



Use the following information to answer the next question.

*Superphénix*, a nuclear “breeder” reactor in Creys-Mépieu, France, was shut down for repair in the late 1980s after a liquid sodium leak. The liquid sodium was used as a coolant in the reactor.

18. The liquid sodium used in the reactor can be produced by the electrolysis of molten  $\text{NaCl}_{(l)}$ . During this electrolysis,
- Cl<sup>-</sup> ox in anode: loses e<sup>-</sup> e<sup>-</sup> move to Na<sup>+</sup>*
- ☒ A. chloride ions migrate toward the anode
- B. oxygen gas is produced at the cathode
- C. sodium ions are reduced at the anode
- D. the pH around the cathode increases

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

### VOLTAIC AND ELECTROLYTIC CELLS

Voltaic cells and electrolytic cells are compared in the following chart:

	Voltaic Cell	Electrolytic Cell
Spontaneous?	Yes	No
$E_{\text{net}}^{\circ}$	Positive	Negative
Direction of electron flow	Anode to cathode	Anode to cathode
Anode	Oxidation	Oxidation
Cathode	Reduction	Reduction
Direction of ion flow	Anions to anode; cations to cathode	Anions to anode; cations to cathode
Salt bridge	Required	Usually not required

Use the following information to answer the next question.

When a car is started, the starter motor draws a current from the battery. The battery recharges while the car is running.

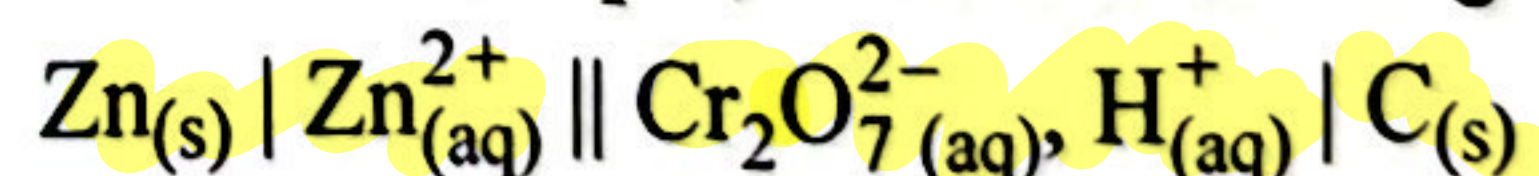
19. A car is started and then left running to recharge the battery. In these two processes, the battery
- A. acts as an electrolytic cell only
- B. acts as a voltaic cell only
- C. first acts as an electrolytic cell, then as a voltaic cell
- ☒ D. first acts as a voltaic cell, then as an electrolytic cell
20. An electrolytic cell differs from a voltaic cell in that an electrolytic cell
- ~~A. is spontaneous~~ *nope*
- ☒ B. consumes electricity *✓*
- ~~C. has an anode and a cathode~~ *same both*
- ~~D. has a positive  $E_{\text{net}}^{\circ}$  value~~

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

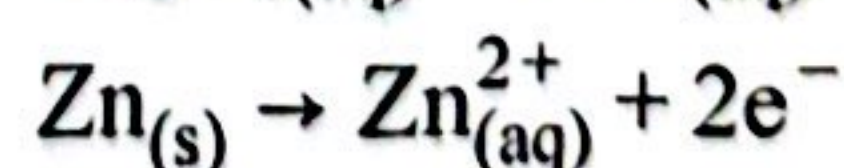
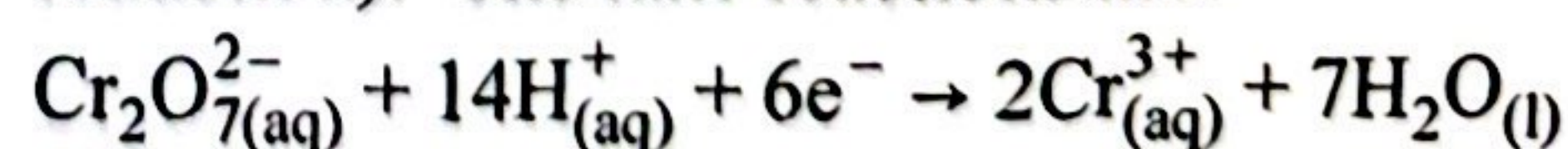
### ELECTROCHEMICAL CELL HALF-REACTION EQUATIONS

Cells are designated by line notation: anode | anode electrolyte || cathode electrolyte | cathode

For this example, use the following cell:



By examining the "Table of Selected Standard Electrode Potentials" in your Data Booklet, it can be seen that  $\text{Zn}_{(s)}$  is the strongest reducing agent (undergoing oxidation), and that  $\text{Cr}_2\text{O}_7^{2-}{}_{(aq)}$  with  $\text{H}^+{}_{(aq)}$  is the strongest oxidizing agent (undergoing reduction). The half-reactions are:

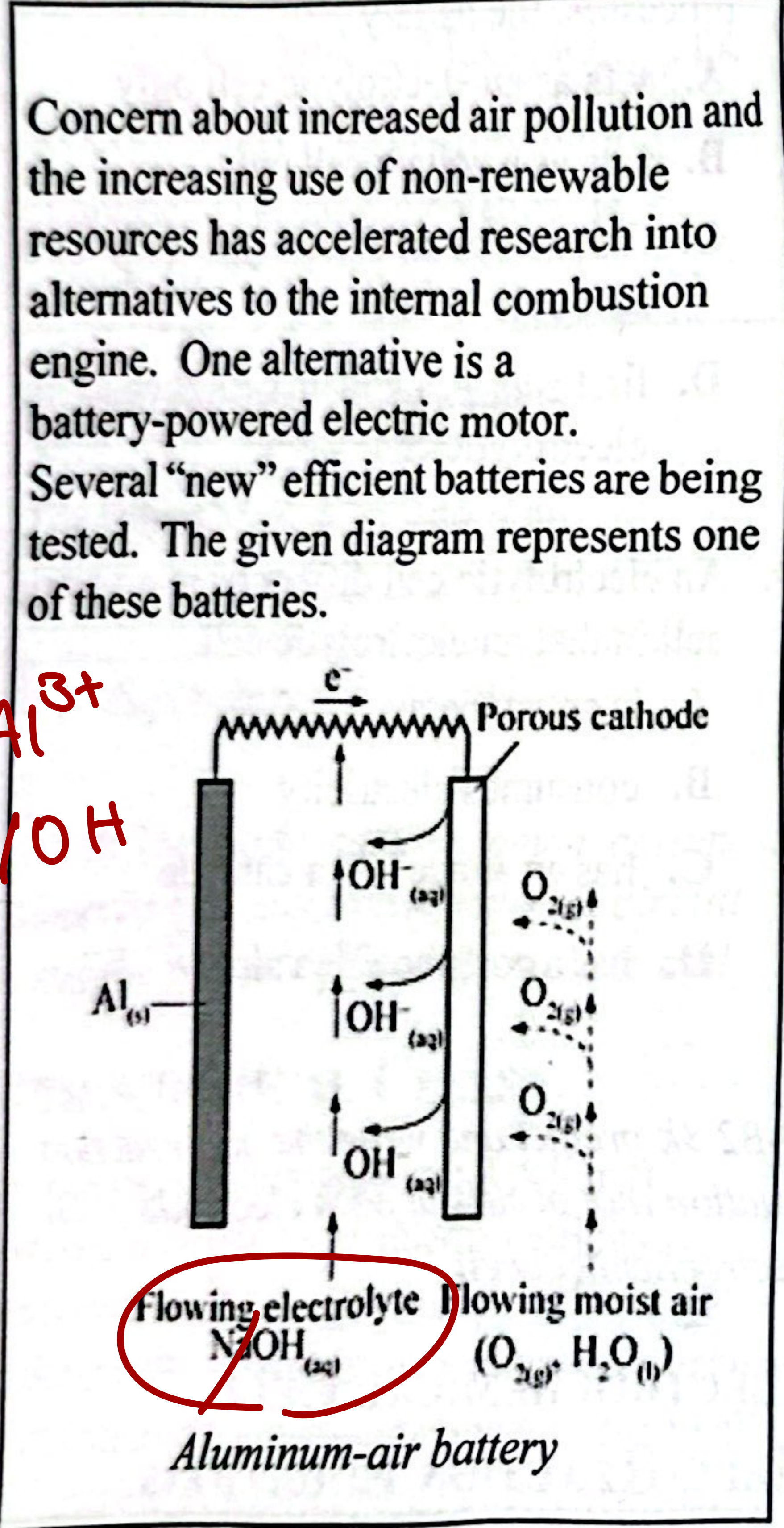






Since the  $\text{Cr}_2\text{O}_7^{2-}$  half-reaction does not contain a solid electrode to serve as a conductor, an inert electrode,  $\text{C}_{(s)}$ , is used in this half-cell. Since this half-reaction is a reduction half-reaction, by definition the cathode is  $\text{C}_{(s)}$ . For the Zn half-reaction, the  $\text{Zn}_{(s)}$  can serve as an electrode. Since this half-reaction is an oxidation half-reaction, by definition the anode is  $\text{Zn}_{(s)}$ .

Use the following information to answer the next question.

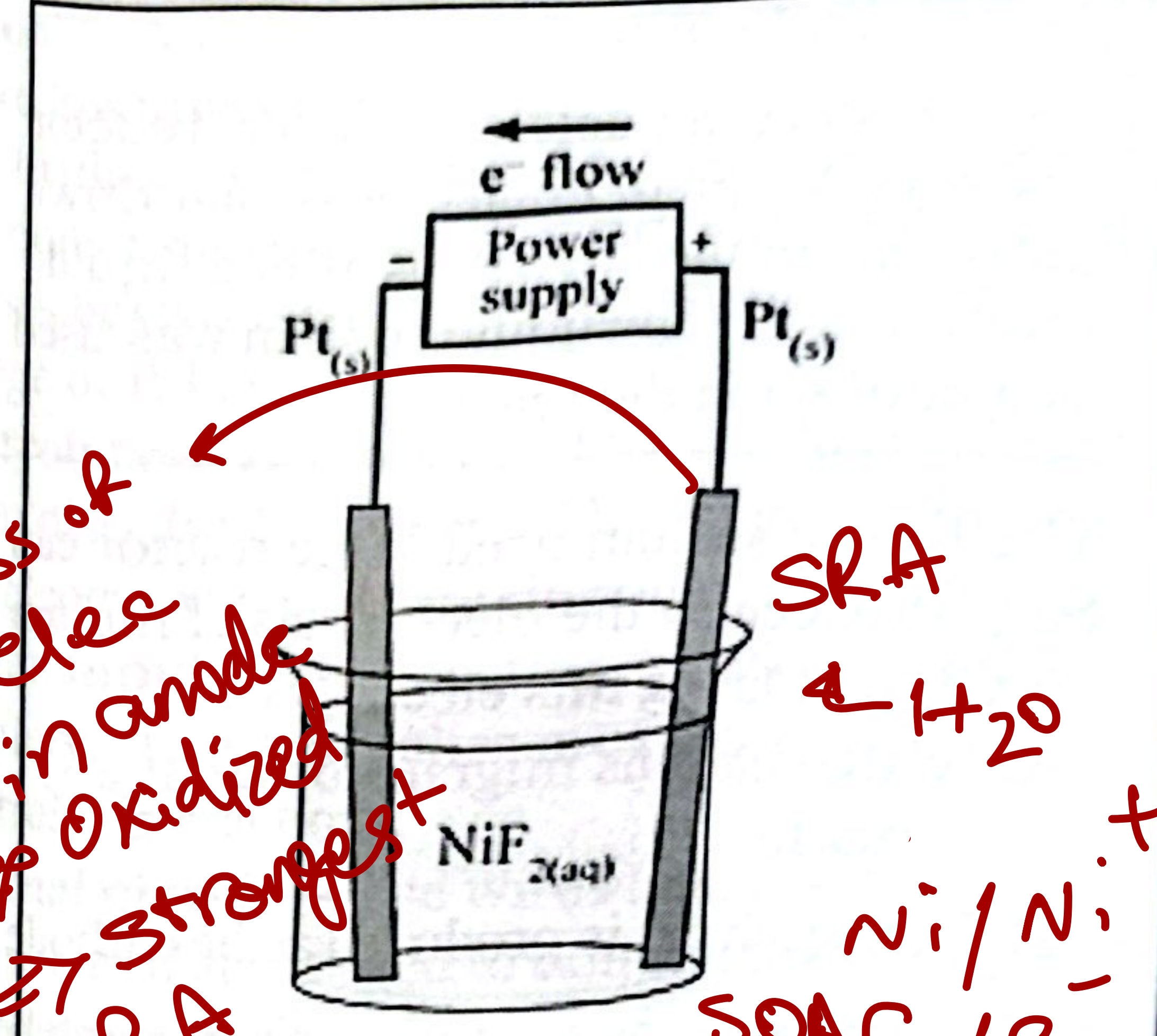


$\text{Al}/\text{Al}^{3+}$   
 $\text{H}_2\text{O}/\text{OH}^-$

21. The reduction half-reaction for this aluminum-air battery is

- A.  $\text{Na}^+_{(aq)} + e^- \rightarrow \text{Na}_{(s)}$  *no Na just*
- B.  $\text{O}_{2(g)} + 4\text{H}^+_{(aq)} + 4e^- \rightarrow 2\text{H}_2\text{O}_{(l)}$  *not acidic*
- C.  $2\text{H}_2\text{O}_{(l)} + 2e^- \rightarrow \text{H}_{2(g)} + 2\text{OH}^-_{(aq)}$  *no H2 gas*
- D.  $\text{O}_{2(g)} + 2\text{H}_2\text{O}_{(l)} + 4e^- \rightarrow 4\text{OH}^-_{(aq)}$

Use the following information to answer the next question.



The electrode on the right is the i, and the half-reaction occurring at its surface is ii.

22. This statement is completed by the information in table

A.	<u>i</u>	<u>ii</u>
	<u>anode</u>	$2\text{F}^-_{(aq)} \rightarrow 2e^- + \text{F}_{2(g)}$
B.	<u>i</u>	<u>ii</u>
	<u>anode</u>	$2\text{H}_2\text{O}_{(l)} \rightarrow \text{O}_{2(g)} + 4\text{H}^+_{(aq)} + 4e^-$
C.	<u>i</u>	<u>ii</u>
	cathode	$\text{Ni}^{2+}_{(aq)} + 2e^- \rightarrow \text{Ni}_{(s)}$
D.	<u>i</u>	<u>ii</u>
	cathode	$2\text{H}_2\text{O}_{(l)} + 2e^- \rightarrow \text{H}_{2(g)} + 2\text{OH}^-_{(aq)}$





30-B2.4k recognize that predicted reactions do not always occur

## EXCEPTIONS TO PREDICTED REDOX REACTIONS

Usually, redox reactions can be predicted by finding the strongest oxidizing agent (SOA) and strongest reducing agent (SRA), writing their half-reactions, multiplying the half-reactions to make electrons gained equal to electrons lost, and adding the half-reactions.

In some cases a reaction other than the predicted reaction will occur. This is particularly notable for the electrolysis of aqueous chloride compounds. It is referred to as the *chloride anomaly*.

