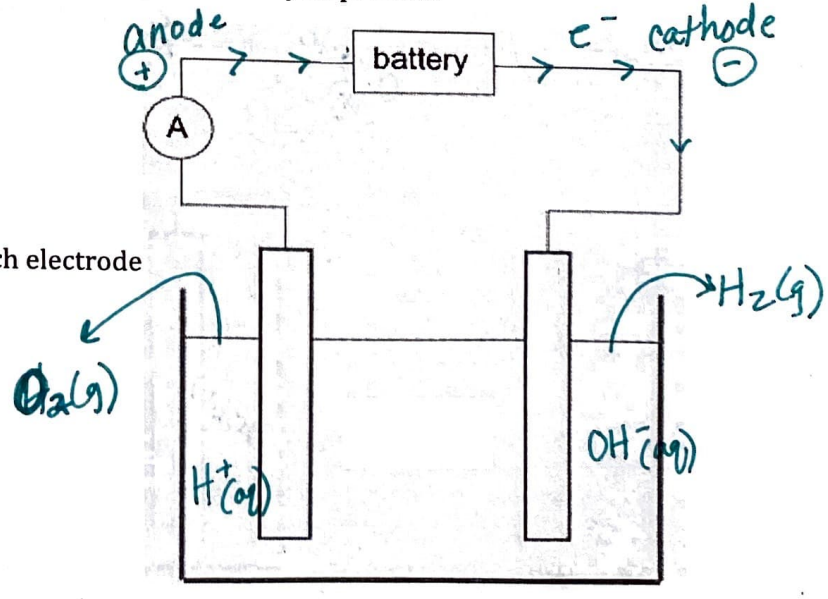
Practice Problems: Electrolytic Cells **Name** Answer Key

Write answers on a separate sheet of paper.

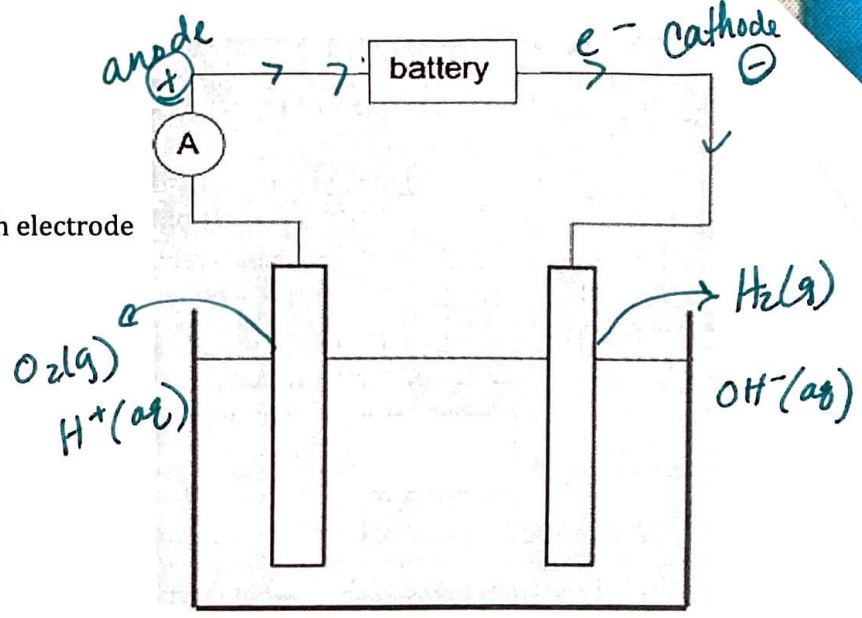
1. Consider the following unbalanced redox reactions:
   a. $I_2(s) + Cl^-(aq) \rightarrow I^-(aq) + Cl_2(g)$
   b. $MnO_4^-(aq) + Br^-(aq) \rightarrow Br_2(l) + Mn^{2+}(aq)$ (acidic conditions)

   For each un-balanced redox reaction above:
   i. indicate if the reaction is spontaneous or not
   ii. write out the half-reaction at the anode
   iii. write out the half-reaction at the cathode
   iv. write out the balanced redox reaction
   v. calculate the cell potential
   vi. indicate if it is a redox reaction that can take place in a voltaic cell or electrolytic cell

