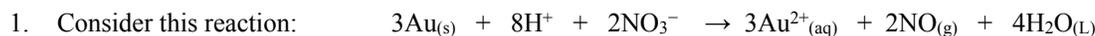


- *This is practice - Do NOT cheat yourself of finding out what you are capable of doing. Be sure you follow the testing conditions outlined below.*
- *DO NOT USE A CALCULATOR. You may use ONLY your periodic table and formula sheets.*
- *Work only for 1.5 minutes per question (51 min total). Time yourself. It is important that you practice working for speed. Then when time is up, continue working, with a calculator if you must, but use a different color writing utensil so you will have a sense of how many you were able to complete under the time constraints and without the calculator.*



Mark each of the following statements (about the reaction above) as true or false.

- T F $\text{Au}_{(s)}$ is reduced during the reaction.
 - T F The oxidation state of nitrogen changes from +6 to +2.
 - T F Hydrogen ions are oxidized when they form $\text{H}_2\text{O}_{(L)}$.
 - T F 8 electrons are lost and 8 electrons are gained during this reaction.
2. Which substance contains the element that is reduced in the following reaction:
- $$\text{Cr}_2\text{O}_7^{2-} + 6\text{S}_2\text{O}_3^{2-} + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 3\text{S}_4\text{O}_6^{2-} + 7\text{H}_2\text{O}$$
- $\text{Cr}_2\text{O}_7^{2-}$
 - $\text{S}_2\text{O}_3^{2-}$
 - H^+
 - Cr^{3+}
 - $\text{S}_4\text{O}_6^{2-}$
3. What is the oxidation number of manganese in the KMnO_4
- +1
 - +2
 - +5
 - +4
 - +7
4. How many electrons appear in the following *half-reaction* when it is balanced.
- $$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$$
- 1
 - 2
 - 3
 - 4
 - 6
5. The more _____ the value of E°_{red} , the greater the driving force for the substance's own oxidation.
- positive
 - negative
 - exothermic
 - endothermic
 - extensive
6. How many electrons appear in the following *half-reaction* when it is balanced.
- $$\text{S}_4\text{O}_6^{2-} \rightarrow 2\text{S}_2\text{O}_3^{2-}$$
- 6
 - 2
 - 4
 - 1
 - 3
7. The *half-reaction* occurring at the anode in the balanced reaction shown below is
- $$3\text{MnO}_4^- + 24\text{H}^+ + 5\text{Fe} \rightarrow 3\text{Mn}^{2+} + 5\text{Fe}^{3+} + 12\text{H}_2\text{O}$$
- $\text{MnO}_4^- + 8\text{H}^+ + 5e^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$
 - $\text{MnO}_4^- + 5e^- \rightarrow \text{Mn}^{2+}$
 - $\text{Fe} \rightarrow \text{Fe}^{3+} + 3e^-$
 - $\text{Fe} + 3e^- \rightarrow \text{Fe}^{3+}$
 - $\text{MnO}_4^- + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} + 5e^-$
8. The oxidation *half-reaction* occurring in the standard hydrogen electrode is
- $\text{H}_2 \rightarrow 2\text{H}^+ + 2e^-$
 - $2\text{H}^+ + 2\text{OH}^- \rightarrow \text{H}_2\text{O}$
 - $\text{O}_2 + 4\text{H}^+ + 4e^- \rightarrow 2\text{H}_2\text{O}$
 - $2\text{H}^+ + 2e^- \rightarrow \text{H}_2$
 - $2\text{H}^+ + \text{Cl}_2 \rightarrow 2\text{HCl}$

