

Recitation Worksheet 13: Electrochemistry

Name:

Key

MyID:

Instructions:

1. Please enter your first and last name as it appears on the eLC classlist (do not use a nickname).
2. Your UGA myID is a combination of letters and numbers (example: Dr. Abdelrahman MyID is ema88805).
Do not use your 81x number.
 - A. If you do not have access to a printer, type your answers in the worksheet PDF and then upload it to **Gradescope** by Friday, April 28th at 11:59 pm. Write your work on separate sheets of paper, convert to a PDF and upload to the "Recitation Worksheet 13 Dropbox" on eLC.
 - B. If you are using an app to annotate the worksheet, make sure the pages are in the correct order and have the same layout as the original or Gradescope will not be able to read it.
 - C. If you have access to a printer, print out the worksheet, write your answer in the answer boxes, and show your work on it when appropriate. Then convert it to a PDF and upload to **Gradescope** by Friday, April 28th at 11:59 pm. You do not need to upload anything to eLC. The pages must be in the correct order and have the same layout as the original, or Gradescope will not be able to read it.
 - D. There is a **Gradescope App** available for both iOS and Android devices that allows you to scan and submit your printed work or you can submit your fillable PDF directly. Detailed instructions on how to access and use the app can be found on your CHEM 1212 class eLC page under content → Welcome module → Gradescope → Gradescope new mobile app.
3. Answers must be written in the corresponding answer box, or no credit will be awarded.
4. The instructions for uploading worksheets to Gradescope can be found in the Content area of eLC in the Welcome Module.

1. Which assignment of oxidation number is **INCORRECT** for the underlined element? Select all that apply. Insert letters without spaces in the answer box, example **ABCD**.

BE

- A. $K_2Cr_2O_7$; +6
- ☒ B. NH_3 ; +3
- ☒ C. $H_2PO_2^-$; +1
- D. SeO_3^{2-} ; +4
- ☒ E. $Cu(NO_3)_2$; +2

(A) $K = 1+$, $Cr = ?$, and $O = 2-$ ($K_2Cr_2O_7$)
 $2(1+) + 2Cr + 7(2-) = 0$
 $\therefore 2Cr + (12-) = 0$
 $2Cr = 12 \therefore Cr = 6+$

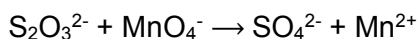
(B) $N = ?$, $H = 1+$ (NH_3)
 $N + (3 \times 1+) = 0$
 $N = 3-$

(C) $H = 1+$, $P = ?$, and $O = 2-$ ($H_2PO_2^-$)
 $(2 \times 1+) + P + (2 \times 2-) = 1-$
 $(2) + P + (-4)$
 $P + (-2) = -1$
 $P = +1$

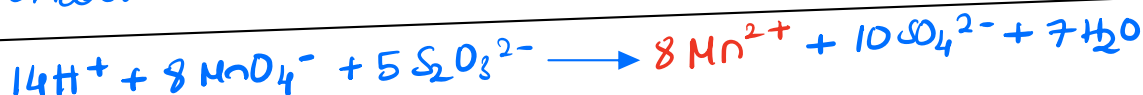
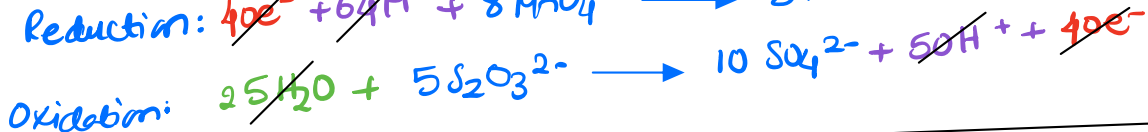
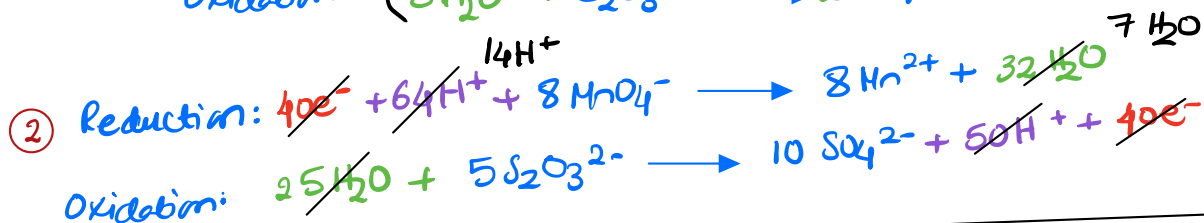
(D) $Se = ?$, $O = 2-$ (SeO_3^{2-})
 $Se + (3 \times 2-) = -2$
 $Se + (-6) = -2$
 $Se = +4$

