NOTES

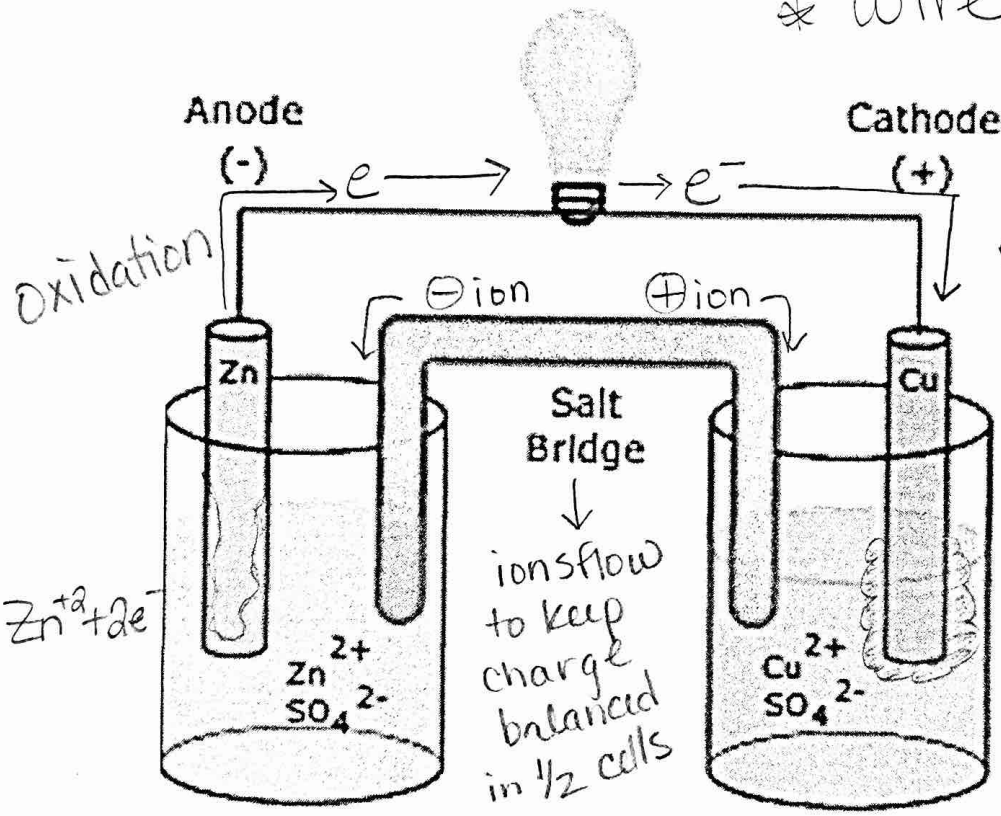
Voltaic Cells = Batteries

SPONTANEOUS!

Chemical energy → electrical energy

Components

* Site of OX (anode)
* Site of Red (cathode)
* Electrolyte (salt bridge) → flow of ions
* 2 Solutions
* wire



$Zn^0 \rightarrow Zn^{+2} + 2e^-$

$Cu^{+2} + 2e^- \rightarrow Cu^0$

* more reactive metal is oxidized

An Ox

oxidation happens at the anode

→ more reactive metal (higher on Table J)

→ mass decreases

electron flow

Red Cat

reduction happens at the cathode

→ less reactive metal (lower on Table J)

