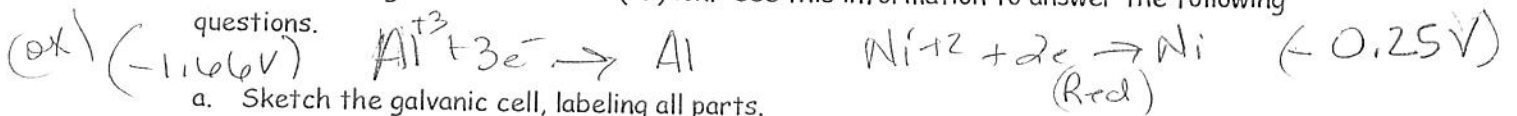
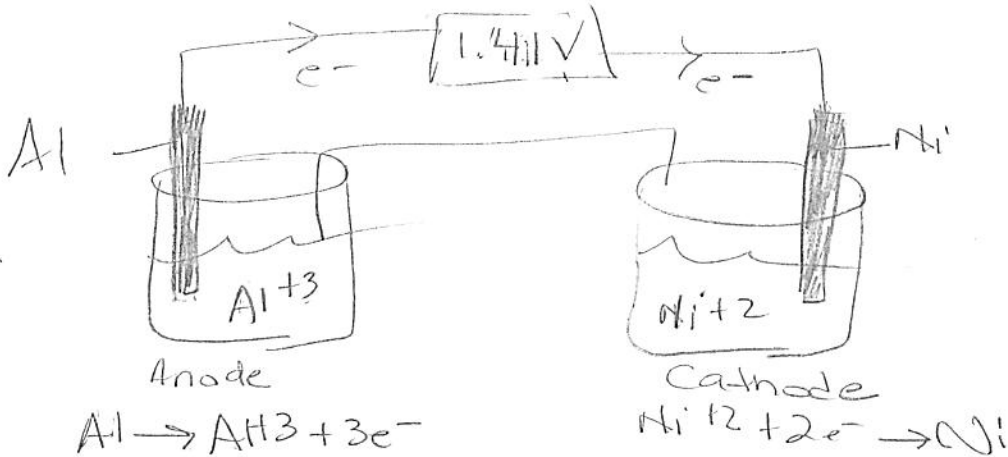




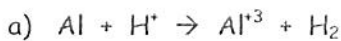
3. A galvanic cell is composed of two half cells, one containing aluminum and aluminum ion and the other containing nickel and nickel (II) ion. Use this information to answer the following questions.



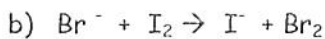
- Sketch the galvanic cell, labeling all parts.
- Identify the metal at the anode and cathode.
- Write the reaction that occurs at the anode and the cathode.
- Show the direction of the electron flow in the cell.
- Calculate the standard cell potential for the galvanic cell.



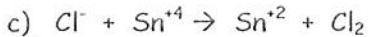
4. For the following calculate the net voltages and determine if the reaction will be spontaneous or not as written. Please note these equations are NOT balanced.



$E_{cell}^{\circ} = 1.66V$  - spontaneous

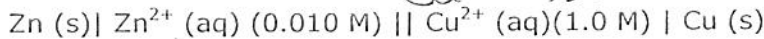
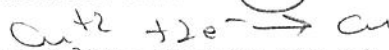


$E_{cell}^{\circ} = -0.531$  Nonspontaneous



$E_{cell}^{\circ} = -1.21$  Nonspontaneous

5. For the voltaic cell shown below in line notation calculate  $\Delta G$  in Joules.



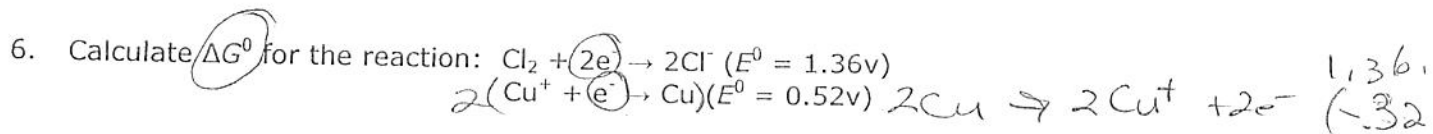
$E_{cell}^{\circ} = 1.100$

$E_{cell} = E_{cell}^{\circ} - \frac{0.0592}{n} \log \frac{[Cu^{2+}]}{[Zn^{2+}]}$