9. $6\text{H}^+ + 5\text{H}_2\text{O}_2 + 2\text{MnO}_4^- \rightarrow 5\text{O}_2 + 2\text{Mn}^{2+} + 8\text{H}_2\text{O}$
 According to the balanced equation above, how many moles of the permanganate ion are required to react completely with 25.0 ml of 0.100 M hydrogen peroxide?
- 0.000500 mol
 - 0.00100 mol
 - 0.00500 mol
 - 0.00625 mol
 - 0.0500 mol
10. A 1.0 L sample of an aqueous solution contains 0.10 mol of NaCl and 0.10 mol of CaCl₂. What is the minimum number of moles of AgNO₃ that must be added to the solution in order to precipitate all of the Cl⁻ as AgCl_(s)? (Assume that AgCl is insoluble.)
- 0.10 mol
 - 0.20 mol
 - 0.30 mol
 - 0.40 mol
 - 0.60 mol
11. When 400. milliliter of 0.10-molar sodium chloride is added to 200. milliliters of 0.10-molar aluminum chloride the number of moles of Pb²⁺ that must be added to precipitate out all of the Cl⁻ would be
- 0.020 moles
 - 0.040 moles
 - 0.050 moles
 - 0.060 moles
 - 0.10 moles
12. A 20.0-milliliter sample of 0.200-molar K₂CO₃ solution is added to 30.0 milliliters of 0.300-molar Ba(NO₃)₂ solution. Barium carbonate precipitates. The concentration of barium ion, Ba²⁺, in solution after reaction is
- 0.00500 M
 - 0.00900 M
 - 0.0500 M
 - 0.100 M
 - 0.00250 M
13. How many moles of solid Ba(NO₃)₂ should be added to 300. milliliters of 0.20-molar Fe(NO₃)₃ to increase the concentration of the NO₃⁻ ion to 1.0-molar? (Assume that the volume of the solution remains constant.)
- 0.060 mole
 - 0.12 mole
 - 0.24 mole
 - 0.30 mole
 - 0.40 mole
14. Given the following half reactions:
- $$\text{Sn}^{4+} + 2\text{e}^- \rightarrow \text{Sn}^{2+} \quad E^\circ = 0.15 \text{ V}$$
- $$\text{Fe}^{3+} + 1\text{e}^- \rightarrow \text{Fe}^{2+} \quad E^\circ = 0.77 \text{ V}$$
- Determine the standard cell potential (E°_{cell}) for the voltaic cell based on the reaction
- $$\text{Sn}^{2+} + 2\text{Fe}^{3+} \rightarrow 2\text{Fe}^{2+} + \text{Sn}^{4+}$$
- +0.47 V
 - +0.62 V
 - +0.92 V
 - +1.39 V
 - +1.69 V
15. When this reaction is balanced, the coefficient on the Sn²⁺ is.
- $$\text{Sn}^{4+} + \text{Cr} \rightarrow \text{Cr}^{3+} + \text{Sn}^{2+}$$
- 1
 - 2
 - 3
 - 4
 - 6
16. Identify the redox reactions below.
- $\text{K}_2\text{CrO}_4 + \text{BaCl}_2 \rightarrow \text{BaCrO}_4 + 2\text{KCl}$
 - $\text{Pb}^{2+} + 2\text{Br}^- \rightarrow \text{PbBr}_2$
 - $\text{Cu} + \text{S} \rightarrow \text{CuS}$
- III only
 - II only
 - I only
 - I and III
 - II and III

17. Which one of the following reactions is a redox reaction?

- $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
- $\text{Pb}^{2+} + 2\text{Cl}^- \rightarrow \text{PbCl}_2$
- $\text{AgNO}_3 + \text{HCl} \rightarrow \text{HNO}_3 + \text{AgCl}$
- $2\text{HC}_2\text{H}_3\text{O}_2 + \text{Ca}(\text{OH})_2 \rightarrow 2\text{H}_2\text{O} + \text{Ca}(\text{C}_2\text{H}_3\text{O}_2)_2$
- None of the above are redox reactions.

18. Which substance is the reducing agent (the substance that caused reduction i.e. the substance that is oxidized) in the following reaction?



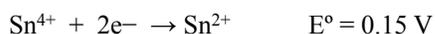
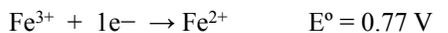
- HNO_3
- S
- NO_2
- Fe_2S_3
- H_2O

19. What is the coefficient of the bromide ion when the following equation is balanced?



- 1
- 2
- 4
- 5
- 10

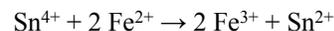
20. Consider the half reactions shown below to answer the following question:



Which of the following reactions will occur spontaneously as written?

