General Chemistry II Jasperse
Electrochemistry. Extra Practice Problems

Key Equations Given for Test:

$E^\circ_{cell}=E^\circ_{reduction}+E^\circ_{oxidation}$	$\Delta G^\circ = -96.5nE^\circ_{cell}$ (ΔG° in kJ)
$E_{cell} = E^\circ - [0.0592/n]\log Q$	$\log K = nE^\circ/0.0592$
Mol $e^- = [A \bullet$ time (sec)/96,500]	time (sec)= mol e • 96,500/current (in A)
$t = (t_{1/2}/0.693)\ln(A_o/A_t)$	$\ln(A_o/A_t) = 0.693 \bullet t/t_{1/2}$
$E = \Delta mc^2$ (m in kg, E in J, c = $3x10^8$ m/s)	

Oxidation Numbers

1. What is the **oxidation number of chromium** in the ionic compound ammonium dichromate, $(NH_4)_2Cr_2O_7$?

   a. +3
   b. +4
   c. +5
   d. +6
   e. +7

   1. Treat Polyatomics as a "package" (helpful shortcut)
   2. Sum Oxidation Numbers
   3. Memorize rules/priorities.
   4. Suggestion: Write "individual" charges above, "sum" below

2. What is the **oxidation number of carbon** in the ionic compound potassium carbonate, K_2CO_3?

   a. +3
   b. +4
   c. +5
   d. +6
   e. +7

   Memorize priority in ox. # rules:
   elements >
   simple obvious ions >
   G1,G2 metal cation charge >
   H (+1) > O (-2) > halogen (-1) >
   deduce the rest.

3. What are the **oxidation numbers for nickel, sulfur, and oxygen** in $Ni_2(SO_4)_3$?

   a. Ni +3; S +6; O -2
   b. Ni +2; S +4; O -2
   c. Ni +3; S +4; O -2
   d. Ni +2; S +2; O -2
   e. Ni +2; S +4; O -1

4. When hydrogen reacts with calcium metal, **what are the oxidation numbers** of the calcium and hydrogen in the CaH_2 product?

   $$Ca(s) + H_2(g) \rightarrow CaH_2(s)$$

   a. –2 and +1
   b. +1 and –2
   c. +2 and –1
   d. 0 and 0
   e. +2 and –2

   Memorize priority in ox. # rules:
   elements >
   simple obvious ions >
   G1,G2 metal cation charge >
   H (+1) > O (-2) > halogen (-1) >
   deduce the rest.

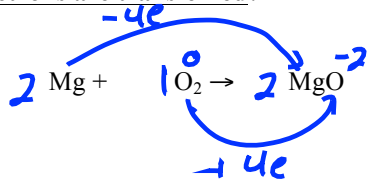
5. What **are the original and final oxidation numbers for iron** in the smelting of iron from iron oxide?

   $$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

   a. +2 → 0
   b. +3 → 0
   c. 0 → +2
   d. +6 → 0
   e. No change

Balancing Redox; Electrons Transferred; Oxidizing Agents; Reducing Agents

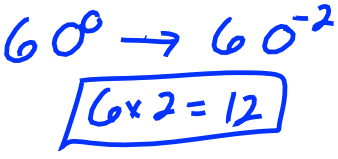
6. Balance the following reaction. **How many electrons are transferred**?

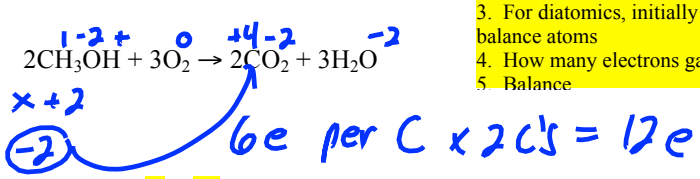


a. 3 b. 4 c. 6 d. 8 e. 2

7. Methanol fuel cells use the following reaction. **How many electrons are transferr**

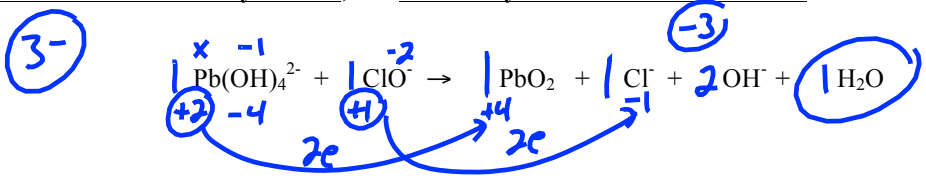
1. Identify Ox #'s for everything
2. Identify which change
3. For diatomics, initially use a "1" coefficient, and balance atoms
4. How many electrons gained/lost?
5. Balance





a. 3
b. 6
c. 8
d. 12
e. 2

8. **What is the coefficient for hydroxide**, and **how many electrons are transferred** after balancing the reaction?



