**Dougherty Valley HS AP Chemistry**
**Electrochemistry – NChO Extra* Practice**

# WORKSHEET #7*

**Name:** **Date:** **Period:** **Seat #:**

**Directions**: Any worksheet that is labeled with an * means it is suggested extra practice. We do not always have time to assign every possible worksheet that would be good practice for you to do. You can do this worksheet when you have extra time, when you finish something early, or to help you study for a quiz or a test. If and when you choose to do this Extra Practice worksheet, please do the work on binder paper. You will include this paper stapled into your Rainbow Packet when you turn it in, even if you didn't do any of this. We want to make sure we keep it where it belongs so you can do it later if you want to (or need to). If you did the work on binder paper you can include that in your Rainbow Packet after this worksheet. If we end up with extra class time then portions of this may turn into required work. If that happens you will be told which problems are turned into required. Remember there is tons of other extra practice on the class website…and the entire internet! See me if you need help finding practice on a topic you are struggling with.

## 1998

40. For this reaction, $E°_{cell} = 0.79$ V.

$6I^-(aq) + Cr_2O_7^{2-}(aq) + 14H^+$
$\rightarrow 3I_2(aq) + 2Cr^{3+}(aq) + 7H_2O(aq)$

Given that the standard reduction potential for $Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq)$ is 1.33 V, what is $E°_{red}$ for $I_2(aq)$?

a) +0.54 V b) -0.54 V
c) +0.18 V d) -0.18 V

41. What is the product formed at the anode in the electrolysis of 1.0 M $NaNO_3(aq)$?

a) $H_2(g)$ b) $NO_2(g)$
c) $O_2(g)$ d) $Na(s)$

42. Which of these ions is the best reducing agent?

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

a) $Fe^{3+}$ b) $Fe^{2+}$
c) $Cu^{2+}$ d) $Cu^+$

43. $Zn(s) + Cl_2(g, 1\ atm)$
$\rightleftharpoons Zn^{2+}(aq, 1\ M) + 2Cl^-(aq, 1\ M)$

An electrochemical cell based on this reaction has a cell voltage, E°, of 2.12 V. Which change could make the cell voltage greater than 2.12 V?

a) add more Zn(s)
b) add more $Cl^-(aq)$ ions
c) decrease the concentration of $Zn^{2+}(aq)$ ions
d) decrease the partial pressure of $Cl_2$

## 1997

43. What is the function of $H_2O_2$ in this reaction?

$6H^+ + 2MnO_4^- + 5H_2O_2 \rightarrow 2Mn^{2+} + 5O_2 + 8H_2O$

a) catalyst b) reducing agent
c) oxidizing agent d) inhibitor

44. How much hydrogen is produced from the electrolysis of water in the same time that 2.2 L of oxygen is formed?

a) 0.14 L b) 1.1 L
c) 2.2 L d) 4.4 L

45. Which of these changes will cause the value of the potential for this half-reaction to be less negative? (E° = -0.28 V for the reaction.)

$Co^{2+}(aq) + 2e^- \rightarrow Co(s)$

a) increasing the amount of solid Co
b) decreasing the amount of solid Co
c) increasing the concentration of $Co^{2+}(aq)$
d) decreasing the concentration of $Co^{2+}(aq)$

## 1996

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

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

a) $V^{2+} + 2Ag \rightarrow V + 2Ag^+$
b) $V^{2+} + Pb \rightarrow V + Pb^{2+}$
c) $2Ag^+ + Pb^{2+} \rightarrow 2Ag + Pb$
d) $2Ag^+ + Pb \rightarrow 2Ag + Pb^{2+}$

44. It is possible to produce chlorine gas by electrolyzing any of these chlorine-containing compounds under the proper conditions. Which compound will require the smallest number of coulombs to produce one mole of chlorine?

a) $Ca(OCl)_2$ b) $NaClO_2$
c) $KClO_3$ d) $Mg(ClO_4)_2$

## 1994

46. If solid nickel metal were added to separate aqueous solutions each containing 1M concentrations of $Ag^+$, $Cd^{2+}$, and $Sn^{2+}$ ions, how many 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 |

a) zero b) one
c) two d) three

48. Solutions of $Ag^+$, $Cu^{2+}$, $Fe^{3+}$ and $Ti^{4+}$ are electrolyzed with a constant current until 0.10 mol of metal is deposited. Which will require the greatest length of time?