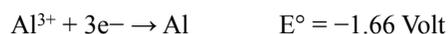
- $\text{Sn}^{4+} + \text{Fe}^{3+} \rightarrow \text{Sn}^{2+} + \text{Fe}^{2+}$
- $3\text{Fe} + 2\text{Cr}^{3+} \rightarrow 2\text{Cr} + 3\text{Fe}^{2+}$
- $\text{Sn}^{2+} + \text{Fe} \rightarrow \text{Sn}^{4+} + \text{Fe}^{2+}$
- $3\text{Sn}^{4+} + 2\text{Cr} \rightarrow 2\text{Cr}^{3+} + 3\text{Sn}^{2+}$
- $3\text{Fe}^{2+} \rightarrow \text{Fe} + 2\text{Fe}^{3+}$

21. In the oxidation-reduction reaction

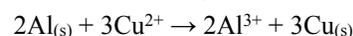


- Sn^{4+} is reduced and Fe^{2+} is oxidized.
- Sn^{4+} is oxidized and Fe^{2+} is reduced.
- Sn^{4+} is oxidized and Fe^{3+} is reduced.
- Fe^{3+} is reduced and Sn^{2+} is oxidized.
- Fe^{2+} is oxidized and Fe^{3+} is reduced.

22. Given the standard reduction potentials



Calculate the standard voltage for the reaction



- 1.22 Volts
- +1.22 Volts
- +2.00 Volts
- +2.30 Volts
- +4.34 Volts

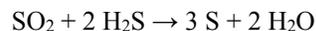
23. Given the standard electrode (reduction) potentials:



What would be the E° for a cadmium-silver cell?

- 0.0 V
- 0.4 V
- 0.5 V
- 1.2 V
- 2.0 V

24. In the reaction



- sulfur is oxidized and hydrogen is reduced
- sulfur is reduced and there is no oxidation
- sulfur is oxidized and there is no reduction
- sulfur is reduced and hydrogen is oxidized
- sulfur is both reduced and oxidized

25. Which ion, in solution, can be oxidized by chemical means but also can be reduced by a different chemical reaction?

- Fe^{2+}
- F^-
- CO_3^{2-}
- NO_3^-
- Al^{3+}

26. Zinc reacts with dilute acid to produce H_2 and Zn^{2+} but silver does not liberate hydrogen gas from an acid. This information enables one to predict that which of the following reactions has a positive emf ?

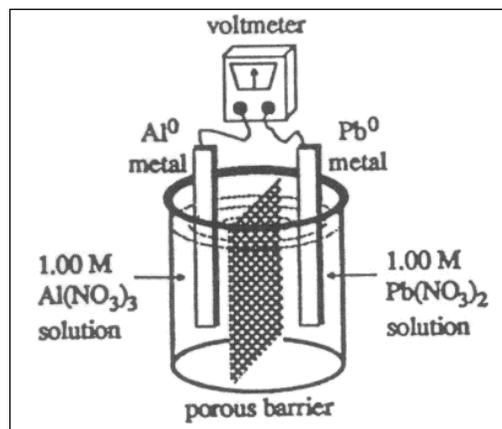
- $\text{H}_{2(\text{g})} + \text{Zn}^{2+} \rightarrow 2\text{H}^+ + \text{Zn}$
- $2\text{Ag} + \text{Zn}^{2+} \rightarrow 2\text{Ag}^+ + \text{Zn}$
- $2\text{Ag}^+ + \text{Zn} \rightarrow 2\text{Ag} + \text{Zn}^{2+}$
- $2\text{Ag} + 2\text{H}^+ \rightarrow \text{H}_{2(\text{g})} + 2\text{Ag}^+$
- $2\text{Ag}^+ + \text{Zn}^{2+} \rightarrow 2\text{Ag} + \text{Zn}$

27. If solid nickel metal were added to separate aqueous solutions each containing 1M concentrations of Ag^+ , Cd^{2+} , Sn^{2+} , and Ni^{2+} ions, which metals would plate out, based on the given standard reaction potentials?

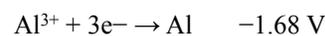
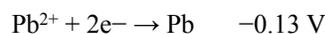
Standard Reduction Potentials	
Ag^+/Ag	0.799 V
Sn^{2+}/Sn	-0.141 V
Ni^{2+}/Ni	-0.236 V
Cd^{2+}/Cd	-0.400 V

- Ag
 - Sn
 - Ni
 - Cd
- I only
 - IV only
 - I and II only
 - III and IV only
 - I, II, and III

28. What voltage will be produced by the electrochemical cell?



Reduction Potentials



- +4.78 V
- +1.55 V
- +1.81 V
- +2.97 V
- +1.29 V

29. Which statements below are true about the voltaic cell in the previous problem.

- electrons move right to left through the external circuit
 - the lead electrode will gain mass
 - the porous barrier maintains the charge balance in both half-cells
- I only
 - II only
 - I and II only
 - II and III only
 - I, II, and III

