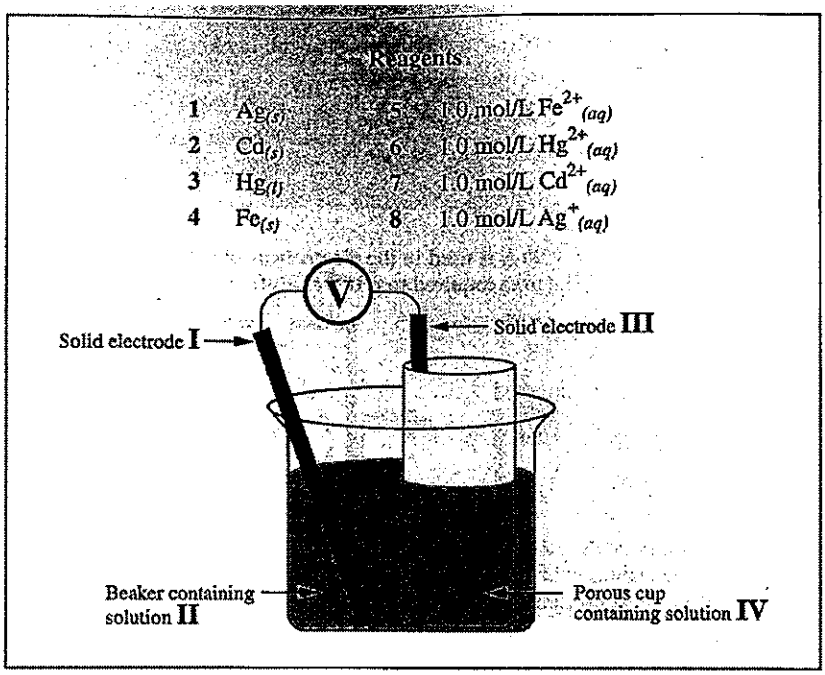
As the cell operates, the

A. $[OH^-_{(aq)}]$ increases
B. mass of $Zn_{(s)}$ increases
C. mass of $Ag_2O_{(s)}$ decreases
D. mass of $Zn(OH)_{2(s)}$ decreases

Numerical Response

The voltage generated by the silver oxide cell is 1.59 V.

Use the following to answer the next question



Numerical Response

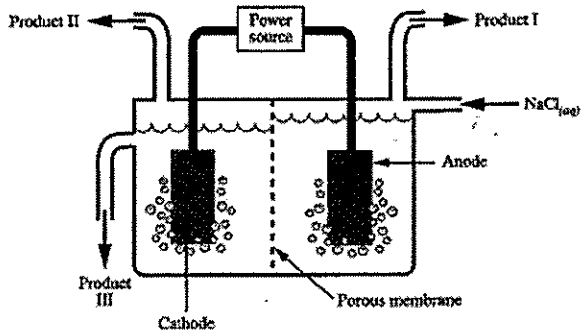
What reagents are required in order for the cell to produce a voltage of 1.25 V?

Electrode I ___1___ (Record in first column)
Solution II ___8___ (Record in second column)
Electrode III ___4___ (Record in third column)
Solution IV ___5___ (Record in fourth column)

4518
or
1845

Use the following information to answer the next two questions.

This simplified diagram shows the cell used in the Chlor-Alkali process at the Dow Chemical Plant in Fort Saskatchewan, Alberta.



The cell shown in the diagram is

A. electrolytic
B. voltaic
C. galvanic
D. acid–base

Products I, II, and III from this cell, respectively, are

A. $Cl_{2(g)}$, $H_{2(g)}$, and $HCl_{(aq)}$
B. $H_{2(g)}$, $Cl_{2(g)}$, and $NaOH_{(aq)}$
C. $HCl_{(g)}$, $Cl_{2(g)}$, and $NaOH_{(aq)}$
D. $Cl_{2(g)}$, $H_{2(g)}$, and $NaOH_{(aq)}$

Use the following information to answer the next 3 questions

The chlor-alkali process used by Dow Chemical in Fort Saskatchewan uses sodium chloride from underground deposits. The sodium chloride is dissolved in water and then pumped into electrolytic cells where a current is passed through the solution to form yellow chlorine gas, colorless hydrogen gas and aqueous sodium hydroxide. The ionic equation is

$$2\,H_2O_{(l)} + 2\,Na^+_{(aq)} + 2\,Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2\,Na^+_{(aq)} + 2\,OH^-_{(aq)} + 2\,H_{2(g)}$$

In the electrolysis of $NaCl_{(aq)}$, the cathode half-reaction is

m: 78.1 g
M: 70.90
n: 1.101 mol

n: 2.203 mol
M: 2.02 g/mol
m: ? g

A. $Na^+_{(aq)} + e^- \rightarrow Na_{(s)}$

B. $2\,Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2\,e^-$

C. $2\,H_2O_{(l)} \rightarrow O_{2(g)} + 4\,H^+_{(aq)} + 4\,e^-$

D. $2\,H_2O_{(l)} + 2\,e^- \rightarrow H_{2(g)} + 2\,OH^-_{(aq)}$

Numerical Response

If the mass of the element formed at the anode is 78.1 g, the mass of element formed at the cathode is 4.45 g.
(Record your answer to three digits on the answer sheet.)

When the electric current is switched off, the

A. pH stops decreasing

B. concentration of the $Na^+_{(aq)}$ stops changing

C. concentration of the $Cl^-_{(aq)}$ stops decreasing

D. concentration of the $H_2O_{(l)}$ stops decreasing