→ mass increases

NOTES

Electrolytic Cells

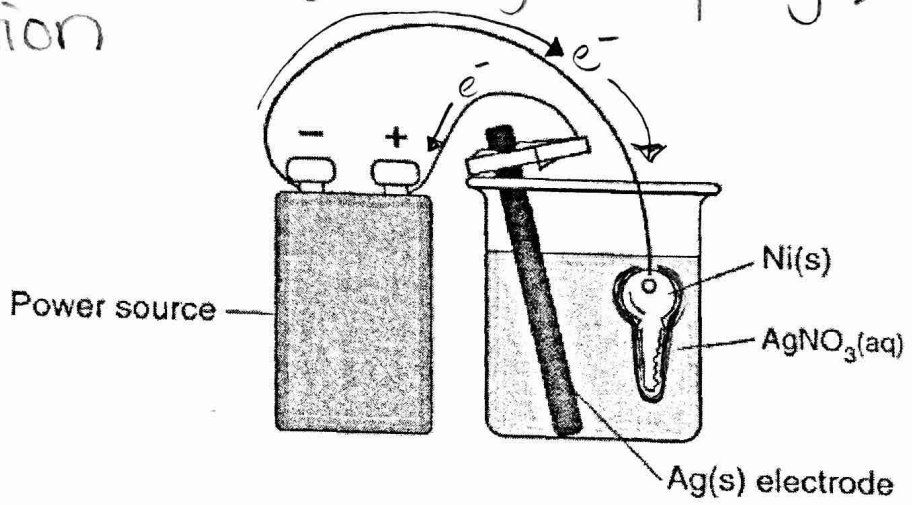
* electrolysis
* electroplating

NON-SPONTANEOUS

Electrical energy ⟶ chemical energy

Components:

* 2 electrodes (site for ox & red)
*** Power source (battery or "plug")
* 1 Solution
* Wire