(E) $Cu = 2+$, $N = ?$, $O = 2-$ ($Cu(NO_3)_2$)
 $(2+) + 2N + 6(2-) = 0$
 $2N + (10-) = 0$
 $2N = 10 \therefore N = +5$

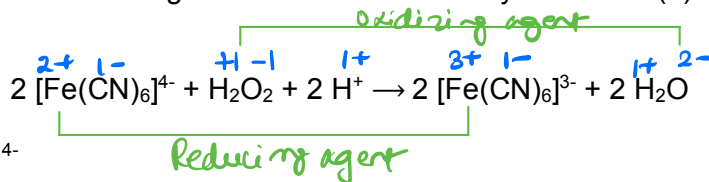
2. What is the coefficient of Mn^{2+} if the reaction below occurs in **acidic solution**?



- C**
- A. 5
B. 7
C. 8
D. 10
E. 14



3. What is the **reducing agent** in the following reaction between hexacyanoferrate (II) complex and hydrogen peroxide in acidic solution?



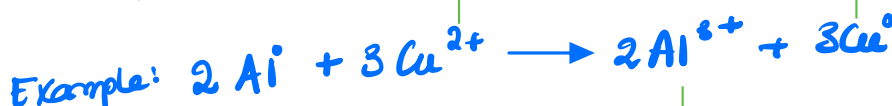
- A**
- A. $[\text{Fe}(\text{CN})_6]^{4-}$
B. H_2O_2
C. H^+
D. $[\text{Fe}(\text{CN})_6]^{3-}$
E. H_2O

Strategy: track the oxidation state of each element & the reducing agent will contain the element that its oxidation state increases

4. Which **statement is true** about redox reactions?

- D**
- A. A half-reaction can occur by itself.
B. A redox reaction in base can include excess H^+ after it has been balanced.
C. Two oxidations can occur instead of one oxidation and one reduction.
D. At least 2 atoms must have their oxidation states change during a redox reaction.
E. None of these statements are true.

- A) the oxidation & reduction half-reactions occur by themselves
- B) the excess H^+ will be balanced by adding OH^- on both sides of the equation
- C) should be one oxidation & one reduction



In this example the oxidation states of both Al & Cu change

5. Which of the following equations are an oxidation-reduction reaction? Select all that apply. Insert letters without spaces in the answer box, example **ABCD**.

BCD

- A. $\text{H}_2\text{CO}_3(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
 B. $2 \text{Li}(\text{s}) + 2 \text{H}_2\text{O}(\text{l}) \rightarrow 2 \text{LiOH}(\text{aq}) + \text{H}_2(\text{g})$
 C. $\text{C}_3\text{H}_8(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 3 \text{CO}_2(\text{g}) + 5 \text{H}_2\text{O}(\text{g})$
 D. $4 \text{Ag}(\text{s}) + \text{PtCl}_4(\text{aq}) \rightarrow 4 \text{AgCl}(\text{s}) + \text{Pt}(\text{s})$
 E. $2 \text{HClO}_4(\text{aq}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow 2 \text{H}_2\text{O}(\text{l}) + \text{Ca}(\text{ClO}_4)_2(\text{aq})$
 F. $3 \text{Cu}(\text{NO}_3)_2(\text{aq}) + 2 \text{Na}_3\text{PO}_4(\text{aq}) \rightarrow 6 \text{NaNO}_3 + \text{Cu}_3(\text{PO}_4)_2$

Types of Redox Reactions

① Single displacement

② Combustion

6. For the reaction:

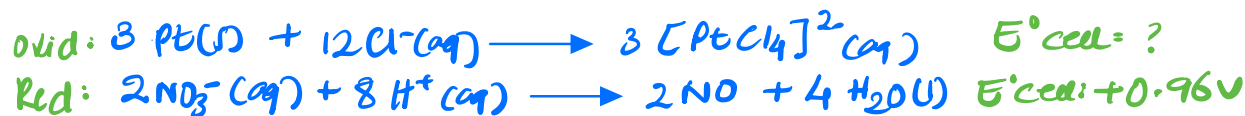


What is E° for the reduction half-cell reaction of $[\text{PtCl}_4]^{2-}$ to Pt in acidic solution? Please refer to the table below for additional information. Keep your answers to two significant figures.

