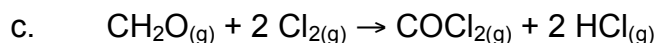
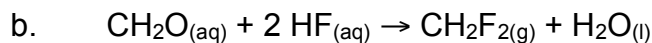
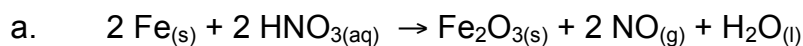


## Conference Problems: Week of April 14<sup>th</sup>

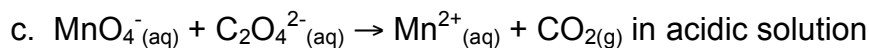
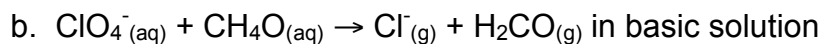
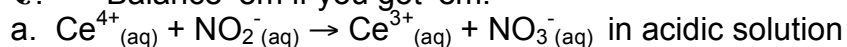
☼. Identify the oxidation number for the underlined element in each of the following species:

- a. Fe<sub>2</sub>O<sub>3</sub>    b. XeF<sub>3</sub>    c. NH<sub>4</sub><sup>+</sup>    d. Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup>    e. C<sub>2</sub>H<sub>2</sub>Cl<sub>4</sub>    f. H<sub>3</sub>PO<sub>4</sub>

★. Identify which of the following are redox reactions by determining oxidation states of all non-H/non-O atoms and identifying oxidants and reductants, where relevant.



☼. Balance 'em if you got 'em.



- §. Consider the following reduction potentials
- $$\text{NO}_2^- + 2 \text{H}^+ + \text{e}^- \rightarrow \text{NO} + \text{H}_2\text{O} \quad E_{\text{red}}^\circ = 1.19 \text{ V}$$
- $$\text{O}_2 + 4 \text{H}^+ + 4 \text{e}^- \rightarrow 2 \text{H}_2\text{O} \quad E_{\text{red}}^\circ = 1.23 \text{ V}$$

Predict whether  $\text{NO}_2^-$  will spontaneously oxidize water at pH 7. Assume all other species are at 1 M or 1 atm and that the temperature is 25°C.

- ✱. Consider an electrochemical cell at the standard state composed of two half-cells with the following half-reactions:



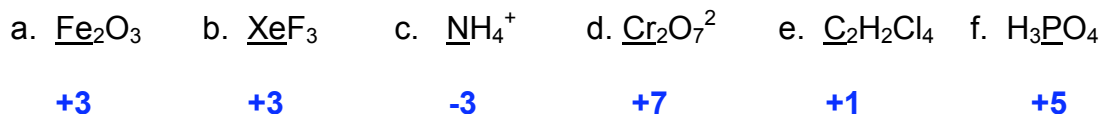
- a. Diagram the cell that functions spontaneously, specifically labeling the anode, the cathode, their contents, and the flow of electrons. What is the cell potential?

- b. How long would it take this cell, operating at a current of 1 A, to oxidize 1.00 g of metal?

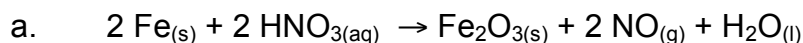
- ⇒. Revisiting a problem from last week, lead(II) carbonate is a weak electrolyte whose solubility is affected by the pH of solution because  $\text{CO}_3^{2-}$  is a basic anion ( $\text{p}K_{\text{b}1} = 3.7$ ,  $\text{p}K_{\text{b}2} = 7.6$ ). The  $K_{\text{sp}}$  of  $\text{PbCO}_3$  is  $1.5 \times 10^{-13}$ . At pH 5.6, the fraction of carbonate present in solution as  $\text{CO}_3^{2-}$  is  $2.5 \times 10^{-6}$ . What concentration of lead will you get when the salt goes to equilibrium in these conditions?

## Solutions

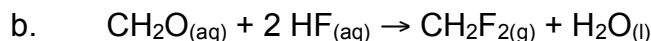
☼. Identify the oxidation number for the underlined element in each of the following species:



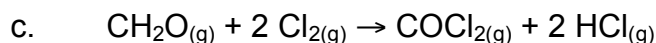
★. Identify which of the following are redox reactions by determining oxidation states of all non-H/non-O atoms and identifying oxidants and reductants, where relevant.