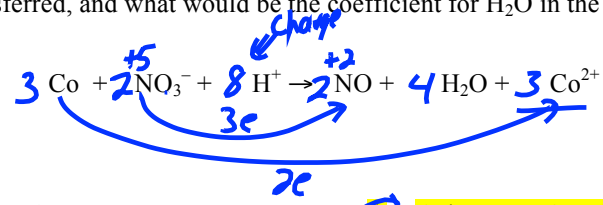
a. 2 OH^- and 2 electrons
b. 3 OH^- and 4 electrons
c. 1 OH^- and 2 electrons
d. 2 OH^- and 4 electrons
e. None of the above

1. Balance redox atoms first
2. Balance charges
3. Balance spectator atoms

-Between balancing charges and atoms, either can be done first

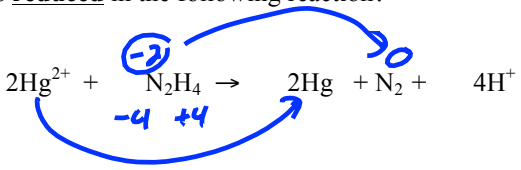
9. Cobalt is one of many metals that can be oxidized by nitric acid. Balance the following the reaction. How many electrons are transferred, and what would be the coefficient for H_2O in the balanced reaction?

6 e



a. 3 electrons; 2 H_2O
b. 6 electrons; 6 H_2O
c. 4 electrons; 2 H_2O
d. 6 electrons; 4 H_2O
e. None of the above

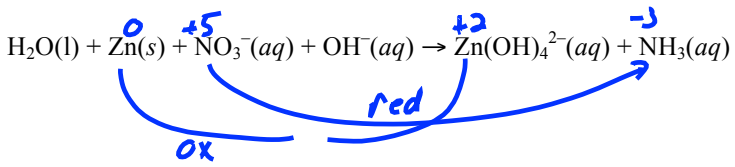
10. What was **oxidized** and what was **reduced** in the following reaction?



Determine all oxidation numbers and see which change!
"Oxidized" Ox # increased
"Reduced" Ox # reduced

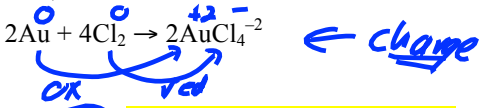
a. Hg^{2+} was oxidized; N_2H_4 was reduced
b. Hg^{2+} was reduced; N_2H_4 was oxidized
c. Hg^{2+} was oxidized; N_2H_4 was oxidized
d. Hg^{2+} was reduced; N_2H_4 was reduced
e. None of the above

11. The following reaction occurs in basic solution. **Identify the oxidizing agent**. Note the reaction equation is not balanced.



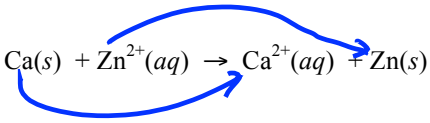
a. $Zn(s)$
b. $NO_3^-(aq)$ (the nitrogen)
c. $OH^-(aq)$
d. $H_2O(l)$ (the oxygen)
e. $NH_3(aq)$ (the nitrogren)

12. For the following reaction, which statement, A–D, is *not* correct? If more than one is *not* correct, respond E.



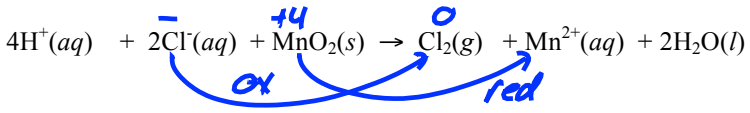
a. Au is the reducing agent. T
b. Cl_2 is the oxidizing agent T
c. Au is oxidized. T
d. The equation is fully balanced.
e. More than one statement is *not* correct.

13. Which substance is the **reducing agent** in the following reaction?



a. $Ca(s)$
b. $Zn^{2+}(aq)$
c. $Ca^{2+}(aq)$
d. $Zn(s)$
e. None of the above

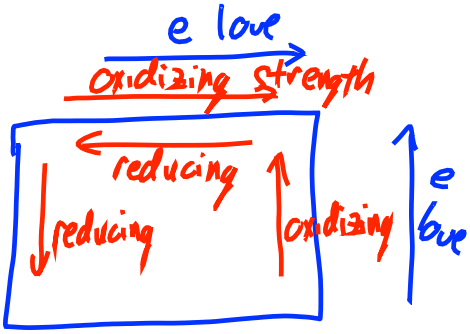
14. Which substance is the **reducing agent** in the following reaction?



a. $H^+(aq)$
b. $Cl^-(aq)$
c. $MnO_2(s)$
d. $Cl_2(g)$
e. $Mn^{2+}(aq)$

15. Which one of the following items does *not* characterize an oxidizing agent?

a. An oxidizing agent gains electrons. T
b. An oxidizing agent causes another species to be oxidized. T
c. The oxidation number of an oxidizing agent decreases. T
d. A good oxidizing agent is a metal in a high oxidation state, such as Mn^{7+}. T
e. An example of a good oxidizing agent is an alkali metal, such as Na.



16. Which of the following statements about electrochemical cells is true?

a. Reduction occurs at the anode F
b. An element with a high love for electrons is likely to be easily oxidized
c. Oxidation occurs at the anode
d. Only oxidation half–reactions are useful
e. none of the above

Spontaneous Voltaic Electrochemical Cells

17. Which statement about a voltaic cell is ***not* correct**?

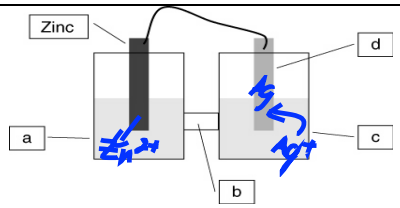
a. Chemical species can have their oxidation number decreased at the cathode.
b. Reduction occurs at the cathode.
c. Usually the cathode is a metal strip.
d. Oxidation occurs at the anode.
e. Elemental metal is routinely converted to metal cations at the cathode

18. Which statement regarding voltaic cells is ***not correct***?

a. Reduction occurs at the cathode.
b. Anions move through the barrier/bridge toward the electrode where oxidation is occurring.
c. The electrode where reduction is occurring is represented by a positive sign.
d. Electrons flow in the external circuit from the cathode to the anode.
e. Electrons flow in the external circuit toward the electrode represented by a positive sign.