0.76

V

Half-reaction	E° (V)	Half-reaction	E° (V)
$\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$	2.87	$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$	0.40
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	1.99	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	0.34
$\text{Co}^{3+} + \text{e}^- \rightarrow \text{Co}^{2+}$	1.82	$\text{Hg}_2\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Hg} + 2\text{Cl}^-$	0.27
$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$	1.78	$\text{AgCl} + \text{e}^- \rightarrow \text{Ag} + \text{Cl}^-$	0.22
$\text{Ce}^{4+} + \text{e}^- \rightarrow \text{Ce}^{3+}$	1.70	$\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{SO}_3 + \text{H}_2\text{O}$	0.20
$\text{PbO}_2 + 4\text{H}^+ + \text{SO}_4^{2-} + 2\text{e}^- \rightarrow \text{PbSO}_4 + 2\text{H}_2\text{O}$	1.69	$\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$	0.16
$\text{MnO}_4^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{MnO}_2 + 2\text{H}_2\text{O}$	1.68	$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	0.00
$\text{IO}_4^- + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{IO}_3^- + \text{H}_2\text{O}$	1.60	$\text{Fe}^{3+} + 3\text{e}^- \rightarrow \text{Fe}$	-0.036
$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	1.51	$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	-0.13
$\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}$	1.50	$\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$	-0.14
$\text{PbO}_2 + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{Pb}^{2+} + 2\text{H}_2\text{O}$	1.46	$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.23
$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$	1.36	$\text{PbSO}_4 + 2\text{e}^- \rightarrow \text{Pb} + \text{SO}_4^{2-}$	-0.35
$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$	1.33	$\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}$	-0.40
$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$	1.23	$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	-0.44
$\text{MnO}_2 + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O}$	1.21	$\text{Cr}^{3+} + \text{e}^- \rightarrow \text{Cr}^{2+}$	-0.50
$\text{IO}_3^- + 6\text{H}^+ + 5\text{e}^- \rightarrow \frac{1}{2}\text{I}_2 + 3\text{H}_2\text{O}$	1.20	$\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr}$	-0.73
$\text{Br}_2 + 2\text{e}^- \rightarrow 2\text{Br}^-$	1.09	$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.76
$\text{VO}_2^+ + 2\text{H}^+ + \text{e}^- \rightarrow \text{VO}^{2+} + \text{H}_2\text{O}$	1.00	$2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$	-0.83
$\text{AuCl}_4^- + 3\text{e}^- \rightarrow \text{Au} + 4\text{Cl}^-$	0.99	$\text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}$	-1.18
$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{NO} + 2\text{H}_2\text{O}$	0.96	$\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$	-1.66
$\text{ClO}_2 + \text{e}^- \rightarrow \text{ClO}_2^-$	0.954	$\text{H}_2 + 2\text{e}^- \rightarrow 2\text{H}^-$	-2.23
$2\text{Hg}^{2+} + 2\text{e}^- \rightarrow \text{Hg}_2^{2+}$	0.91	$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	-2.37
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	0.80	$\text{La}^{3+} + 3\text{e}^- \rightarrow \text{La}$	-2.37
$\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow 2\text{Hg}$	0.80	$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$	-2.71
$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	0.77	$\text{Ca}^{2+} + 2\text{e}^- \rightarrow \text{Ca}$	-2.76
$\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{O}_2$	0.68	$\text{Ba}^{2+} + 2\text{e}^- \rightarrow \text{Ba}$	-2.90
$\text{MnO}_4^- + \text{e}^- \rightarrow \text{MnO}_4^{2-}$	0.56	$\text{K}^+ + \text{e}^- \rightarrow \text{K}$	-2.92
$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$	0.54	$\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$	-3.05
$\text{Cu}^+ + \text{e}^- \rightarrow \text{Cu}$	0.52		



$$E^\circ_{\text{cell}} = E^\circ_{\text{red}}(\text{Cathode}) - E^\circ_{\text{red}}(\text{Anode})$$

