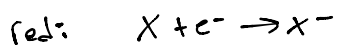
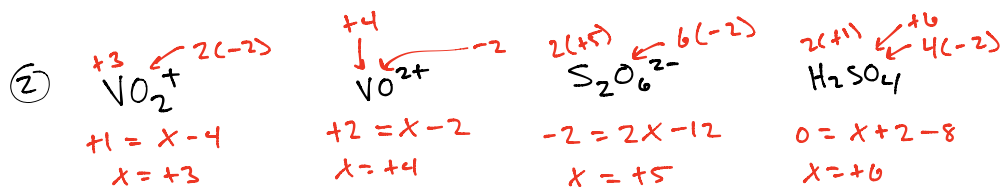
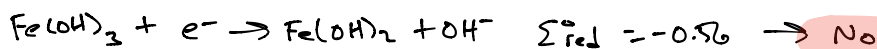
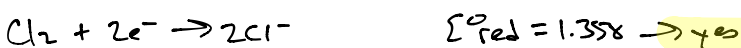
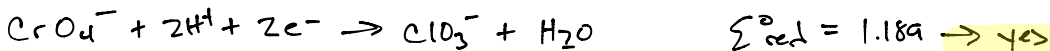
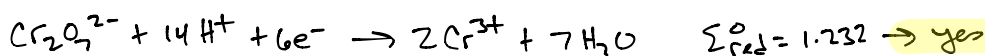
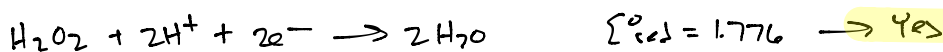
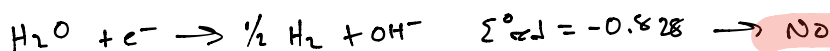
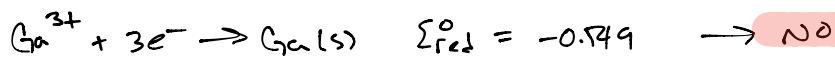


1. The key to this problem is recognizing that I^- must be the reducing agent; so it must be oxidized if the reaction happens:

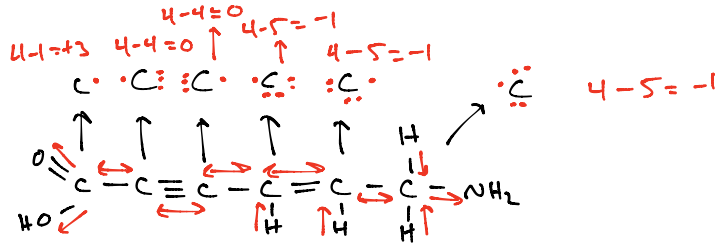
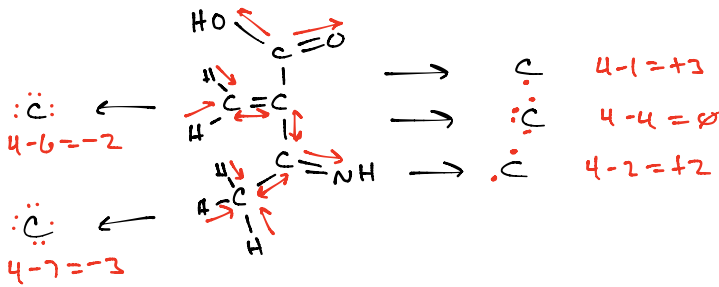
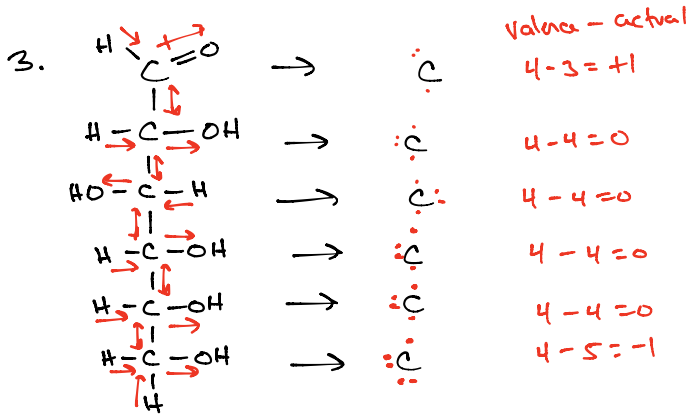
• $\sum \sigma_{\text{oxidation}} + \sum \sigma_{\text{reduction}} > 0$ for a spontaneous reaction



- So, IF $\sum \sigma_{\text{red}} > 0.536 \text{ V}$, the reaction will be spontaneous