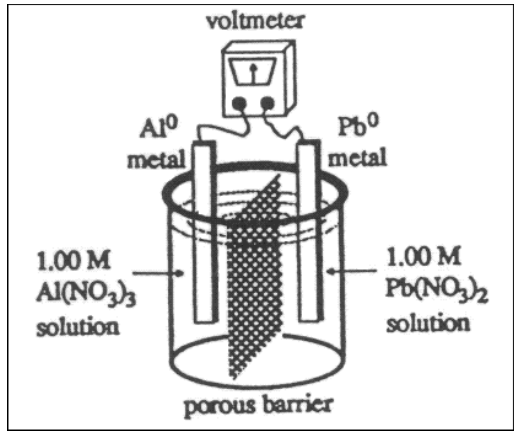
a) $Ag^+$ b) $Cu^{2+}$
c) $Fe^{3+}$ d) $Ti^{4+}$

## 1993

67. How many grams of cobalt metal will be deposited when a solution of cobalt(II) chloride is electrolyzed with a current of 10. amperes for 109 minutes?

a) 0.66 b) 4.0
c) 20 d) 40

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




Reduction Potentials

$Pb^{2+} + 2e^- \rightarrow Pb$ -0.13 V

$Al^{3+} + 3e^- \rightarrow Al$ -1.68 V

a) 2.97V b) 1.55V
c) -1.81V d) -2.97V

## 1992

59. A spoon is made the cathode in an electroplating apparatus containing a $AgNO_3$ solution. How many grams of Ag will be plated on the spoon if a current of 2.00 A is passed through the apparatus for 1.90 min.?

a) 0.255 g b) 0.150 g
c) 0.128 g d) 0.0638 g

60. A cell is set up using the following reactions:

$Zn \mid Zn^{2+}$ (0.5M) $\parallel Ni^{2+}$ (0.1 M) $\mid Ni$

$Ni^{2+} + 2e^- \rightarrow Ni$ E° = -0.250 V

$Zn^{2+} + 2e^- \rightarrow Zn$ E° = -0.763 V

What is the voltage of the cell?

a) -0.513 V b) -1.013 V
c) 0.492 V d) 0.513 V

## Answers:

| | |
|---|---|
| 1998 | 40 a, 41 c, 42 d, 43 c |
| 1997 | 43 b, 44 d, 45 c |
| 1996 | 43 d, 44 a |
| 1994 | 46 c, 48 d |
| 1993 | 67 c, 66 b |
| 1992 | 59 a, 60 c |

# Ch 21 – Electron Transfer Reactions

Selected NChO Problems

| | |
|---|---|
| 1998-40. | **(A) +0.54 V**<br>E°(cell) = $E°_{red}$(reduction) - $E°_{red}$(oxidation)<br>0.79 = 1.33 – x<br>x = 1.33 – 0.79 = +0.54 volts |
| 1998-41. | **(C) $O_2$(g)**<br>anode = oxidation (lose electrons "LeO") (+ electrode removes electrons from chemicals)<br>$NO_3^-$ and $H_2O$ are at the + electrode<br>$NO_3^-$ is already oxidized. O in $H_2O$ is oxidized. $2H_2O + 4e^- \rightarrow O_2 + 4H^+$ |
| 1998-42. | **(D) $Cu^+$**<br>best **reducing agent** = most easily **oxidized = smallest E° value**<br>look at **product** in equation with the **smaller E°** value<br>$Cu^{2+} + e^- \rightarrow$ **$Cu^+$** |
| 1998-43. | **(C) decrease the concentration of $Zn^{2+}$(aq) ions**<br>anything that drives this reaction **forward** will **increase** the cell's voltage… Le Châtelier's<br>~~add more Zn(s)~~ [adding a **solid** will not shift the equilibrium]<br>~~add more $Cl^-$ ions~~ [this will increase the **reverse** reaction]<br>decrease $Zn^{2+}$ ions [this is the answer]<br>~~decrease pressure of $Cl_2$(g)~~ [this will increase the **reverse** reaction] |
| 1997-43. | **(B) reducing agent**<br>if $H_2O_2$ is getting **reduced** it is an **oxidizing agent**<br>if $H_2O_2$ is getting **oxidized**, it is a **reducing agent**<br>$H_2O_2 \rightarrow O_2$ is **oxidation**; O's oxidation number changes from 1- to 0<br>a **catalyst** would have been written over the arrow.<br>I don't know how you would recognize an **inhibitor**.<br>catalyst and inhibitor are "distractors" …wrong answers |
| 1997-44. | **(D) 4.4 L**<br>Know that electrolysis of water is **$2H_2O$(l) ® $2H_2$(g) + $O_2$(g)**<br>if you make 2.2 L of $O_2$, you make **twice as much** $H_2$(g) |
| 1997-45. | **(C) increasing the concentration of $Co^{2+}$(aq)**<br>be careful of convoluted wording… "less negative = more positive"<br>anything that drives this reaction **forward** will **increase** the E° of the half-reaction.<br>increasing or decreasing the **solid Co** will have no effect |
| 1996-43. | **(D) 2 $Ag^+$ + Pb ® 2 Ag + $Pb^{2+}$**<br>when the E° values are written in decreasing order (most + to most -) the<br>**upper-left & lower-right rule** applies… $Ag^+$ reacts with Pb° and V°, etc.<br>$Pb^{2+}$ reacts with V° but not Ag° |
| 1996-44. | **(a) $Ca(OCl)_2$**<br>You need to look at the half-reaction forming $Cl_2$(g).<br>The reaction with the **least number of moles of $e^-$'s** is the answer.<br>$Ca(OCl)_2$ — Cl has oxidation # of +1 in $OCl^-$ — $2 OCl^- + 2e^- \rightarrow Cl_2 + 2 O^{2-}$<br>$NaClO_2$ — Cl has oxidation # of +3 in $ClO_2^-$ — $2 ClO_2^- + 6e^- \rightarrow Cl_2 + 4 O^{2-}$<br>$KClO_3$ — Cl has oxidation # of +5 in $ClO_3^-$ — $2 ClO_3^- + 10e^- \rightarrow Cl_2 + 6 O^{2-}$<br>$Mg(ClO_4)_2$ — Cl has oxidation # of +7 in $ClO_4^-$ — $2 ClO_4^- + 14e^- \rightarrow Cl_2 + 8 O^{2-}$ |
| 1994-46. | **(C) two**<br>The chart is in order of decreasing reduction potentials (E° values) so the upper-left lower-right rule applies. $Ag^+$ reacts with Ni°, $Sn^{2+}$ reacts with Ni° |