$E_{cell} = 1.100 - \frac{0.0592}{2} \log \frac{[1.0]}{[0.01]}$

$E_{cell} = 1.100 + 0.0592 = 1.1592$

$\Delta G = -nFE$   
 $= -(2)(96,500)(1.1592)$   
 $\Delta G = -223,725 J$   
 $\Delta G = -224 KJ$



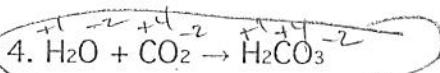
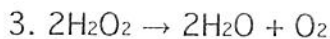
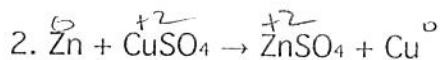
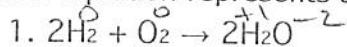
$$\Delta G^\circ = -nFE^\circ$$

$$E^\circ_{\text{cell}} = 0.84$$

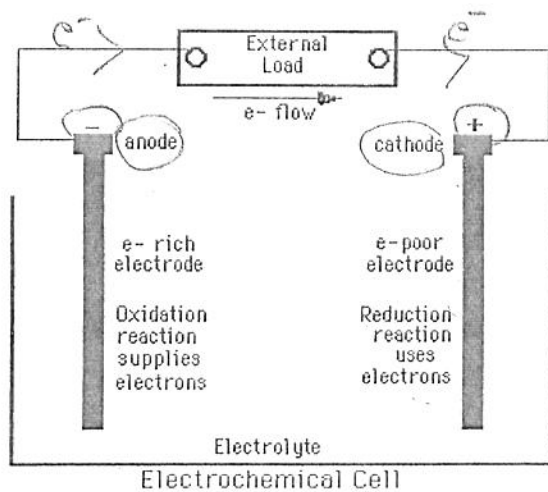
$$\Delta G^\circ = -(2 \text{ mol } e^-)(96,500 \text{ J}) (0.84 \text{ V})$$

$$\Delta G = -162,120 \text{ J} = -160 \text{ KJ} = \Delta G^\circ$$

7. Which equation represents a reaction that is not Redox?



8.



Note: E.P.A.  
 \* Electrolytic cell  
 • Anode + (red)  
 • Cathode - (black)

A galvanic cell

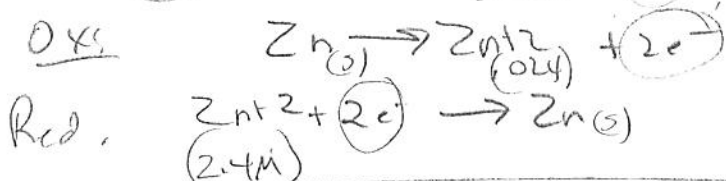
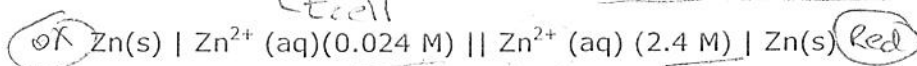
This electrochemical cell can best be described as:

- a) A complete electrolytic cell
- b) A galvanic cell missing at least one essential component
- c) A complete galvanic cell
- d) An electrolytic cell missing at least one essential component

• 2 containers  
 • salt bridge

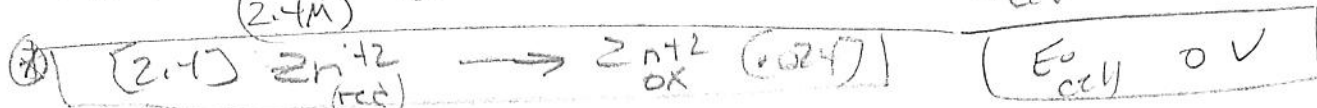
Problems using the Nerst Equation:

1. Calculate the EMF of the following cell NOT at standard conditions:



$E_{\text{cell}}^{\circ} = 0.763 \text{ V}$

$E_{\text{cell}}^{\circ} = -0.763 \text{ V}$

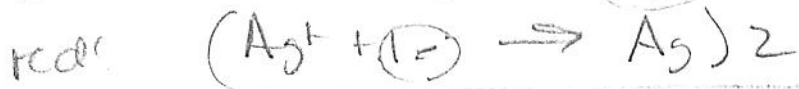


$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0592}{n} \log \left[ \frac{[\text{Zn}^{2+}]_{\text{OK}}}{[\text{Zn}^{2+}]_{\text{red}}} \right]$