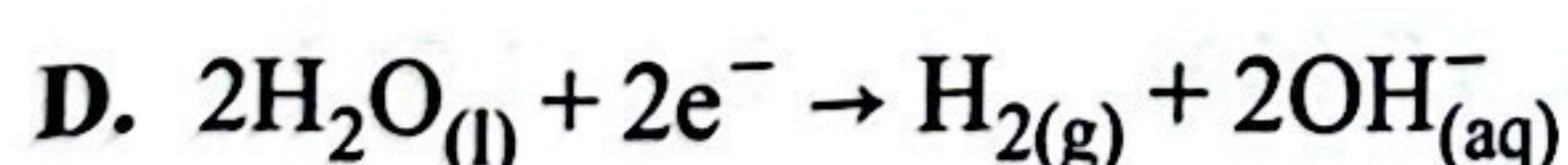
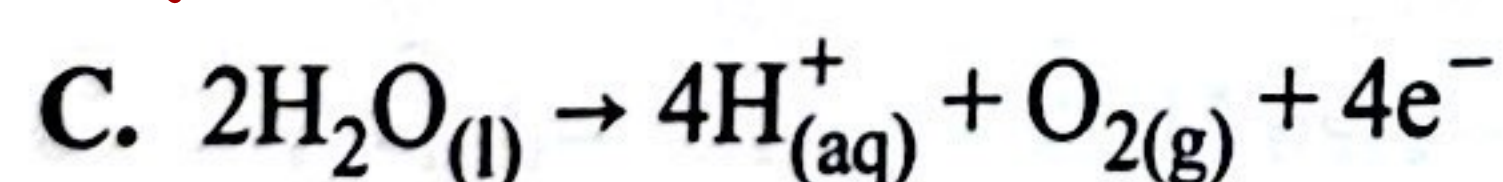
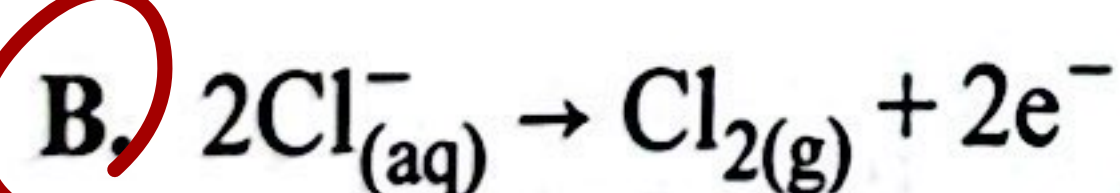
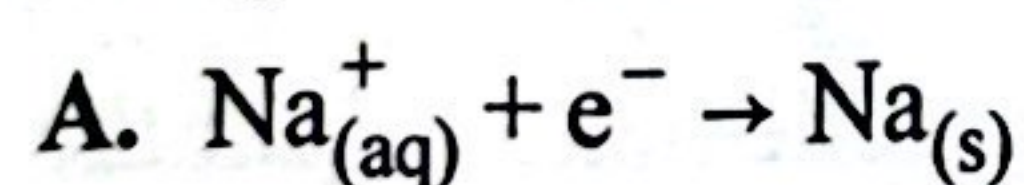
In the electrolysis of aqueous sodium chloride (brine), the SOA and SRA both appear to be  $\text{H}_2\text{O(l)}$ . The predicted half-reactions are  $2\text{H}_2\text{O(l)} + 2\text{e}^- \rightarrow \text{H}_{2(\text{g})} + 2\text{OH}^-_{(\text{aq})}$  at the cathode and

$2\text{H}_2\text{O(l)} \rightarrow \text{O}_{2(\text{g})} + 4\text{H}^+_{(\text{aq})} + 4\text{e}^-$  at the anode.

When this electrolysis is performed, the predicted cathode half-reaction does occur at the cathode, but at the anode, the half-reaction that actually occurs is  $2\text{Cl}^-_{(\text{aq})} \rightarrow \text{Cl}_{2(\text{g})} + 2\text{e}^-$

The chloride anomaly can be accounted for using a more complete version of the theory. You are not required to explain the anomaly, but you should know that it occurs.

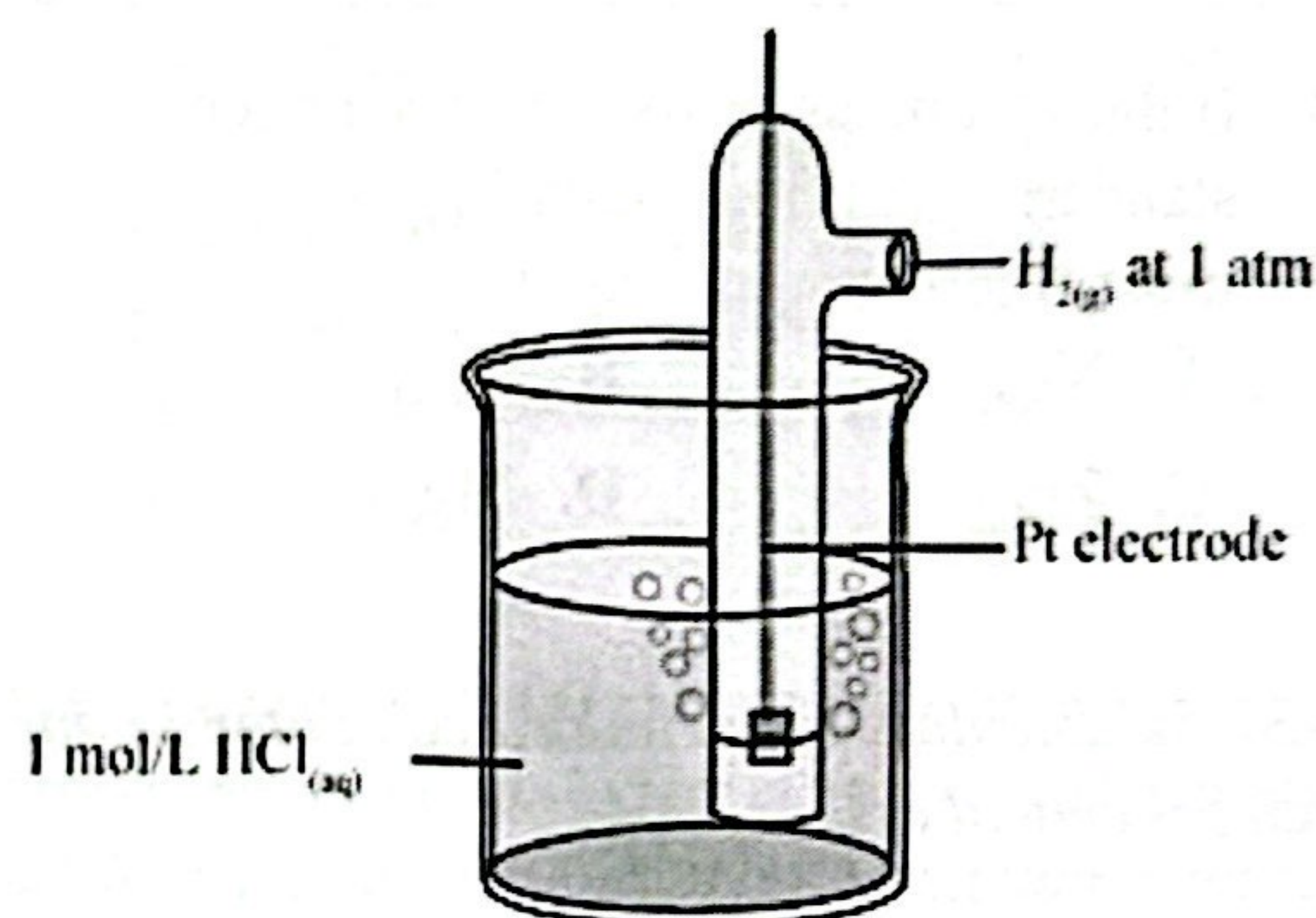
23. Many companies obtain caustic soda from the electrolysis of brine ( $\text{NaCl}_{(\text{aq})}$ ). The **primary** reaction that occurs at the anode during the electrolysis of  $\text{NaCl}_{(\text{aq})}$  is



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

## STANDARD REDUCTION POTENTIALS

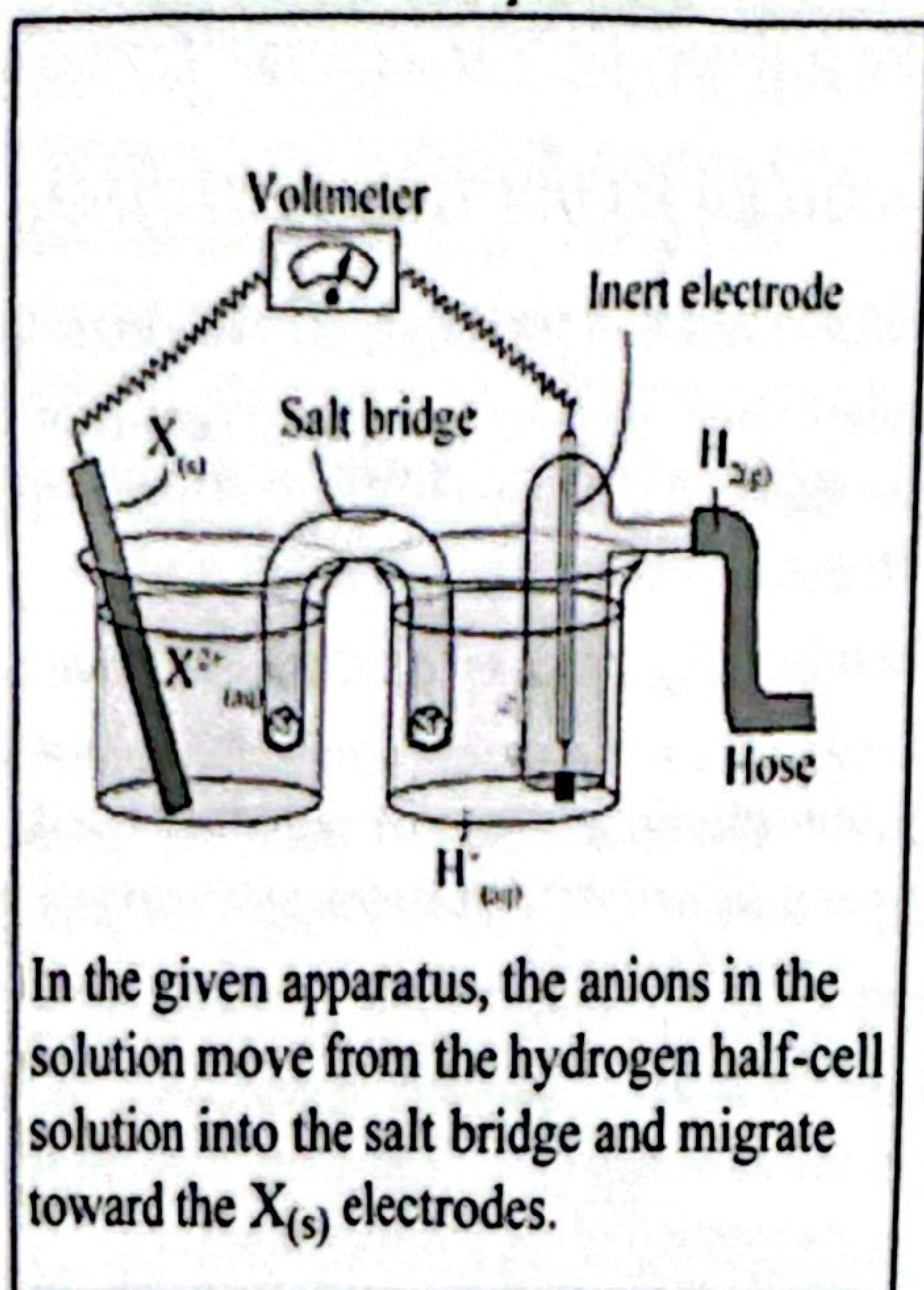
Standard reduction potentials,  $E_r^\circ$  (also known as standard electrode potentials), are measured relative to the standard hydrogen half-cell. The standard hydrogen half-cell contains a solution of  $1.00 \text{ mol/L H}^+_{(\text{aq})}$ , an inert platinum electrode, and hydrogen gas at 1 atm pressure bubbling over the platinum electrode. The half-cell is maintained at a temperature of  $25^\circ\text{C}$ . The hydrogen half-cell is assigned an  $E_r^\circ$  value of  $0.00 \text{ V}$ . It may serve as either the anode or the cathode depending on what is in the other half-cell. The given diagram shows a representation of the hydrogen half-cell:



A second half-cell, such as the  $\text{Zn}_{(\text{s})}/\text{Zn}^{2+}_{(\text{aq})}$  half-cell, is then wired to the standard hydrogen half-cell, and a salt bridge is inserted. The standard reduction potential of the second half-cell relative to the standard hydrogen half-cell can be found by measuring the resulting voltage.



Use the following information to answer the next question.



24. If the voltmeter reads +0.45 V under standard conditions, then  $X_{(s)}$  is most likely composed of

A.  $Ni_{(s)}$       **B.  $Fe_{(s)}$**   
C.  $Zn_{(s)}$       D.  $Mg_{(s)}$

30-B2.6k calculate the standard cell potential for electrochemical cells

## STANDARD CELL POTENTIAL

The standard cell potential,  $E^\circ_{net}$ , is also known as the standard cell voltage or standard cell potential (difference). It is calculated by using the formula  $E^\circ_{net} = E^\circ_{r(cathode)} - E^\circ_{r(anode)}$ , where  $E^\circ_r$  is the standard reduction potential (also known as standard electrode potential). Recall that the cathode is where reduction takes place (where the strongest oxidizing agent reacts), and the anode is where oxidation takes place (where the strongest reducing agent reacts).

Once you have determined the strongest oxidizing agent and strongest reducing agent, you will use the anode and cathode half-reactions and you use  $E^\circ_r$  values from your table to find  $E^\circ_{net}$ .

If one of the half-reactions is multiplied by a coefficient, its  $E^\circ_r$  remains the same.

### Example

Determine the cell potential when a cell is made using a standard silver half-cell and a standard zinc half-cell.

From the Standard Electrode Potentials table,

- i.  $Zn_{(s)} \rightleftharpoons Zn^{2+}_{(aq)} + 2e^-$ ,  $E^\circ_r = -0.76$  V  
ii.  $Ag_{(s)} \rightleftharpoons e^- + Ag^+_{(aq)}$ ,  $E^\circ_r = +0.80$  V

$Ag^+_{(aq)}$  is the strongest oxidizing agent and  $Zn_{(s)}$  is the strongest reducing agent, so  $Zn_{(s)}$  is the anode.

