$Ag^+ + 1e \rightarrow Ag_{(s)}$ ← I

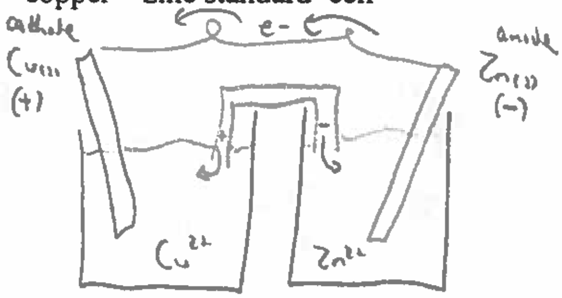
KEY

Electrochemical Cells

Draw the following electrochemical cells and calculate the cell potential. Label the following:

- anode
- cathode
- positive terminal
- negative terminal
- direction of electron flow
- direction of cation flow
- direction of anion flow
- reduction half-reaction
- oxidation half-reaction
- net reaction

1. copper – zinc standard cell

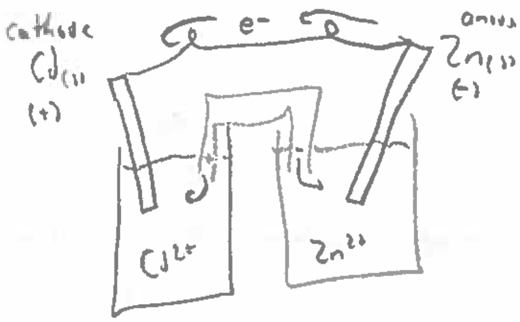


cat: $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

an: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$

$E^\circ_{cell} = +0.34V - {}^-0.76V = \boxed{+1.10V}$

2. zinc – cadmium standard cell

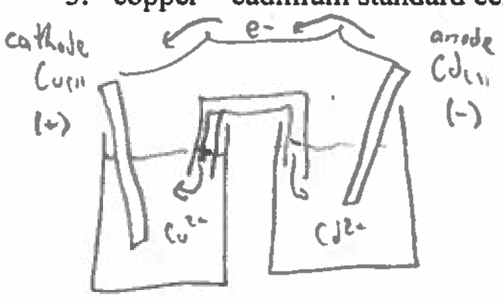


cat: $Cd^{2+}_{(aq)} + 2e^- \rightarrow Cd_{(s)}$

an: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$

$E^\circ_{cell} = -0.40V - -0.76V = \boxed{+0.36V}$

3. copper – cadmium standard cell



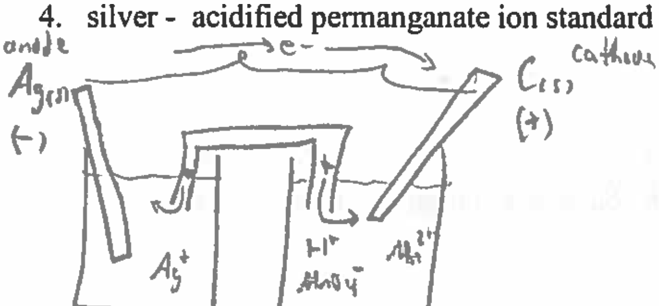
cat: $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

an: $Cd_{(s)} \rightarrow Cd^{2+}_{(aq)} + 2e^-$

$E^\circ_{cell} = +0.34V - {}^-0.28V$

$\boxed{= +0.62V}$

4. silver - acidified permanganate ion standard cell



cat: $MnO_{4\,(aq)}^- + 8H^+_{(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$

an: $Ag^+_{(aq)} + 1e^- \rightarrow Ag_{(s)}$

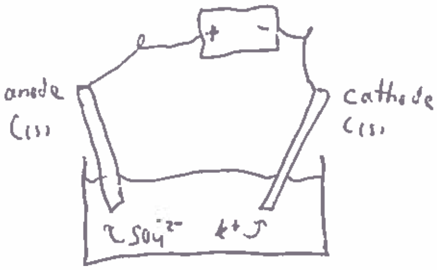
$E^\circ_{cell} = +1.51V - +0.80V$

$\boxed{= +0.71V}$

Electrolytic Cells

For each of the following, draw the cell, determine the half-reactions, net reaction and the minimum voltage required.

1. An aqueous solution of potassium sulphate is electrolyzed.



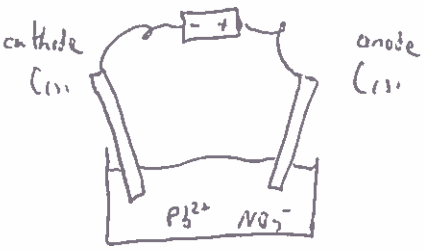
cat: $2[2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}]$

an: $2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$

$6H_2O_{(l)} \rightarrow 2H_{2(g)} + 4OH^-_{(aq)} + O_{2(g)} + 4H^+_{(aq)}$

$E^\circ_{cell} = -0.83\text{ V} - 1.23\text{ V} = \boxed{-2.06\text{ V}}$

2. An aqueous solution of lead (II) nitrate is electrolyzed.



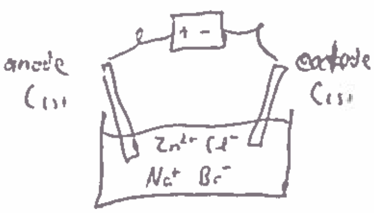
cat: $2[Pb^{2+}_{(aq)} + 2e^- \rightarrow Pb_{(s)}]$

an: $2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$

$2Pb^{2+}_{(aq)} + 2H_2O_{(l)} \rightarrow 2Pb_{(s)} + O_{2(g)} + 4H^+_{(aq)}$

$E^\circ_{cell} = \boxed{-1.36\text{ V}}$

3. A solution of aqueous sodium bromide and aqueous zinc chloride are mixed in an electrolytic cell using inert electrodes.