$$0.201 \text{ V} = 0.96 \text{ V} - E^\circ_{\text{red}}(\text{Anode})$$

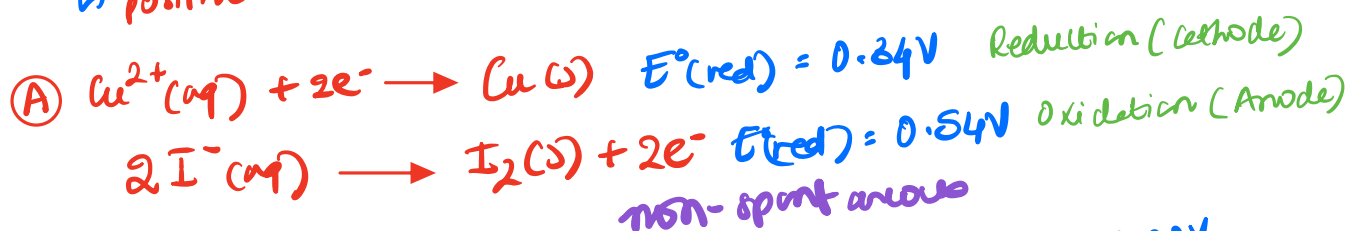
$$\therefore E_{\text{anode}} = 0.759 \sim 0.76 \text{ V}$$

7. Which of the reactions do you predict to be **non-spontaneous** in the forward direction? Assume all the reactants and products in their standard states. Please refer to the table in question 6 for additional information.

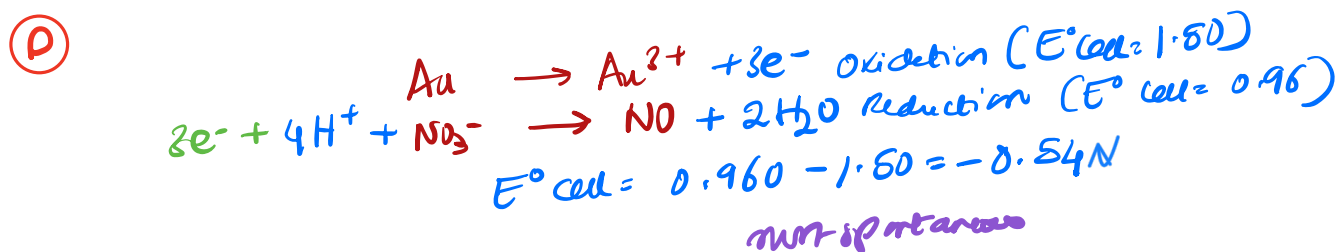
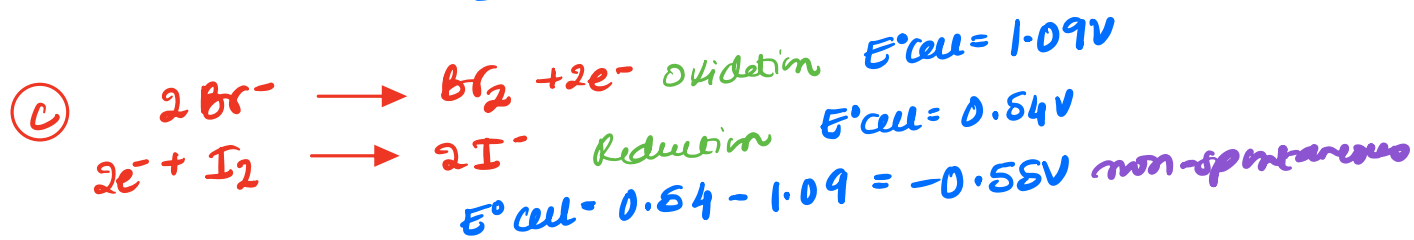
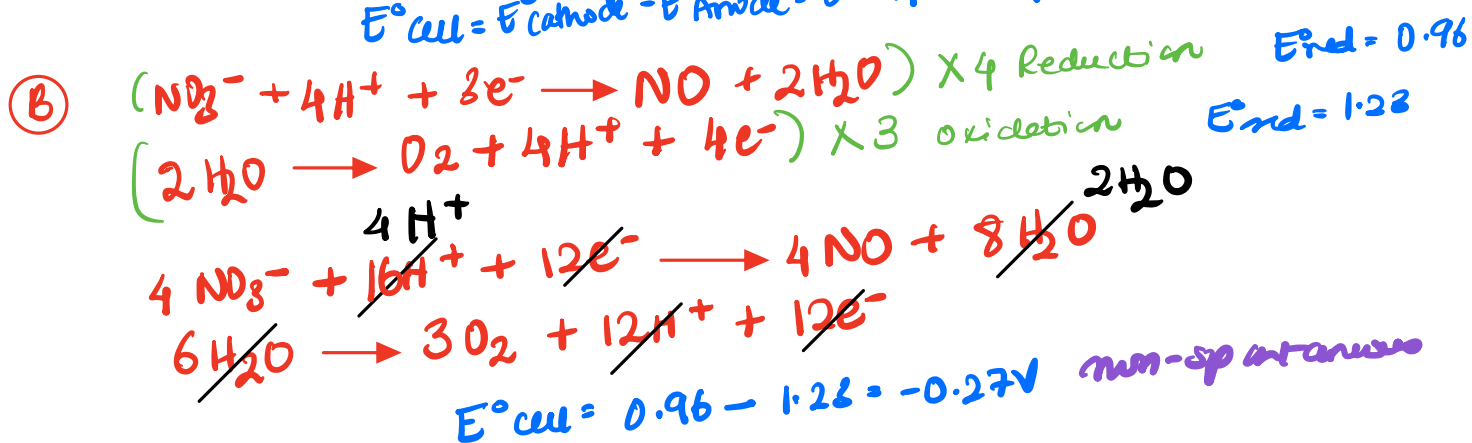
E