2. Oxidation takes place at what electrode in a voltaic cell?
3. Oxidation takes place at what electrode in an electrolytic cell?
4. An electrolytic cell and a voltaic cell can look very similar except for the presence or absence of what?
5. In a voltaic cell, electrons flow out of which electrode?
6. In an electrolytic cell, electrons flow out of which electrode?
7. Consider an electrolytic nickel-cadmium cell.
   i. Identify the anode and cathode.
   ii. Write out the oxidation half-reaction, the reduction half-reaction, and the net redox reaction.
   iii. Calculate the cell potential.
8. Consider the electrolysis of KI(aq).
   a. Identify and write out the half-reactions that occur at the anode and cathode.
   b. Write out the redox reaction that occurs in the electrolysis process.
   c. Label the electrolytic cell:
      *Direction of electron flow
      *Label the anode and cathode
      *Charge of anode and cathode
      *Draw ions and products at each electrode



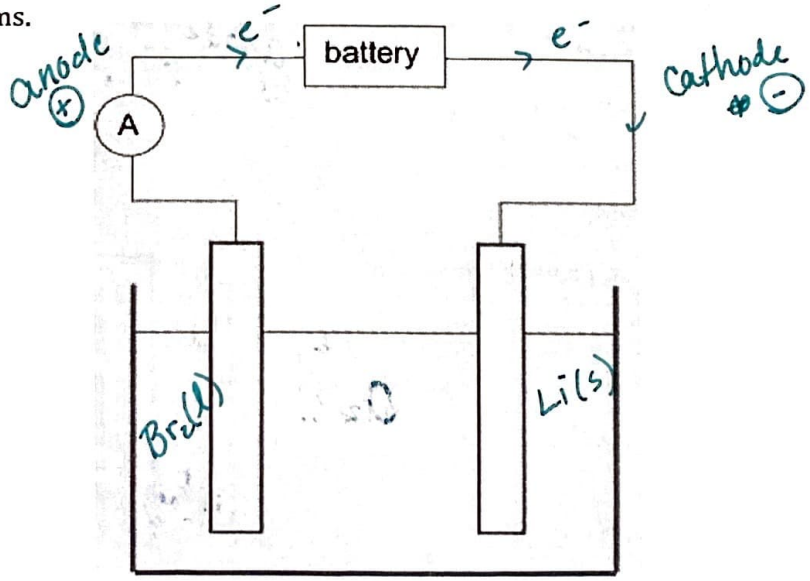
9. Consider the electrolysis of NaCl(aq).
   a. Identify and write out the half-reactions that occur at the anode and cathode.
   b. Write the redox reaction that occurs in the electrolytic cell.
   c. Label the electrolytic cell:
      *Direction of electron flow
      *Label the anode and cathode
      *Charge of anode and cathode
      *Draw ions and products at each electrode



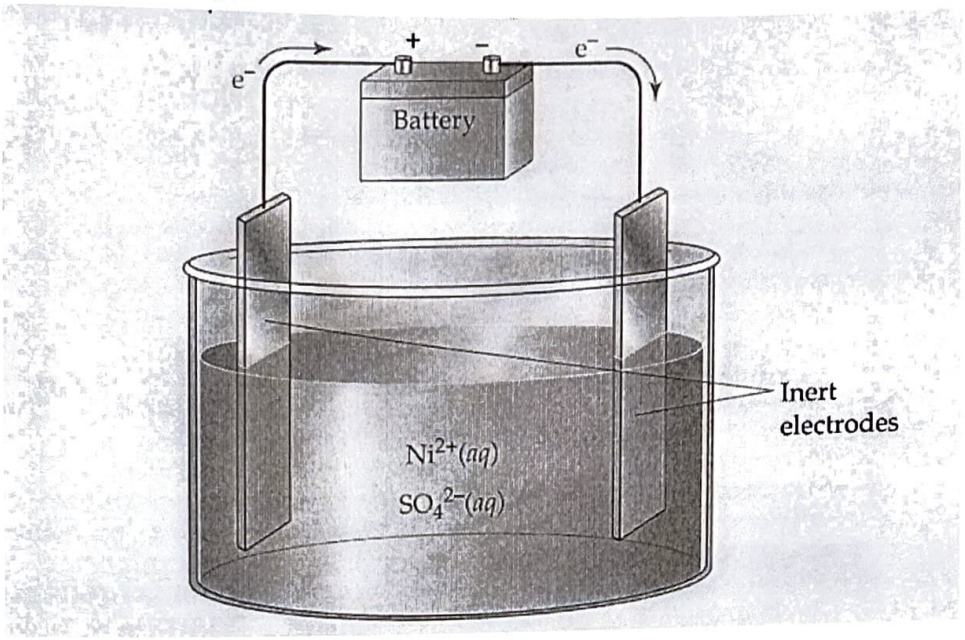
10. Predict whether the following reaction would take place in a voltaic cell or an electrolytic cell.
   a. $2\ Ag^+(aq) + H_2SO_3(aq) + H_2O(l) \rightarrow 2\ Ag(s) + SO_4^{2-}(aq) + 4H^+(aq)$
   b. $FeI_2(aq) \rightarrow Fe(s) + I_2(s)$