We are using a power source to force a non-spontaneous reaction

↓

* less reactive metal is oxidized

Anode = Oxidation $\xrightarrow{e^- \text{ flow}}$ Cathode = reduction

$Ag^0 \longrightarrow Ag^{+1} + 1e^-$

→ mass decreases

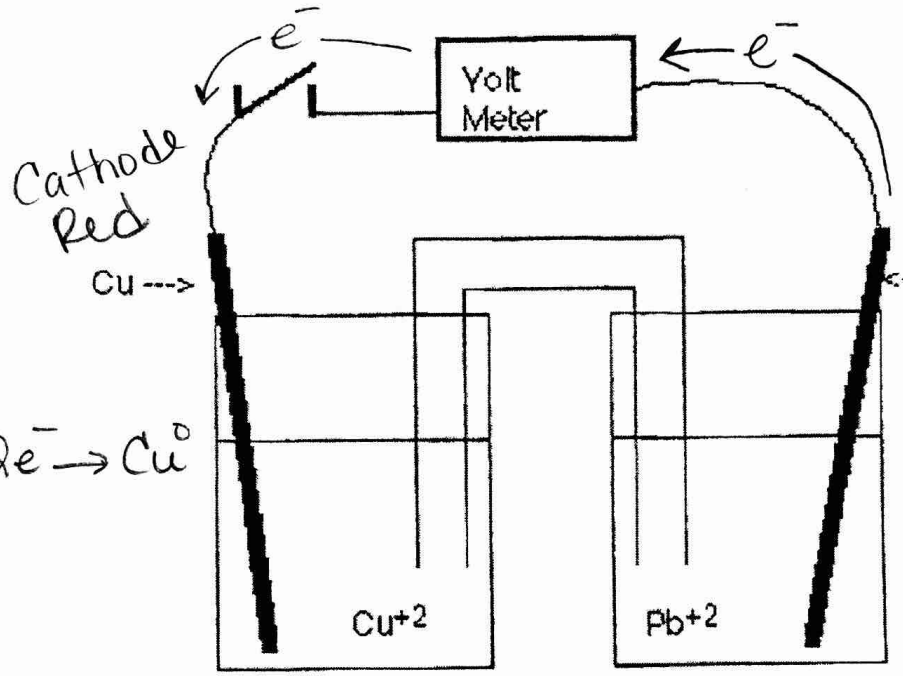
$Ag^{+1} + 1e^- \longrightarrow Ag^0$

→ mass increases

Electrochemical Cells Practice- Voltaic Cells Name: _____

Directions: For each of the given cell diagrams

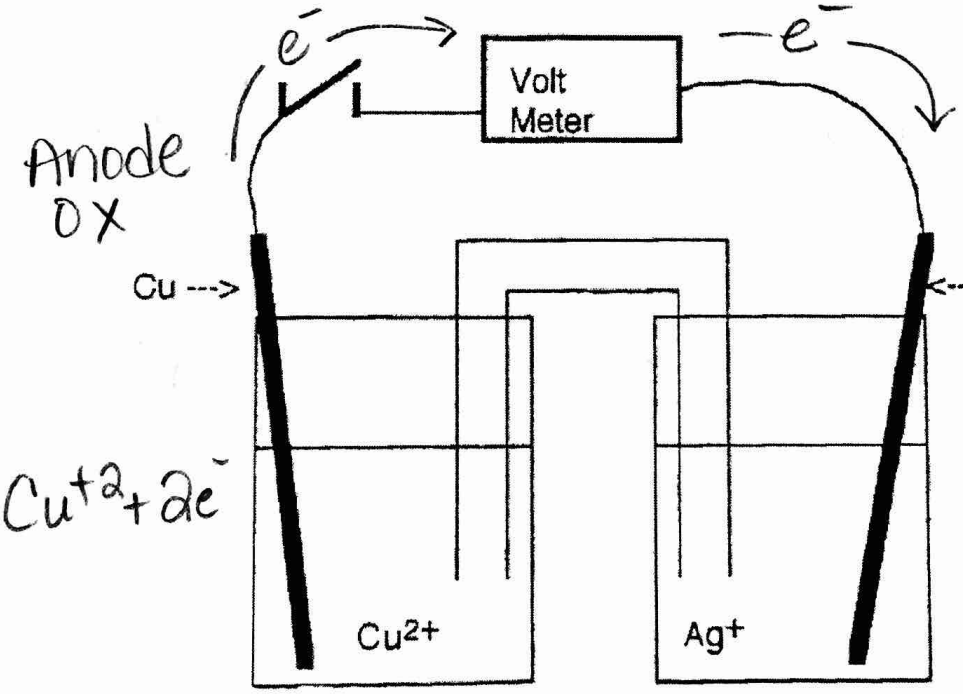
1. Determine which of the 2 metals will be oxidized
2. Label the cathode and the anode
3. Write the chemical half reactions for each cell
4. Write a balanced net ionic redox equation
5. Show the direction of electron flow



$Cu^{+2} + 2e^- \rightarrow Cu^0$

$Pb^0 \rightarrow Pb^{+2} + 2e^-$

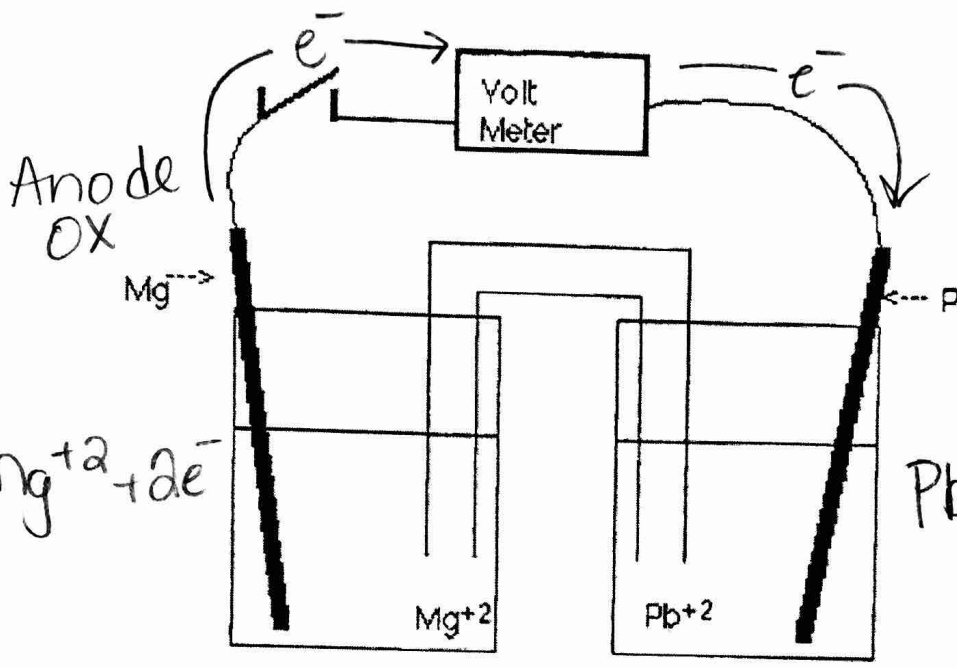
$Cu^{+2} + Pb^0 \rightarrow Cu^0 + Pb^{+2}$