$E_{\text{cell}} = 0 - \left( \frac{0.0592}{2} \right) \log \left( \frac{0.024}{2.4} \right)$

$E_{\text{cell}} = 0 - 0.0296 \log 0.01 = \boxed{0.0592 \text{ V}}$

2. Calculate the potential at 25°C for the following voltaic cell.



$E_{\text{cell}}^{\circ}$   
 -0.337

+0.799

$\boxed{0.462 \text{ V}}$

$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0592}{n} \log \frac{[\text{Cu}^{2+}]}{[\text{Ag}^{+}]^2}$

$= 0.462 - \frac{0.0592}{2} \log \frac{(0.024)}{(0.0048)^2}$

$= 0.462 - 0.0296 (3.018) = \boxed{0.373 \text{ V}}$

$$\frac{nFE^\circ}{RT} = \ln K$$



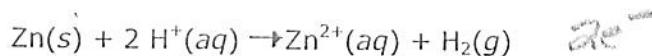
$$E_{\text{cell}}$$

$$0,76$$

$$0$$

$$0,76\text{V}$$

3. Calculate the equilibrium constant at  $25^\circ\text{C}$  for the reaction between zinc metal and acid.



$$\ln K = \frac{(2)(96,500\text{J/V}\cdot\text{mole})(0,76\text{V})(\text{mol}\cdot\text{K})}{(8,31\text{J})298\text{K}}$$

$$\ln K = 59,2316$$

$$K = 5,3 \times 10^{25}$$

4. What value of  $E$  do you expect for the reaction  $\text{Pb(s)} \rightarrow \text{Pb}^{2+} + 2e^-$  in  $0,015\text{M}$   $\text{Pb}^{2+}$  solution?  $E^\circ = +0,13\text{V}$



$$E_{\text{cell}} = 0,13 - \frac{0,0592}{2} \log \frac{1}{0,015}$$

$$E_{\text{cell}} = 0,13 + 0,054 = 0,18\text{V} = E_{\text{cell}}$$

Estimating voltage for a single cell

Name \_\_\_\_\_

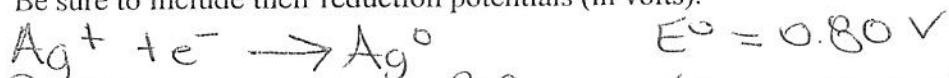
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Period \_\_\_ Date \_\_\_/\_\_\_/\_\_\_

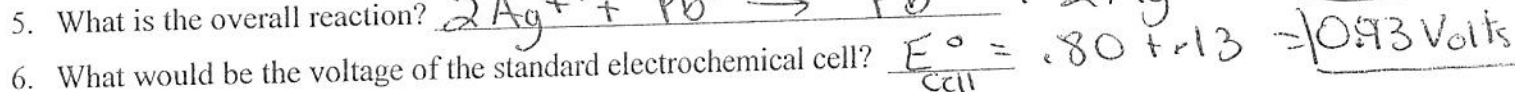
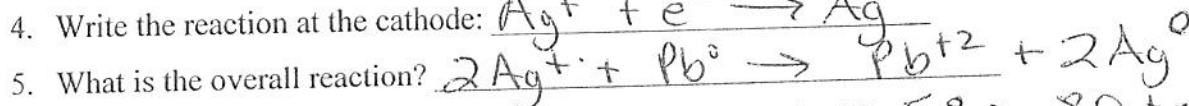
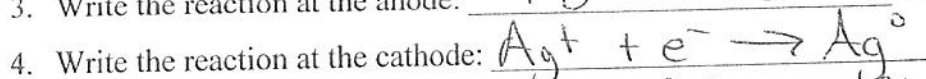
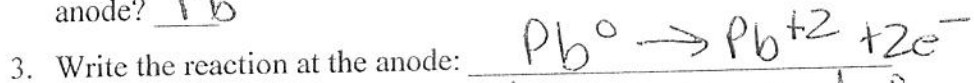
## 21 • Electron Transfer Reactions