30. Which of these ions is the best at being oxidized?

Standard Reduction Potentials, E°	
$\text{Fe}^{3+} + e^- \rightarrow \text{Fe}^{2+}$	+0.77 V
$\text{Cu}^{2+} + e^- \rightarrow \text{Cu}^+$	+0.15 V

- Fe^{3+}
- Fe^{2+}
- Cu^{2+}
- Cu^+
- None of them can act as reducing agents

31. Use these reduction potentials to determine which one of the reactions below is spontaneous.

Reaction	Reduction Potentials, E°
$\text{Ag}^+ + e^- \rightarrow \text{Ag}$	0.800 V
$\text{Pb}^{2+} + 2e^- \rightarrow \text{Pb}$	- 0.126 V
$\text{V}^{2+} + 2e^- \rightarrow \text{V}$	- 1.18 V

- $\text{V}^{2+} + 2 \text{Ag} \rightarrow \text{V} + 2 \text{Ag}^+$
 - $\text{V} + \text{Pb}^{2+} \rightarrow \text{V}^{2+} + \text{Pb}$
 - $2 \text{Ag}^+ + \text{Pb}^{2+} \rightarrow 2 \text{Ag} + \text{Pb}$
- II only
 - III only
 - I and II only
 - II and III only
 - I, II, and III

32. For this reaction, $E^\circ_{\text{cell}} = 0.79 \text{ V}$.



Given that the standard reduction potential for $\text{Cr}_2\text{O}_7^{2-} \rightarrow 2\text{Cr}^{3+}$ is 1.33 V, what is E°_{red} for $\text{I}_2(\text{aq})$?

- +0.54 V
- 0.54 V
- +0.18 V
- 0.18 V
- +2.12 V

33. All of the following may get oxidized EXCEPT

- NO_2^-
- Na
- Fe^{2+}
- MnO_4^-
- Br^-

34. For this reaction,



Choose the true statement from the following list.

- The Cl oxidation state is reduced from +8 to +6
- This is not an oxidation-reduction reaction.
- H_2O_2 caused reduction.
- Hydrogen is reduced from +2 to +1
- Potassium is reduced during the reaction

1. true or false?

