

Electrochemistry Problems

- 1) Given the E° for the following half-reactions:



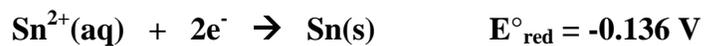
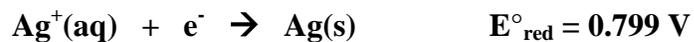
What is E° for the reaction:



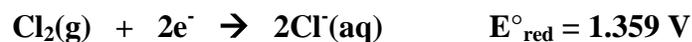
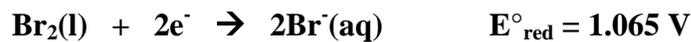
- 2) How many Faradays are required to produce 21.58 g of silver from a silver nitrate solution?
- 3) A current of 2.75 amperes is used to electrolyze a solution of copper(II) sulfate. How long will it take to deposit 10.47 grams of copper?
- 4) A voltaic cell consists of a copper electrode in a solution of copper(II) ions and a palladium electrode in a solution of palladium(II) ions. The palladium is the cathode and its reduction potential is 0.951 V.
- Write the half-reaction that occurs at the anode.
 - If E° is 0.609 V, what is the potential for the oxidation half-reaction?
 - What is K_{eq} for this reaction?
- 5) 5.77 g of zinc is deposited at the cathode when a current of 7.1 amperes passes through an electrolytic cell for 40. minutes. What is the oxidation state of the zinc in the aqueous solution?

6) For each pair of species, choose the better reducing agent.

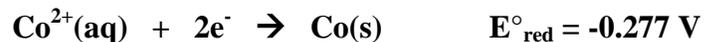
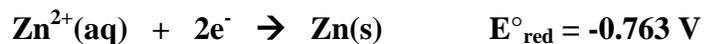
(a) Ag(s) or Sn(s), given:



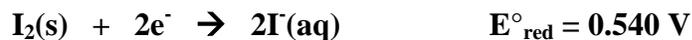
(b) Br⁻(aq) or Cl⁻(aq), given:



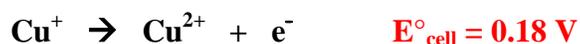
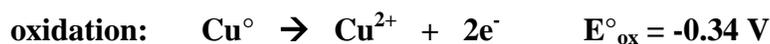
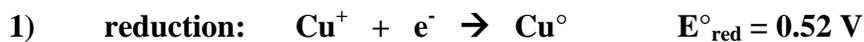
(c) Zn(s) or Co(s), given:



(d) Au(s) or I⁻(aq), given:

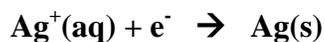


Solutions



2) $m = 21.58 \text{ g Ag}$

$$1 \text{ F} = 1 \text{ mol e}^-$$



$$n_{\text{F}} = 21.58 \text{ g Ag} \times 1 \text{ mol Ag} / 107.90 \text{ g Ag} \times 1 \text{ mol e}^- / 1 \text{ mol Ag} \times 1 \text{ F} / 1 \text{ mol e}^-$$

$$n_{\text{F}} = 0.200 \text{ F}$$

3) $I = 2.75 \text{ A} = 2.75 \text{ C/s}$

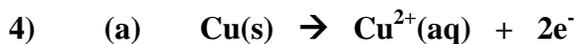
$$m = 10.47 \text{ g Cu}$$



$$10.47 \text{ g Cu} = 2.75 \text{ C/s} \times t \times 1 \text{ mol Cu} / 2 \text{ mol e}^- \times 1 \text{ mol e}^- / 96500 \text{ C} \times$$

$$63.55 \text{ g Cu} / 1 \text{ mol Cu}$$

$$t = 11600 \text{ s} \times 1 \text{ hr} / 3600 \text{ s} = 3.22 \text{ hr}$$



$$E^\circ_{\text{cell}} = E^\circ_{\text{red}} + E^\circ_{\text{ox}}$$

$$0.609 \text{ V} = 0.951 \text{ V} + E^\circ_{\text{ox}}$$

$$E^\circ_{\text{ox}} = -0.340 \text{ V}$$

(c) $\log K_{\text{eq}} = n \times E^\circ / 0.0592$

$$\log K_{\text{eq}} = 2 \times 0.609 \text{ V} / 0.0592 = 20.6$$

$$K_{\text{eq}} = 10^{20.6} = 3.98 \times 10^{20}$$

5) $m = 5.77 \text{ g Zn}$

$$I = 7.1 \text{ A} = 7.1 \text{ C/s}$$

$$t = 40. \text{ min} \times 60 \text{ s/1 min} = 2.4 \times 10^3 \text{ s}$$



$$5.77 \text{ g Zn} = 7.1 \text{ C/s} \cdot 2.4 \times 10^3 \text{ s} \cdot 1 \text{ mol Zn} / x \text{ mol e}^- \cdot 1 \text{ mol e}^- / 96500 \text{ C}$$

$$\cdot 65.39 \text{ g Zn} / 1 \text{ mol Zn}$$

$$x = 2, \text{ therefore } \text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn(s)}$$

- 6) (a) **Sn(s) because it is more difficult to reduce than Ag(s).**

The more difficult it is to reduce a species, the more readily its products are oxidized. This inverse relationship between reducing and oxidizing agents is similar to the inverse relationship between the strengths of conjugate acids and bases.

- (b) **Br⁻(aq) because Cl⁻(aq) is the better oxidizing agent making Br⁻(aq) the better reducing agent.**

- (c) **Zn(s) because it is a weaker oxidizing agent than Co(s).**

- (d) **Au(s) is the better oxidizing agent, therefore I⁻(aq) is the better reducing agent.**