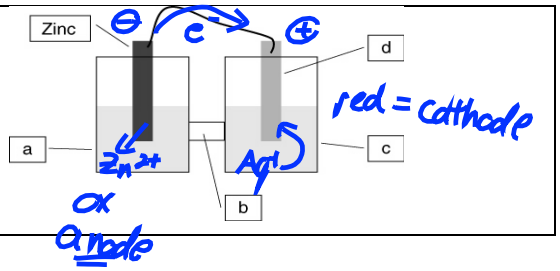
19. A voltaic cell is constructed based on the oxidation of zinc metal and the reduction of silver metal. Solutions of silver nitrate and zinc nitrate also were used. **Locate the silver and the silver nitrate on the diagram**.

a) silver = b; silver nitrate = a
b) silver = d; silver nitrate = b
c) silver = d; silver nitrate = c
d) silver = d; silver nitrate = a



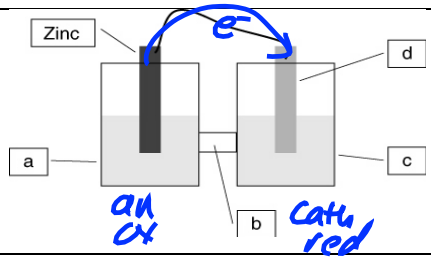
20. A voltaic cell is constructed based on the oxidation of zinc metal and the reduction of silver metal. Solutions of silver nitrate and zinc nitrate also were used. **Locate the zinc nitrate on the diagram, and identify the anode**.

a) Zinc nitrate = a; anode = d
b) Zinc nitrate = a; anode = Zinc
c) Zinc nitrate = c; anode = d
d) Zinc nitrate = c; anode = Zinc



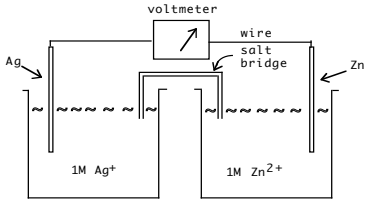
21. A voltaic cell is constructed based on the oxidation of zinc metal and the reduction of silver metal. Solutions of silver nitrate and zinc nitrate also were used. **Which statement is true** regarding the direction of electron flow through the external wire?

a) Electrons flow from left to right, from the Zinc
b) Electrons flow from right to left, to the Zinc
c) The zinc electrode will get larger as more zinc forms.
d) Anions will flow through the "bridge" from the zinc side to the silver side



22. For the cell shown, the standard reduction potentials are +0.80 V for Ag^+ and –0.76 V for Zn^{2+}. Based on the reduction potentials, the _____ electrode is where the reduction will occur and it is called the _____.

a. Ag, cathode
b. Ag, anode
c. Zn, cathode
d. Zn, anode
e. none of the above



Cell Potentials

23. What is E° for the following balanced reaction?

$E°(cell) = E°(red) + E°(ox)$
$E°(ox) = -E°(red)$

$$Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$$

Half-reaction	*Standard Reduction Potential*
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	–0.763
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	–0.126

a. +0.637 V
b. –0.637 V
c. +1.274 V
d. –0.889 V
e. +0.889 V

24. What is E° for the following balanced reaction?

$$Al(s) + Fe^{3+}(aq) \rightarrow Al^{3+}(aq) + Fe(s)$$

Half-reaction	*Standard Reduction Potential*
$Fe^{3+}(aq) + 2e^- \rightarrow Fe(s)$	+0.771
$Al^{3+}(aq) + 2e^- \rightarrow Al(s)$	–1.660

a. +1.280 V
b. –2.431 V
c. +2.431 V
d. –0.889 V
e. +0.889 V

25. What is E° for the following balanced reaction?

$$4H^+(aq) + Fe(s) + NO_3^-(aq) \rightarrow Fe^{3+}(aq) + NO(aq) + 2H_2O(l)$$

	Standard Reduction Potential
$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	E° = +0.960 V
$Fe^{3+} + 3e^- \rightarrow Fe$	E° = +0.771 V

a. +0.189 V
b. –0.189 V
c. +1.731 V
d. –1.731 V
e. None of the above

26. Given the electrochemical reaction shown, if the standard reduction potential of $Ag^+ \rightarrow Ag$ is +0.80 V, and the standard reduction potential of $Cu^{2+} \rightarrow Cu$ is +0.34V, what is E° for the following?

Cu / Cu^{2+}(aq) // Ag^+(aq) / A E° = ??? V

a. +1.26 V
b. +0.85 V
c. +0.46 V
d. -0.37 V
e. none of the above

27. Given the electrochemical reaction shown, if the standard reduction potential of $Ni^{+2} \rightarrow Ni$ is -0.26 V, and the standard reduction potential of $Al^{3+} \rightarrow Al$ is -1.66V, what is E° for the following?

