Practice problems #10

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

$$Cr_2O_7^{2-} + 3\ Sn^{2+} + 14\ H^+ \rightarrow 2\ Cr^{3+} + 3\ Sn^{4+} + 7\ H_2O$$

(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 $Cr_2O_7^{2-}$ 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: $[Cr_2O_7^{2-}] = [Sn^{2+}] = [Cr^{3+}] = [Sn^{4+}] = 1.00M$ and pH = 1.00; what is E_{cell} =? (assume, T =298K)

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

a) C in ethanol, C_2H_6O, b) C in $C_3H_3^+$

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

$$I_2 + IO_3^- + Cl^- <==> ICl_2^-$$

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

Table:

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

a) If $E°_{Pd2+/Pd} = 0.99$ V, and $E°_{Mg2+/Mg} < 0$ 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).

$E°_{Mg2+/Mg}$ = __________ $E°_{Co2+/Co}$ = __________

Table:

$E°_{reduction}$

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

(1) a) answer: $Cr_2O_7^{2-}$

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

$$Sn^{2+} \rightarrow Sn^{4+} + 2e^-$$

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

d) answer: $$\Delta E = \Delta E° - \frac{0.0592V}{n} \log \frac{[Cr^{3+}]^2[Sn^{4+}]^3}{[Cr_2O_7^{2-}][Sn^{2+}]^3[H^+]^{14}}$$

$$= 1.22V - \frac{0.0592V}{6} \log \frac{1^2 1^3}{1\ 1^3\ (10^{-1})^{14}} = 1.22 - 0.138 = 1.082\ V$$

(2) answer: a) 2x + 6(+1)+(-2) =0 => x =-2 b) 3x +3(+1) =+1 => x = -2/3

(3) answer:

$$\underset{0}{I_2} + \underset{+5\ -2}{IO_3^-} + \underset{-1}{Cl^-} <==> \underset{+1\ -1}{ICl_2^-}$$

so: oxidation ½ rxn: $I_2 \rightarrow 2\ ICl_2^- + 2e^-$ => $I_2 + 4\ Cl^- \rightarrow 2\ ICl_2 + 2e^-$

reduction: $IO_3^- + 4\ e^- <===> ICl_2^-$
=> $IO_3^- + 4\ e^- + 6\ H^+ + 2\ Cl^- <===> ICl_2 + 3\ H_2O$

so 2x(1) + (2):
$2I_2 + 10\ Cl^- + IO_3^- + 6\ H^+ \rightarrow 5\ ICl_2^- + 3\ H_2O$

(4) Answer: $E°_{Mg2+/Mg} = -2.37$ V; $E°_{Co2+/Co} = -0.28$ V

$E°_{reduction}$

Pd^{2+}/Pd (s) 0.99V
| 1.27
$Co^{2+}/Co(s)$ .99-1.27 = -0.28
| 2.09
$Mg^{2+}/Mg(s)$ 0.99-3.36=-2.37

3.36