Red arrows indicate where e^- go (to more electronegative atoms)

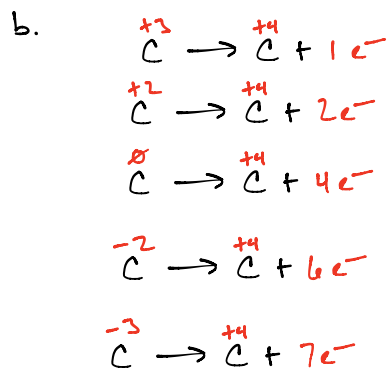


4. In combustion reactions, ALL carbon ends up as CO_2 (C has a +4 ox. state)

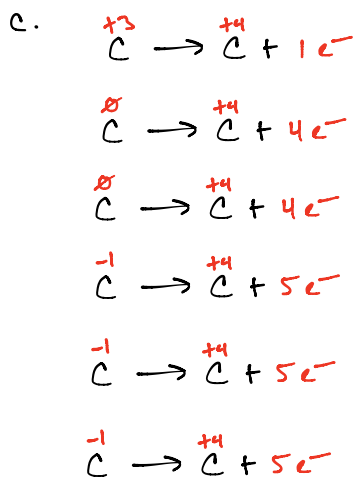
a. 4 carbons @ $\overset{0}{\text{C}}$, 1 @ $\overset{-1}{\text{C}}$, 1 @ $\overset{+1}{\text{C}}$



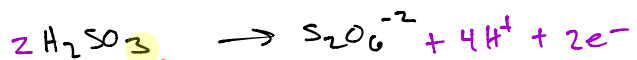
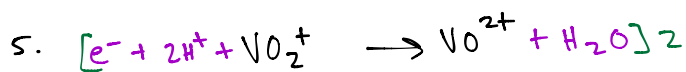
Total $24e^-$
Transferred to O_2



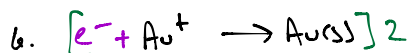
Total = 20e⁻ transferred to O₂



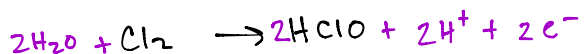
Total 24e⁻ transferred to O₂



↑ note that this is a typo in the sheet seen in class



$E^\circ = 1.692 \text{ V}$



$E^\circ = -1.611 \text{ V}$



b. Oxidizing Agent (gets reduced) = Au⁺

c. Reducing Agent (gets oxidized) = Cl₂

$$d. \Sigma^{\circ} = \Sigma_{red}^{\circ} + \Sigma_{ox}^{\circ} = 1.692 - 1.611 = 0.081 \text{ V}$$

$$\Delta G^{\circ} = -n F \Sigma^{\circ} = -2(96485)(0.081) = -15,630.6 \text{ J/mol}$$

↑
2 (because 2e⁻ in both reactions)

$$e. \Delta G^{\circ} = -RT \ln K$$

$$-15630.6 \text{ J/mol} = 8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} (298.15 \text{ K}) \ln K$$

$$\ln K = 6.31$$

$$K = 547.6$$

These are used interchangeably

← [H⁺] = [CH₃O⁺]

$$f. \Delta G = \Delta G^{\circ} + RT \ln Q$$

$$\Delta G = -15630.6 + 8.314 (50 + 273.15) \ln 2 \times 10^{-3}$$

$$\Delta G = -15630.6 - 16696.6$$

$$\Delta G = -32327.2 \text{ J/mol}$$

$$Q = \frac{[\text{HClO}]^2 [\text{H}^+]^2}{[\text{AO}^+]^2 [\text{Cl}_2]} = \frac{(0.01)^2 (10^{-7})^2}{(10 \times 10^{-6})^2 (5 \times 10^{-4})}$$

$$Q = 2 \times 10^{-3}$$

K > Q, so this should end up being spontaneous!

g. Same approach as f, except [H⁺] is the unknown:

$$\Delta G = 0 = \Delta G^{\circ} + RT \ln Q$$

$$15,630.6 = 8.314 (273.15 + 50) \ln Q$$

$$\ln Q = 5.8178$$

$$Q = 336.24$$

$$Q = \frac{[\text{HClO}]^2 [\text{H}^+]^2}{[\text{AO}^+]^2 [\text{Cl}_2]} = \frac{(0.01)^2 (X)^2}{(10 \times 10^{-6})^2 (5 \times 10^{-4})} = 336.24$$

$$X^2 = 1.68 \times 10^{-9}$$

$$X = [\text{H}^+] = 4.1 \times 10^{-5}$$

$$\text{pH} = 4.39$$