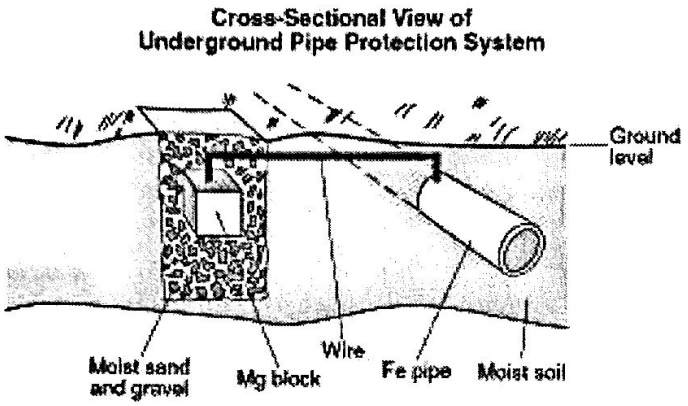
Name _____ **Electrochemical Cells Practice Problems**

1. Which statement describes electrolysis?
   (1) Chemical energy is used to produce an electrical change.
   (2) Chemical energy is used to produce a thermal change.
   (3) Electrical energy is used to produce a chemical change.
   (4) Thermal energy is used to produce a chemical change.

Base your answers to questions 2 and 3 on the information below.

Underground iron pipes in contact with moist soil are likely to corrode. This corrosion can be prevented by applying the principles of electrochemistry. Connecting an iron pipe to a magnesium block with a wire creates an electrochemical cell. The magnesium block acts as the anode and the iron pipe acts as the cathode. A diagram of this system is shown below.



2. State the direction of the flow of electrons between the electrodes in this cell.

3. Explain, in terms of reactivity, why magnesium is preferred over zinc to protect underground iron pipes. Your response must include *both* magnesium and zinc.

4. Given the balanced equation representing a reaction occurring in an electrolytic cell:
   $2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$
   Where is Na(l) produced in the cell?
   (1) at the anode, where oxidation occurs (2) at the anode, where reduction occurs
   (3) at the cathode, where oxidation occurs (4) at the cathode, where reduction occurs

Base your answers to questions 5 through 7 on the information below.

A flashlight can be powered by a rechargeable nickel-cadmium battery. In the battery, the anode is Cd(s) and the cathode is $NiO_2(s)$. The unbalanced equation below represents the reaction that occurs as the battery produces electricity. When a nickel-cadmium battery is recharged, the reverse reaction occurs.

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

5. Balance the equation for the reaction that produces electricity, using the smallest whole-number coefficients.

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

6. Determine the change in oxidation number for the element that makes up the anode in the reaction that produces electricity.

7. Explain why Cd would be above Ni if placed on Table *J*.

8. Which energy conversion occurs in a voltaic cell?
   (1) chemical energy to electrical energy (2) chemical energy to nuclear energy
   (3) electrical energy to chemical energy (4) nuclear energy to electrical energy

9. Which metal is more active than Ni and *less* active than Zn?
   (1) Cu (2) Cr (3) Mg (4) Pb

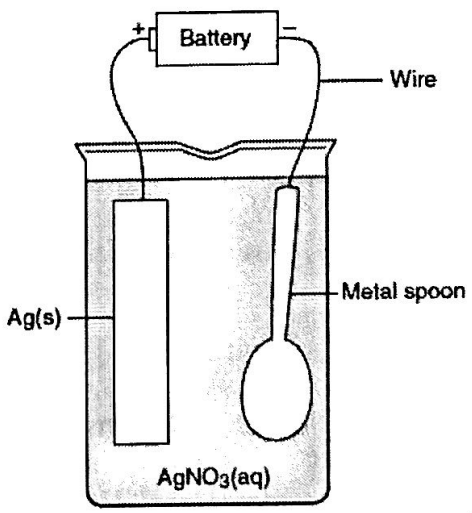
10. Reduction occurs at the cathode in
    (1) electrolytic cells, only (2) voltaic cells, only
    (3) both electrolytic cells and voltaic cells (4) neither electrolytic cells nor voltaic cells

11. Given the balanced equation representing the reaction occurring in a voltaic cell:
    $Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$
    In the completed external circuit, the electrons flow from
    (1) Pb(s) to Zn(s) (2) $Pb^{2+}(aq)$ to $Zn^{2+}(aq)$
    (3) Zn(s) to Pb(s) (4) $Zn^{2+}(aq)$ to $Pb^{2+}(aq)$

Base your answers to questions 12 and 13 on the information below.

Electroplating is an electrolytic process used to coat metal objects with a more expensive and less reactive metal. The diagram below shows an electroplating cell that includes a battery connected to a silver bar and a metal spoon. The bar and spoon are submerged in $AgNO_3(aq)$.