$Cu^0 \rightarrow Cu^{+2} + 2e^-$

$Ag^{+1} + 1e^- \rightarrow Ag^0$

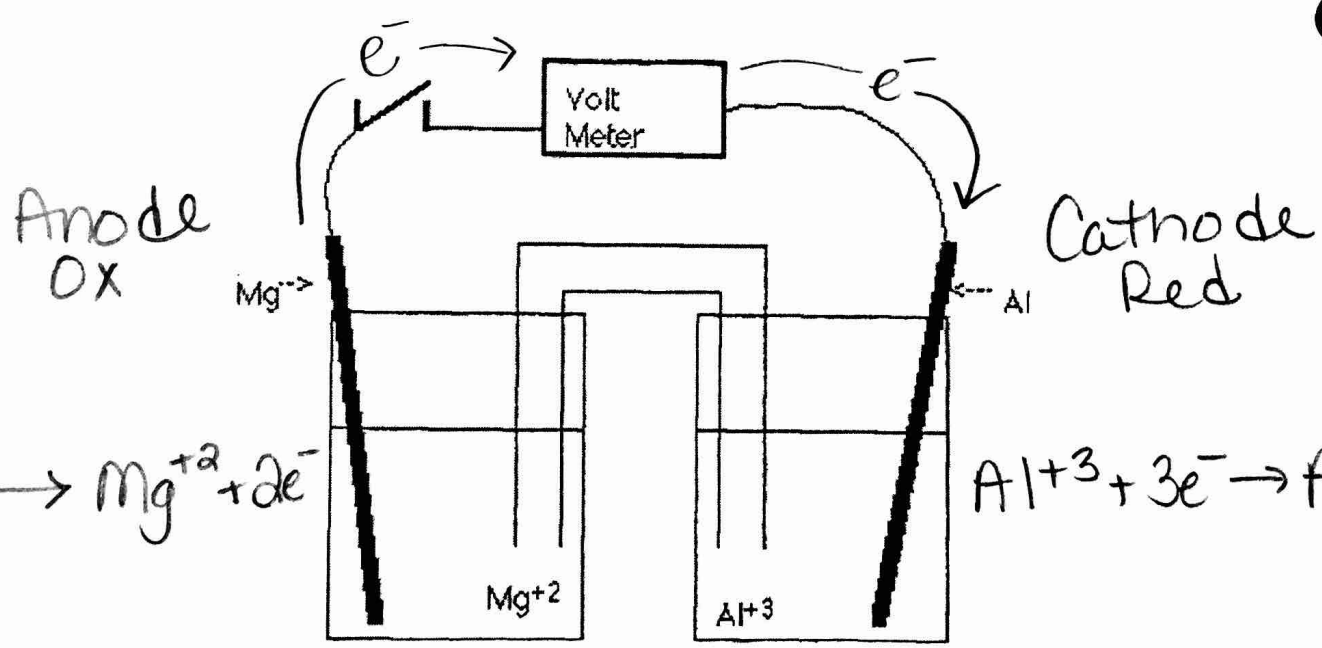
$Cu^0 + 2Ag^{+1} \rightarrow Cu^{+2} + 2Ag^0$