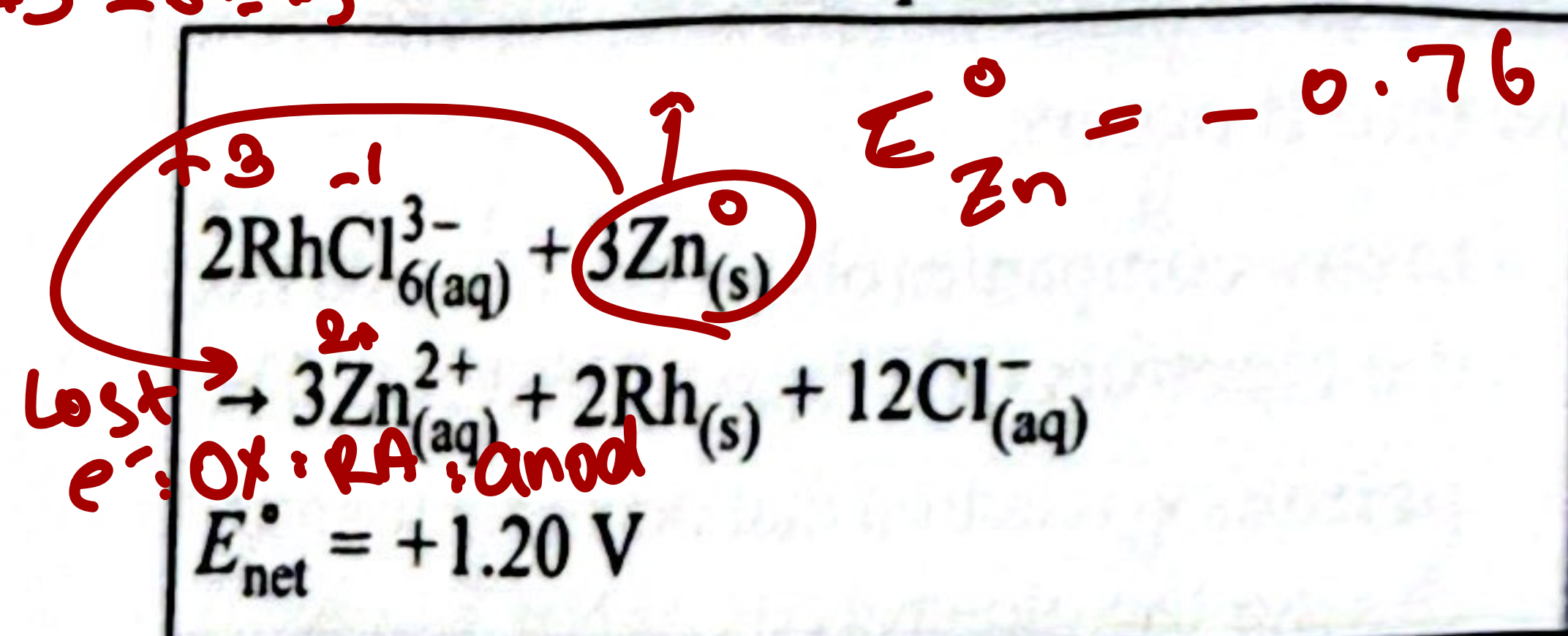
$$\begin{aligned} E^\circ_{net} &= E^\circ_{r(cathode)} - E^\circ_{r(anode)} \\ &= 0.80 \text{ V} - (-0.76 \text{ V}) \\ &= 1.56 \text{ V} \end{aligned}$$

The cell potential is 1.56 V.

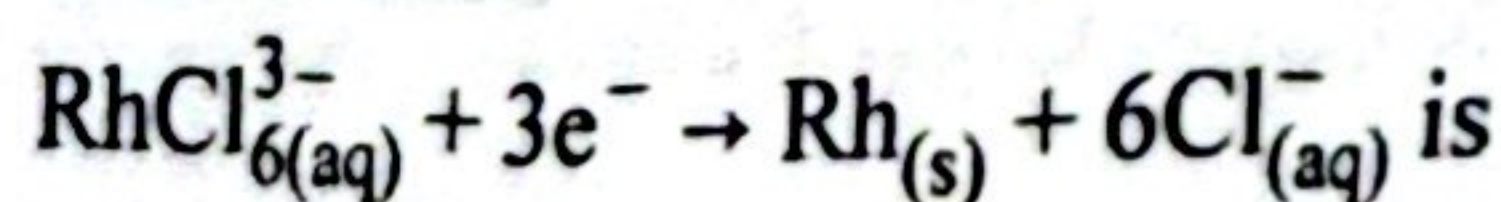
Note that the net reaction is  $Zn_{(s)} + 2Ag^+_{(aq)} \rightarrow Zn^{2+}_{(aq)} + 2Ag_{(s)}$ .

Use the following information to

answer the next question.



25. The standard electrode potential for the half-reaction



- A. -1.96 V      B. -0.44 V  
**C. +0.44 V**      D. +1.96 V

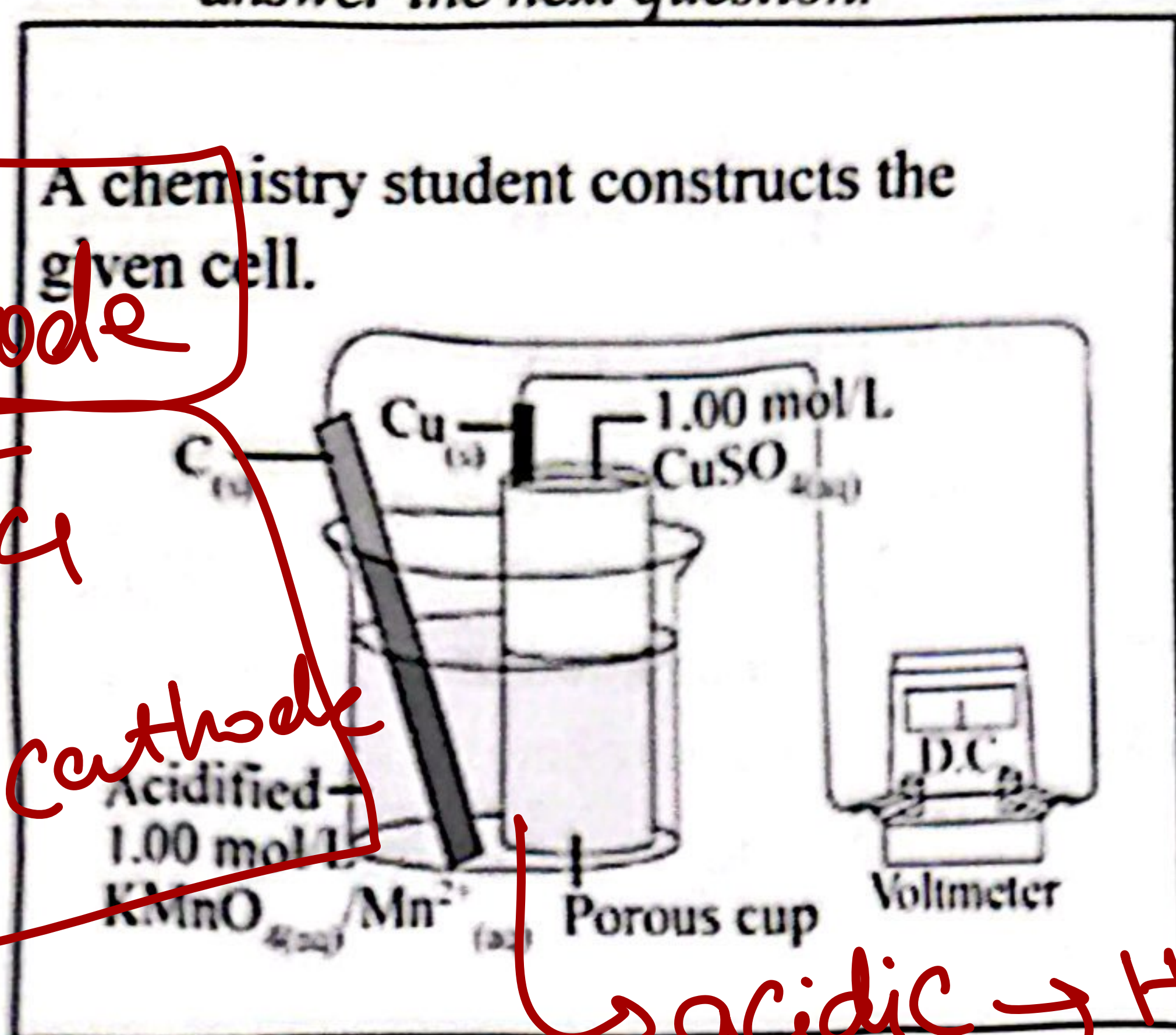
$$1.20 = x - (-0.76)$$

$$1.20 + 0.76 = x$$





Use the following information to answer the next question.



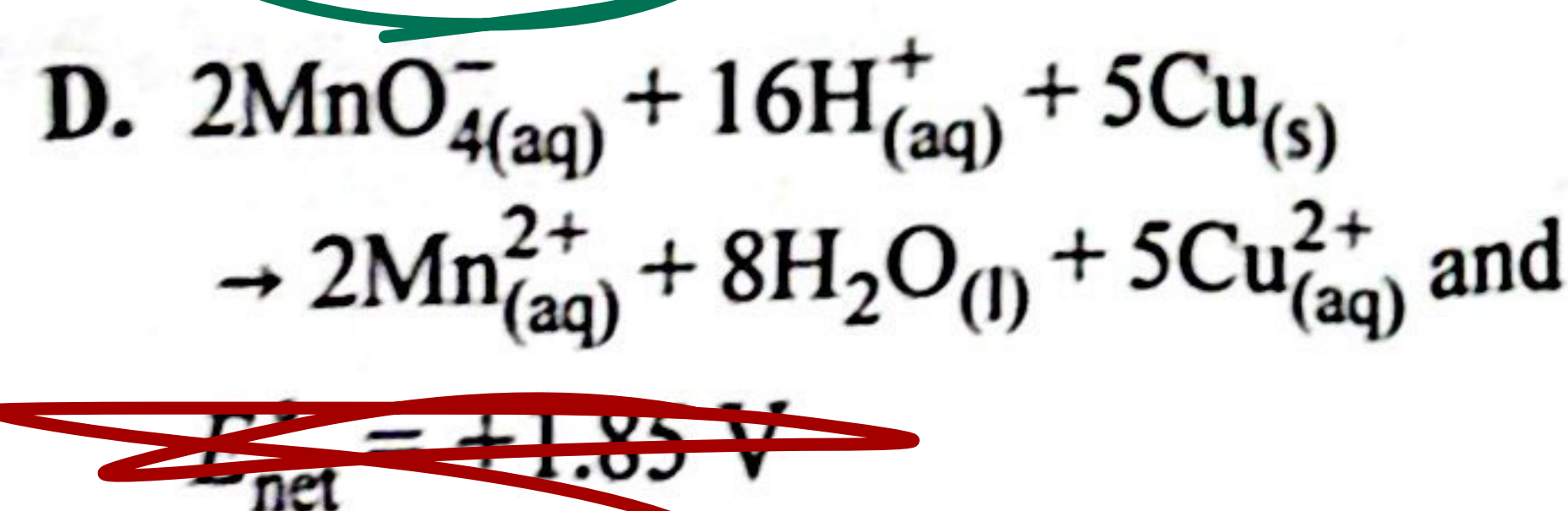
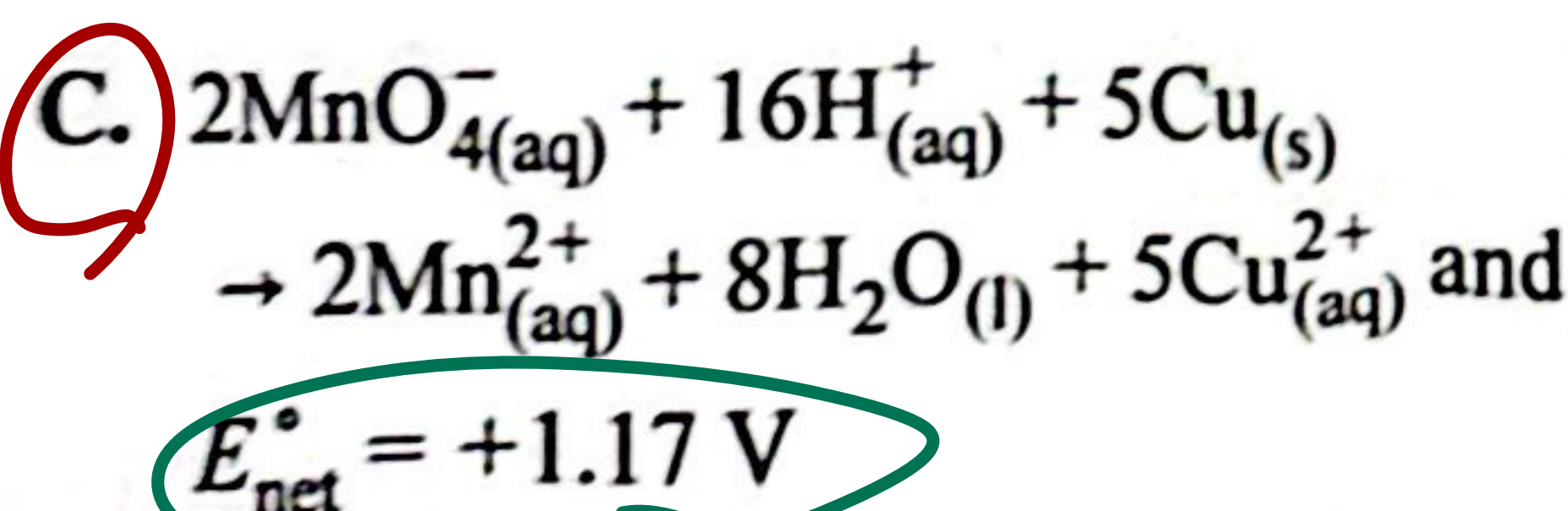
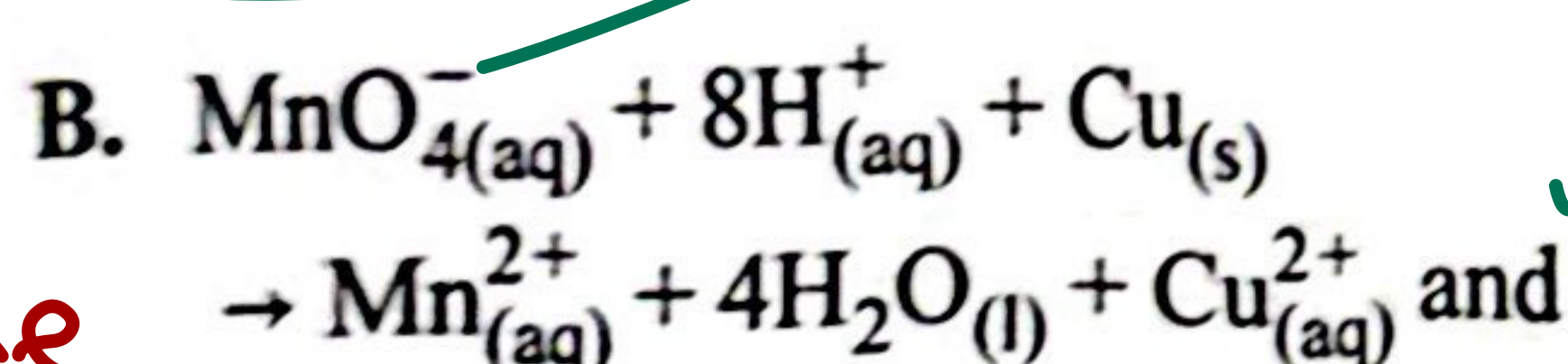
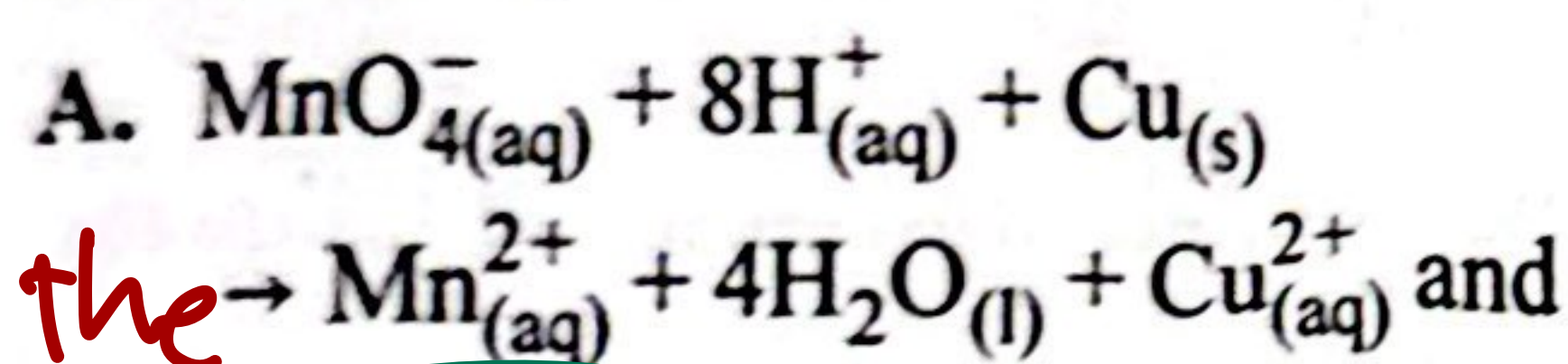
SRA  
Cu: Anode