### ELECTROCHEMICAL CELL WORKSHEET

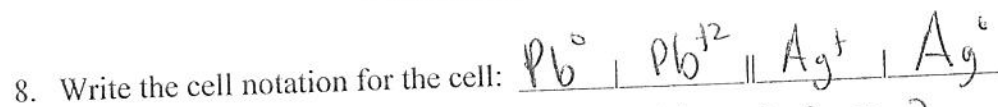
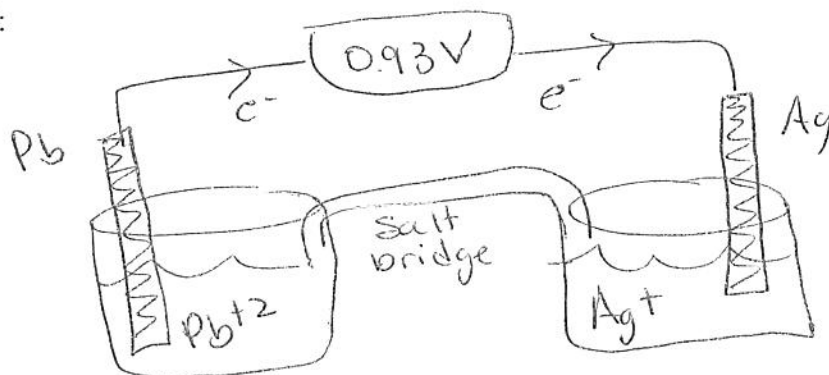
Consider the reduction potential chart. Find and copy the reduction equations for  $\text{Ag}^+ \rightarrow \text{Ag}^0$  and  $\text{Pb}^{2+} \rightarrow \text{Pb}^0$ . Be sure to include their reduction potentials (in volts).



- Which metal ion has the greater reduction potential? Ag
- If these two metals (and their solutions) were used to create a galvanic cell, which metal would be the anode? Pb



7. Sketch the cell:



9. How many moles of electrons are involved in this reaction?  $n = 2$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0592}{n} \log Q$$

10. Find and copy down the Nernst Equation: \_\_\_\_\_

11. If a new cell is set up with the  $[\text{Ag}^+] = 0.50 \text{ M}$  and the  $[\text{Pb}^{2+}] = 2.0 \text{ M}$ , the cell voltage will be \_\_\_\_\_

(greater less).

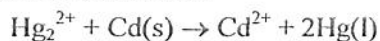
12. Use the Nernst equation to calculate the cell voltage with these new concentrations.

$$E_{\text{cell}} = 0.93 - \left( \frac{0.0592}{2} \right) \log \frac{(2.0)}{(0.5)^2}$$

$$E_{\text{cell}} = 0.93 - 0.2368$$

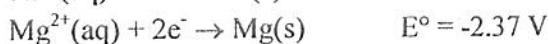
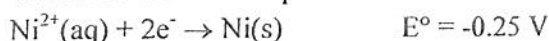
$$E_{\text{cell}} = 0.69 \text{ V}$$

1. Which of the following is the correct cell notation for the reaction



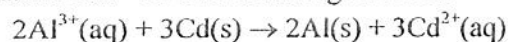
- a)  $\text{Cd}^{2+} | \text{Cd} || \text{Hg}_2^{2+} | \text{Hg}$   
b)  $\text{Cd}^{2+} | \text{Hg}_2^{2+} || \text{Cd} | \text{Hg}$   
c)  $\text{Cd} | \text{Cd}^{2+} || \text{Hg}_2^{2+} | \text{Hg}$   
d)  $\text{Cd}^{2+} | \text{Hg} || \text{Hg}_2^{2+} | \text{Cd}$   
e)  $\text{Hg} | \text{Cd} || \text{Hg}_2^{2+} | \text{Cd}^{2+}$

2. Given the two half reactions and their potentials, which net reaction is spontaneous?



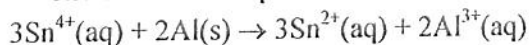
- a)  $\text{Ni(s)} + \text{Mg}^{2+}(\text{aq}) \rightarrow \text{Mg(s)} + \text{Ni}^{2+}(\text{aq})$   
b)  $\text{Ni}^{2+}(\text{aq}) + \text{Mg(s)} \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{Ni(s)}$   
c)  $\text{Ni(s)} + \text{Mg(s)} \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{Ni}^{2+}(\text{aq})$   
d)  $\text{Mg}^{2+}(\text{aq}) + \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Mg(s)} + \text{Ni(s)}$   
e)  $\text{Mg}^{2+}(\text{aq}) + \text{Mg(s)} \rightarrow \text{Ni(s)} + \text{Ni}^{2+}(\text{aq})$