$Mg^0 \rightarrow Mg^{+2} + 2e^-$

$Pb^{+2} + 2e^- \rightarrow Pb^0$

$Mg^0 + Pb^{+2} \rightarrow Mg^{+2} + Pb^0$



$Mg^0 \rightarrow Mg^{+2} + 2e^-$

$Al^{+3} + 3e^- \rightarrow Al^0$

$3Mg^0 + 2Al^{+3} \rightarrow 3Mg^{+2} + 2Al^0$

Electrochemical Cells Practice- Comparing Cells

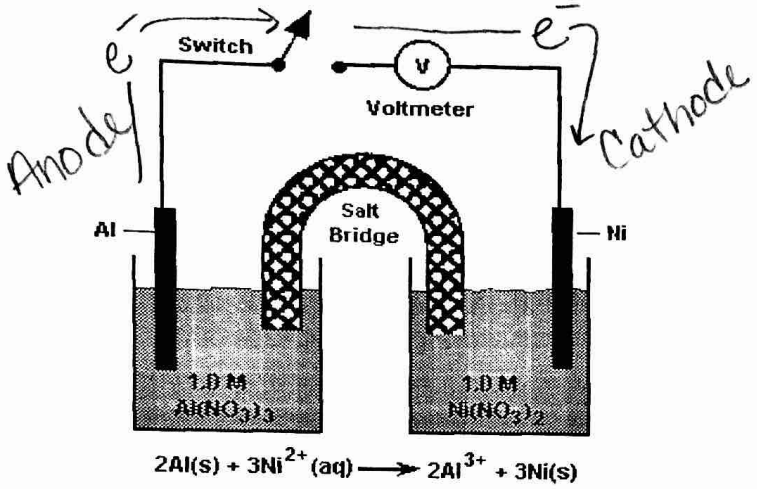
Name: _____

Compare the two types of electrochemical cells by filling in the applicable information for each below.
*** Remember:

→ The anode is always oxidized and the cathode is always reduced.

→ Electrons always flow from anode to cathode (alphabetical order – A → C)

Voltaic Cells



Parts:

Spontaneous ✓ or Requires Energy ____?

1 OR 2 Solutions? 2

Salt Bridge? yes

1 or 2 cells? 2

Table J:

Naturally occurring so electrons must spontaneously flow from anode to cathode. The most reactive metal (most likely to be oxidized) is actually being oxidized.

Anode: Al (Al or Ni) Acts as a ____ (+ or -)

Cathode: Ni (Al or Ni) Acts as a ____ (+ or -)

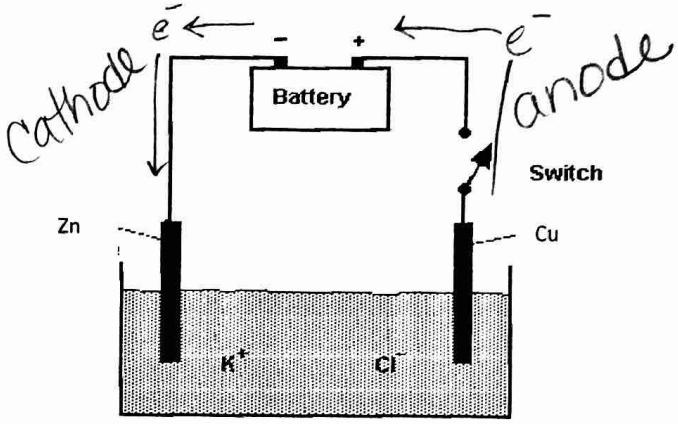
→ Label the anode and cathode in the cell above

Electron Flow: *(from* Al (anode) *to* Ni (cathode)*)*

Energy Conversion: *(from* chem *to* electric*)*

Used for:
Battery; Fuel Cells

Electrolytic Cells



Parts:

Spontaneous ____ or Requires Energy ✓?

1 OR 2 Solutions? 1

Salt Bridge? no

1 or 2 cells? 1

Table J:

Not naturally occurring so electrons must be forced to flow. This means the most reactive metal (most likely to be oxidized) is actually reduced. So the location on Table J must be opposite.