MnO4-

SOA: Cathode

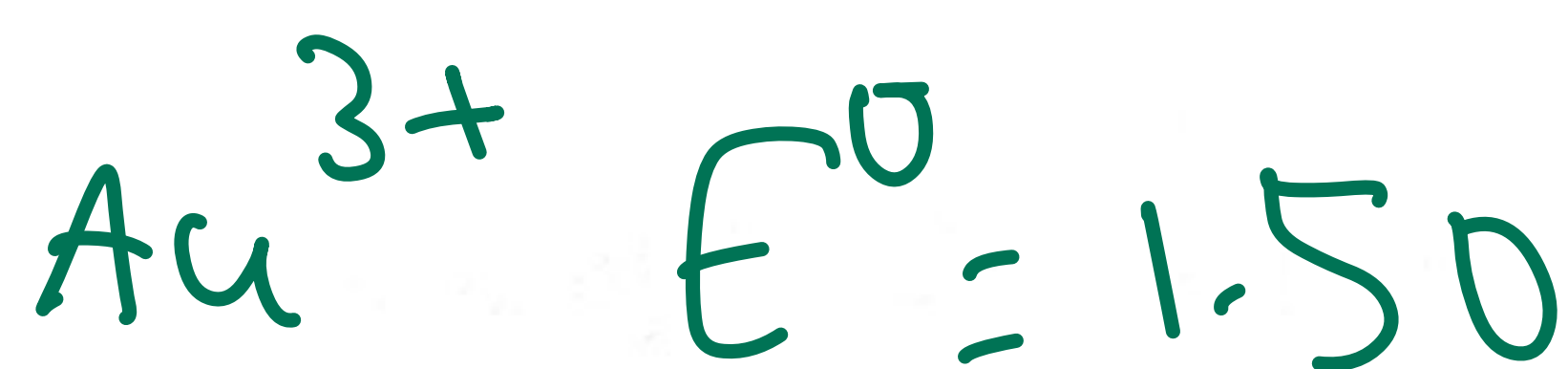
acidic → H+

26. The net equation and predicted voltage for the operating cell are



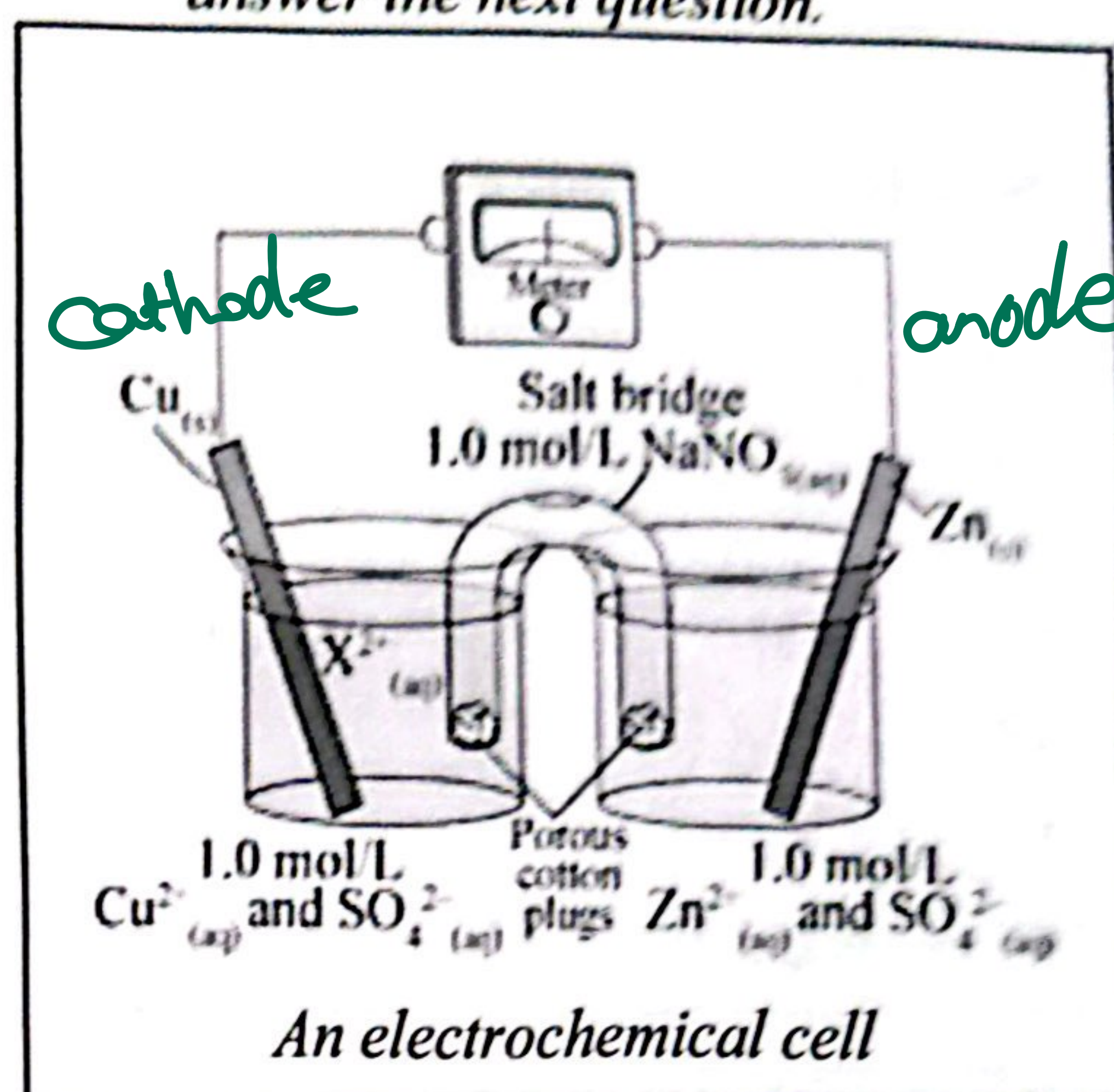
### Numerical Response

27. Under standard conditions, hydrogen gas reacts with  $\text{Au}^{3+}$  ions to produce  $\text{Au}(\text{s})$ . The net cell potential for the reaction is + 1.50 V.



$1.50 - 0 = 1.50$

Use the following information to answer the next question.



### Numerical Response

28. A student attempted to replicate a traditional Daniell cell by setting up the given electrochemical cell. Under standard conditions, the electrical potential of the cell should be +/- 1.10 V. (Record your answer as a three-digit number.)

$0.34 - (-0.76)$

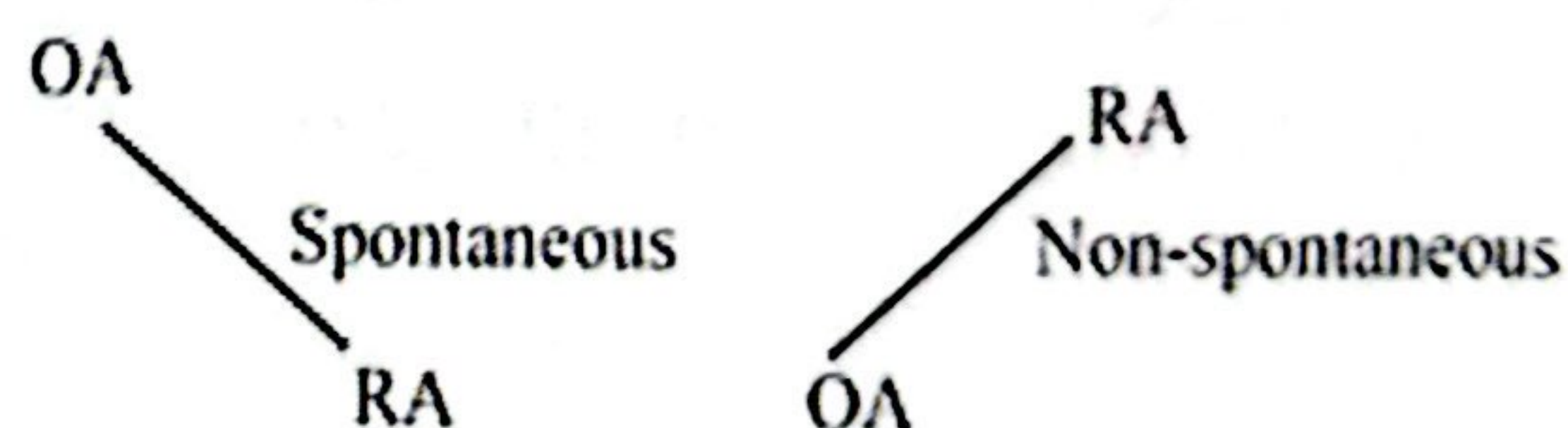
30-B2.7k predict the spontaneity or nonspontaneity of redox reactions, based on standard cell potential, and the relative positions of half-reaction equations on a standard reduction potential table

### SPONTANEITY OF REDOX REACTIONS IN ELECTROCHEMICAL CELLS

Redox spontaneity may be predicted from the spontaneity rule "If the OA is above the RA, the reaction is spontaneous; if the OA is not above the RA, the reaction is non-spontaneous":

Oxidizing agent: OA

Reducing agent: RA



Redox spontaneity may also be predicted by calculating the standard cell potential,  $E_{\text{net}}^\circ$ . If  $E_{\text{net}}^\circ$  is positive the reaction is spontaneous. If  $E_{\text{net}}^\circ$  is negative the reaction is non-spontaneous. If you check several examples, you will see that the two rules are really the same.

voltic cells

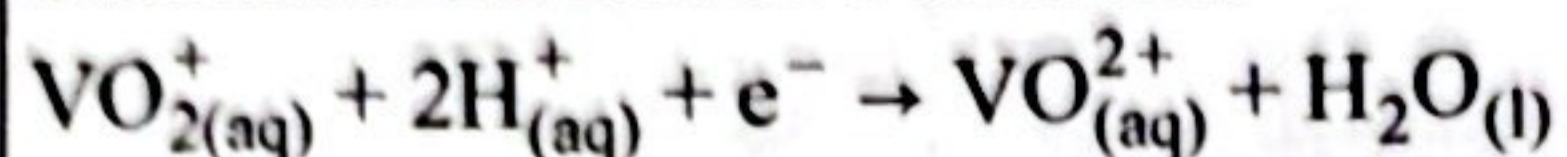
electrolytic cells



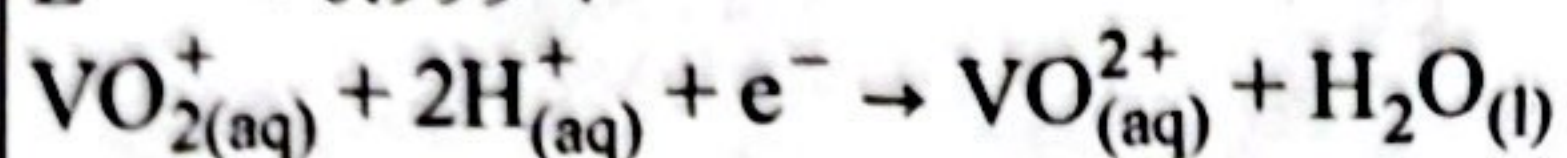


Use the following information to answer the next question.

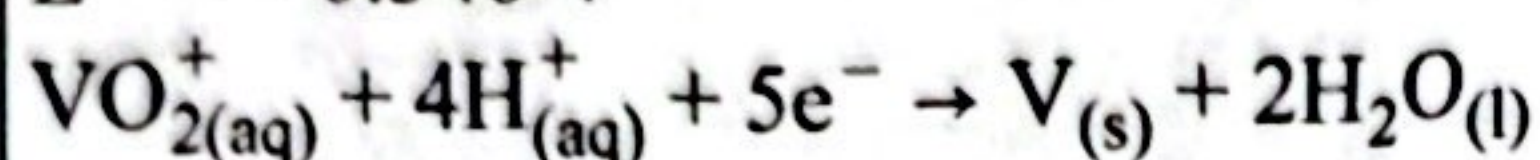
#### Standard Electrode Potentials



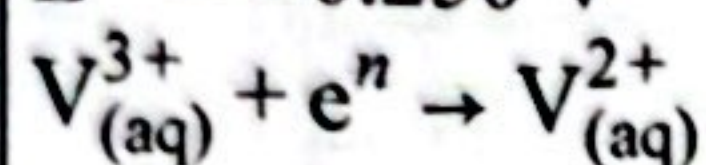
$$E^\circ = +0.999 \text{ V}$$



$$E^\circ = +0.340 \text{ V}$$

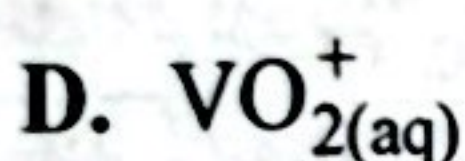
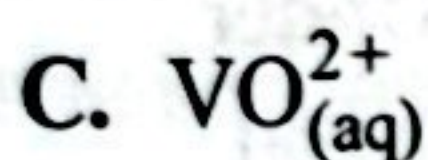
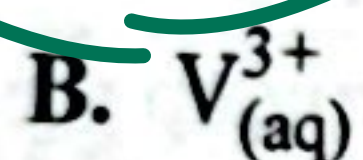
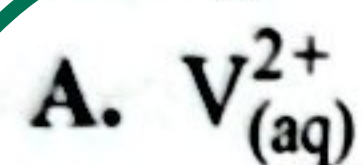


$$E^\circ = -0.250 \text{ V}$$



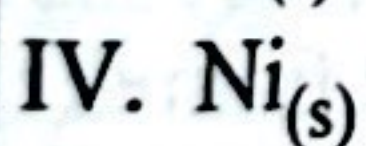
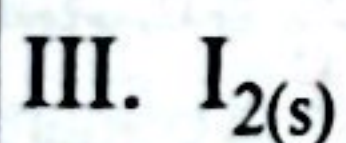
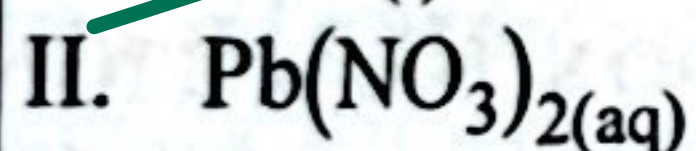
$$E^\circ = -0.255 \text{ V}$$

29. Which of the following substances is the strongest reducing agent?



Use the following information to answer the next question.