Al / Al^{3+}(aq) // Ni^{+2}(aq) / Ni E° = ??? V

a. +1.26 V
b. +0.85 V
c. +0.46 V
d. +1.40 V
e. none of the above

28. Given the electrochemical reaction shown, if the standard reduction potential of $Zn^{2+} \rightarrow Zn$ is –0.76 V, **what is the standard reduction potential of $Mg^{2+} \rightarrow Mg$?**

$Mg / Mg^{2+}(aq) // Zn^{2+}(aq) / Zn$ $\quad E^\circ = +1.61$ V

a. –0.85 V
b. +0.85 V
c. +2.37 V
d. –2.37 V
e. none of the above

1. Solve for magnesium E°(ox)
2. To get E°(red), reverse the +/- sign

? x ox $\quad -.76$ red $= 1.61$

$x = 1.61 + .76 = 2.37$ ← ox potential

$Mg^{2+} \rightarrow Mg \quad -2.37$

29. Given the electrochemical reaction shown, if the standard reduction potential of $Cu^{2+} \rightarrow Cu$ is +0.34 V, **what is the standard reduction potential of $Sn^{2+} \rightarrow Sn$?**

$Sn / Sn^{2+}(aq) // Cu^{2+}(aq) / Cu$ $\quad E^\circ = +0.48$ V

a. –0.14 V
b. +0.14 V
c. +0.37 V
d. –0.37 V
e. none of the above

red, x, ox $\quad +.34 = +.48$

$x = +.48 - .34 = +.14$ $\quad E^\circ_{ox} = +.14$

so $E^\circ_{red} = -.14$

30. Identify the strongest reducing agent based on the following half-reactions. The standard reduction potentials are listed.

in order

+1.22 V	$MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$
+0.61 V	$Hg_2SO_4(s) + 2e^- \rightarrow 2Hg(l) + SO_4^{2-}(aq)$
–0.95 V	$SnO_2(s) + 2H_2O() + 4e^- \rightarrow Sn(s) + 4OH^-(aq)$
–1.48 V	$Cr(OH)_3(s) + 3e^- \rightarrow Cr(s) + 3OH^-(aq)$

Oxidizing agents $\quad$ Reducing agents

a. Cr
b. MnO_2
c. Hg_2SO_4
d. Sn
e. Hg

31. Identify the strongest oxidizing agent from the following half-reactions. The standard reduction potentials are listed.

in order

+1.22 V	$MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$
+0.61 V	$Hg_2SO_4(s) + 2e^- \rightarrow 2Hg(l) + SO_4^{2-}(aq)$
–0.95 V	$SnO_2(s) + 2H_2O() + 4e^- \rightarrow Sn(s) + 4OH^-(aq)$
–1.48 V	$Cr(OH)_3(s) + 3e^- \rightarrow Cr(s) + 3OH^-(aq)$

Oxidizing agents $\quad$ Reducing agents

a. Cr
b. MnO_2
c. Hg_2SO_4
d. Sn
e. Hg

32. In one episode of the TV sitcom, *Gilligan's Island,* the "professor" constructed voltaic cells to use as substitutes for their radio's dead batteries. **Which scraps of metal from their damaged boat, the *Minnow*, could best be used to create a 1.5 V voltaic cell**? (Assume that coconuts make great beakers and that seawater is a terrific electrolyte!)

Metal/Metal ion	E°_{red}
lead/lead(II) (fishing weights)	–0.126
iron/iron(II) (the anchor)	–0.44
silver/silver(I) (Mrs. Howell's brooch)	–0.799
aluminum/aluminum(III) (the boat's flagpole)	–1.677

red $\quad Pb^{2+} \rightarrow Pb \quad +.126$ ox

$Fe^{2+} \rightarrow Fe \quad +.44$

$Ag^{+} \rightarrow Ag \quad +.799$

$Al^{3+} \rightarrow Al \quad +1.677$ ox

ox

a. silver anode and lead cathode
b. aluminum anode and lead cathode
c. iron anode and aluminum cathode
d. aluminum anode and silver cathode
e. lead cathode and silver anode

1.677 ox ✓
−0.126 red ✓
1.551 V°
~ 1.5 V

To get the correct approximate voltage, choose an oxidation and redution pair that sum appropriately.
Cathode: reduction element (lead)
Anode: oxidation (aluminum)

Standard Reduction Potentials (volts) in Aqueous Solution		
	$Pb^{4+} + 2e^- \rightarrow Pb^{2+}$	+1.80
	$Au^{3+} + 3e^- \rightarrow Au$	+1.50
	$Fe^{3+} + 3e^- \rightarrow Fe$	+0.771
	$I_2 + 2e^- \rightarrow 2\ I^-$	+0.535
	$Pb^{2+} + 2e^- \rightarrow Pb$	–0.124
	$Al^{3+} + 3e^- \rightarrow Al$	–1.66
	$Mg^{2+} + 2e^- \rightarrow Mg$	–2.37
	$K^+ + e^- \rightarrow K$	–2.93

Oxidizing agents Reducing Agents

33. What is the standard cell potential for a voltaic cell using the Pb^{2+}/Pb and Mg^{2+}/Mg half-reactions? Which metal is the cathode? (Use the Standard Reduction Potentials table shown above)

a. –2.25 V, Pb is the cathode
b. +2.25 V, Mg is the cathode
c. –2.25 V, Mg is the cathode
d. +2.25 V, Pb is the cathode
e. –2.49 V, Mg is the cathode

1. which red is better?
*Make that the reduction/cathode
2. Use the other as oxidation/cathode