 - a. F Gold is oxidized from 0 to 2+.
 - b. F Nitrogen is reduced from +5 to +2.
 - c. F Hydrogen's oxidation state does not change, thus it is neither oxidized nor reduced.
 - d. F Actually it is 6 electrons transferred during this reaction.
2. a LEO says GER or OIL RIG The element that is reduced in this reaction is chromium, and the dichromate ion contains this element.
3. e $(+1) + (\text{Mn}) + (4 \times -2) = 0$, thus the Mn must be +7
4. e You can determine this by oxidation numbers, Cr is +6 on left and changes to +3 on the right, but twice thus $6e^-$ transferred OR you can do it by balancing the charge with the correct number of electrons. As written, the charge on left is $(2^-) + (14^+) \neq (6^+)$, thus you need $6e^-$ more on the left, you can get that with 6 electrons.
5. b If a substance has a negative $E^\circ_{\text{reduction}}$, then it will have a positive, and thus favorable, $E^\circ_{\text{oxidation}}$.
6. b The easiest way to balance these electrons is by balancing the charge. $2^- \neq (2 \times 2^-)$, thus adding $2e^-$ to the left side will balance the charge. It is a bit trickier to try to use oxidation numbers since the S on the left has an oxidation number of +2.5, and the oxidation number of +2. This is a change of oxidation state of $+1/2$, four times, thus 2 electrons!
7. c Remember "An Ox" and thus we are trying to pull out the oxidation half reaction. Looking at the oxidation numbers, you can see that iron moves from an oxidation state of 0 to +3, which is oxidation.
8. a The standard hydrogen electrode has either oxidation or reduction occurring, depending on what other half reaction it is paired with. $\text{H}_2 \rightarrow 2\text{H}^+ + 2e^-$ OR $2\text{H}^+ + 2e^- \rightarrow \text{H}_2$, and the first one listed is the oxidation.
9. b Change to millimoles: $25 \text{ ml} \times 0.1 \text{ M} = 2.5 \text{ mmol}$, then "stoich" over to permanganate. $2.5 \text{ mmol} \times \frac{2\text{MnO}_4^-}{5\text{H}_2\text{O}_2} = 1 \text{ mmol}$ and of course, 1 mmol is 0.001 mol
10. c This is one of those problems that you need to combine the Cl^- ions together since the precipitate reaction is $\text{Ag}^+ + \text{Cl}^- \rightarrow \text{AgCl}$. Calculate mole of NaCl: $1 \text{ L} \times 0.1 \text{ M} = 0.1 \text{ mol}$ of NaCl and NaCl gives us only one Cl^- per NaCl. But for the CaCl_2 , a "buy 1 get 2" substance, for 0.1 mol of CaCl_2 we get 0.2 mole of Cl^- . Thus we have a total of 0.3 mol of Cl^- which will require 0.3 mol of Ag^+ to precipitate out all of the Cl^- from solution.
11. c This problem is very similar to the last, we need to combine the total moles of chloride ions, and then consider how many lead ions would be needed to precipitate out the chloride ions.
Write a balanced net ionic equation. $\text{Pb}^{2+} + 2 \text{Cl}^- \rightarrow \text{PbCl}_2$
Then calculate the total millimol of Cl^- . For NaCl: $400 \text{ ml} \times 0.1 \text{ M} = 40 \text{ mmol}$ NaCl which is 40 mmol of Cl^-
For AlCl_3 : $200 \text{ ml} \times 0.1 \text{ M} = 20 \text{ mmol}$ AlCl_3 which is 60 mmol of Cl^-
Add the total Cl^- together to get 100 mmol of Cl^- which when your "stoich" over to the lead ions required, you see that you only need 50 mmol of Pb^{2+} which is of course 0.05 mol of Pb^{2+} .
12. d For this problem, you must realize that barium ions and carbonate ions combine in a 1 to 1 ratio as shown by this net ionic precipitation reaction: $\text{Ba}^{2+} + \text{CO}_3^{2-} \rightarrow \text{BaCO}_3$. The question is indirectly telling you that the carbonate ion limits, because they tell you that barium ions are left in solution. So lets calculate the amount of moles of each ion. For carbonate: $20 \text{ ml} \times 0.2 \text{ M} = 4 \text{ mmol}$ CO_3^{2-} and for barium ion: $30 \text{ ml} \times 0.3 \text{ M} = 9 \text{ mmol}$ Ba^{2+} . Since 4 mmol of CO_3^{2-} requires only 4 mmol of Ba^{2+} to go with it, thus there must be 5 mmol of Ba^{2+} left over, swimming in a TOTAL volume of 50 ml, thus 0.1 M for left over Ba^{2+} .
13. a This may be the trickiest problem on the practice test, so buckle your seat belt. This problem is similar to #'s 10 and 11 because of the need to add together the nitrate ions. We want to make a solution that will be 1 M for nitrate ions. Since it is 300 ml of solution, we will need 300 mmole of NO_3^- ions to achieve a 1 M solution.
Calculate the NO_3^- ions from the $\text{Fe}(\text{NO}_3)_3$ solution: $300 \text{ ml} \times 0.2 \text{ M} = 60 \text{ mmol}$ $\text{Fe}(\text{NO}_3)_3$, but this is a "buy 1 get 3" substance, thus we actually have 180 mmol of NO_3^- ions. To make up our 300 mmol of NO_3^- ions, we will need 120 mmol more of NO_3^- ions. But wait, before you choose (b) there is one more twist. The source of the nitrate ions you are dumping into this solution is $\text{Ba}(\text{NO}_3)_2$ and you can see that $\text{Ba}(\text{NO}_3)_2$ is a "buy 1 get 2" substance, so you only need 60 mmol of $\text{Ba}(\text{NO}_3)_2$ to provide the 120 mmol of NO_3^- ions required, and of course 60 mmol is 0.06 mol of $\text{Ba}(\text{NO}_3)_2$.
Whew!
14. b Flip the reduction reaction and the add $E^\circ_{\text{oxidation}} + E^\circ_{\text{reduction}} = E^\circ_{\text{cell}}$. $(-0.15 \text{ V}) + (+0.77 \text{ V}) = +0.62 \text{ V}$
15. c You can balance this by inspection by balancing the charge.
16. a The first reaction listed is an overall precipitation reaction, the second reaction listed is net ionic precipitation reaction, and precipitation reactions are never redox reactions.
17. e (a) and (d) are acid base reactions and (b) and (c) are precipitation. None of these are redox reactions.
18. d Since we are looking for the reducing agent, remember that this means we are looking for the substance that contains the element that is being oxidized. The sulfur changes from -2 on the left to 0 on the right, thus the sulfur is oxidized and the Fe_2S_3 is the reducing agent.