Four chemical species are given:



30. Which of the given species will react spontaneously with  $\text{Sn}_{(\text{aq})}^{2+}$ ?

A. I and II

B. III only

C. IV only

D. III and IV

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

## FARADAY'S LAW AND ELECTROCHEMICAL CELLS

Electric current, symbolized  $I$  and measured in amperes (A), is the rate of flow of electric charge.

Electric charge is measured in coulombs (C).

It takes a very large number of electrons

$(6.25 \times 10^{18})$  to produce 1 C of charge.

$$1 \text{ A} = 1 \text{ C/s}$$

Faraday's constant is  $F$ , the number of coulombs per mole of  $\text{e}^-$ .

$$F = 9.65 \times 10^4 \text{ C/mol}$$

To find moles of electrons, use the following formula:

$$n_{\text{e}^-} = \frac{I \times t}{F}$$

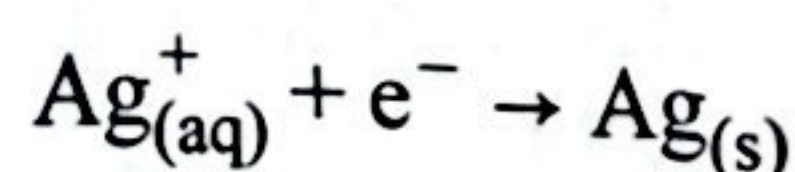
$$= \frac{I \times t}{9.65 \times 10^4 \text{ C/mol}}$$

This equation can be used to relate current and time to mass and number of moles of substances in an electrolysis reaction.

### Example

If a current of 1.50 A is applied for 100 min to a silver nitrate solution, what mass of silver metal could be plated on a steel bar?

### Solution



$\text{e}^-$	$\text{Ag}_{(\text{s})}$
$n_1$	$n_2$
1.50 A	$m = ?$
100 min	

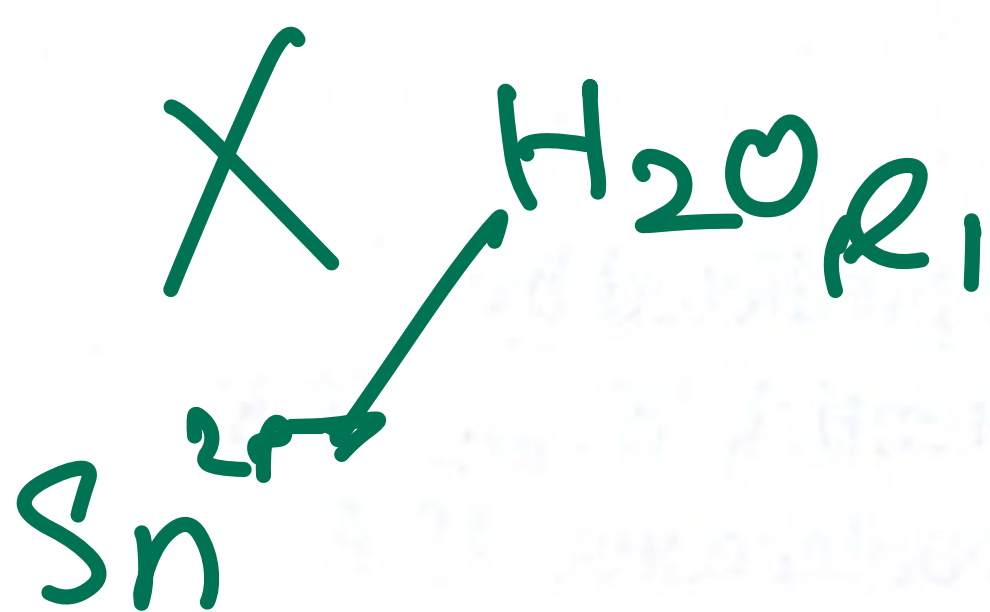
### Step 1

Calculate the number of moles of electrons.

$$n_1 = \frac{I \times t}{9.65 \times 10^4 \text{ C/mol}}$$

$$= \frac{1.50 \text{ C/s} \times 100 \text{ min} \times 60 \text{ s/min}}{9.65 \times 10^4 \text{ C/mol}}$$

$$= 0.0933 \text{ mol of } \text{e}^-$$





**Step 2**

Calculate the number of moles of silver by using the number of moles of electrons.

$$n_2 = 0.0933 \text{ mol of } e^- \times \frac{1 \text{ mol Ag}_{(s)}}{1 \text{ mol of } e^-}$$

$$= 0.0933 \text{ mol of Ag}_{(s)}$$

**Step 3**

Convert the number of moles into mass by using the molar mass.

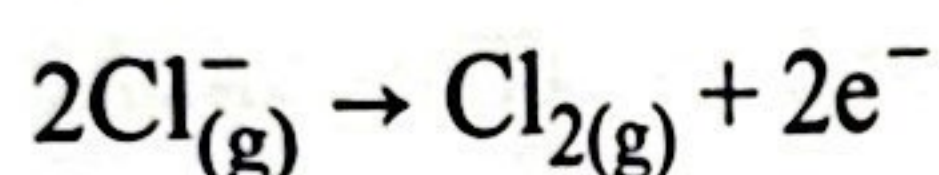
$$m = nM$$

$$m = 0.0933 \text{ mol} \times 107.87 \text{ g/mol}$$

$$= 10.1 \text{ g}$$

**Example**

Calculate the current required to produce 100 g of chlorine,  $\text{Cl}_{2(g)}$ , in a time of 90.0 min from the electrolysis of  $\text{NaCl}_{(aq)}$ .

**Solution**

$\text{Cl}_{2(g)}$	$e^-$
$n_1$	$n_2$
100 g	90 min
	$I = ?$

**Step 1**

Calculate the number of moles of  $\text{Cl}_{2(g)}$ .

$$n_1 = \frac{100 \text{ g}}{70.90 \text{ g/mol}}$$

$$= 1.41 \text{ mol of } \text{Cl}_{2(g)}$$

**Step 2**

Calculate the number of moles of electrons from the number of moles of  $\text{Cl}_{2(g)}$ .

$$n_2 = 1.41 \text{ mol } \text{Cl}_{2(g)} \times \frac{2 \text{ mol of } e^-}{1 \text{ mol } \text{Cl}_{2(g)}}$$

$$= 2.82 \text{ mol}$$

**Step 3**

Calculate the current.

$$n_{e^-} = \frac{I \times t}{F}$$

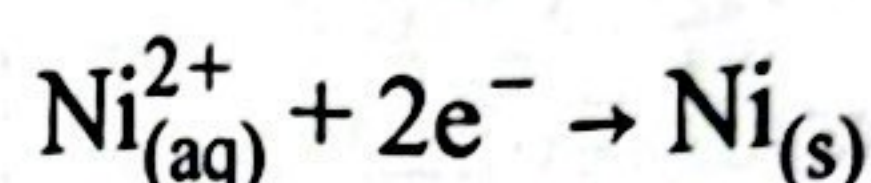
$$I = \frac{n_{e^-} \times F}{t}$$

$$= \frac{2.82 \text{ mol} \times (9.65 \times 10^4 \text{ C/mol})}{90.0 \text{ min} \times 60 \text{ s/min}}$$

$$= 50.4 \text{ C/s or } 50.4 \text{ A}$$

**Example**

Calculate the time required to produce 23.0 g of nickel from the electrolysis of  $\text{NiCl}_{2(aq)}$  by using a current of 15.0 A.

**Solution**

$e^-$	$\text{Ni}_{(s)}$
$n_2$	$n_1$
15.0 A	23.0 g
$t = ?$	

**Step 1**

Calculate the number of moles of  $\text{Ni}_{(s)}$ .

$$n_1 = \frac{23.0 \text{ g}}{58.69 \text{ g/mol}}$$

$$= 0.392 \text{ mol of } \text{Ni}_{(s)}$$

**Step 2**

Calculate the number of moles of electrons by using the number of moles of  $\text{Ni}_{(s)}$ .

$$n_2 = 0.392 \text{ mol } \text{Ni}_{(s)} \times \frac{2 \text{ mol } e^-}{1 \text{ mol } \text{Ni}_{(s)}}$$

$$= 0.784 \text{ mol } e^-$$

**Step 3**

Calculate the time.

$$n_{e^-} = \frac{I \times t}{F}$$

$$t = \frac{n_{e^-} \times F}{I}$$

$$= \frac{0.784 \text{ mol} \times (9.65 \times 10^4 \text{ C/mol})}{15.0 \text{ C/s}}$$

$$= 5.04 \times 10^3 \text{ s}$$

$$= 84.0 \text{ min}$$





Use the following information to answer the next question.

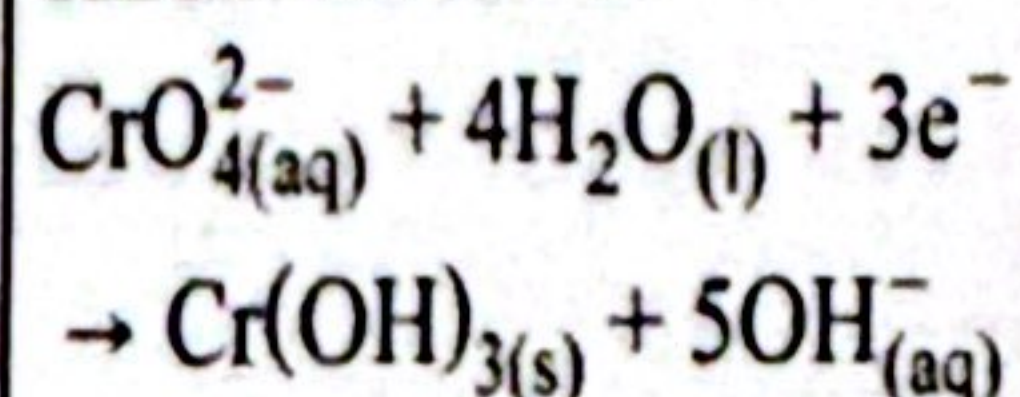
A 6.00 A current is applied to an electrolytic cell for 2.0 h in order to plate an object in silver.

### Numerical Response

31. What mass of silver could be plated onto the object during this time? \_\_\_\_\_ g (Record your answer as a three-digit number.)

Use the following information to answer the next question.

In some industrial processes, sodium chromate is added to water coolants. When the coolant is drained, the chromate ions can be removed through an electrolysis process that uses an iron anode. The products of the electrolysis are aqueous iron(II) ions and solid chromium(III) hydroxide, a recoverable pollutant. The half-reaction involving the chromate ion is



### Numerical Response

32. The electrolytic cell used to produce the  $\text{Cr}(\text{OH})_3(\text{s})$  runs at a current of 3.00 A for 48.0 h. What mass of the iron anode is consumed? \_\_\_\_\_ g





## ANSWERS AND SOLUTIONS

### ELECTROCHEMICAL CHANGES

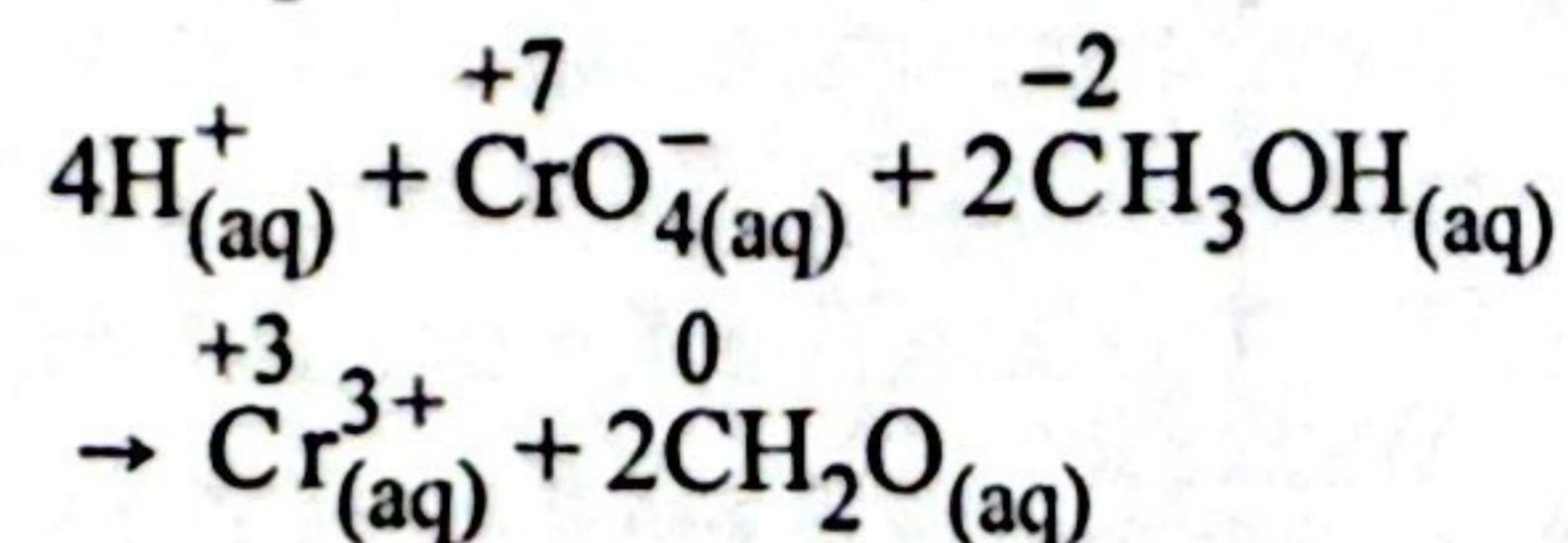
1. B	8. C	15. C	22. B	29. A
2. D	9. B	16. 6631	23. B	30. D
3. B	10. D	17. A	24. B	31. 48.3
4. B	11. 5371	18. A	25. C	32. 150
5. 1167	12. C	19. D	26. C	
6. A	13. B	20. B	27. 1.50	
7. D	14. A	21. D	28. 1.10	

1. B

When zinc metal is oxidized, it loses electrons and forms zinc ions,  $\text{Zn}_{(\text{aq})}^{2+}$ . (The relevant oxidation reaction is the reverse of the reduction at  $E^\circ = -0.76 \text{ V}$  on the Table of Selected Standard Electrode Potentials.)

2. D

Identify the oxidation numbers. (H and O do not change in this case, so they can be ignored.)



The C in  $\text{CH}_3\text{OH}_{(\text{aq})}$  has gone from  $-2$  to  $0$ , so it has lost two electrons. This means it has been oxidized, making it the reducing agent.