2.37 − .124 = 2.25

34. What is the standard cell potential for a voltaic cell using the Al^{3+}/Al and Fe^{3+}/Fe half-reactions? Which metal is the anode? (Use the Standard Reduction Potentials table shown above)

a. –2.43 V, Al is the anode
b. +2.43 V, Al is the anode
c. –0.89 V, Fe is the anode
d. +0.89 V, Fe is the anode
e. None of the above

35. Using the *Table of Standard Reduction Potentials* table shown above, **which is the strongest oxidizing agent**?

a. Pb^{4+}
b. Pb^{2+}
c. K^+
d. K
e. Al

36. Using the *Table of Standard Reduction Potentials* table shown above, **which is the strongest reducing agent**?

a. Pb^{4+}
b. Pb^{2+}
c. K^+
d. K
e. Al

37. Use the *Table of Standard Reduction Potentials* table, **which species would react with Fe**?

a. Pb^{4+} only
b. Au^{3+} only
c. I_2 and Pb^{2+}
d. Both Pb^{4+} and Au^{3+}
e. Both Pb^{2+} and Au

1. Is it reducing agent or oxidizing agent?
2. A reducing agent will only react favorably with oxidizing agents higher on the chart

38. Use the *Table of Standard Reduction Potentials* table, **which species would react with Al^{3+}**?

a. Pb only
b. Au^{3+} only
c. Fe and Pb
d. Both Mg^{+2} and K^+
e. Both Mg and K

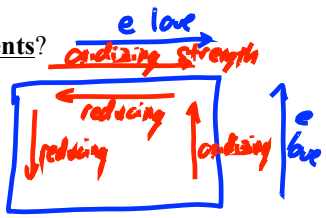
39. Using the *Table of Standard Reduction Potentials* table shown above, what is the standard cell potential for an electrochemical cell that has iron (Fe) and magnesium (Mg) electrodes? Also, identify the cathode.

a. +3.14 V with Fe as the cathode
b. +3.14 V with Mg as the cathode
c. –3.14 V with Fe as the cathode
d. –3.14 V with Mg as the cathode
e. +1.60 V with Fe as the cathode

Predictable Patterns in Oxidation and Reduction Strength. (Should be able to recognize from periodic table, but without looking at a table with reduction potentials)

40. Glancing at a periodic table, where do you expect to find **elements that are good oxidizing agents**?

a. on the right (except for the noble gases)
b. in the middle left
c. in the top left
d. at the bottom
e. in the transition metals



41. Glancing at a periodic table, where do you expect to find **elements that are good reducing agents**?

a. in groups 16 and 17
b. on the left
c. in the middle
d. at the bottom
e. in group 17

42. Based on the periodic table and general patterns of activity, which is the correct ranking of the halogens as oxidizing agents? (you should be able to answer without looking at a reduction-potential table.)

F_2^0 F^- Cl_2^0 Cl^- Br_2^0 Br^- I_2^0 I^-

a. F_2 (strongest oxidant) > Cl_2 > Br_2 > I_2 (weakest oxidant)
b. I_2 (strongest oxidant) > Br_2 > Cl_2 > F_2 (weakest oxidant)
c. Cl_2 (strongest oxidant) > F_2 > Br_2 > I_2 (weakest oxidant)
d. Br_2 (strongest oxidant) > I_2 > Cl_2 > F_2 (weakest oxidant)

43. Based on the periodic table and general patterns of activity, which is the correct ranking of the following metals as reducing agents? (Atomic numbers shown)

G2 Mg (12) G1 K (19) TM Au (79) TM Fe (26)

gold precious

a. Mg (strongest reducing agent) > K > Fe > Au (weakest oxidizing agent)
b. K (strongest reducing agent) > Mg > Fe > Au (weakest oxidizing agent)
c. Au (strongest reducing agent) > Mg > Fe > K (weakest oxidizing agent)
d. Fe (strongest reducing agent) > Au > Mg > K (weakest oxidizing agent)

reducing

44. Based on the periodic table and general patterns of activity, which of the following would react with metallic **calcium**?

KBr G1 NaI G1 $FeCl_2$ T $NiBr_2$ T

reducing agent G2

a. KBr and NaI only
b. $FeCl_2$ only
c. $NiBr_2$ only
d. both $FeCl_2$ and $NiBr_2$

1. Calcium is a reducing agent
2. Each formula has a metal cation, potential oxidizing agents.
3. Reduction potential (for elemental metals):
G1 > G2 > Al > Transition Metals > Hydrogen > "Precious metals"
4. Calcium, being group 2, can reduce anything except Group 1 cations

$Ca + 2K^+ \not\rightarrow Ca^{2+}$

$Ca + Fe^{2+} \rightarrow Fe + Ca^{2+}$ $+2K$

45. Based on the periodic table and general patterns of activity, which of the following would react with metallic **sodium**?