11. In the electrolysis of water, identify the gas produced at the anode and identify the gas produced at the cathode.

12. Consider the electrolysis of molten lithium bromide, LiBr(l).
   i. Write the half-reactions that take place at the anode and the cathode.
   ii. What voltage is required for this electrolysis purpose?
   iii. Will the mass of the cathode increase or decrease?
   iv. Correctly label the diagram below with the direction of electron flow and the location $Br_2(l)$ forms.



13. Consider the picture of the following electrochemical cell.



a. Is the cell electrolytic or voltaic?

b. What side is the anode located (left or right)?

c. Write the half-reaction that takes place at the anode and cathode.

d. Calculate the cell potential.

a) $\overset{0}{I_2(s)} + \overset{-1}{Cl^-(aq)} \rightarrow \overset{-1}{I^-(aq)} + \overset{0}{Cl_2(g)}$

i. nonspontaneous
→ I_2 has lower E°_{cell} so will be preferentially oxidized

ii anode = oxidation

$2Cl^-_{(aq)} \rightleftharpoons Cl_2(g) + 2e^- \quad E^\circ = +1.36\ V$

iii cathode = reduction

$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq) \quad E^\circ = +0.54 V$

iv. $2Cl^-(aq) + I_2(s) \rightleftharpoons Cl_2(g) + 2I^-(aq)$

v. $E^\circ_{cell} = +0.54 - (+1.36)$
$= -0.82V$

vi. electrolytic

b) $\overset{+7}{Mn}\overset{-2}{O_4}^-(aq) + \overset{-1}{Br^-}(aq) \rightarrow \overset{0}{Br_2}(l) + \overset{+2}{Mn^{2+}}(aq)$

i spontaneous
→ MnO_4^- has a higher E°_{cell} so will be preferentially reduced

ii anode = oxidation

$2Br^-(aq) \rightarrow {}^1\!/_2 Br_2(l) + 2e^-$

iii cathode = reduction

$8H^+_{(aq)} + 5e^- + MnO_4^-(aq) \rightarrow Mn^{2+} + 4H_2O(l)$

iv. $5Br^-_{(aq)} + 8H^+_{(aq)} + MnO_4^-(aq) \rightarrow 5/2\ Br_2(l) + Mn^{2+}_{(aq)} + 4H_2O(l)$

$10\ Br^-(aq) + 16H^+(aq) + 2MnO_4^-(aq) \rightarrow 5Br_2(l) + 2Mn^{2+}(aq) + 8H_2O_{(l)}$

v. $E^\circ_{cell} = +1.51 - (1.09) = \boxed{+0.42V}$

vi = voltaic

anode!

3. anode! always!
4. salt bridge – in voltaic, not in electrolytic
5. anode
6. cathode
7. Nickel / Cadmium electrolytic

$Ni(s) \rightleftharpoons Ni^{2+}(aq) + 2e^- \quad E^\circ = -0.26V$ oxidation @ anode

~~$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$~~ $E^\circ = -0.403V$ reduction @ cathode

~~$Cd(s) \rightleftharpoons$~~

i: anode: Ni
cathode: Cd

ii oxidation: $Ni(s) \rightleftharpoons Ni^{2+}(aq) + 2e^-$
reduction: $Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$

net redox reaction: $Ni(s) + Cd^{2+}(aq) \rightarrow Ni^{2+}(aq) + Cd(s)$

iii $E^\circ_{cell} = -0.403 - (-0.26) = \boxed{-0.143V}$

8. $KCl(aq) \rightarrow K^+(aq) + Cl^-(aq)$

↓ to cathode reduction ↓ to anode oxidation

a) possible rxns @ cathode:

$K^+(aq) + e^- \rightleftharpoons K(s) \quad E^\circ = -2.93V$

$H_2O(l) + e^- \rightleftharpoons \frac{1}{2} H_2(g) + OH^-(aq) \quad E^\circ = -0.83V$ ← winner, winner chicken dinner!