Fe (o.s. = 0) is reductant,  $\text{HNO}_3$  (N o.s. = +5) is oxidant.  
In  $\text{Fe}_2\text{O}_3$ , Fe has o.s. = +3, and in NO, N has o.s. = +2.

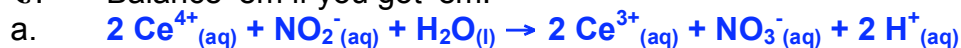


Not a redox reaction. Carbon has o.s. = 0 in reactants and products.

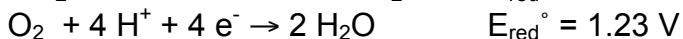
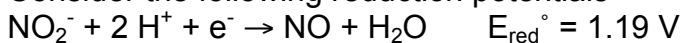


$\text{CH}_2\text{O}$  is the reductant (C has o.s. = 0) and  $\text{Cl}_2$  is oxidant (Cl o.s. = 0)  
In  $\text{COCl}_2$  carbon has o.s. = +4 and chlorine has o.s. = -1.

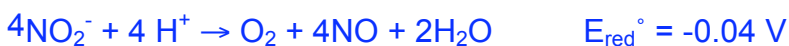
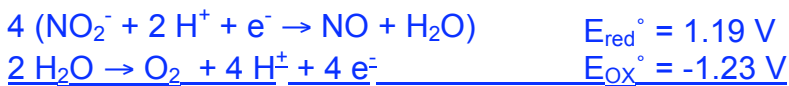
☼. Balance 'em if you got 'em.



§. Consider the following reduction potentials

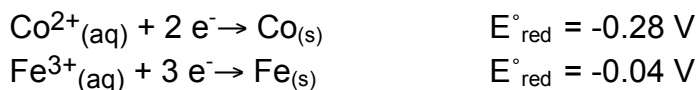


Predict whether  $\text{NO}_2^{-}$  will spontaneously oxidize water at pH 7. Assume all other species are at 1 M or 1 atm and that the temperature is 25°C.

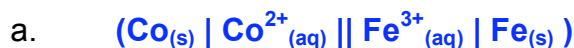


$E = E^{\circ} - (0.0592 \text{ V}/4)\log(1/(10^{-7})^4) = -0.44 \text{ V}$ . Not spontaneous!

✱. Consider an electrochemical cell at the standard state composed of two half-cells with the following half-reactions:



a. Diagram the cell that functions spontaneously, specifically labeling the anode, the cathode, their contents, and the flow of electrons. What is the cell potential?



**This is short hand for saying that the anode (on the left) will contain a solid cobalt strip immersed in  $\text{Co}^{2+}$  solution and the cathode (on the right) will contain a solid iron strip immersed in  $\text{Fe}^{3+}$  solution. The metal strips will be connected by a wire conducting electrons from the anode to the cathode. The two solutions will be linked by a salt bridge to conduct ions.**



b. How long would it take this cell, operating at a current of 1 A, to oxidize 1.00 g of metal?

$$\begin{array}{l} 1 \text{ A is } 1 \text{ C/s. } 1.00 \text{ g Co}/(58.93 \text{ g/mol}) = 1.70 \times 10^{-2} \text{ mol} \\ [(1.70 \times 10^{-2} \text{ mol Co})(2 \text{ mol e}^{-}/\text{mol Co})(96485 \text{ C/mol e}^{-})] / (1 \text{ C/s}) = 3280 \text{ s} \end{array}$$

⇒. Revisiting a problem from last week, lead(II) carbonate is a weak electrolyte whose solubility is affected by the pH of solution because  $\text{CO}_3^{2-}$  is a basic anion ( $\text{pK}_{\text{b}1} = 3.7$ ,  $\text{pK}_{\text{b}2} = 7.6$ ). The  $K_{\text{sp}}$  of  $\text{PbCO}_3$  is  $1.5 \times 10^{-13}$ . At pH 5.6, the fraction of carbonate present in solution as  $\text{CO}_3^{2-}$  is  $2.5 \times 10^{-6}$ . What concentration of lead will you get when the salt goes to equilibrium in these conditions?

$$\begin{array}{l} [\text{Pb}^{2+}][\text{CO}_3^{2-}] = s(s \cdot 2.5 \times 10^{-6}) = 1.5 \times 10^{-13} \\ s^2 = 6.0 \times 10^{-8} \\ [\text{Pb}^{2+}] = s = 2.4 \times 10^{-4} \end{array}$$