I_2^0 Y I^- X $FeCl_2$ T $NiBr_2$ T

a. I_2 only
b. I^- only
c. $NiBr_2$ only
d. I_2, $FeCl_2$ and $NiBr_2$

1. Sodium is a reducing agent, group 1
2. Iodize is already in reduced form, so can't react with reducing agent.
3. Reduction potential (for elemental metals):
G1 > G2 > Al > Transition Metals > Hydrogen > "Precious metals"
4. Calcium, being group 1, can reduce anything except Group 1 cations
5. Iodine is a strong oxidizing agent, and like any elemental halogen would react with pretty much any elemental metal

Ranking Relative Activity, Based on Observed Reactivity or Lack Thereof

46. Given the following laboratory observation, which of the following statements is **NOT TRUE**?

$Sn + 2AgBr \rightarrow 2Ag + SnBr_2$
$2Ag + SnBr_2 \rightarrow$ No Reaction

1. Reaction means the reacting cation is stronger oxidant and reacting metal is stronger reducing agent than the product cation/metal.
2. No reaction means the reacting cation is weaker oxidizing agent and reacting metal is weaker reducing agent than the product cation/metal would be.

a. Sn is a stronger reducing agent than Ag
b. Ag^+ is a stronger oxidizing agent than Sn^{2+}
c. The reduction potential for Ag^+ is more positive than the reduction potential for Sn^{2+}
d. Sn^{2+} is a stronger oxidizing agent than Ag^+
e. none of the above

47. Given the following laboratory observation, which of the following statements is **NOT TRUE**?

$Zn + CuBr_2 \rightarrow Cu + ZnBr_2$
$Cu + ZnBr_2 \rightarrow$ No Reaction

a. Zn is a stronger reducing agent than Cu
b. Cu^{+2} is a stronger oxidizing agent than Zn^{2+}
c. Cu is a stronger reducing agent than Zn
d. The fact that copper doesn't react with $ZnBr_2$ proves that copper loves/attracts/holds electrons more than does zinc.
e. none of the above

48. Given the following laboratory observation, which of the following statements is **NOT TRUE**?

$Mg + NiBr_2 \rightarrow Ni + MgBr_2$
$Ni + MgBr_2 \rightarrow$ No Reaction

1. Reaction means the reacting cation is stronger oxidant and reacting metal is stronger reducing agent than the product cation/metal.
2. No reaction means the reacting cation is weaker oxidizing agent and reacting metal is weaker reducing agent than the product cation/metal would be.

a. Ni loves electrons more than Mg. That's why Mg gives electrons to Ni^{2+}
b. Mg loves electrons less than Ni. That's why Mg^{2+} doesn't take electrons from Mg
c. When a redox reaction does NOT occur (equation 2), it means that the reduced form of nickel is a weaker reducing agent than the reduced form of Mg
d. When a redox reaction DOES occur (equation 1), it means that the reduced form of Mg is a stronger reducing agent than the reduced form of Ni
e. When a redox reaction does NOT occur (equation 2), it means that the reduced form of nickel is a stronger reducing agent than the reduced form of Mg

49. Which of the following correctly ranks the "activity" (strength as reducing agents) of the elements Ag, Au, and Sn, given the following observed reactivity information?

$Sn + 2AgBr \rightarrow SnBr_2 + 2Ag$
$3Sn + 2AuBr_3 \rightarrow 3SnBr_2 + 2Au$
$3Ag + AuBr_3 \rightarrow 3AgBr + Au$

a. Sn > Ag > Au
b. Sn > Au > Ag
c. Au > Ag > Sn
d. Ag > Au > Sn

50. Which of the following correctly ranks the "activity" (strength as reducing agents) of the elements Cu, Cd, and Zn, given the following observed reactivity information?

$Zn + CuBr_2 \rightarrow Cu + ZnBr_2$
$Cd + ZnBr_2 \rightarrow$ No Reaction
$Cu + CdBr_2 \rightarrow$ No Reaction

1. Reaction means the reacting cation is stronger oxidant and reacting metal is stronger reducing agent than the product cation/metal.
2. No reaction means the reacting cation is weaker oxidizing agent and reacting metal is weaker reducing agent than the product cation/metal would be.

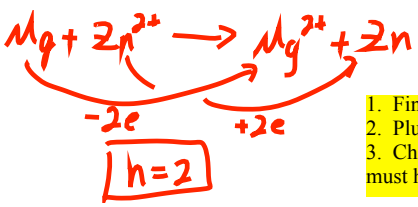
a. Zn > Cu > Cd
b. Zn > Cd > Cu
c. Cd > Cu > Zn
d. Cu > Cd > Zn

Free Energy and Equilibrium. Key Equation: $\Delta G^\circ = -96.5nE^\circ_{cell}$ (ΔG° in kJ/mol)

51. Given the electrochemical reaction shown, what is the standard free energy change ΔG° if $E^\circ = +1.61$ V?

$$Mg / Mg^{2+}(aq) // Zn^{2+}(aq) / Zn \qquad E^\circ = +1.61 \text{ V}$$

a. -311 kJ/mol
b. +311 kJ/mol
c. -155 kJ/mol
d. +155 kJ/mol
e. none of the above



$\Delta G = -96.5(2)(1.61) = -311$

1. Find "n"
2. Plug and Chug
3. Check: a positive E° means you must have a negative delta G

52. The oxidation of hydrogen by oxygen is one of the most-used reactions in fuel-cell technology. The overall reaction, which is given below, has a ΔG° value of –474 kJ/mol. What is the standard cell potential for this fuel cell?