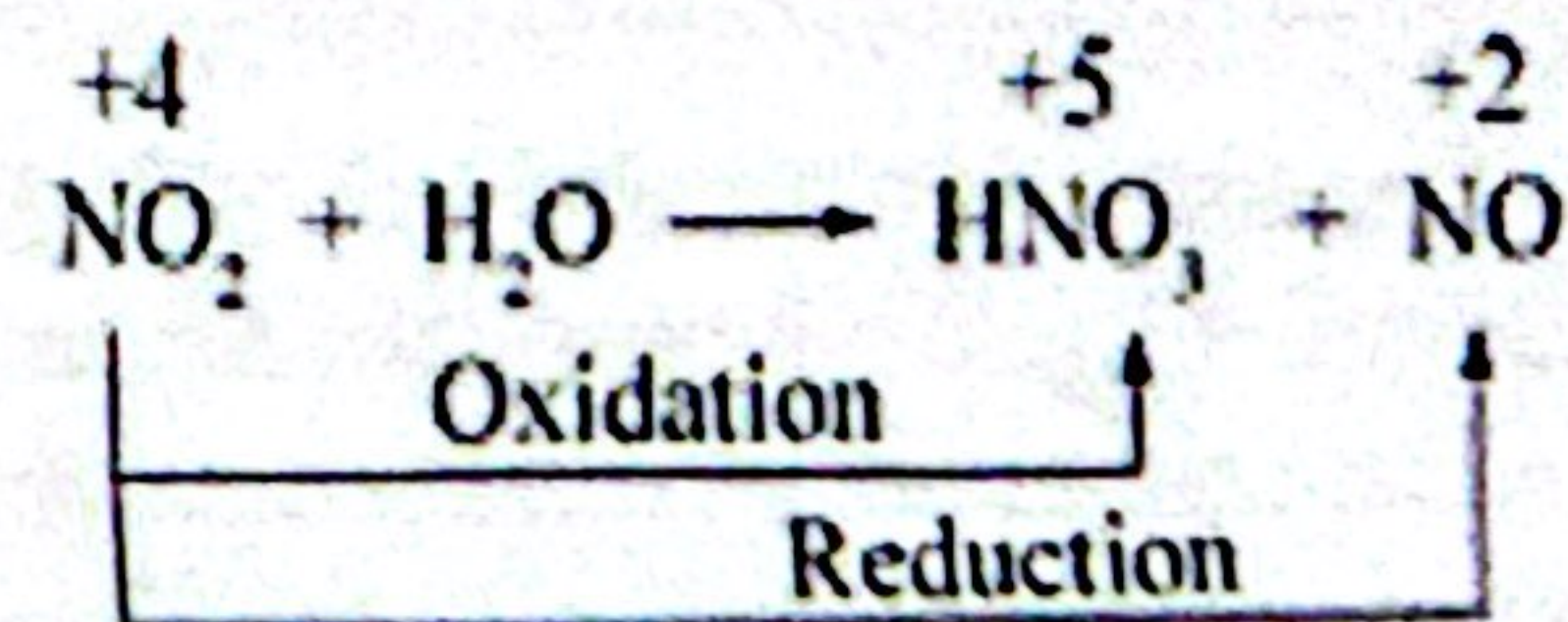
4. B

In a disproportionation (also called an auto-oxidation), the oxidation number of one atom in a reactant both increases and decreases over the course of an oxidation-reduction reaction.

In reaction B, the oxidation number of chlorine changes from  $0$  (in  $\text{Cl}_{2(\text{aq})}$ ) to  $+1$  (in  $\text{HOCl}_{(\text{aq})}$ ) and to  $-1$  (in  $\text{Cl}_{(\text{aq})}^-$ ).

In reactions A, C, and D, the oxidation numbers of two reactant entities change. This renders these reactions as regular oxidation-reduction processes.

5. 1167





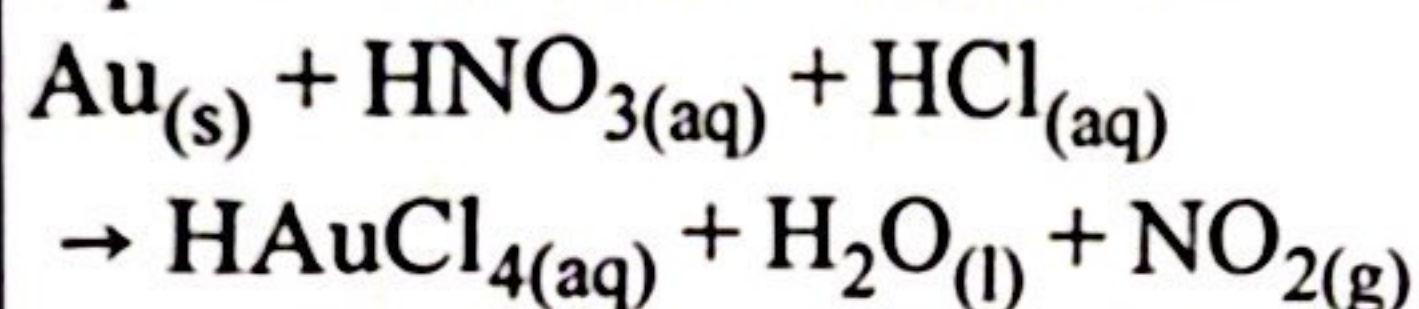


## UNIT TEST — ELECTROCHEMICAL CHANGES

Use the following information to answer the next question.

In order to hide their golden Nobel Prize-winning medals during the Second World War, two scientists at the Niels Bohr Institute in Denmark asked a friend to “dissolve” the gold, store it in a solution, and recover it at the end of the war.

One way to “dissolve” gold is to react it with *aqua regia*, a mixture of nitric and hydrochloric acids. The unbalanced equation for this reaction is:

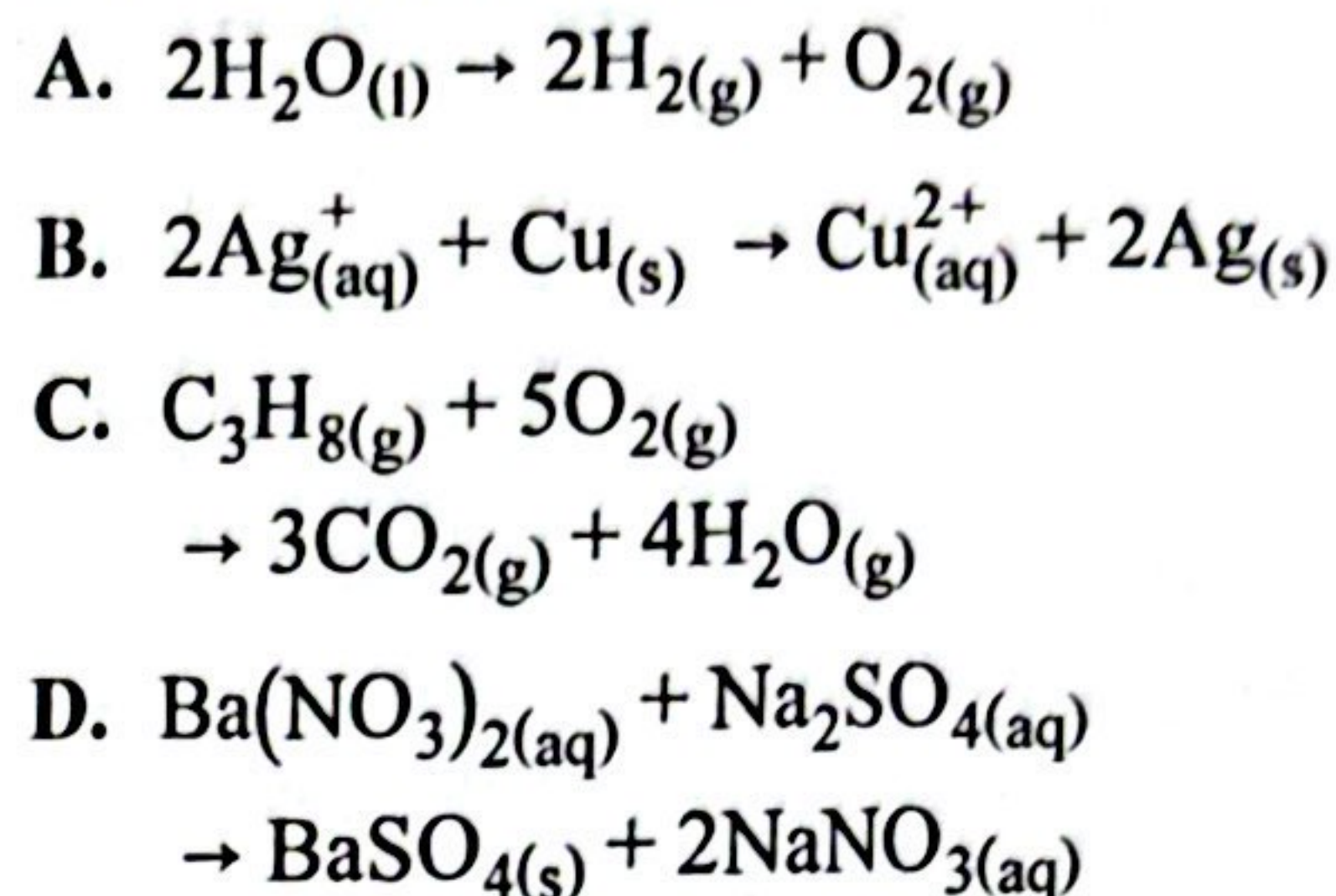


- The atom that undergoes reduction in this reaction is  
 A. Au                                  B. Cl  
 C. N                                    D. H
- If a block of refined copper was selected to serve as a sacrificial anode, and if it was bolted onto the iron hull of a ship, one would expect the  
 A. copper to oxidize  $\text{Cl}_{(aq)}^-$   
 B. copper to oxidize before the iron  
 C. copper to corrode at a faster rate than the iron  
 D. copper to remain unchanged and the iron to continue to oxidize

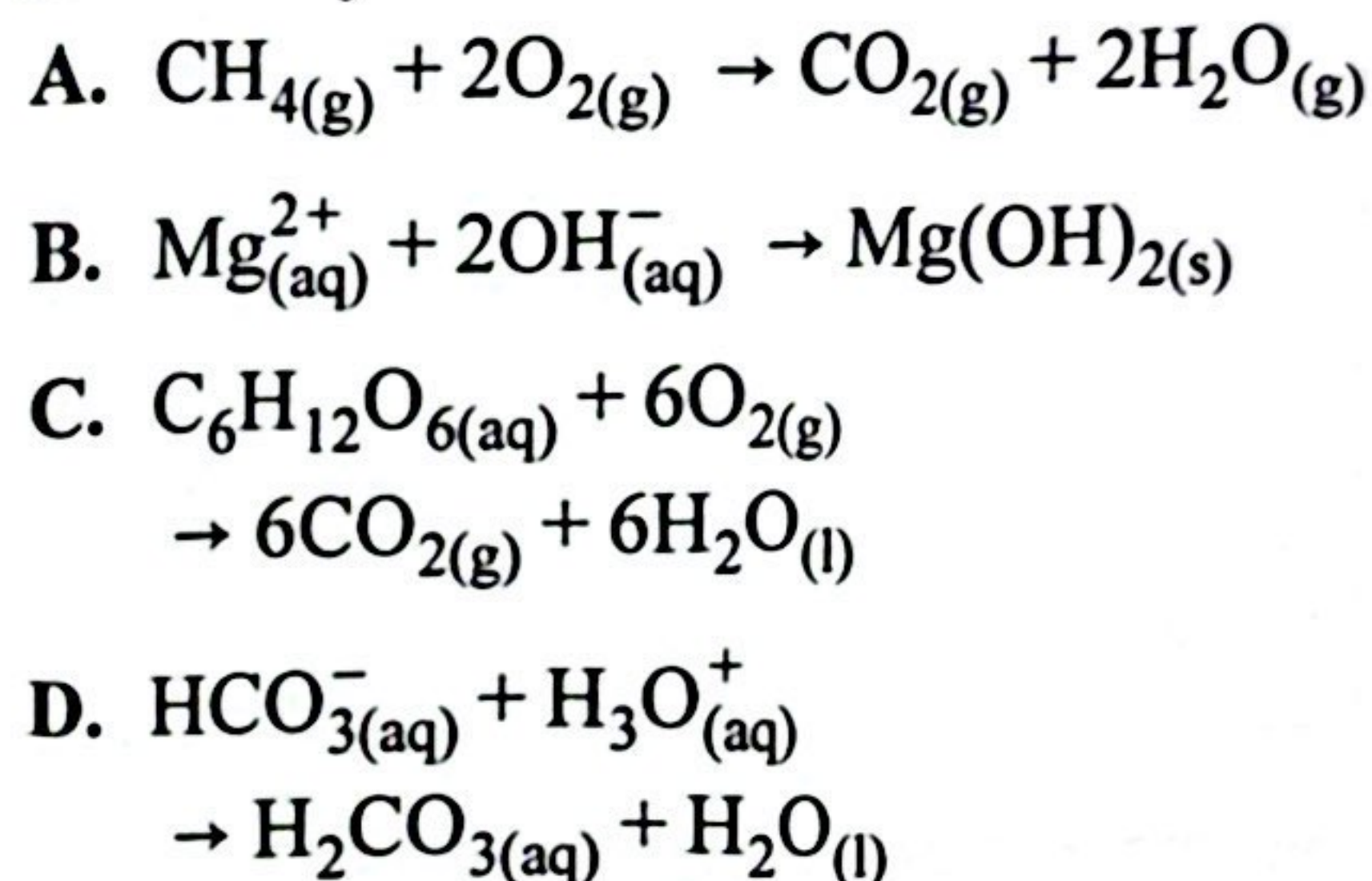
### Numerical Response

- In the reaction  
 $2\text{HNO}_{3(aq)} + 3\text{SO}_{2(g)} + 2\text{H}_2\text{O}_{(l)} \rightarrow 3\text{H}_2\text{SO}_{4(aq)} + 2\text{NO}_{(g)}$ , the oxidation number of nitrogen changes from + \_\_\_\_\_ to + \_\_\_\_\_, while the oxidation number of sulfur changes from + \_\_\_\_\_ to + \_\_\_\_\_. (Record your answer as a four-digit number.)

- Which of the following reaction equations is **not** a redox reaction?



- Oxidation-reduction reactions occur in biological systems. One net oxidation-reduction reaction that occurs in the body is



Use the following information to answer the next question.

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. silver  
 C. lead                              D. tin





Use the following information to answer the next question.

The given data concern the reactions of metals  $W$ ,  $X$ ,  $Y$ , and  $Z$  and their ions.

