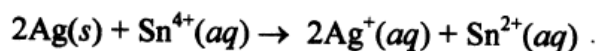
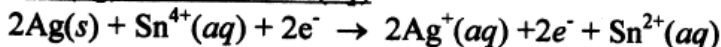
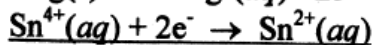


2. Refer to Sections 4.4 and 18.1.

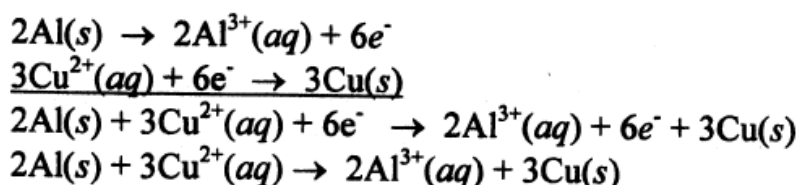
Remember that the oxidation is always shown on the left side of the cell notation. For part (c), note that the Pt serves as an inert electrode and is not part of the chemical equation.



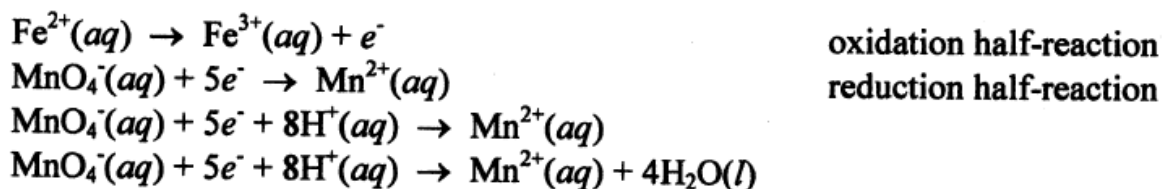
Multiply the oxidation half-reaction by two, then add the two half-reactions together.



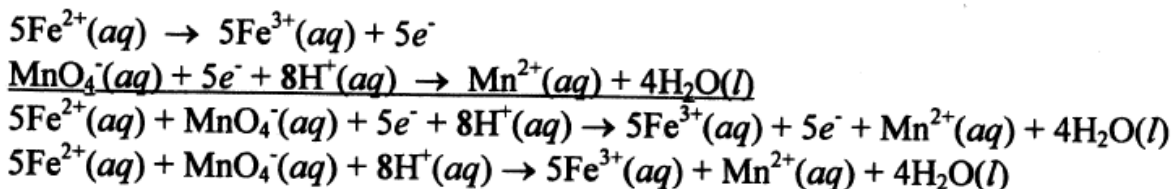
Multiply the oxidation half-reaction by two and the reduction half-reaction by three, then add the two half-reactions together.



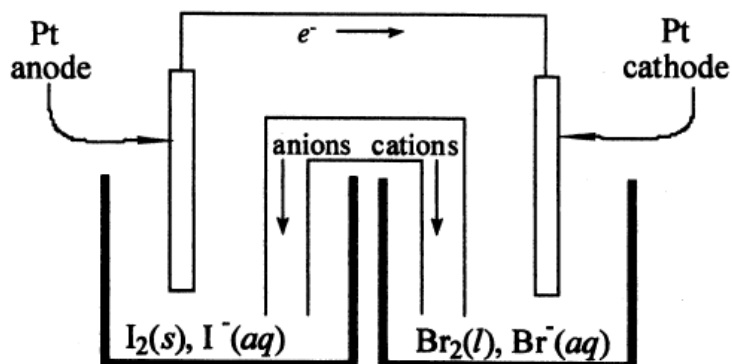
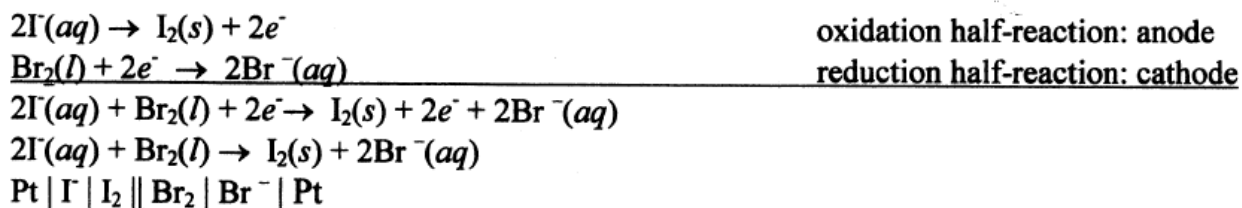
c. Note: The cell notation should be:



Balance electrons by multiplying the oxidation equation by 5, then add the two half-reactions together.



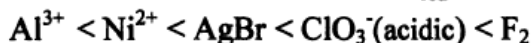
6. Refer to Section 18.1 and Example 18.1.



10. Refer to Section 18.2 and Example 18.2.

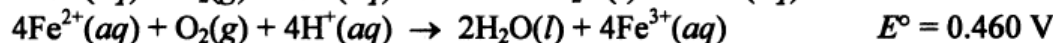
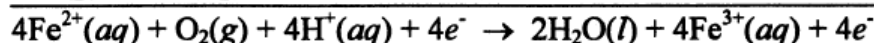
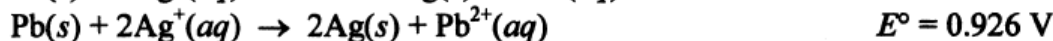
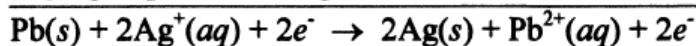
	Al^{3+}	Ni^{2+}	AgBr	ClO_3^- (acidic)	F_2
E°_{ox} (V)	-1.68	-0.236	0.073	1.458	2.889

The species with the highest E°_{red} is the strongest oxidizing agent, thus:

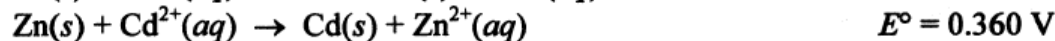
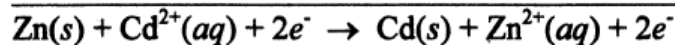
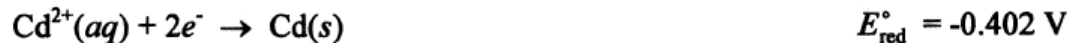


16. Refer to Section 18.2 and Example 18.3.

Write the half-reactions and the associated redox potentials. When combining the half-reactions, add the potentials to calculate E° . Bear in mind that **multiplying a half-reaction by a factor does not change that E° value.**