An Electroplating Cell



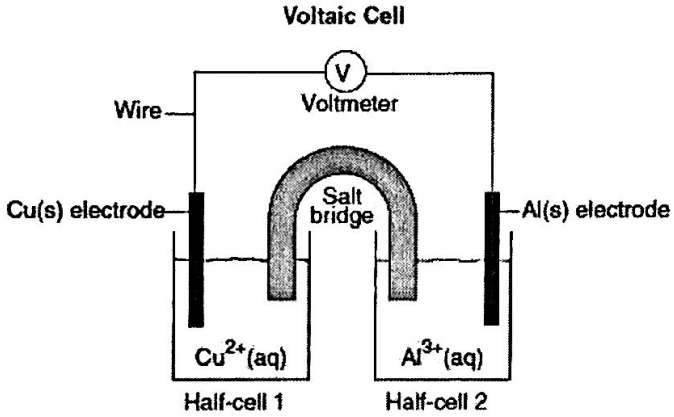
12. Explain why $AgNO_3$ is a better choice than AgCl for use in this electrolytic process.

13. Explain the purpose of the battery in this cell.

14. Which energy conversion occurs during the operation of a voltaic cell?
    (1) Chemical energy is spontaneously converted to electrical energy.
    (2) Chemical energy is converted to electrical energy only when an external power source is provided.
    (3) Electrical energy is spontaneously converted to chemical energy.
    (4) Electrical energy is converted to chemical energy only when an external power source is provided.

15. Based on Reference Table *J*, identify *one* metal that does *not* react spontaneously with HCl.

Base your answers to questions 16 through 18 on the diagram below. The diagram shows a voltaic cell with copper and aluminum electrodes immediately after the external circuit is completed.



16. Balance the redox equation using the smallest whole-number coefficients.

_____$Cu^{2+}(aq)$ + _____$Al(s)$ → _____$Cu(s)$ + _____$Al^{3+}(aq)$

17. As this voltaic cell operates, the mass of the Al(s) electrode decreases. Explain, in terms of particles, why this decrease in mass occurs.

18. Explain the function of the salt bridge.

19. Which conversion of energy always occurs in a voltaic cell?
    (1) light energy to chemical energy (2) electrical energy to chemical energy
    (3) chemical energy to light energy (4) chemical energy to electrical energy

20. Which process occurs at the anode in an electrochemical cell?
    (1) the loss of protons (2) the loss of electrons
    (3) the gain of protons (4) the gain of electrons

21. Examine the reaction: $Fe(s) + 2HNO_3(aq) \rightarrow Fe(NO_3)_2(aq) + H_2(g)$
    Explain in terms of activity why this reaction is spontaneous.

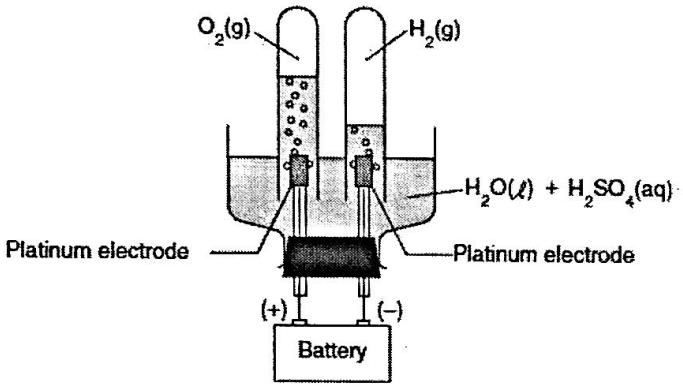
22. Because tap water is slightly acidic, water pipes made of iron corrode over time, as shown by the balanced ionic equation below:

$2Fe + 6H^{+} \rightarrow 2Fe^{3+} + 3H_2$

Explain, in terms of chemical reactivity, why copper pipes are *less* likely to corrode than iron pipes.

Base your answers to questions 23 and 24 on the information and diagram below.

The apparatus shown in the diagram consists of two inert platinum electrodes immersed in water. A small amount of an electrolyte, H_2SO_4, must be added to the water for the reaction to take place. The electrodes are connected to a source that supplies electricity.



23. What type of electrochemical cell is shown?

24. What particles provided by the electrolyte allow an electric current to flow?

Name Key **Electrochemical Cells Practice Problems**

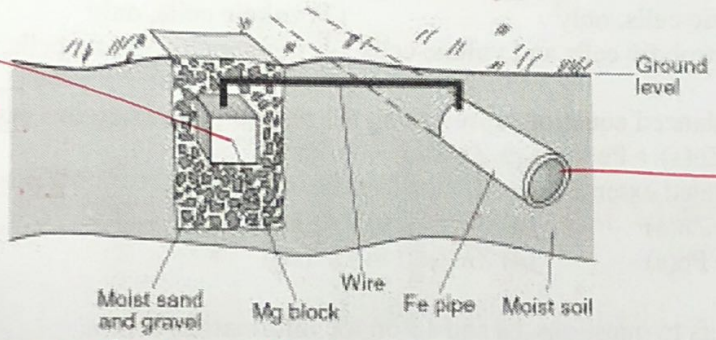
1. Which statement describes electrolysis? uses electricity to create a chemical change.