Anode: Cu (Zn or Cu) Acts as a ____ (+ or -)

Cathode: Zn (Zn or Cu) Acts as a ____ (+ or -)

→ Label the anode and cathode in the cell above.

Electron Flow: *(From* Cu *to* Zn*)*

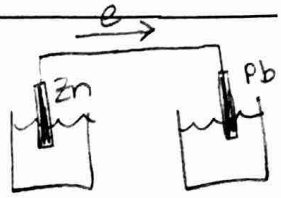
Energy Conversion: *(from* electric *to* chem*)*

Used for:
Battery Chargers; Electroplating

Electrochemical Cells Regents Practice

Name: KEY

1. 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 $Pb^{2+}(aq)$

2. A student collects the materials and equipment below to construct a voltaic cell.

- two 250-mL beakers
- wire and a switch
- one strip of magnesium
- one strip of copper
- 125 mL of 0.20 M $Mg(NO_3)_2(aq)$
- 125 mL of 0.20 M $Cu(NO_3)_2(aq)$

Which additional item is required for the construction of the voltaic cell?

(1) an anode
(2) a battery
(3) a cathode
(4) a salt bridge

3. A voltaic cell spontaneously converts chemical energy to

(1) electrical energy
(2) geothermal energy
(3) mechanical energy
(4) nuclear energy

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

$2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$ $Na^{+1} \rightarrow Na^{0}(e)$ reduced → cathode

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

5. 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

6. 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.

7. Which energy conversion occurs during the operation of an electrolytic cell?

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

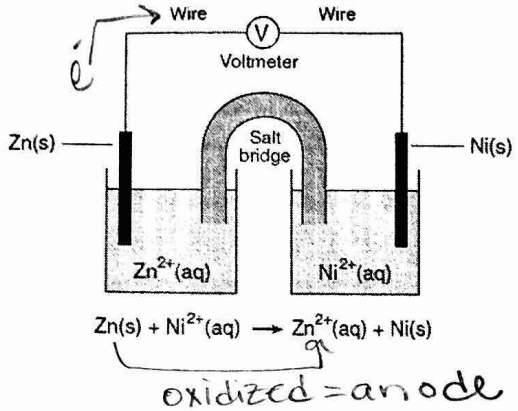
8. What occurs at one of the electrodes in both an electrolytic cell and a voltaic cell?

(1) Oxidation occurs as electrons are gained at the cathode.
(2) Oxidation occurs as electrons are lost at the anode.
(3) Reduction occurs as electrons are gained at the anode.
(4) Reduction occurs as electrons are lost at the cathode.

9. Which statement describes one characteristic of an operating electrolytic cell?

(1) It produces electrical energy.
(2) It requires an external energy source.
(3) It uses radioactive nuclides.
(4) It undergoes a spontaneous redox reaction.

10. The diagram below represents an operating electrochemical cell and the balanced ionic equation for the reaction occurring in the cell.

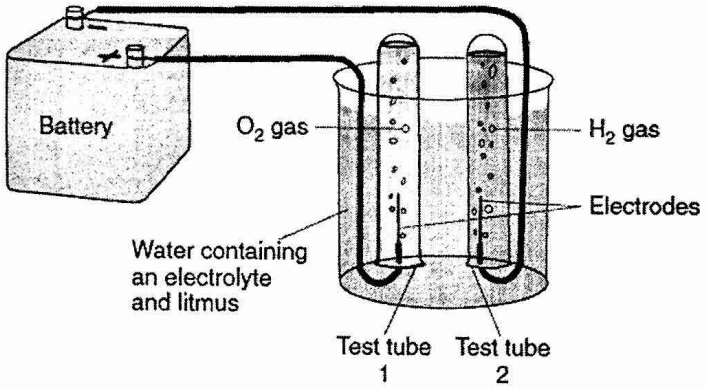


Which statement identifies the part of the cell that conducts electron and describes the direction of electron flow as the cell operates?