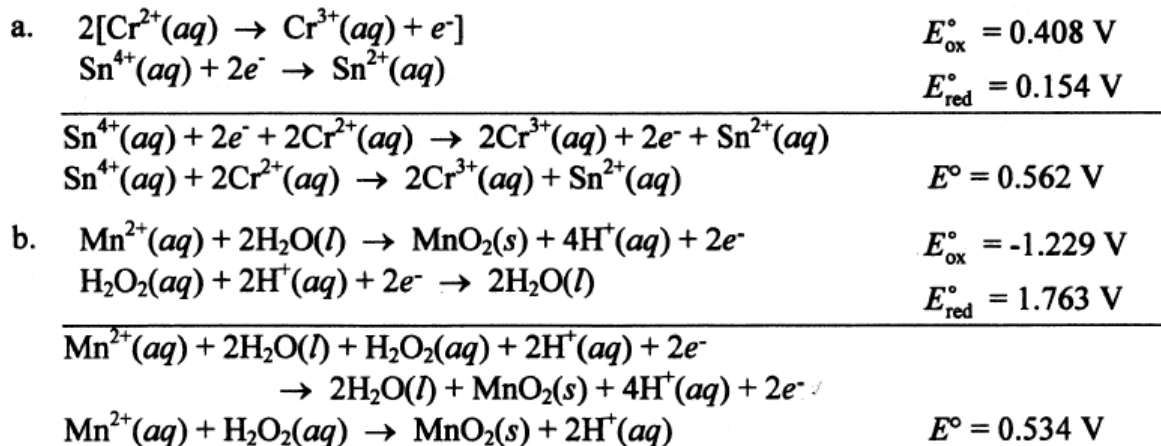


c. Reverse the Zn-Zn²⁺ reaction and change the sign of E° to obtain a positive E° (cell).



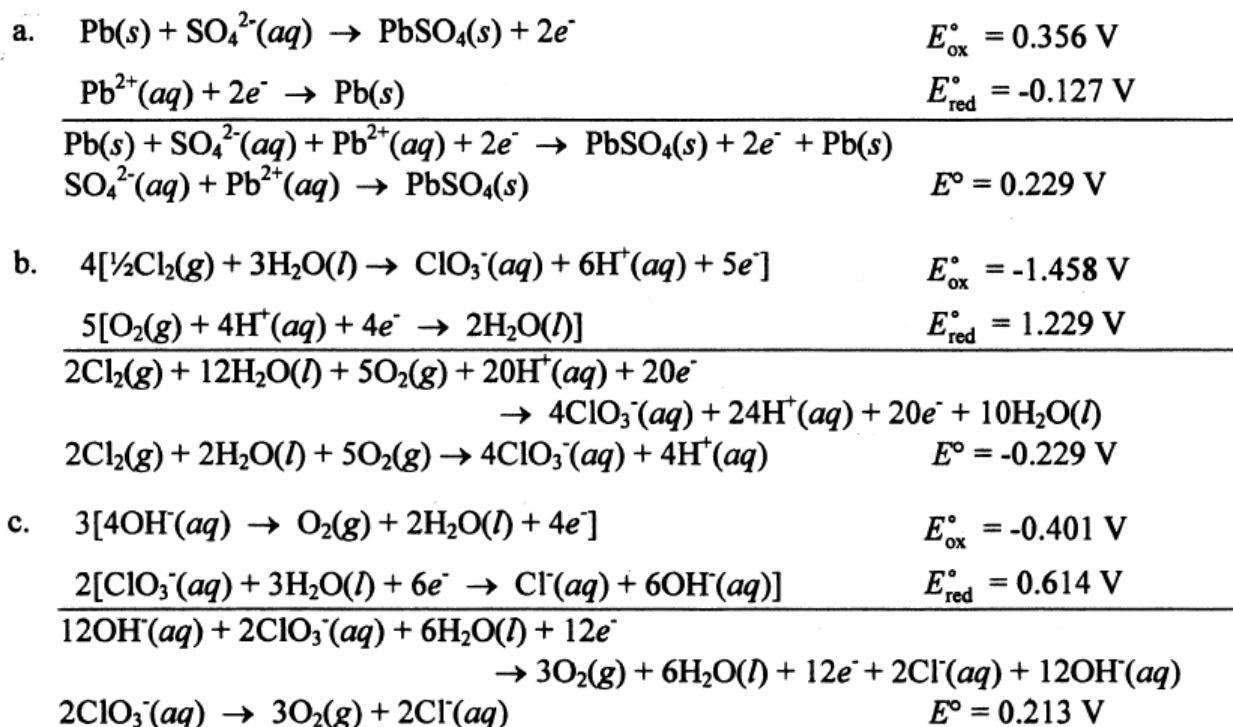
18. Refer to Section 18.2, Table 18.1, and Example 18.3.