3. Calculate  $E^\circ$  for the following reaction:



- a) -2.06 V                      d) -4.52 V  
b) +4.52 V                      e) -1.26 V  
c) +2.06 V

4. The standard cell potential for



is  $E^\circ = 1.81 \text{ V}$ . What is  $E_{\text{cell}}$  when

$$[\text{Sn}^{4+}] = 1.0,$$

$$[\text{Sn}^{2+}] = 1.0 \times 10^{-2}, \text{ and}$$

$$[\text{Al}^{3+}] = 1.5 \times 10^{-3} \text{ at } 298 \text{ K.}$$

- a) 1.70 V                      d) 1.86 V  
b) 1.76 V                      e) 1.93 V  
c) 1.81 V

5. How long would it take to deposit 1.36 g of copper from an aqueous solution of copper(II) sulfate by passing a current of two amperes through the solution?

- a) 2070 sec                      d) 736 sec  
b)  $1.11 \times 10^{-5}$  sec                      e) 1030 sec  
c) 2570 sec

6. If a current of 6.0 amps is passed through a solution of  $\text{Ag}^+$  for 1.5 hours, how many grams of silver are produced?

- a) 0.60 g                      d) 3.0 g  
b) 36 g                      e) 1.0 g  
c) 0.34 g

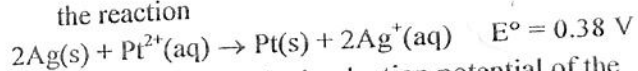
7. Under acidic conditions the bromate ion is reduced to the bromide ion. Write the balanced half-reaction for this process.

- a)  $\text{BrO}_3^- + 6\text{H}^+ + 6\text{e}^- \rightarrow \text{Br}^- + 3\text{H}_2\text{O}$   
b)  $2\text{BrO}_3^- + 6\text{H}^+ \rightarrow \text{Br}_2^- + 6\text{H}_2\text{O} + 3\text{e}^-$   
c)  $\text{BrO}_3^- + 6\text{H}_2\text{O} + 10\text{e}^- \rightarrow \text{Br}_2^- + 12\text{H}^+ + 3\text{O}_2$   
d)  $2\text{BrO}_3^- + 6\text{H}_2\text{O} \rightarrow 2\text{Br}^- + 12\text{H}^+ + 6\text{O}_2 + 8\text{e}^-$   
e)  $2\text{BrO}_3^- + 6\text{H}^+ \rightarrow \text{Br}_2^- + 3\text{H}_2\text{O} + 3\text{e}^-$

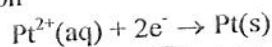
8. Which of the following reactions is NOT a redox reaction?

- a)  $2\text{HgO(s)} \rightarrow 2\text{Hg(l)} + \text{O}_2(\text{g})$   
b)  $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightarrow 2\text{HBr(g)}$   
c)  $2\text{HCl(aq)} + \text{Zn(s)} \rightarrow \text{H}_2(\text{g}) + \text{ZnCl}_2(\text{aq})$   
d)  $\text{H}_2\text{CO}_3(\text{aq}) \rightarrow \text{H}_2\text{O(l)} + \text{CO}_2(\text{g})$   
e)  $2\text{KClO}_3 \rightarrow 2\text{KCl(s)} + 3\text{O}_2(\text{g})$

9. Using data from the reduction potential table and the reaction



calculate the standard reduction potential of the half-reaction



- a) -1.18 V                      d) 1.18 V  
b) -0.40 V                      e) 2.00 V  
c) 0.40 V

10. Predict the product at the anode when electric current is passed through a solution of KI.

- a)  $\text{I}_2(\text{l})$                       d)  $\text{K(s)}$   
b)  $\text{K}^+(\text{aq})$                       e)  $\text{O}_2(\text{g})$   
c)  $\text{H}_2(\text{g})$