- A. $\text{Cu}^{2+}(\text{aq}) + 2\text{I}^{-}(\text{aq}) \rightarrow \text{Cu}(\text{s}) + \text{I}_2(\text{s})$
 B. $4\text{NO}_3^{-}(\text{aq}) + 4\text{H}^{+}(\text{aq}) \rightarrow 3\text{O}_2(\text{g}) + 4\text{NO}(\text{g}) + 2\text{H}_2\text{O}(\text{l})$
 C. $2\text{Br}^{-}(\text{aq}) + \text{I}_2(\text{s}) \rightarrow \text{Br}_2(\text{aq}) + 2\text{I}^{-}(\text{aq})$
 D. $\text{Au}(\text{s}) + \text{NO}_3^{-}(\text{aq}) \rightarrow \text{Au}^{3+}(\text{aq}) + \text{NO}(\text{g})$ (in acidic solution)
 E. All of the above reactions are non-spontaneous in the forward direction.

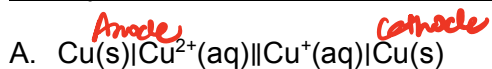
Strategy: Determine the E°_{cell} for each reaction. If the E°_{cell} for the reaction is positive \therefore the reaction is spontaneous



$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} = 0.34 - 0.54 = -0.20\text{V}$$



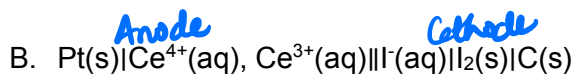
8. Calculate the E°_{cell} for the following reactions. Please refer to the table in question 6 for additional information. Keep your answers to two decimal places.



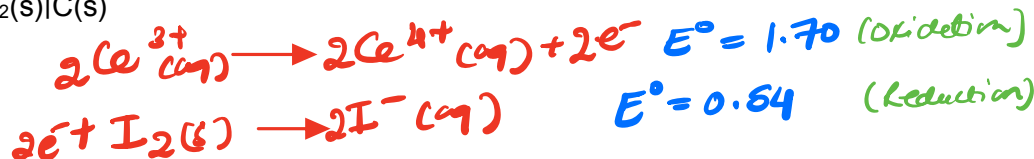
0.18 V

(Anode = Oxidation)
(Cathode = Reduction)

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = 0.52 - 0.34 = 0.18 \text{ V}$$



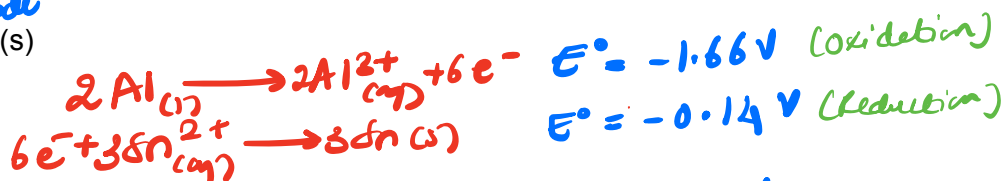
-1.16 V



$$E^\circ_{\text{cell}} = 0.54 - 1.70 = -1.16 \text{ V}$$



1.52 V



$$E^\circ_{\text{cell}} = -0.14 - (-1.66) = 1.52 \text{ V}$$

9. Which statements below is **FALSE** regarding standard cell potentials? Select all that apply. Insert letters without spaces in the answer box, example **ABCD**.

