The following may help you to answer many questions including review question 4.

4. Answer the questions using the following half reactions:

		E^o (V)
	$Cl_2 + 2e^- \rightarrow 2\,Cl^-$	1.36
	$O_2 + 4\,H^+ + 4\,e^- \rightarrow 2\,H_2O$	1.23
Oxidizing agents are	$ClO_2 + e^- \rightarrow ClO_2^-$	0.95
on the left side of the	$Ag^+ + e^- \rightarrow Ag\,(s)$	0.80
table. They increase	$Cu^{2+} + 2e^- \rightarrow Cu\,(s)$	0.34
in strength as E^o	$2\,H^+ + 2\,e^- \rightarrow H_2$	0.00
becomes more positive	$Co^{2+} + 2\,e^- \rightarrow Co\,(s)$	– 0.28
(going up the table).	$Cd^{2+} + 2\,e^- \rightarrow Cd\,(s)$	– 0.40
	$Cr^{3+} + 3\,e^- \rightarrow Cr\,(s)$	– 0.76

Reducing agents are on the right side of the table. They increase in strength as E^o becomes less positive (going down the table).

The way the table is organized makes it convenient to use. All the reactions are written as reduction half reactions and they are listed in order from the most positive reduction potential, $E^o = 1.36$ V, to the most negative reduction potential $E^o = -0.76$ V.

The oxidizing agents are all on the left side of the table.

Cl_2, O_2, ClO_2, Ag^+, Cu^{2+}, H^+, Co^{2+}, Cd^{2+} and Cr^{3+} are all oxidizing agents.

The reducing agents are all on the right side of the table.

Cl^-, H_2O, ClO_2^-, Ag, Cu, H_2, Co, Cd, Cr are all reducing agents.

You have learned that the more negative the ΔG^o for a reaction, the more spontaneous the reaction. The following equation shows the relationship between ΔG^o and E^o.

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

In this equation, n is the number of moles of electrons transferred in the reaction, F is Faraday's constant and E^o is the cell potential. Because of the negative sign in the equation, when $E^o > 0$ then $\Delta G^o < 0$ and the reaction is spontaneous. Thus, E^o must be positive for a reaction to be spontaneous. The more positive E^o, the more negative ΔG^o, and thus the more spontaneous the reaction.

Using the table above and this information, you should be able to figure out the following questions.

Which is the strongest oxidizing agent? The oxidizing agents are all on the left side of the table. The strongest oxidizing agent has the most positive reduction potential. In this table, the reaction, $Cl_2 + 2e^- \rightarrow 2\,Cl^-$ has the most positive reduction potential, +1.36 V. Therefore, Cl_2 is the strongest oxidizing agent.

4. a) Which is the strongest reducing agent?

The reducing agents are all on the right side of the arrow. To answer this question you consider only the right side of each of these reactions. The trick is that the reducing agents increase in strength as you go down the table (the strength increases as E^o becomes less positive). This is reasonable if you think about what a reducing agent does. A reducing agent loses electrons to another substance that is reduced by gaining electrons. So the reducing agent is oxidized (loses electrons) which is the reverse of the reduction reaction shown in the table. That is, a large negative reduction potential, implies a large positive oxidation potential. The strongest reducing agent has the most positive oxidation potential.

Reduction: $Cr^{3+} + 3\ e^- \rightarrow Cr\ (s)$ $\quad E^o = -0.76$ V

Oxidation: $Cr\ (s) \rightarrow Cr^{3+} + 3\ e^-$ $\quad E^o = +0.76$ V

b) Is Cr^{3+} capable of oxidizing Ag (s)?

First you need to figure out how to write the oxidation and reduction half reactions. The question tells you how to write these reactions. The question tells you Ag(s) is oxidized and therefore Cr^{3+} must be reduced.

From the table we have the following reduction half reactions.

$$Ag^+ + e^- \rightarrow Ag\ (s) \qquad E^o = 0.80\ V$$
$$Cr^{3+} + 3\ e^- \rightarrow Cr\ (s) \qquad E^o = -0.76\ V$$

We have to reverse one of these reactions so that we have one oxidation reaction and one reduction reaction. Which one do we reverse? We reverse the silver reaction, because in this question, Ag(s) is oxidized, it loses electrons. **When we reverse a reaction we have to reverse the sign of E^o.**

$$Ag\ (s) \rightarrow Ag^+ + e^- \qquad E^o = -0.80\ V$$
$$Cr^{3+} + 3\ e^- \rightarrow Cr\ (s) \qquad E^o = -0.76\ V$$

What happens to E^o when we balance the electrons lost and gained? That is, when we multiply the silver reaction by a factor of three, does this change the value of E^o?

$$3\ Ag\ (s) \rightarrow 3\ Ag^+ + 3e^- \qquad E^o = \ ?$$
$$Cr^{3+} + 3\ e^- \rightarrow Cr\ (s) \qquad E^o = \ ?$$

Then add the reactions to obtain the overall reaction and add the half reaction potentials to obtain the cell potential. From the sign of the cell potential, E^o, figure out if the reaction is spontaneous or not.