(3) (1) Chemical energy is used to produce an electrical change.
(2) Chemical energy is used to produce a thermal change.
(3) Electrical energy is used to produce a chemical change.
(4) Thermal energy is used to produce a chemical change.

Base your answers to questions 2 and 3 on the information below.

Underground iron pipes in contact with moist soil are likely to corrode. This corrosion can be prevented by applying the principles of electrochemistry. Connecting an iron pipe to a magnesium block with a wire creates an electrochemical cell. The magnesium block acts as the anode and the iron pipe acts as the cathode. A diagram of this system is shown below.

Cross-Sectional View of Underground Pipe Protection System



anode (higher on Table J)

cathode

anode to cathode

2. State the direction of the flow of electrons between the electrodes in this cell.

Mg to Fe

3. Explain, in terms of reactivity, why magnesium is preferred over zinc to protect underground iron pipes. Your response must include *both* magnesium and zinc.

Mg is more active than Zn. so it is more likely to oxidize + will better protect the iron pipes.

4. Given the balanced equation representing a reaction occurring in an electrolytic cell:

$$2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$$

Where is Na(l) produced in the cell?

(4) (1) at the anode, where oxidation occurs (2) at the anode, where reduction occurs
(3) at the cathode, where oxidation occurs (4) at the cathode, where reduction occurs

Na^+ undergoes reduction
$Na^+ + e^- \rightarrow Na$
→ reduction is always @ cathode.

Base your answers to questions 5 through 7 on the information below.

A flashlight can be powered by a rechargeable nickel-cadmium battery. In the battery, the anode is Cd(s) and the cathode is $NiO_2(s)$. The unbalanced equation below represents the reaction that occurs as the battery produces electricity. When a nickel-cadmium battery is recharged, the reverse reaction occurs.

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

5. Balance the equation for the reaction that produces electricity, using the smallest whole-number coefficients.

_____ $Cd(s)$ + _____ $NiO_2(s)$ + 2 $H_2O(l)$ → _____ $Cd(OH)_2(s)$ + _____ $Ni(OH)_2(s)$

6. Determine the change in oxidation number for the element that makes up the anode in the reaction that produces electricity.

Cd^0 to Cd^{+2}

7. Explain why Cd would be above Ni if placed on Table *J*.

Since Cd oxidizes in the presence of Ni^{+4} Therefore Cd is more active + would be higher on Table J

8. Which energy conversion occurs in a voltaic cell?

(1)

(1) chemical energy to electrical energy (2) chemical energy to nuclear energy
(3) electrical energy to chemical energy (4) nuclear energy to electrical energy

→ electrolytic

9. Which metal is more active than Ni and *less* active than Zn? (Table J)

(2)

(1) Cu (2) Cr (3) Mg (4) Pb

* must be in between

10. Reduction occurs at the cathode in

(3)

(1) electrolytic cells, only (2) voltaic cells, only
(3) both electrolytic cells and voltaic cells (4) neither electrolytic cells nor voltaic cells

11. Given the balanced equation representing the reaction occurring in a voltaic cell: spontaneous

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

In the completed external circuit, the electrons flow from

(3)

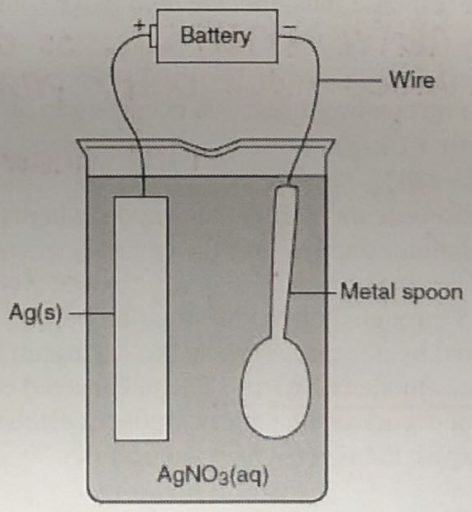
(1) Pb(s) to Zn(s) (2) $Pb^{2+}(aq)$ to $Zn^{2+}(aq)$
(3) Zn(s) to Pb(s) (4) $Zn^{2+}(aq)$ to $Pb^{2+}(aq)$

Zn is oxidized (anode)

Base your answers to questions 12 and 13 on the information below.