$X^{2+} + Z \rightarrow X + Z^{2+}$	Spontaneous reaction
$W^{+} + Z \rightarrow$	Non-spontaneous reaction
$X^{2+} + Y \rightarrow$	Non-spontaneous reaction

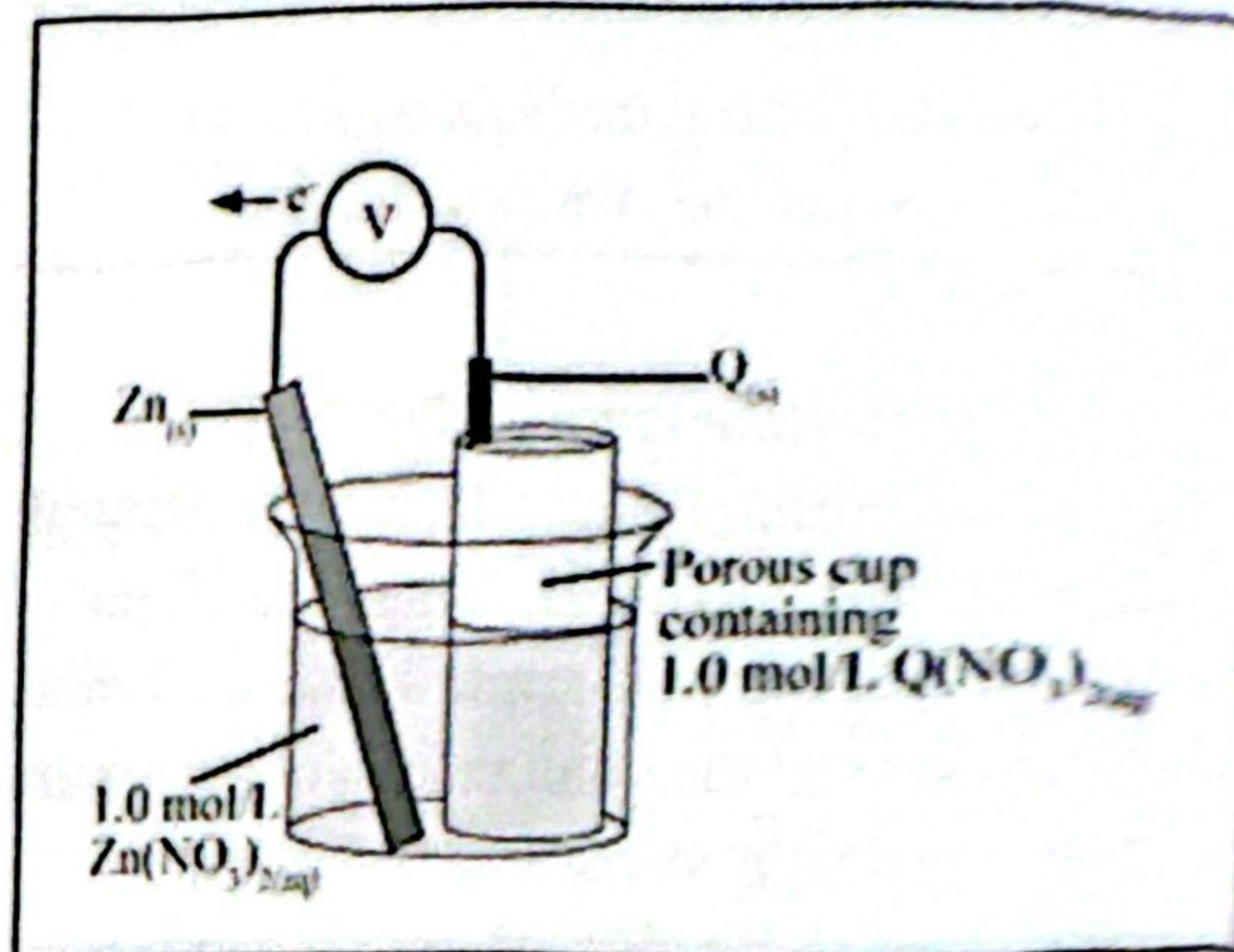
The following key lists the metals and their ions:

1.  $W^{+}$
2.  $W$
3.  $X^{2+}$
4.  $X$
5.  $Y^{3+}$
6.  $Y$
7.  $Z^{2+}$
8.  $Z$

### Numerical Response

7. When the oxidizing agents given in the key are arranged in order from strongest to weakest, the order is \_\_\_\_\_, \_\_\_\_\_, \_\_\_\_\_, and \_\_\_\_\_. (Record your answer as a four-digit number.)

Use the following information to answer the next question.



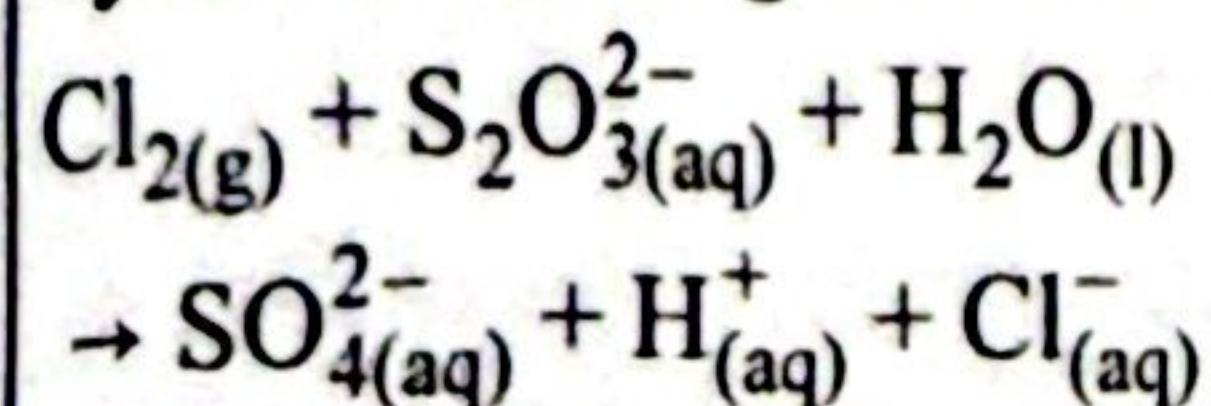
8. Given that the reading on the voltmeter for this cell is +1.74 V, which of the following statements is true?
- A.  $Zn_{(s)}$  is a weaker reducing agent than  $Q_{(s)}$ .
  - B.  $Q_{(aq)}^{2+}$  would react spontaneously with  $Cu_{(s)}$ .
  - C. The reduction potential of  $Q_{(aq)}^{2+}$  is -0.98 V.
  - D. The reduction potential of  $Q_{(aq)}^{2+}$  is +2.50 V.



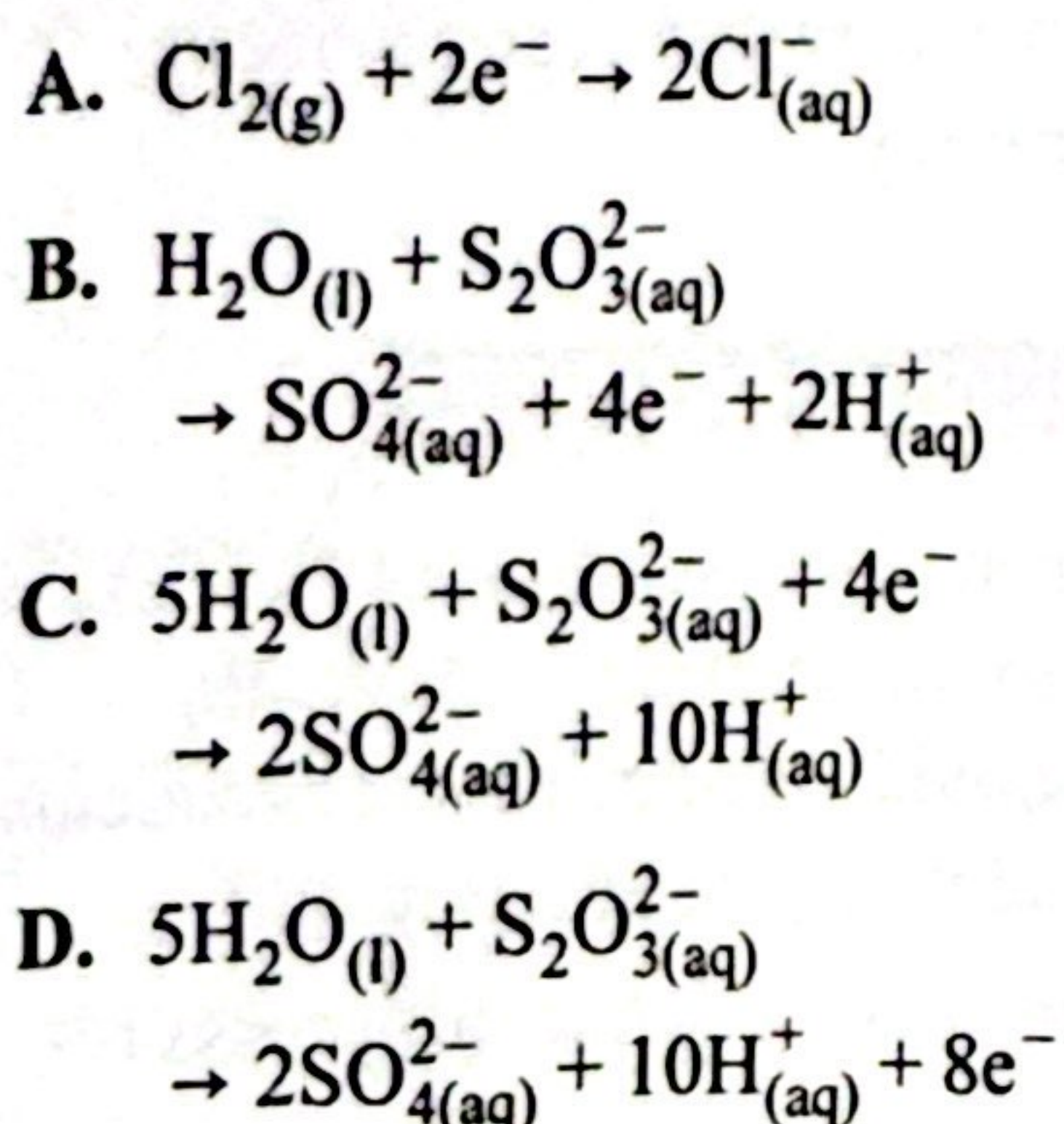


Use the following information to answer the next question.

To prevent it from contaminating the air, chlorine gas can be reacted, as represented by the following unbalanced equation:



9. The balanced oxidation half-reaction for this change is

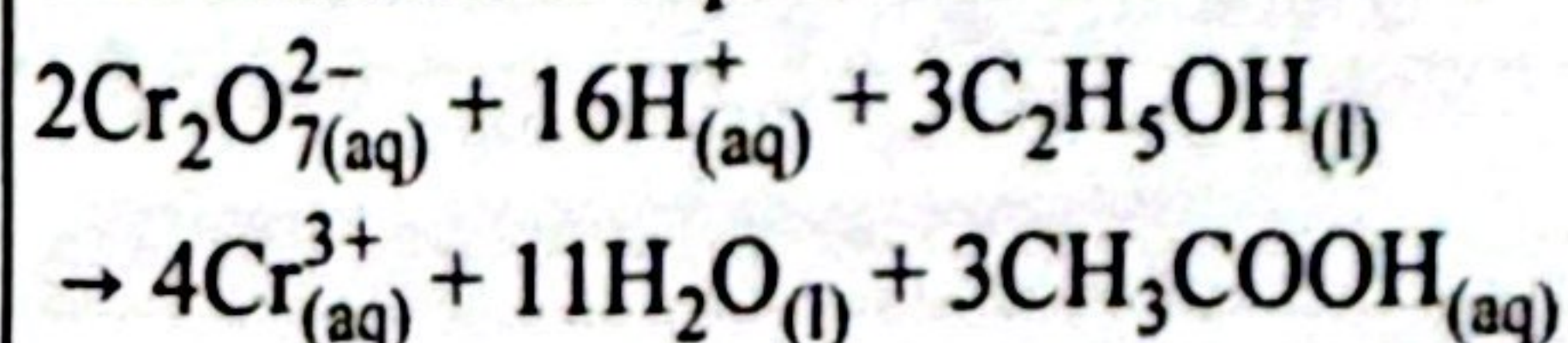


10. In an experiment, a student used 11.33 mL of  $\text{H}_2\text{O}_{2(l)}$  to titrate a 17.00 mL sample of acidified  $8.0 \times 10^{-3} \text{ mol/L}$   $\text{KMnO}_{4(aq)}$ . If  $\text{Mn}_{(aq)}^{2+}$  is one of the products, then the concentration of the  $\text{H}_2\text{O}_{2(l)}$  is

- A.  $1.2 \times 10^{-2} \text{ mol/L}$
- B.  $1.5 \times 10^{-2} \text{ mol/L}$
- C.  $3.0 \times 10^{-2} \text{ mol/L}$
- D.  $6.0 \times 10^{-2} \text{ mol/L}$

Use the following information to answer the next question.

Older breathalyzer machines involved the subject's breath passing through a solution of potassium dichromate. If alcohol was present, the colour of the solution would change from orange to green as  $\text{Cr}_{(aq)}^{3+}$  was produced. The colour change in the solution could be detected by photometry. The reaction is represented as follows:

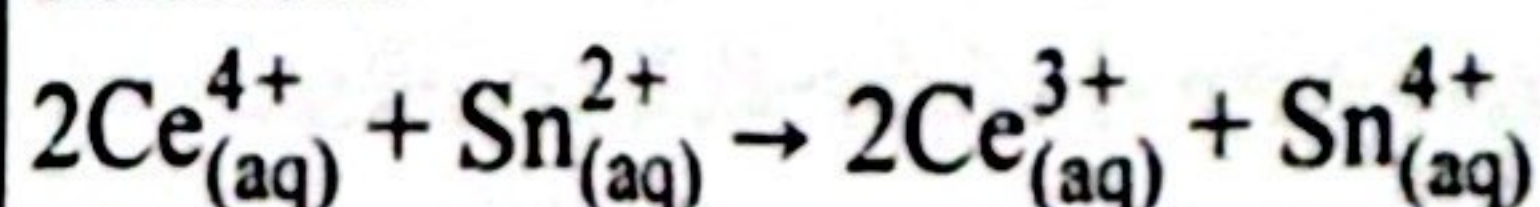


### Numerical Response

11. A 10.0 mL breath sample reacts with 2.70 mL of  $0.01062 \text{ mol/L}$   $\text{Cr}_2\text{O}_{7(aq)}^{2-}$ . The concentration of ethanol,  $\text{C}_2\text{H}_5\text{OH}_{(l)}$ , in the breath sample is \_\_\_\_\_ mmol/L. (Record your answer as a three-digit number.)

Use the following information to answer the next question.

The following equation represents a redox titration.



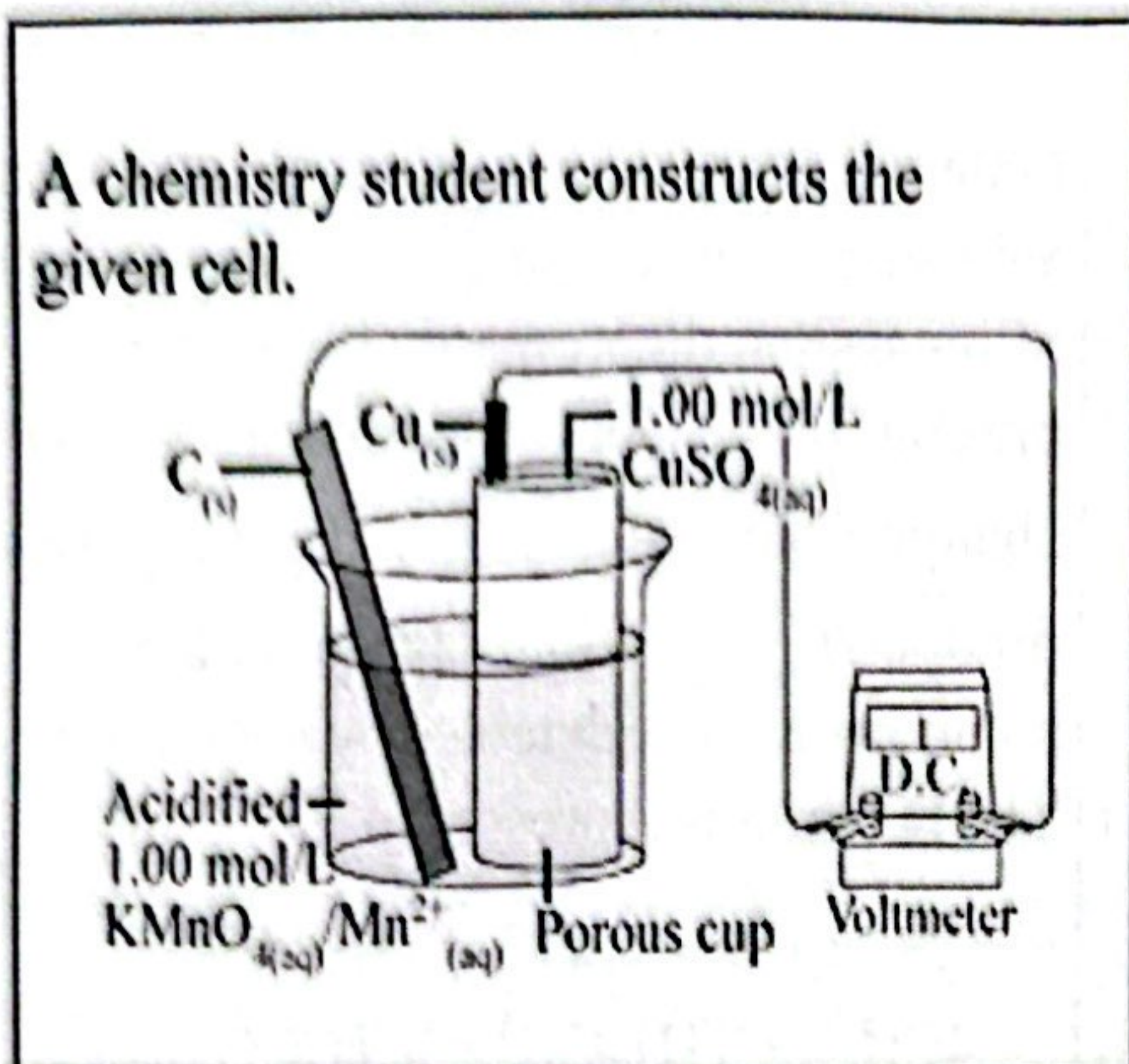
### Numerical Response

12. If 25.0 mL of  $0.0550 \text{ mol/L}$   $\text{Sn}_{(aq)}^{2+}$  is titrated with  $0.1050 \text{ mol/L}$   $\text{Ce}_{(aq)}^{4+}$ , the expected titrant volume is \_\_\_\_\_ mL. (Record your answer as a three-digit number.)