Write the half-reactions and the associated redox potentials. When combining the half-reactions, add the potentials to calculate E° . Bear in mind that **multiplying a half-reaction by a factor does not change that E° value.**



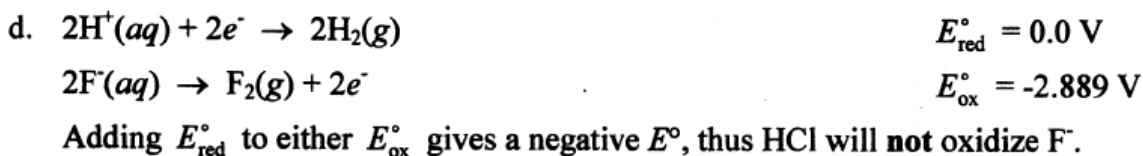
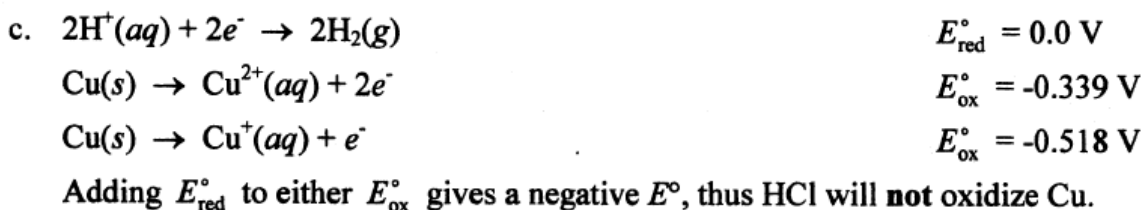
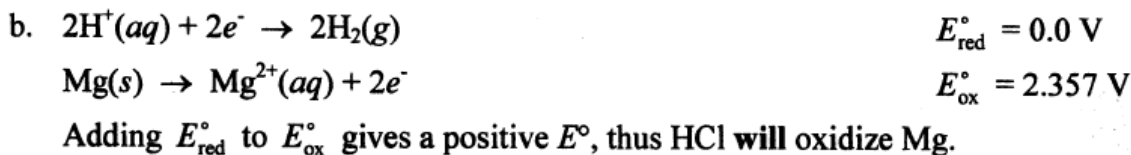
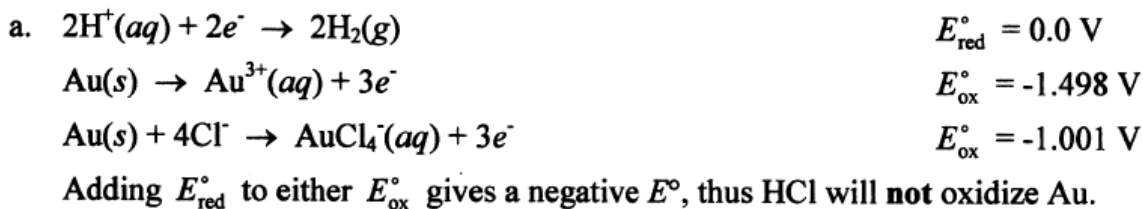
20. Refer to Section 18.2.

Write the half-reactions and the associated redox potentials. When combining the half-reactions, add the potentials to calculate E° . Bear in mind that **multiplying a half-reaction by a factor does not change that E° value.**

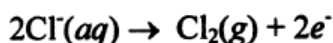
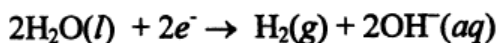


30. Refer to Sections 18.2 and 18.3 and Example 18.4.

If HCl is oxidizing one of the listed species, then the HCl itself must be reduced. Since the Cl^- is already in a reduced state, the likely species to be reduced is H^+ .