Electroplating is an electrolytic process used to coat metal objects with a more expensive and less reactive metal. The diagram below shows an electroplating cell that includes a battery connected to a silver bar and a metal spoon. The bar and spoon are submerged in $AgNO_3(aq)$.

An Electroplating Cell



12. Explain why $AgNO_3$ is a better choice than AgCl for use in this electrolytic process.

AgCl is insoluble and will not form an aqueous solution.

13. Explain the purpose of the battery in this cell.

The battery provides electrical energy.

spontaneous.

14. Which energy conversion occurs during the operation of a voltaic cell?

(1)

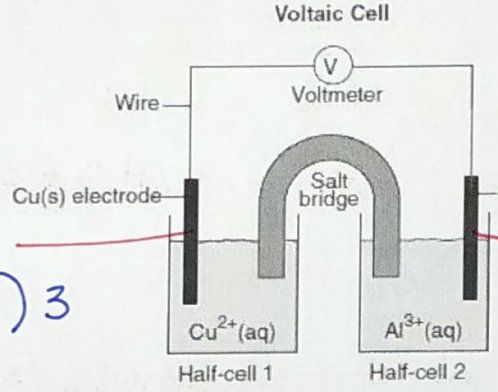
(1) Chemical energy is spontaneously converted to electrical energy.
(2) Chemical energy is converted to electrical energy only when an external power source is provided.
(3) Electrical energy is spontaneously converted to chemical energy.
(4) Electrical energy is converted to chemical energy only when an external power source is provided.

metal is below H^+

15. Based on Reference Table *J*, identify *one* metal that does *not* react spontaneously with HCl.

One metal that does not react w/ HCl spontaneously is Cu (or Ag or Au)

Base your answers to questions 16 through 18 on the diagram below. The diagram shows a voltaic cell with copper and aluminum electrodes immediately after the external circuit is completed.



Cathode

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

Anode (higher on Table J)

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

16. Balance the redox equation using the smallest whole-number coefficients.

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

17. As this voltaic cell operates, the mass of the Al(s) electrode decreases. Explain, in terms of particles, why this decrease in mass occurs.

Al is oxidized into Al^{3+} ions. This decreases the mass of Al solid.

18. Explain the function of the salt bridge.

The salt bridge maintains a neutral cell by allowing for the movement of ions.

19. Which conversion of energy always occurs in a voltaic cell?

(4)

(1) light energy to chemical energy
(2) electrical energy to chemical energy
(3) chemical energy to light energy
(4) chemical energy to electrical energy

20. Which process occurs at the anode in an electrochemical cell?

(2)

(1) the loss of protons
(2) the loss of electrons
(3) the gain of protons
(4) the gain of electrons

AN OX ← loss of e- always

21. Examine the reaction: $Fe(s) + 2HNO_3(aq) \rightarrow Fe(NO_3)_2(aq) + H_2(g)$
Explain in terms of activity why this reaction is spontaneous.

Iron is more active than hydrogen. Therefore iron will oxidize in the presence of H^+.

22. Because tap water is slightly acidic, water pipes made of iron corrode over time, as shown by the balanced ionic equation below:

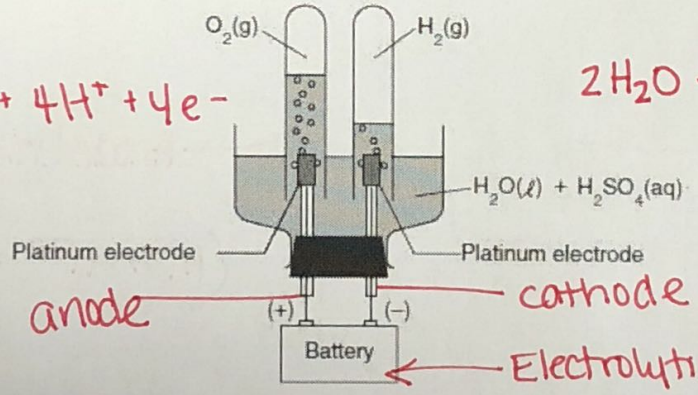
$2Fe + 6H^{+} \rightarrow 2Fe^{3+} + 3H_2$

Explain, in terms of chemical reactivity, why copper pipes are *less* likely to corrode than iron pipes.

Copper pipes are less likely to corrode since Copper is less active than hydrogen.

Base your answers to questions 23 and 24 on the information and diagram below.

The apparatus shown in the diagram consists of two inert platinum electrodes immersed in water. A small amount of an electrolyte, H_2SO_4, must be added to the water for the reaction to take place. The electrodes are connected to a source that supplies electricity.



$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

anode

cathode (*always connected to neg. terminal of battery)

Electrolytic

23. What type of electrochemical cell is shown?

Electrolytic.

24. What particles provided by the electrolyte allow an electric current to flow?

The electrolyte provides freely moving ions.