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \qquad \Delta G^\circ = -474 \text{ kJ/mol}$$

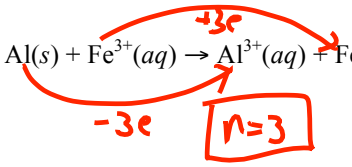
2×2 = 4
n = 4

a. 2.46 V
b. 4.91 V
c. 1.23 V
d. 2.46 V
e. 1.50 V

$\Delta G = -96.5\, n\, E^\circ$
$-474 = -96.5(4)E^\circ$
$E^\circ = 1.23$

53. What is ΔG° for the following balanced reaction, if $E^\circ = +2.431$ V?

$$Al(s) + Fe^{3+}(aq) \rightarrow Al^{3+}(aq) + Fe(s) \qquad E^\circ = +2.431 \text{ V}$$



a. -704 kJ/mol
b. +704 kJ/mol
c. -235 kJ/mol
d. -469 kJ/mol
e. none of the above

$\Delta G = -96.5\, n\, E^\circ$
$\Delta G = -96.5(3)(2.431)$
$= -704$

54. The oxidation of methanol, as described by the equation below, has a ΔG° value of –937.9 kJ/mol. What is the standard cell potential for a methanol fuel cell?

$$2CH_3OH + 3O_2 \rightarrow 2CO_2 + 4H_2O \qquad \Delta G^\circ = -937.9 \text{ kJ/mol}$$

C: -2 → +4
6e per Carbon
×2 = 12 electrons
2 per ea
6 oxygens
12 electrons
n = 12

a. 0.405 V
b. 9.72 V
c. 0.810 V
d. –2.43 V
e. –9.72 V

$\Delta G = -96.5\, n\, E^\circ$
$-937.9 = -96.5(12)E^\circ$
$E = 0.81$

55. For the following reaction, all of the reactants and products are in their standard states/standard 1.0M concentrations. Which of the following statements must be true?

$$Zn(s) + SnBr_2(aq) \rightarrow ZnBr_2(aq) + Sn(s) \qquad E^\circ = +0.60 \text{ V}$$

favorable

a. The reaction would be **product-favored** as written T
b. ΔG° for the reaction as written is **positive**
c. Zinc is undergoing reduction F
d. none of the above

ΔG neg favorable

K Values and Voltage Key Equation: $\log K = nE^\circ/0.0592$

56. When a voltaic cell reaches **equilibrium**,

a. $E = 0$
b. $E_{cell} = 0$
c. $E_{cell} = K$
d. $E = K$
e. $E_{cell} = Q$

57. Electrochemical cell potentials can be used to determine equilibrium constants that would be otherwise difficult to determine because concentrations are small. **What is K for the following balanced reaction, if $E^\circ = +0.0218$ V**?

Answer Key Error

$3Zn(s) + 2Cr^{3+}(aq) \rightarrow 3Zn^{+2}(aq) + Cr(s)$ $\quad E^\circ = +0.0218$ V

3 x 2 = 6, n = 6

$\log K = \frac{6(.0218)}{.0592} = 2.209$

$K = 10^{2.209} = 162$

a. 1.3×10^{-3}
b. 2.3×10^{-4}
c. 2.2
d. 0.37

162

58. What is E° for the following balanced reaction, if $K = 4.38 \times 10^{10}$?

n = 2

$Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{+2}(aq) + Fe(s)$ $\quad K = 4.38 \times 10^{10}$

a. -0.578 V
b. +0.866 V
c. -0.315 V
d. +0.315 V

$\log 4.38 \times 10^{10} = \frac{2(E^\circ)}{.0592}$

$10.64 = \frac{2E}{.0592} \Rightarrow E = 0.315$

Nonstandard Concentrations and Cell Potential. Key Equation: $E_{cell} = E^\circ - [0.0592/n]\log Q$

59. The value of E° for the following reaction is 1.10 V. What is the value of E_{cell} when the concentration of Cu^{2+} is 1.0 M and the concentration of Zn^{2+} is 0.025 M?

n = 2

$Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$ $\quad E^\circ = 1.10$ V
1.0 M $\quad$ 0.025 M

a. 1.40 V
b. 0.95 V
c. 1.15 V
d. 0.80 V

$E = 1.10 - \left(\frac{.0592}{2}\right)\log\left(\frac{.025}{1}\right) = 1.10 - \left(\frac{.0592}{2}\right)(-1.602) = 1.10 + .047$

$= 1.15$

60. The value of E° for the following reaction is 1.260 V. What is the value of E_{cell} given the concentrations shown?

6e⁻, n = 6

$2Al(s) + 3Cd^{2+}(aq) \rightarrow 3Cd(s) + 2Al^{3+}(aq)$ $\quad E^\circ = 1.260$ V
0.1 M $\quad$ 0.6 M

a. 1.235 V
b. 1.285 V
c. 1.15 V
d. 1.37 V

$E = 1.260 - \left(\frac{.0592}{6}\right)\log\frac{(.6)^2}{(.1)^3} = 1.260 - \left(\frac{.0592}{6}\right)(2.56) = 1.235$

360

61. The value of E° for the following reaction is 0.189 V. What is the value of E_{cell} given the concentrations shown?

n = 3

$4H^+(aq) + Fe(s) + NO_3^-(aq) \rightarrow Fe^{3+}(aq) + NO(aq) + 2H_2O(l)$ $\quad E^\circ = 0.189$ V
0.1 M $\quad$ 0.6 M $\quad$ 1.0 M $\quad$ 0.5 M

a. -0.215 V
b. 0.112 V
c. 0.189 V
d. 0.266 V

$E = .189 - \frac{.0592}{3}\log\frac{(1)(.5)}{(.1)^4(.6)} = .189 - \frac{.0592}{3}\log 8333 = .189 - \frac{.0592}{3}(3.921)$