BCE

A. E°_{cell} is positive for spontaneous reactions. True (E°_{cell} is negative for non-spontaneous reactions).

B. Electrons will flow from the positive electrode to the negative electrode in a galvanic cell. False (from negative to positive)

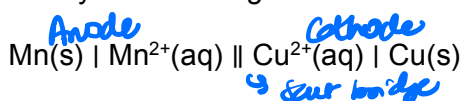
C. E°_{cell} is the difference in voltage between the anode and the cathode, $E^\circ_{\text{cell}} = E^\circ_{\text{cell}}(\text{anode}) - E^\circ_{\text{cell}}(\text{cathode})$. False. $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$

D. The electrode potential of the standard hydrogen electrode is exactly zero. True

E. The electrode in any half-cell with a greater tendency to undergo reduction is negatively charged relative to the standard hydrogen electrode and therefore has a negative E° . False

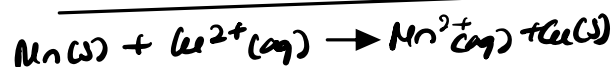
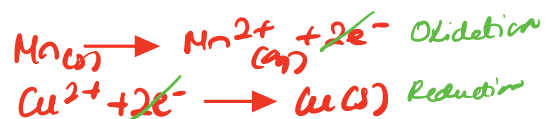
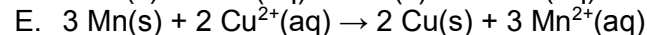
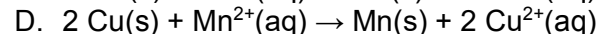
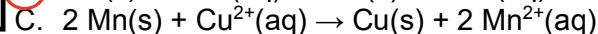
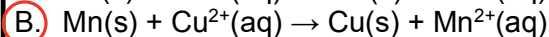
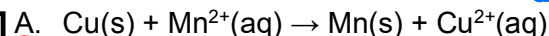
↳ positively charged

10. What is the redox reaction represented by the following cell notation?



Anode = Oxidation
Cathode = Reduction

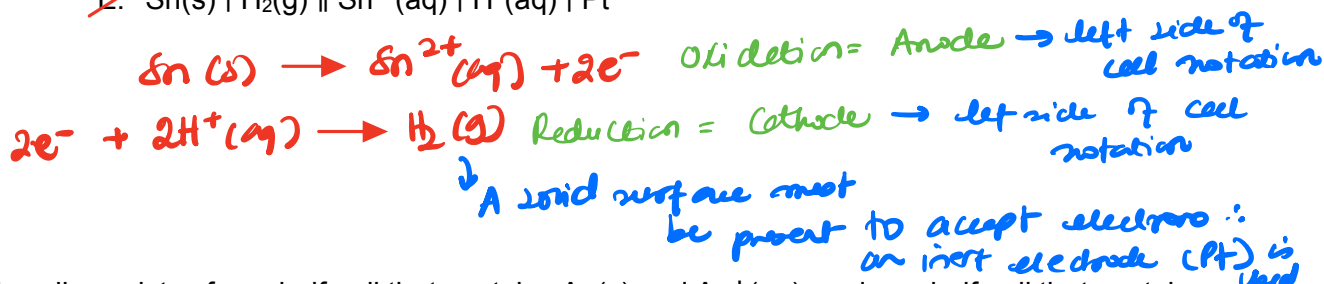
B



11. Which of the choices represents the correct cell notation for $\text{Sn(s)} + 2 \text{H}^+(\text{aq}) \rightarrow \text{Sn}^{2+}(\text{aq}) + \text{H}_2(\text{g})$?

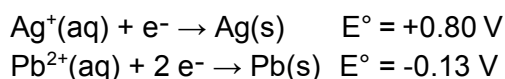
D

- A. $\text{H}^+(\text{aq}) \mid \text{H}_2(\text{g}) \mid \text{Pt} \parallel \text{Sn(s)} \mid \text{Sn}^{2+}(\text{aq})$
- B. $\text{H}_2(\text{g}) \mid \text{H}^+(\text{aq}) \mid \text{Pt} \parallel \text{Sn}^{2+}(\text{aq}) \mid \text{Sn(s)}$
- C. $\text{Sn}^{2+}(\text{aq}) \mid \text{Sn(s)} \parallel \text{H}_2(\text{g}) \mid \text{H}^+(\text{aq}) \mid \text{Pt}$
- D. $\text{Sn(s)} \mid \text{Sn}^{2+}(\text{aq}) \parallel \text{H}^+(\text{aq}) \mid \text{H}_2(\text{g}) \mid \text{Pt}$**
- E. $\text{Sn(s)} \mid \text{H}_2(\text{g}) \parallel \text{Sn}^{2+}(\text{aq}) \mid \text{H}^+(\text{aq}) \mid \text{Pt}$



12. A galvanic cell consists of one half-cell that contains Ag(s) and $\text{Ag}^+(\text{aq})$, and one half-cell that contains Pb(s) and $\text{Pb}^{2+}(\text{aq})$. What species are produced at the electrodes under standard conditions?

