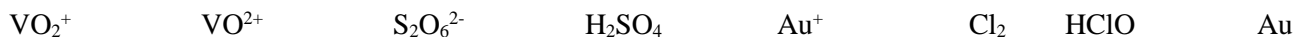


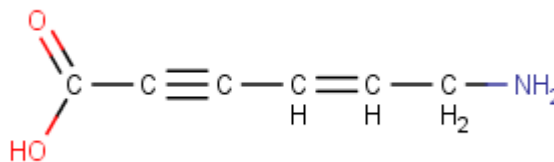
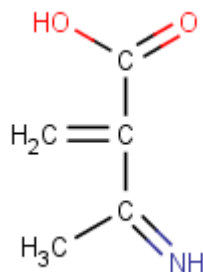
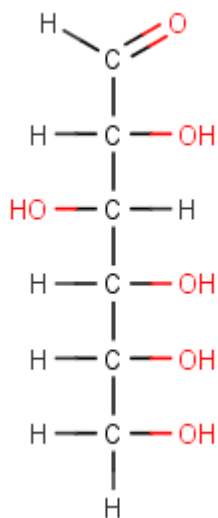
1. Which of the following molecules can be reduced by KI?



2. Assign oxidation states to each atom in the following molecules:

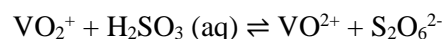


3. Determine the oxidation state or EACH carbon (you need the Lewis structure approach for this).

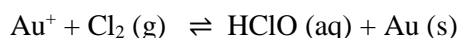


4. Combustion reactions are redox reactions. Determine the total number of electrons that are passed from carbon to oxygen when each of the molecules in problem 2 is combusted. Note that  $\text{NH}_3$  is a product in the combustion of the second and third molecules.

5. Balance this redox reaction and determine  $E^\circ$ .



6. Consider the following reaction:



- Balance the reaction
- Identify the Oxidizing Agent
- Identify the Reducing Agent
- Determine  $E^\circ$  and  $\Delta G^\circ$
- Determine  $K$  at standard conditions.
- If  $[\text{Au}^+] = 10 \mu\text{M}$ ,  $[\text{Cl}_2] = 5 \mu\text{M}$ , and  $[\text{HClO}] = 10 \text{mM}$ , determine if the reaction is spontaneous at neutral pH and  $50^\circ\text{C}$ .
- If  $[\text{Au}^+] = 10 \mu\text{M}$ ,  $[\text{Cl}_2] = 5 \mu\text{M}$ , and  $[\text{HClO}] = 10 \text{mM}$ , determine the pH that will make this reaction non-spontaneous at  $50^\circ\text{C}$ .

Recall that

$$\Delta G^\circ = -nFE^\circ$$

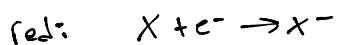
$$\Delta G = \Delta G^\circ + RT \ln Q$$

$$R = 8.314 \text{ J/mol K}$$

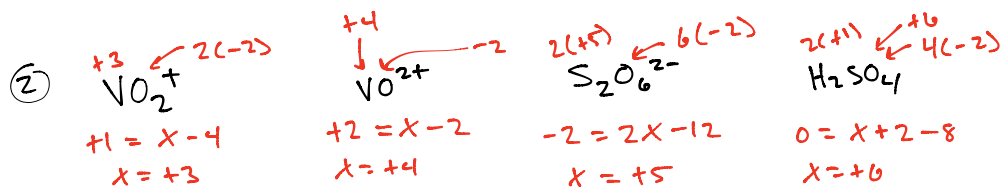
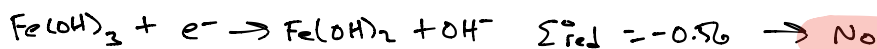
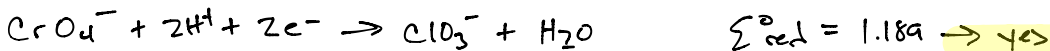
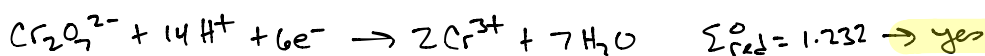
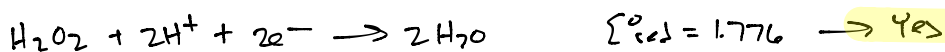
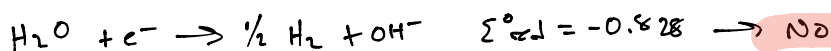
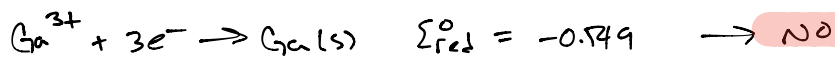
$$F = 96485 \text{ J/mol V}$$