Possible rxns @ anode:

$2Cl^-(aq) \rightleftharpoons Cl_2(g) + 2e^- \quad E^\circ = +1.36V$

$H_2O(l) \rightleftharpoons \frac{1}{2} O_2(g) + 2H^+(aq) + 2e^- \quad E^\circ = +1.23V$ ←

b) $2H_2O(l) + H_2O(l) \rightarrow H_2(g) + \frac{1}{2}O_2(g) + 2OH^-(aq) + 2H^+(aq) \Rightarrow \boxed{2H_2O(l) \rightarrow H_2(g) + \frac{1}{2}O_2(g)}$

$3H_2O(l) \rightarrow H_2(g) + \frac{1}{2}O_2(g) + 2H_2O(l) +$ ~~[illegible]~~

1. $NaCl(aq) \rightarrow Na^+(aq) + Cl^-(aq)$

Na^+ → to cathode, reduction; Cl^- → to anode, oxidation

a) Possible rxns @ cathode:

$Na^+(aq) + e^- \rightleftharpoons Na(s)$ $E^\circ = -2.71\,V$

$H_2O(l) + e^- \rightleftharpoons \frac{1}{2} H_2(g) + OH^-(aq)$ $\boxed{E^\circ = -0.83\,V}$ winner!

Possible rxns @ anode:

$Cl^-(aq) \rightleftharpoons \frac{1}{2} Cl_2(g) + e^-$ $E^\circ = +1.36$

$H_2O(l) \rightleftharpoons \frac{1}{2} O_2(g) + 2H^+(aq) + 2e^-$ $\boxed{E^\circ = 1.23\,V}$ winner!

b) $H_2O(l) \rightarrow H_2(g) + \frac{1}{2} O_2(g)$

10. $\overset{+1}{2Ag^+}(aq) + \overset{+1\ +4\ -2}{H_2SO_3}(aq) + H_2O(l) \rightarrow \overset{0}{2Ag}(s) + \overset{+6\ -2}{SO_4^{2-}}(aq) + 4H^+(aq)$

a)

$2Ag^+ + 2e^- \rightleftharpoons 2Ag(s)$ reduction $E^\circ = +0.80\,V$ prefers to be reduced

$H_2O(l) + SO_{3\,(aq)}^{2-} \rightleftharpoons SO_{4\,(aq)}^{2-} + 2e^- + 2H^+(aq)$ oxidation $E^\circ = +0.17\,V$ prefers to be oxidized

yes!

b) $\overset{+2\ -1}{FeI_2}(aq) \rightarrow \overset{0}{Fe}(s) + \overset{0}{I_2}(s)$

$Fe^{2+} + 2e^- \rightarrow Fe(s)$ reduction $E^\circ = -0.45\,V$ prefers to be oxidized

$2I^-(aq) \rightarrow I_2(s) + 2e^-$ oxidation $E^\circ = +0.54\,V$ prefers to be reduced

no!!!

11. anode = oxidation = $O_2(g)$

cathode = reduction = $H_2(g)$

$LiBr(l) \rightarrow Li^+(l) + Br^-(l)$

↓ to cathode reduction ↓ to anode oxidation

i. $Li^+_{(aq)} + e^- \rightleftharpoons Li(s)$ cathode $E^\circ = -3.04\,V$

$2Br^- \rightleftharpoons Br_2(l) + 2e^-$ anode $E^\circ = +1.09\,V$

ii $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= -3.04 - 1.09$

$= -4.13\,V$

4.13 V necessary

iii cathode = reduction = buildup of $Li(s)$
mass will increase

iv. on image

13. a. electrolytic

b. left

c. anode: SO_4^{2-} oxidation

$SO_{3(aq)}^{2-} + H_2O(l) \rightleftharpoons SO_4^{2-}(aq) + 2e^- + 2H^+(aq)$ $E^\circ = +0.17\,V$

cathode: Ni^{2+} reduction

$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$ $E^\circ = -0.26\,V$

d. $E^\circ_{cell} = -0.26 - 0.17$

$= -0.43\,V$