B

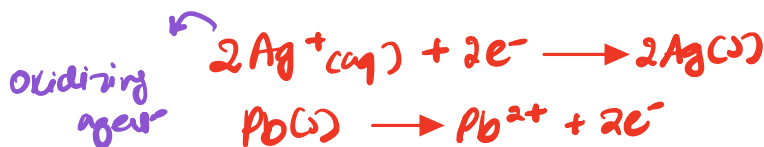


Ag^+ has a higher reduction potential than Pb^{2+} \therefore will be the oxidizing agent

$\therefore \text{Ag}^+$ acts as the cathode (reduction) half reaction

\downarrow
You can also think about Ag^+ has the oxidizing agent which gets reduced at the cathode

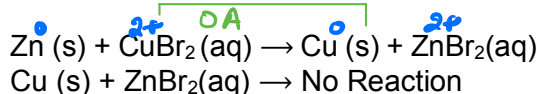
- A. Ag(aq) is formed at the cathode and, Pb(s) is formed at the anode.
- B. Ag(s) is formed at the cathode, and $\text{Pb}^{2+}(\text{aq})$ is formed at the anode.**
- C. Pb(s) is formed at the cathode, and $\text{Ag}^+(\text{aq})$ is formed at the anode.
- D. $\text{Pb}^{2+}(\text{aq})$ is formed at the cathode, and Cu(s) is formed at the anode.



* * Remember oxidizing agents get reduced & reducing agents are oxidized in a redox reaction * *

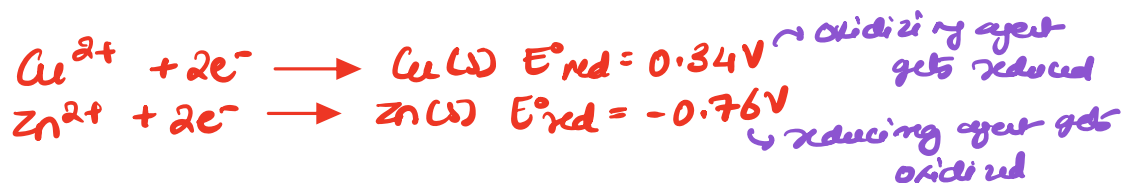
13. Given the following laboratory observation, which of the following statements is FALSE?

C

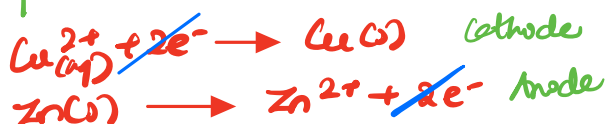


Cu^{2+} = oxidizing agent \therefore has a higher reduction potential \therefore undergoes reduction

- A. Zn is a stronger reducing agent than Cu. **true**
- B. Cu^{2+} is a stronger oxidizing agent than Zn^{2+} . **true**
- C. Cu is a stronger reducing agent than Zn. **false****
- D. The fact that Cu doesn't react with ZnBr_2 proves that copper attracts electrons more than does Zn. **true**
- E. None of the above.

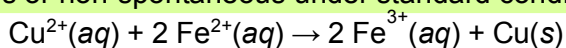


Rewriting the equation

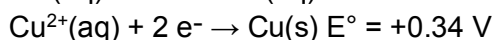
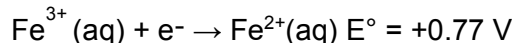


14. Calculate the standard cell potential for the galvanic cell reaction given below and determine whether this reaction is spontaneous or non-spontaneous under standard conditions.