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

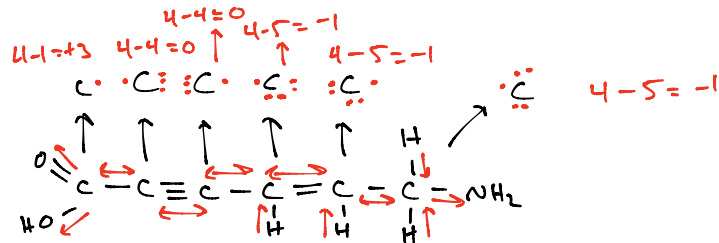
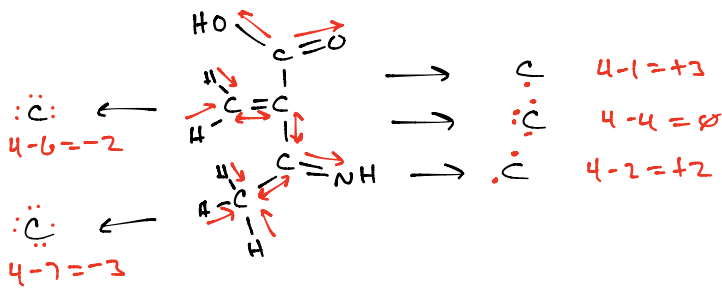
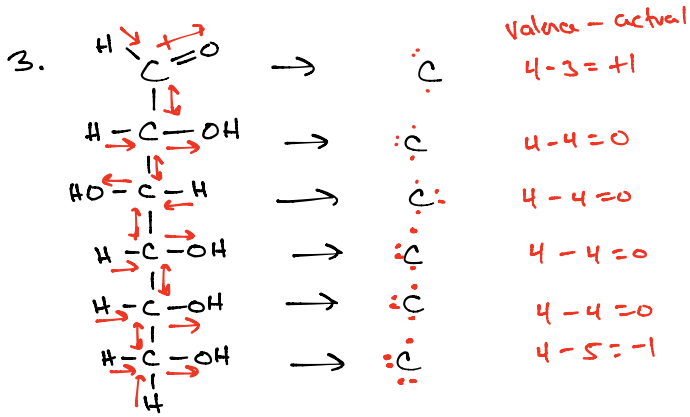
•  $\sum \sigma_{\text{oxidation}} + \sum \sigma_{\text{reduction}} > 0$  for a spontaneous reaction



- 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)

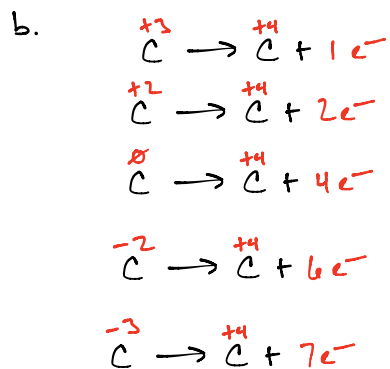


4. In combustion reactions, ALL carbon ends up as  $\text{CO}_2$  (C has a +4 ox. state)

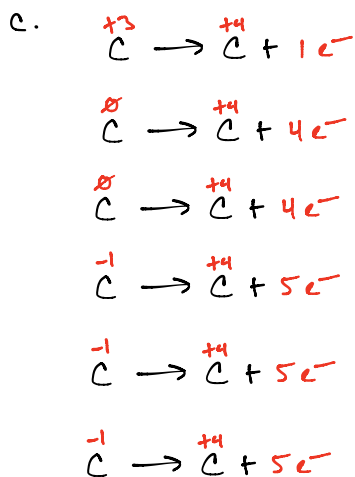
a. 4 carbons @  $\overset{0}{\text{C}}$ , 1 @  $\overset{-1}{\text{C}}$ , 1 @  $\overset{+1}{\text{C}}$



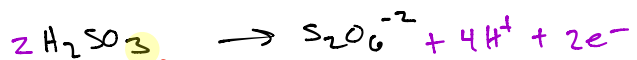
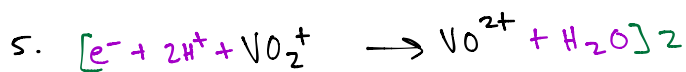
Total  $24e^-$   
Transferred to  $\text{O}_2$



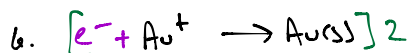
Total = 20e<sup>-</sup> transferred to O<sub>2</sub>



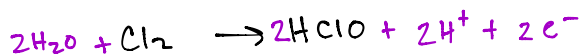
Total 24e<sup>-</sup> transferred to O<sub>2</sub>



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



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



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



b. Oxidizing Agent (gets reduced) = Au<sup>+</sup>

c. Reducing Agent (gets oxidized) = Cl<sub>2</sub>

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

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

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

$[H^{+}] = [CH_3O^{+}]$

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

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

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

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

$$Q = \frac{[HClO]^{2} [H^{+}]^{2}}{[AO^{+}]^{2} [Cl_2]} = \frac{(0.01)^{2} (10^{-7})^{2}}{(10 \times 10^{-6})^{2} (5 \times 10^{-4})}$$

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

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

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

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

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

$$\ln Q = 5.8178$$

$$Q = 336.24$$

$$Q = \frac{[HClO]^{2} [H^{+}]^{2}}{[AO^{+}]^{2} [Cl_2]} = \frac{(0.01)^{2} (X)^{2}}{(10 \times 10^{-6})^{2} (5 \times 10^{-4})} = 336.24$$

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

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

$$pH = 4.39$$