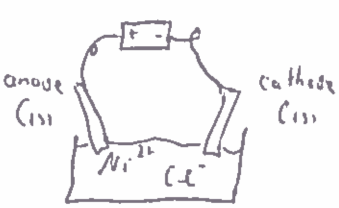
cat: $Zn^{2+} + 2e^- \rightarrow Zn_{(s)}$

an: $2Br^- \rightarrow Br_2 + 2e^-$

$Zn^{2+}_{(aq)} + 2Br^-_{(aq)} \rightarrow Zn_{(s)} + Br_{2(l)}$

$E_{cell} = \boxed{-1.83\text{ V}}$

4. An aqueous solution of nickel (II) chloride is electrolyzed.



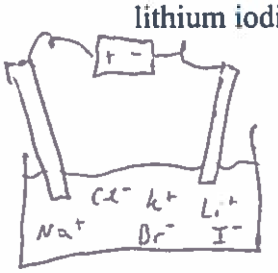
cat: $Ni^{2+} + 2e^- \rightarrow Ni_{(s)}$

an: $2Cl^- \rightarrow Cl_{2(g)} + 2e^-$

$Ni^{2+}_{(aq)} + 2Cl^-_{(aq)} \rightarrow Ni_{(s)} + Cl_{2(g)}$

$E^\circ_{cell} = \boxed{-1.62\text{ V}}$

5. Electricity is passed through an aqueous solution containing sodium chloride, potassium bromide, and lithium iodide.



cat: $2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}$

an: $2I^- \rightarrow I_{2(s)} + 2e^-$

$2H_2O_{(l)} + 2I^- \rightarrow H_{2(g)} + 2OH^-_{(aq)} + I_{2(s)}$

$E^\circ_{cell} = \boxed{-1.37\text{ V}}$

Quantitative Study of Electrolysis

1. Over a period of 30.0 minutes, a nickel-cadmium battery supplied a current of 0.268 A to a calculator. Calculate the charge supplied by the battery. $Q = 482\ C$

2. How many moles of electrons were supplied by an electrochemical cell producing a current of 0.200 A for a period of one hour? $n_e = 0.00746$ mol

3. A transistor radio is turned on for two hours. The battery caused 0.500 mol of electrons to flow. What was the average current supplied to the radio? $I = 6.70\ A$

4. How many moles of electrons pass through a bulb when 8600 C of charge are supplied? $n_e = 0.08912$ mol

5. If a 5.00 A current flows from an electrochemical cell and 0.280 mol of electrons leave the cell, how long was the cell in operation? $t = 5.40 \times 10^3$ s or 90.1 min

6. How long does it take a 0.500 A current to produce a charge of 5800 C? $t = 3.22$ h

7. Determine the amperage involved when 0.500 mol of electrons flow through a wire for 90.0 min. $I = 8.94\ A$

8. An electrolytic cell containing molten chromium(III) chloride operated for 45.0 minutes. It was found that the mass of molten chromium metal formed was 1.56 g. Calculate the average current of the cell. $I = 3.22\ A$

9. How long must a 0.250 A current run through an electrolytic cell containing a solution of iron(III) nitrate, so that 8.37 g of iron(II) ions are produced? $t = 16.1$ h

10. When a 7.50 A current is passed through molten nickel(II) chloride for 1.40 hours, what mass of solid nickel will collect on the cathode? $m_{Ni} = 11.5$ g

11. What is the average current required to produce 8.25 g of iodine at the anode of an electrolytic cell containing a solution of tin(II) iodide, if the cell operates for 150 minutes? $I = 0.697\ A$

12. If a 2.80 A current runs for 4.00 hours through a solution of zinc sulphate, what mass of zinc solid will be produced? $m_{Zn} = 13.7$ g

13. If 9.72 g of magnesium was formed in an electrolytic cell using a current of 0.600 A, how long did the cell operate? $t = 35.7$ h

14. Determine the mass of magnesium deposited at the cathode of a molten $MgCl_2$ electrolytic cell if 10.0 A flow through the cell for 9.65 h. $m = 43.8$ g

15. An electroplating firm wishes to plate 12.7 g of copper from a $Cu(NO_3)_2(aq)$ solution onto a pair of baby shoes. If a 2.00 A current is used, calculate the time required. At which electrode would the shoes be attached? cathode $t = 5.36$ h

16. If 76 g of fluorine are required, what current would have to flow for 10 h to produce the fluorine from molten NaF? At which electrode could this reaction occur? anode
$I = 10.72... = \boxed{11 A}$

17. If a current plates out 13.5 g of aluminum, what mass of magnesium would be plated out in the same time by the same current?

$Al^{3+} + 3e^- \rightarrow Al_{(l)}$ $\quad$ $n_e = 1.5011...$ $\quad \times \frac{3}{1}$ $\quad m = 13.5$ g, $M =$, $n = 0.500...$ mol

$Mg^{2+} + 2e^- \rightarrow Mg_{(l)}$ $\quad$ $n_e = 1.5011...$ mol $\quad \times \frac{1}{2}$ $\quad n = 0.750...$ mol, $M =$

$\boxed{m = 18.2 g}$

$[H_3O^+]$	pH	
1 mol/L	0	10×
0.1 mol/L	1	