(1) Electrons flow through the salt bridge from the Ni(s) to the Zn(s).
(2) Electrons flow through the salt bridge from the Zn(s) to the Ni(s).
(3) Electrons flow through the wire from the Ni(s) to the Zn(s).
(4) Electrons flow through the wire from the Zn(s) to the Ni(s).

Base your answers to questions 11 through 13 on the information below.

The diagram below shows a system in which water is being decomposed into oxygen gas and hydrogen gas. Litmus is used as an indicator in the water. The litmus turns red in test tube 1 and blue in test tube 2.



The oxidation and reduction occurring in the test tubes are represented by the balanced equations below.

Test tube 1: $2H_2O(\ell) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

Test tube 2: $4H_2O(\ell) + 4e^- \rightarrow 2H_2(g) + 4OH^-(aq)$

11. Identify the information in the diagram that indicates this system is an electrolytic cell.

* an external power source (battery)
* only 1 solution

12. Determine the change in oxidation number of oxygen during the reaction in test tube 1.

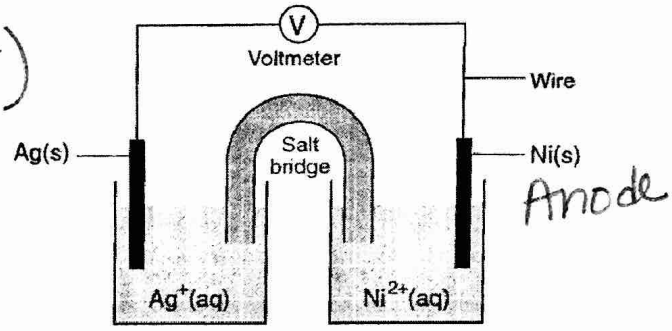
$O^{-2} \rightarrow O^0$ / $-2 \rightarrow 0$

13. Explain, in terms of the products formed in test tube 2, why litmus turns blue in test tube 2.

In test tube 2 OH^- (basic) is a product and litmus turns blue in bases (pH > 8.3)

Base your answers to questions 14 through 16 on the information below.

The diagram below represents an operating voltaic cell at 298 K and 1.0 atmosphere in a laboratory investigation. The reaction occurring in the cell is represented by the balanced ionic equation below.



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

4moles x moles

14. Identify the anode in this cell.

Nickle (higher on Table J)

15. Determine the total number of moles of $Ni^{2+}(aq)$ ions produced when 4.0 moles of Ag^+ (aq) ions completely react in this cell.

2 moles

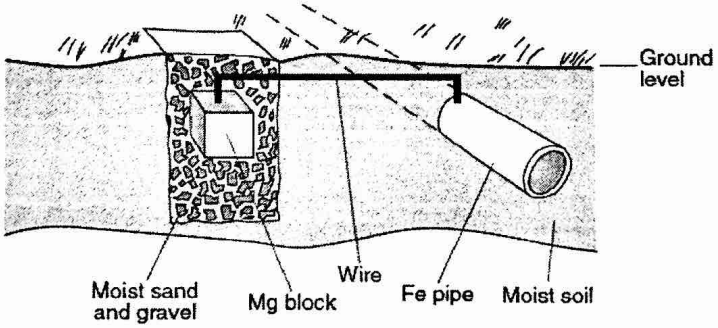
16. Write a balanced half-reaction equation for the reduction that occurs in this cell.

$Ag^{+1} + 1e^- \longrightarrow Ag^0$

Base your answers to questions 17 and 18 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



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

Mg $\longrightarrow$ Fe

18. 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 way more reactive than Fe (6 away on Table J) and Zn is only a little more reactive than F (2 away on Table J)

Base your answers to questions 19and 20 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 $AgNO_3(aq)$.

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

Because AgCl is insoluble and so no Ag^{+1} would be in solution

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

to provide an electrical energy (e^-) to force the reaction to occur

An Electroplating Cell