A



split the reaction into two half cell reactions:



$$E^{\circ}_{\text{cell}} = E^{\circ}_{(\text{red}) \text{ Cathode}} - E^{\circ}_{(\text{red}) \text{ Anode}}$$

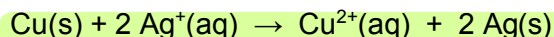
$$= 0.34 - 0.77$$

$= -0.43 \text{ V} \rightarrow$ negative E°_{cell} value indicates that the galvanic cell reaction is non-spontaneous

- A. $E^{\circ} = -0.43 \text{ V}$, nonspontaneous.
B. $E^{\circ} = -0.43 \text{ V}$, spontaneous.
C. $E^{\circ} = +0.43 \text{ V}$, nonspontaneous.
D. $E^{\circ} = +0.43 \text{ V}$, spontaneous.

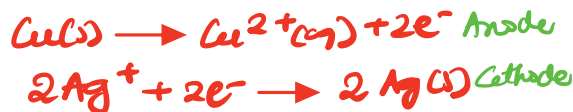
15. A galvanic cell consists of a silver electrode in 1.0 mol/L solution of silver nitrate, a copper electrode in 1.0 mol/L solution of copper(II) nitrate, and a salt bridge. The spontaneous cell reaction is:

CD



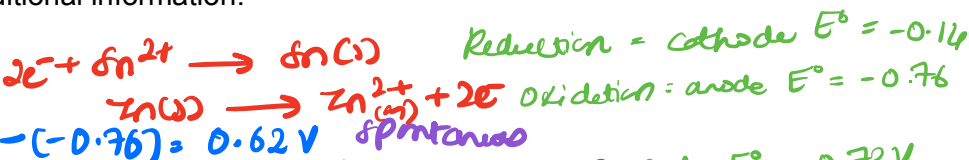
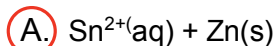
When the two electrodes are connected by a wire, which of the following does not take place? Select all that apply. Insert letters without spaces in the answer box, example ABCD.

- a. Electrons flow in the wire from the copper electrode to the silver electrode. *True (in a galvanic cell, electrons flow from the anode to the cathode)*
b. The silver electrode increases in mass as the cell operates. *True*
c. There is a net movement of silver ions through the salt bridge from the silver half-cell to the copper half-cell. *False (counter ions from the salt bridge to anode & cathode solutions to avoid charge build up)*
d. There is a net movement of copper ions through the salt bridge from the copper half-cell to the silver half-cell. *False*
e. The copper electrode decreases in mass as the cell operates. *True*

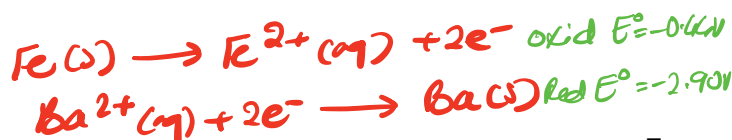
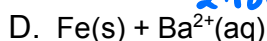
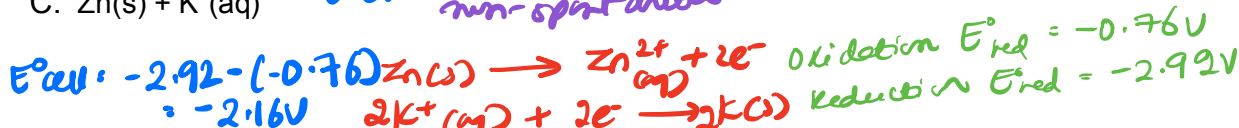
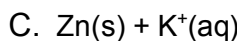
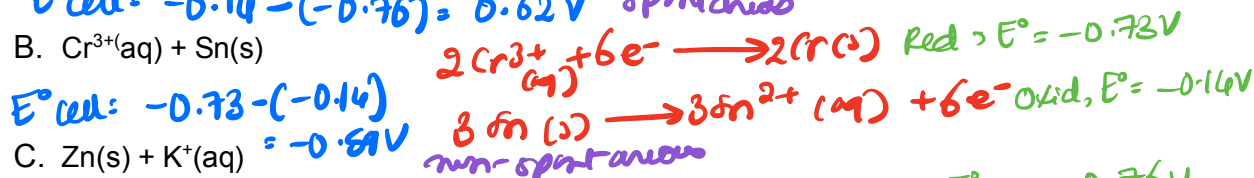
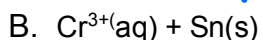


16. Determine which of the following pairs of reactants will result in a spontaneous reaction at 25°C. Please refer to the table in question 6 for additional information.

A



$$E^{\circ}_{\text{cell}} = -0.14 - (-0.76) = 0.62 \text{ V} \quad \text{spontaneous}$$



$$E^{\circ}_{\text{cell}} = -2.90 - (-0.44) = -2.46 \quad \text{non-spontaneous}$$