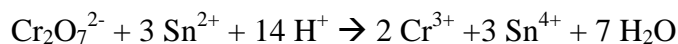


Practice problems #10

1) Consider the balanced redox equation below: [6 points]



(a) Identify the oxidant (or *oxidizing agent*) in the redox reaction above:

(b) Identify the redox couple for the **anode** of a galvanic cell that would use the above redox reaction:

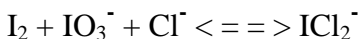
(c) If E°_{cell} for the above galvanic cell is 1.22 V, what is the maximum work you could get from it assuming you start off with 1 mole of $\text{Cr}_2\text{O}_7^{2-}$ and 3 moles Sn^{2+} of and 14 moles of H^+ and are able to maintain a steady voltage?

(d) If the conditions are as follows: $[\text{Cr}_2\text{O}_7^{2-}] = [\text{Sn}^{2+}] = [\text{Cr}^{3+}] = [\text{Sn}^{4+}] = 1.00\text{M}$ and $\text{pH} = 1.00$; what is $E_{\text{cell}} = ?$ (assume, $T = 298\text{K}$)

2) What's the oxidation number of the following elements in the corresponding compound:

a) C in ethanol, $\text{C}_2\text{H}_6\text{O}$, b) C in C_3H_3^+

3) Balance the equation: (hint: chlorine doesn't change its oxidation number)



4) The table below lists the cell potentials for 3 possible galvanic cells as the corresponding redox couples: $\text{Pd}^{2+}/\text{Pd}(\text{s})$, $\text{Co}^{2+}/\text{Co}(\text{s})$, $\text{Mg}^{2+}/\text{Mg}(\text{s})$.

Table:

	$\text{Pd}^{2+}/\text{Pd}(\text{s})$	$\text{Co}^{2+}/\text{Co}(\text{s})$
$\text{Mg}^{2+}/\text{Mg}(\text{s})$	3.36 V	2.09 V
$\text{Co}^{2+}/\text{Co}(\text{s})$	1.27 V	n/a

a) If $E^\circ_{\text{Pd}^{2+}/\text{Pd}} = 0.99 \text{ V}$, and $E^\circ_{\text{Mg}^{2+}/\text{Mg}} < 0 \text{ V}$ (i.e. a negative value), what are the other values: (in your answer, set up a table of E° values, showing their relative positions in the table).

$$E^\circ_{\text{Mg}^{2+}/\text{Mg}} = \underline{\hspace{2cm}}$$

$$E^\circ_{\text{Co}^{2+}/\text{Co}} = \underline{\hspace{2cm}}$$

Table:

$E^\circ_{\text{reduction}}$



Answers: (look at this only after you've done it)

(1) a) answer: $\text{Cr}_2\text{O}_7^{2-}$

b) answer: for the anode, you need the oxidation half rxn:



c) answer: $W_{\text{max}} = -\Delta G^\circ = nFE^\circ$; what's n=? we can just look at one of the substances changing its redox state: $\text{Sn}^{2+} \rightarrow \text{Sn}^{4+} + 2\text{e}^-$; so $3 \text{Sn}^{2+} \rightarrow 3 \text{Sn}^{4+} + 6\text{e}^-$ and n = 6
so $W = (6\text{mol})(96485 \text{ C/mol})(1.22\text{V}) = 706 \text{ kJ of work!}$

$$\begin{aligned} \text{d) answer: } \Delta E &= \Delta E^\circ - \frac{0.0592\text{V}}{n} \log \frac{[\text{Cr}^{3+}]^2[\text{Sn}^{4+}]^3}{[\text{Cr}_2\text{O}_7^{2-}][\text{Sn}^{2+}]^3[\text{H}^+]^{14}} \\ &= 1.22\text{V} - \frac{0.0592\text{V}}{6} \log \frac{1^2 1^3}{1^3 (10^{-1})^{14}} = 1.22 - 0.138 = 1.082 \text{ V} \end{aligned}$$

(2) answer: a) $2x + 6(+1) + (-2) = 0 \Rightarrow x = -2$

b) $3x + 3(+1) = +1 \Rightarrow x = -2/3$

(3) answer:



so: oxidation $\frac{1}{2}$ rxn: $\text{I}_2 \rightarrow 2 \text{ICl}_2^- + 2\text{e}^- \Rightarrow \text{I}_2 + 4 \text{Cl}^- \rightarrow 2 \text{ICl}_2^- + 2\text{e}^-$

reduction: $\text{IO}_3^- + 4\text{e}^- \rightleftharpoons \text{ICl}_2^-$

$\Rightarrow \text{IO}_3^- + 4\text{e}^- + 6\text{H}^+ + 2\text{Cl}^- \rightleftharpoons \text{ICl}_2^- + 3\text{H}_2\text{O}$

so $2x(1) + (2)$:



(4) Answer: $E^\circ_{\text{Mg}^{2+}/\text{Mg}} = -2.37 \text{ V}$; $E^\circ_{\text{Co}^{2+}/\text{Co}} = -0.28 \text{ V}$