58. Refer to Section 18.5.



a. $0.288 \text{ mol. e}^- \times \frac{6.022 \times 10^{23} \text{ e}^-}{1 \text{ mol.}} = 1.37 \times 10^{23} \text{ e}^-$

b. $0.288 \text{ mol. e}^- \times \frac{9.648 \times 10^4 \text{ C}}{1 \text{ mol. e}^-} = 2.20 \times 10^4 \text{ C}$

c. $0.288 \text{ mol. e}^- \times \frac{1 \text{ mol. H}_2}{2 \text{ mol. e}^-} \times \frac{2.016 \text{ g}}{1 \text{ mol. H}_2} = 0.230 \text{ g H}_2$

$0.288 \text{ mol. e}^- \times \frac{1 \text{ mol. Cl}_2}{2 \text{ mol. e}^-} \times \frac{70.90 \text{ g}}{1 \text{ mol. Cl}_2} = 8.08 \text{ g Cl}_2$

60. Refer to Section 18.5.

- a. Calculate the volume and mass of gold to be plated out. Bear in mind that the gold will be plated on both sides of the thin sheet.

$$2(1.5 \text{ in} \times 8.5 \text{ in} \times 0.0020 \text{ in}) = 0.051 \text{ in}^3$$

$$0.051 \text{ in}^3 \times \frac{(2.54 \text{ cm})^3}{(1 \text{ in})^3} \times \frac{19.3 \text{ g Au}}{1 \text{ cm}^3 \text{ Au}} = 16 \text{ g Au}$$

- b. $\text{AuCN}(s) + e^- \rightarrow \text{Au}(s) + \text{CN}^-(aq)$

$$16 \text{ g Au} \times \frac{1 \text{ mol. Au}}{197.0 \text{ g Au}} \times \frac{1 \text{ mol. } e^-}{1 \text{ mol. Au}} \times \frac{9.648 \times 10^4 \text{ C}}{1 \text{ mol. } e^-} = 7.8 \times 10^3 \text{ C}$$

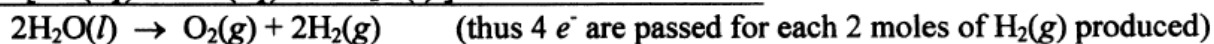
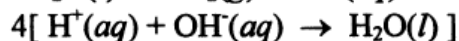
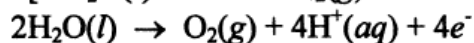
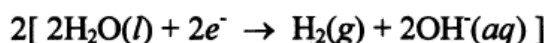
$$7.8 \times 10^3 \text{ C} \times \frac{1 \text{ A} \cdot \text{s}}{1 \text{ C}} \times \frac{1}{7.00 \text{ A}} = 1.1 \times 10^3 \text{ s}$$

$$1.1 \times 10^3 \text{ s} \times \frac{1 \text{ min}}{60 \text{ s}} = 19 \text{ min}$$

64. Refer to Section 18.5 and Chapter 5.

Use the ideal gas equation to calculate the moles of $\text{H}_2(g)$ needed. Write the half-reactions needed to derive the equation for the electrolysis of water. Then use the moles of $\text{H}_2(g)$ and moles of e^- to determine the current and then the time needed for the electrolysis.

$$n = \frac{PV}{RT} = \frac{(0.924 \text{ atm})(10.00 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(295 \text{ K})} = 0.382 \text{ mol. H}_2$$



$$0.382 \text{ mol. H}_2 \times \frac{4 \text{ mol. } e^-}{2 \text{ mol. H}_2} \times \frac{9.648 \times 10^4 \text{ C}}{1 \text{ mol. } e^-} = 7.36 \times 10^4 \text{ C}$$

$$7.36 \times 10^4 \text{ C} \times \frac{1 \text{ A} \cdot \text{s}}{1 \text{ C}} \times \frac{1}{12.0 \text{ A}} = 6.13 \times 10^3 \text{ s}$$

$$6.13 \times 10^3 \text{ s} \times \frac{1 \text{ min}}{60 \text{ s}} \times \frac{1 \text{ hr}}{60 \text{ min}} = 1.70 \text{ hr}$$

72. Refer to Section 18.5.

Water is electrolyzed to form H_2 and O_2 (see p 496). Look for the box containing two squares together (H_2) and two circles together (O_2).
Box (c) represents the electrolysis of water.