| | |
|---|---|
| 1994-48. | **(D) $Ti^{4+}$**<br>You need to visualize the half-reactions for each metal.<br>The reaction with the most electrons involved will require the greatest length of time.<br>$Ag^+ + e^- \rightarrow Ag°$ $Cu^{2+} + 2e^- \rightarrow Cu°$ $Fe^{3+} + 3e^- \rightarrow Fe°$ $Ti^{4+} + 4e^- \rightarrow Ti°$ |
| 1993-67. | **(C) 20**<br>This is a line equation. Remember to begin with "amps x time"<br>$10\,\text{amps} \times 109\,\text{min} \times \frac{60\,\text{sec}}{1\,\text{min}} \times \frac{1\,\text{Couloumb}}{1\,\text{amp} \bullet \text{sec}} \times \frac{1\,\text{mole e}^-}{96{,}500\,\text{C}} \times \frac{1\,\text{mole Co}}{2\,\text{mole e}^-} \times \frac{58.93\,\text{g Co}}{1\,\text{mole Co}} = 19.969\,\text{g Co}$ |
| 1993-66. | **(B) 1.55 V**<br>E°(cell) = E°$_{red}$(reduction) - E°$_{red}$(oxidation)<br>= -0.13 volts –(-1.68 volts) = +1.55 volts<br>Note: The overall reaction is $3Pb^{2+} + 2Al \rightarrow 3Pb + 2Al^{3+}$ (6 moles of electrons are involved) but the coefficients of 2 and 3 are not used for this calculation… not like Hess's Law. |
| 1992-59. | **(A) 0.255 g**<br>This is a line equation. Remember to begin with "amps x time"<br>$2\,\text{amps} \times 1.9\,\text{min} \times \frac{60\,\text{sec}}{1\,\text{min}} \times \frac{1\,\text{Couloumb}}{1\,\text{amp} \bullet \text{sec}} \times \frac{1\,\text{mole e}^-}{96{,}500\,\text{C}} \times \frac{1\,\text{mole Ag}}{1\,\text{mole e}^-} \times \frac{107.9\,\text{g Ag}}{1\,\text{mole Ag}} = 0.2549\,\text{g Ag}$ |
| 1992-60. | **(C) 0.492 V**<br>IF this were a STANDARD cell (everything 1 M) the equation would be:<br>E°(cell) = E°$_{red}$(reduction) - E°$_{red}$(oxidation)<br>-0.250 volts – (-.763 volts) = +0.513 volts (answer D)<br>However…<br>The shortcut cell notation (anode \|\| cathode) shows that the $[Zn^{2+}]$ is only 0.5 M.<br>The overall reaction is: $Zn^{2+} + Ni° \rightarrow Zn° + Ni^{2+}$<br>Since $[Zn^{2+}]$ is reduced, the cell will run at a little lower voltage… best answer is 0.492 volts.<br>You can also calculate this answer using the **Nernst equation**. |