8333

$= .1116$ V

Electrolysis: Key Equations: [Mol e^- = (A • time (sec))/96,500] [time (sec) = mol e • 96,500/current (in A)]

62. Copper metal (63.546 g/mol) is purified by electrolysis. **How much copper metal** (in grams) could be produced from copper(II) oxide by applying a current of 10.0 amps at the appropriate negative potential for 12.0 hours?

a. 284 g
b. 142 g
c. 28.4 g
d. 14.2 g
e. 4.48 g

$2e^- + Cu^{2+} \rightarrow Cu$

$\text{mol } e^- = \frac{(10)(12\text{ h})(60\text{ min}/1\text{ h})(60\text{ s}/1\text{ min})}{96500\text{ C}} = 4.477\text{ mol e}$

$4.477\text{ mol e} \left| \frac{1\text{ mol Cu}}{2\text{ mol e}} \right| \frac{63.546\text{ g}}{1\text{ mol Cu}} = 142\text{ g}$

A, time → mol e^- → mol Cu → g Cu (MW)

Time and Amps to Grams Problems:
1. Use Amps and Time to find moles of electrons
2. Find "n": moles of electrons per mole of metal
3. Using "n", convert from moles e's to moles metal
4. Using formula weight, can convert between moles of metal and grams of metal

63. **How long would it take** to electroplate a flute with 28.3 g of silver (107.87 g/mol) at a constant current of 2.0 amps using $AgNO_3$?

a. 211 min
b. 422 min
c. 844 min
d. 1688 min
e. 105 min

$Ag^+ + 1e^- \rightarrow Ag$

$\text{mol e} = \frac{28.3\text{ g}}{107.87\text{ g}} \left| \frac{1\text{ mol Ag}}{1\text{ mol Ag}} \right| \frac{1\text{ mol } e^-}{1\text{ mol Ag}} = .26234\text{ mol e}$

$s = \frac{(.26234)(96500)}{2} = 1.2666 \times 10^{-4}\text{ s} \left| \frac{1\text{ min}}{60\text{ s}} \right| = 211\text{ min}$

g Ag (MW) → mol Ag → mol e → time

Grams to Time Problems given Amps:
1. Using formula weight, convert grams metal => moles metal
2. Find "n", and convert moles metal => moles electrons
3. Give moles of e's and amps => time (in seconds)
4. Convert seconds to actual time units (min, hours...)

64. **How many grams of aluminum** metal (26.982 g/mol) can be produced by the electrolysis of Al_2O_3 using a current of 100 amperes for 24 hours?

a. 805
b. 2,400
c. 8.1×10^5
d. 2.2×10^{13}
e. 7.5×10^{12}

$Al^{3+} + 3e^- \rightarrow Al$

s, A → mol e → mol Al → g Al

65. If in using a lead-acid battery to start a car, 1.00 gram of Pb (207.2 g/mol) is consumed on the anode, **how long will it take** to recharge the battery, using a current of 0.500 amperes to turn the $PbSO_4$ that was produced back into Pb?

a. 15.5 min
b. 1864 min
c. 31 min
d. 21 min
e. 42 min

$Pb^{2+} + 2e^- \rightarrow Pb$

$\text{mol e} = \frac{1\text{ g Pb}}{207.2\text{ g}} \left| \frac{1\text{ mol Pb}}{1\text{ mol Pb}} \right| \frac{2\text{ mol e}}{1\text{ mol Pb}} = 9.65 \times 10^{-3}\text{ mol e}$

$s = \frac{(9.65 \times 10^{-3})(96500)}{.5} = 1.86 \times 10^3\text{ s} \left| \frac{1\text{ min}}{60\text{ s}} \right| = 31\text{ min}$

1 g Pb → mol Pb → mol e^- → time

66. Chromium often is electroplated on other metals and even plastics to produce a shiny metallic appearance. **How many grams** of chromium (51.996 g/mol) would plate out from a solution of $Cr(NO_3)_3$ when 10 amps of current are passed through the electrolytic cell for 5.36 hours?

a. 17.3 g
b. 34.7 g
c. 52.0 g
d. 104 g
e. 11.6 g

$3e^- + Cr^{3+} \rightarrow Cr$

t, A → mol e → mol Cr (1/3) → Cr (MW) → g

67. Suppose an aluminum (26.982 g/mol) beer can weighs 40.0 g. For how long would a current of 100.0 amp need to be passed through a molten AlF_3 electrolysis cell to produce enough aluminum to replace a discarded beer can?

a. 4.44 hr
b. 23.8 min
c. 1.19 hr
d. 42.7 min
e. 2.38 hr

$3e^- + Al^{3+} \rightarrow Al$

40 g → mol Al → mol e^- (x3) → time (s) → time min

General Chemistry II Jasperse ANSWERS
Electrochemistry. Extra Practice Problems

1. D	35. A
2. B	36. D
3. A	37. D
4. C	38. E
5. B	39. A
6. B	40. A
7. D	41. B
8. A	42. A
9. D	43. B
10. B	44. D
11. B	45. D
12. D	46. D
13. A	47. C
14. B	48. E
15. E	49. A
16. C	50. B
17. E	51. A
18. D	52. C
19. C	53. A
20. A	54. C
21. A	55. A
22. A	56. B
23. A	57. C
24. C	58. D
25. A	59. C
26. C	60. A
27. D	61. B
28. D	62. B
29. A	63. A
30. A	64. A
31. B	65. C
32. B	66. B
33. D	67. C
34. B	