19. e Mn changes from +7 to +2, and Br changes from -1 to 0, twice for $2e^-$, $2 Br^- \rightarrow Br_2$ thus we need to multiply the bromine half reaction by 5 (and the manganese half reaction by 2 to balance the electrons transferred. and $2 \times (2 Br^- \rightarrow Br_2)$ will give us a coefficient of 10 on the bromide ion.
20. d Since we want to choose a spontaneous reaction, we must produce a + E°_{cell} . In (a) two reductions are occurring – not possible, and in (c) two oxidations are occurring – not possible. In (b) when you flip the sign of the $E^\circ_{reduction}$ for Fe^{2+} , you will end up with a negative E°_{cell} . In (d) when you reverse the Cr^{3+} to the oxidation the voltage is $(+0.15) + (+0.74) = +E^\circ_{cell}$
21. a LEO says GER or OIL RIG
22. c Reverse the Al^{3+} reduction reaction and change the sign. $E^\circ_{oxidation} + E^\circ_{reduction} = E^\circ_{cell}$. $(+1.66 V) + (+0.34 V) = +2 V$
23. d If we are making a cell, we would want the voltage to be positive. Remember you need a reduction half reaction AND an oxidation half reaction. Flip the half reaction that will allow you to end up with a positive voltage. Thus flip the Cadmium reduction half-reaction, and $E^\circ_{oxidation} + E^\circ_{reduction} = E^\circ_{cell}$. $(+0.40 V) + (+0.80 V) = +1.2 V$
24. e Sulfur is +4 in SO_2 and -2 in H_2S and it gets both reduced and oxidized to the oxidation state of 0 in S on the product side.
25. a We are looking for some particle that can increase and decrease its oxidation number. The choice is Fe^{2+} which can be oxidized to Fe^{3+} and reduced to Fe.
26. c The information in the question tell us that Zn can get oxidized by acid, and is “active” but the Ag is not oxidized by acid and thus is “less active.” We are then looking for a reaction which would be spontaneous. This occurs when we allow the more active Zn atoms oxidize to ions, and the less active silver atoms get reduced from Ag^+ to Ag
27. c “Plate-out” is the process of metal sticking to the cathode, meaning which elements listed can reduce and “plate out” while the nickel is getting oxidized. Since the oxidation of nickel has a voltage of +0.236 V, (opposite sign of the reduction listed) only reductions that have a voltage of greater than -0.236 will result in a positive voltage, and thus occur.
28. b Reversing the aluminum half reaction will result in a positive voltage for the cell.
 $E^\circ_{oxidation} + E^\circ_{reduction} = E^\circ_{cell}$. $(+1.68 V) + (-0.13 V) = +1.55 V$
29. d Since the lead is being reduced in the previous problem, the electrons must be flowing toward the lead electrode, thus left to right. Since the lead is being reduced, lead will “plate” onto the lead electrode, and it will increase in mass. A porous barrier is the same as a salt bridge.
30. d Remember, a reducing agent is a substance that gets oxidized, thus in this problem, only the Cu^+ and Fe^{2+} can be oxidized. While both oxidations will be negative, the Cu^+ is less negative and therefore better at being oxidized than the Fe^{2+}
31. a The first one listed results in a negative voltage $E^\circ_{oxidation} + E^\circ_{reduction} = E^\circ_{cell}$. $(-0.8 V) + (-1.18 V) = \text{negative}$
The second one listed results in a positive voltage $E^\circ_{oxidation} + E^\circ_{reduction} = E^\circ_{cell}$. $(+1.18 V) + (-0.126 V) = \text{positive}$
and the third one listed is two reductions – not possible
32. a $E^\circ_{oxidation} + E^\circ_{reduction} = E^\circ_{cell}$. $(x) + (1.33 V) = 0.79 V$ Then solve for the x which is $E^\circ_{oxidation}$, but there is a twist, the question asks for the $E^\circ_{reduction}$ of I_2 , so we need to flip the sign of the $E^\circ_{oxidation}$
33. d When looking for a reducing agent, you need a substance with an oxidation number that can get more positive. This is not possible for the Mn in MnO_4^- because +7 is the highest possible oxidation state.
34. c The oxygen in H_2O_2 has an oxidation state of -1 and is oxidized to 0 and is thus a reducing agent. We know the oxygen must be oxidized, not reduced to -2 in water because the Cl is reduced from +7 to +5. Some of the oxygens are -2 on the left in the potassium perchlorate, and they remain -2 on the right in water.