Use the following information to answer the next question.



13. During the operation of this cell,
- electrons flow from the copper electrode to the carbon electrode
  - cations migrate toward the copper electrode
  - the concentration of sulfate ions decreases
  - anions migrate toward the carbon electrode

### Written Response

14. Draw and label a diagram of a voltaic cell that produces a potential greater than 1.00 V. Show that the cell produces a potential greater than 1.00 V by writing out the half-reactions involved and calculating the  $E_{\text{net}}^\circ$ .

Your diagram should include the following labels:

- The anode, the cathode, and the composition metals
- The electrolytes of each half cell
- The directions of electron and ion flow

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15. Which of the following statements about the use of electricity in both voltaic and electrolytic cells is **true**?
- Both voltaic and electrolytic cells produce electricity.
  - Both voltaic and electrolytic cells consume electricity.
  - Voltaic cells produce electricity, while electrolytic cells consume electricity.
  - Voltaic cells consume electricity, while electrolytic cells produce electricity.

### Numerical Response

16. When an aqueous solution of  $\text{Na}_2\text{SO}_4$  undergoes electrolysis, water is the strongest oxidizing and strongest reducing agent, and therefore reacts at both electrodes. The  $E_{\text{net}}^\circ$  for this process is —            V.

Use the following information to answer the next question.

Chromium plating of objects, such as iron car bumpers, to prevent corrosion actually involves the plating of three different metals in three separate electrolytic cells. The first cell contains a solution of a copper salt, the second a solution of nickel salt, and the third a solution of chromium salt.

17. During the nickel stage of the electroplating process, the nickel(II) ions
- gain electrons, and metal is deposited on the anode
  - gain electrons, and metal is deposited on the cathode
  - lose electrons, and metal is deposited on the anode
  - lose electrons, and metal is deposited on the cathode

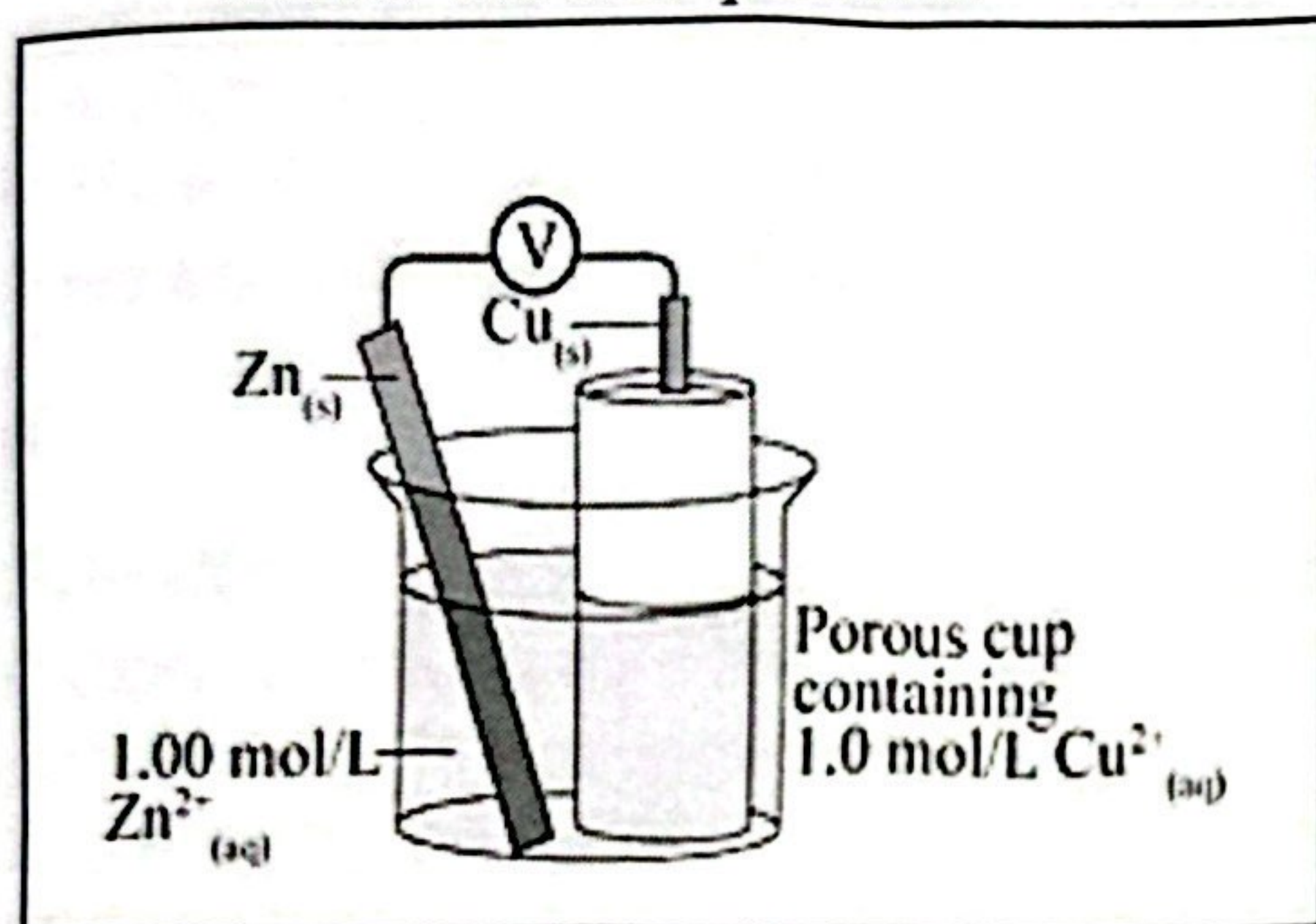




18. Electrolysis of  $\text{MgCl}_{2(aq)}$  will not produce magnesium metal because

- A.  $\text{Cl}^-_{(aq)}$  is a stronger oxidizing agent than  $\text{Mg}^{2+}_{(aq)}$
- B.  $\text{H}_2\text{O}_{(l)}$  is a stronger reducing agent than  $\text{Mg}^{2+}_{(aq)}$
- C.  $\text{H}_2\text{O}_{(l)}$  is a stronger oxidizing agent than  $\text{Mg}^{2+}_{(aq)}$
- D.  $\text{Cl}^-_{(aq)}$  is a stronger reducing agent than  $\text{Mg}^{2+}_{(aq)}$

Use the following information to answer the next question.



19. For this cell, the potential is

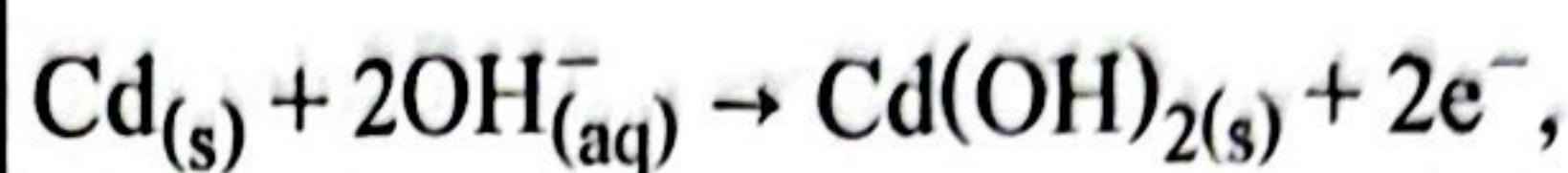
- A. +1.10 V      B. +0.42 V
- C. -0.42 V      D. -1.10 V

20. The net cell equation for the electrolysis of aqueous potassium chloride is

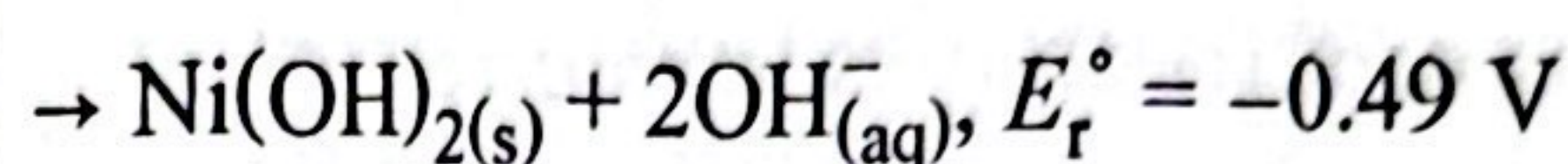
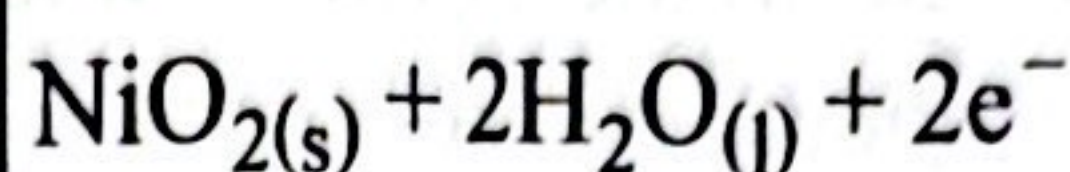
- A.  $2\text{K}^+_{(aq)} + 2\text{Cl}^-_{(aq)} \rightarrow \text{K}_{(s)} + \text{Cl}_{2(g)}$
- B.  $2\text{H}^+_{(aq)} + 2\text{Cl}^-_{(aq)} \rightarrow \text{H}_{2(g)} + \text{Cl}_{2(g)}$
- C.  $2\text{H}_2\text{O}_{(l)} + 4\text{K}^+_{(aq)} \rightarrow 4\text{K}_{(s)} + \text{O}_{2(g)} + 4\text{H}^+_{(aq)}$
- D.  $2\text{H}_2\text{O}_{(l)} + 2\text{Cl}^-_{(aq)} \rightarrow \text{Cl}_{2(g)} + 2\text{H}_{2(g)} + 2\text{OH}^-_{(aq)}$

Use the following information to answer the next question.

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



$$E_r^\circ = ?$$



### Numerical Response

21. On discharging, the electrical potential of a NiCad battery is +1.40 V. The reduction potential for the first half-reaction is \_\_\_\_\_ V.

Use the following information to answer the next question.

Restorers of antique cars often refinish chrome-plated parts by electroplating them. The part is attached to one electrode of an electrolytic cell in which the other electrode is lead. The electrolyte is a solution of dichromic acid,  $\text{H}_2\text{Cr}_2\text{O}_{7(aq)}$ .

22. A metal that will react spontaneously with  $\text{Cr}^{3+}_{(aq)}$  in a chromium-plating solution is

- A. aluminum      B. cadmium
- C. lead      D. tin

23. The reference cell for standard electrode potentials is the

- A. hydroxide half-cell
- B. hydrogen half-cell
- C. oxygen half-cell
- D. water half-cell





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, sulfuric acid flows through a copper ore deposit. Under ideal conditions, the copper metal in the ore reacts with the concentrated sulfuric acid to form copper(II) ions. The resulting copper(II) slurry is transferred to an electrolytic cell where pure copper is produced. (Assume that the sulfuric acid completely ionizes to hydrogen ions and sulfate ions.)

24. What mass of pure copper is produced from the electrolysis of excess copper(II) ions over a 24.0 h period when the cell is operated at 100 A?
- A. 11.4 kg      B. 2.84 kg  
C. 5.49 kg      D. 5.69 kg

Use the following information to answer the next question.

The sodium metal in television picture tubes reacts with oxygen that would otherwise oxidize the tungsten and phosphorus found in the tubes. Tungsten and phosphorus are vital to the function of the picture tubes.

### Numerical Response

25. The mass of sodium that will react when 0.350 mol of electrons is transferred is \_\_\_\_\_ g.





# ANSWERS AND SOLUTIONS — UNIT TEST

1. C	6. A	11. 4.30	16. 2.06	21. 1.89
2. D	7. 5371	12. 26.2	17. B	22. A
3. 5246	8. A	13. A	18. C	23. B
4. D	9. D	14. WR	19. A	24. B
5. C	10. C	15. C	20. D	25. 8.05

1. C

- An atom is reduced if its oxidation number decreases over the course of a reduction-oxidation reaction.
- The oxidation number of gold changes from 0 (in  $\text{Au}_{(s)}$ ) to +3 (in  $\text{HAuCl}_{4(aq)}$ ). The gold atoms are oxidized.
- The oxidation number of nitrogen changes from +5 (in  $\text{HNO}_{3(aq)}$ ) to +4 (in  $\text{NO}_{2(g)}$ ). Clearly, nitrogen atoms are reduced in the reaction.
- The oxidation numbers of hydrogen and chlorine remain unchanged throughout the reaction.

2. D

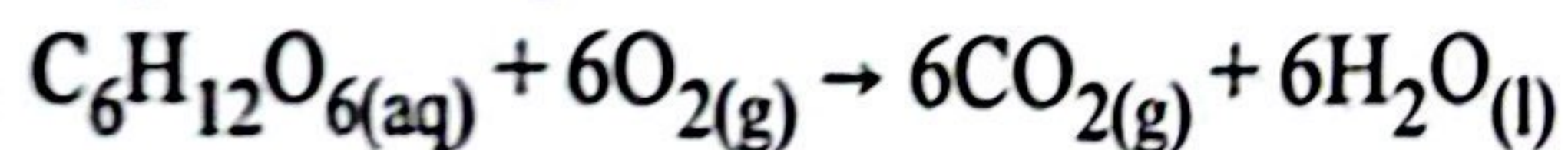
To be effective, a sacrificial anode must be a suitably stronger reducing agent if it is to be oxidized

4. D

This is the only equation in which no elements have undergone a change in oxidation number. This is a double replacement reaction.

5. C

Equation C is correct—it illustrates the cellular respiration of glucose.



Equations B and D are not redox reactions since there is no change in oxidation numbers. Equation A represents the combustion of natural gas.

6. A

For a metal to be a sacrificial anode for iron, it must be a stronger reducing agent than iron. Only  $\text{Mg}_{(s)}$