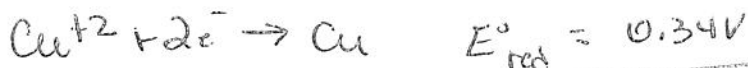
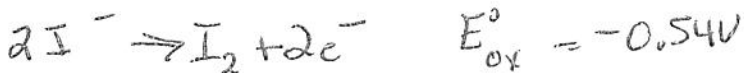
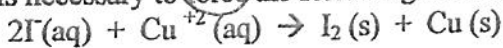
Name: \_\_\_\_\_  
Date: \_\_\_\_\_

Key

$$\# \text{ moles} = \frac{It}{nF}$$

## Electrolysis Problems

1. What voltage is necessary to force the following electrolysis reaction to occur?



$$E_{\text{cell}}^{\circ} = -0.20\text{V}$$

⊖ voltage means NONSPONTANEOUS

Must add 0.20V to force this rxn to occur.

2. How many grams of copper can be reduced by applying a 3.00 A current for 16.2 minutes to a solution containing  $\text{Cu}^{+2}$  ions?

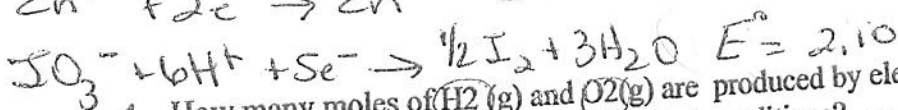
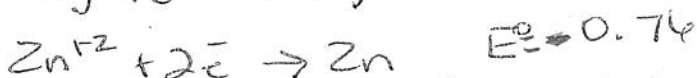
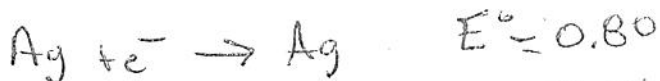
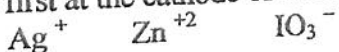
3.00 coulombs/sec

$$\left(\frac{16.2 \text{ min}}{1}\right) \left(\frac{60 \text{ sec}}{1 \text{ min}}\right) = 972 \text{ sec.}$$

$$\# \text{ mol} = \frac{It}{nF} = \left(\frac{3.00 \text{ C}}{\text{sec}}\right) \left(\frac{972 \text{ sec}}{1 \text{ mol Cu}}\right) \left(\frac{1 \text{ mole e}^{-}}{2 \text{ mol e}^{-}}\right) \left(\frac{1}{96,500 \text{ C}}\right) = 0.0151 \text{ mol Cu}$$

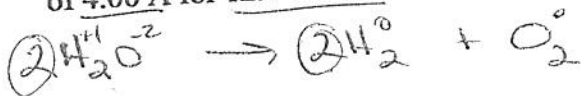
$$\left. \begin{array}{l} 0.0151 \text{ mol} \times 63.5 \\ \hline \text{mol} \\ \hline = 0.959 \text{ g Cu} \end{array} \right\}$$

3. Use your table of reduction potentials to determine which of the following ions would reduce first at the cathode of an electrolytic cell:



4. How many moles of  $\text{H}_2(\text{g})$  and  $\text{O}_2(\text{g})$  are produced by electrolyzing water at a current of 4.00 A for 12.0 minutes assuming ideal conditions?

$$\left(\frac{12 \text{ min}}{1}\right) \left(\frac{60 \text{ sec}}{1 \text{ min}}\right) = 720 \text{ sec}$$



$$\text{moles} = \frac{It}{nF} = \left(\frac{4.00 \text{ C}}{\text{sec}}\right) \left(\frac{720 \text{ sec}}{1 \text{ mol H}_2\text{O}}\right) \left(\frac{1 \text{ mol H}_2}{2 \text{ mol e}^{-}}\right) \left(\frac{1}{96,500 \text{ C}}\right) = 0.0149 \text{ mol H}_2$$

# moles  $\text{O}_2$  is  $\frac{1}{2}$  the # moles of  $\text{H}_2\text{O}$  and  $\text{H}_2$  as shown by the coefficients in the balanced equation so

$$\left(\frac{0.0149 \text{ moles H}_2}{1}\right) \left(\frac{1 \text{ mol O}_2}{2 \text{ mol H}_2}\right) = 0.00746 \text{ moles O}_2$$