Recitation Worksheet (Optional Extra Practice)

Name:

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Textbook:

Chemistry & Chemical Reactivity
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Instructions:

- This recitation worksheet is optional extra practice for 19.2-19.4
- You **do not** need to submit it to Gradescope.
- The answer key has been posted with this worksheet to eLC.
- A periodic table and formula sheet are attached to the end of this worksheet.

1. For the reaction:

   $$3\ Pt(s) + 12\ Cl^-(aq) + 2\ NO_3^-(aq) + 8\ H^+(aq) \longrightarrow 3\ [PtCl_4]^{2-}(aq) + 2\ NO(g) + 4\ H_2O(l) \quad E^\circ_{cell} = 0.201\ V$$

   What is E° for the reduction half-cell reaction of $[PtCl_4]^{2-}$ to Pt in acidic solution? Please refer to the table below for additional information. Keep your answers to two significant figures.

   V

Standard Reduction Potentials at 25°C (298 K) for Many Common Half-reactions

Half-reaction	$\mathscr{E}^\circ$ (V)	Half-reaction	$\mathscr{E}^\circ$ (V)
$F_2 + 2e^- \rightarrow 2F^-$	2.87	$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$	0.40
$Ag^{2+} + e^- \rightarrow Ag^+$	1.99	$Cu^{2+} + 2e^- \rightarrow Cu$	0.34
$Co^{3+} + e^- \rightarrow Co^{2+}$	1.82	$Hg_2Cl_2 + 2e^- \rightarrow 2Hg + 2Cl^-$	0.27
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	1.78	$AgCl + e^- \rightarrow Ag + Cl^-$	0.22
$Ce^{4+} + e^- \rightarrow Ce^{3+}$	1.70	$SO_4^{2-} + 4H^+ + 2e^- \rightarrow H_2SO_3 + H_2O$	0.20
$PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \rightarrow PbSO_4 + 2H_2O$	1.69	$Cu^{2+} + e^- \rightarrow Cu^+$	0.16
$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$	1.68	$2H^+ + 2e^- \rightarrow H_2$	0.00
$IO_4^- + 2H^+ + 2e^- \rightarrow IO_3^- + H_2O$	1.60	$Fe^{3+} + 3e^- \rightarrow Fe$	−0.036
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51	$Pb^{2+} + 2e^- \rightarrow Pb$	−0.13
$Au^{3+} + 3e^- \rightarrow Au$	1.50	$Sn^{2+} + 2e^- \rightarrow Sn$	−0.14
$PbO_2 + 4H^+ + 2e^- \rightarrow Pb^{2+} + 2H_2O$	1.46	$Ni^{2+} + 2e^- \rightarrow Ni$	−0.23
$Cl_2 + 2e^- \rightarrow 2Cl^-$	1.36	$PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$	−0.35
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1.33	$Cd^{2+} + 2e^- \rightarrow Cd$	−0.40
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	1.23	$Fe^{2+} + 2e^- \rightarrow Fe$	−0.44
$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	1.21	$Cr^{3+} + e^- \rightarrow Cr^{2+}$	−0.50
$IO_3^- + 6H^+ + 5e^- \rightarrow \frac{1}{2}I_2 + 3H_2O$	1.20	$Cr^{3+} + 3e^- \rightarrow Cr$	−0.73
$Br_2 + 2e^- \rightarrow 2Br^-$	1.09	$Zn^{2+} + 2e^- \rightarrow Zn$	−0.76
$VO_2^+ + 2H^+ + e^- \rightarrow VO^{2+} + H_2O$	1.00	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	−0.83
$AuCl_4^- + 3e^- \rightarrow Au + 4Cl^-$	0.99	$Mn^{2+} + 2e^- \rightarrow Mn$	−1.18
$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	0.96	$Al^{3+} + 3e^- \rightarrow Al$	−1.66
$ClO_2 + e^- \rightarrow ClO_2^-$	0.954	$H_2 + 2e^- \rightarrow 2H^-$	−2.23
$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	0.91	$Mg^{2+} + 2e^- \rightarrow Mg$	−2.37
$Ag^+ + e^- \rightarrow Ag$	0.80	$La^{3+} + 3e^- \rightarrow La$	−2.37
$Hg_2^{2+} + 2e^- \rightarrow 2Hg$	0.80	$Na^+ + e^- \rightarrow Na$	−2.71
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77	$Ca^{2+} + 2e^- \rightarrow Ca$	−2.76
$O_2 + 2H^+ + 2e^- \rightarrow H_2O_2$	0.68	$Ba^{2+} + 2e^- \rightarrow Ba$	−2.90
$MnO_4^- + e^- \rightarrow MnO_4^{2-}$	0.56	$K^+ + e^- \rightarrow K$	−2.92
$I_2 + 2e^- \rightarrow 2I^-$	0.54	$Li^+ + e^- \rightarrow Li$	−3.05
$Cu^+ + e^- \rightarrow Cu$	0.52		

2. Which of the reactions do you predict to be **spontaneous** in the forward direction? Assume all the reactants and products in their standard states. Please refer to the table in question 1 for additional information.

A. $Cu^{2+}(aq) + 2I^{-}(aq) \longrightarrow Cu(s) + I_2(s)$
B. $4\ NO_3^{-}(aq) + 4\ H^{+}(aq) \longrightarrow 3\ O_2(g) + 4\ NO(g) + 2\ H_2O(l)$
C. $2\ Br^{-}(aq) + I_2(s) \longrightarrow Br_2(aq) + 2\ I^{-}(aq)$
D. $Au(s) + NO_3^{-}(aq) \longrightarrow Au^{3+}(aq) + NO(g)$ (in acidic solution)
E. All of the above reactions are non-spontaneous in the forward direction.

3. Calculate the $E°_{cell}$ for the following reactions. Please refer to the table in question 6 for additional information.

A. $Cu(s)|Cu^{2+}(aq)||Cu^{+}(aq)|Cu(s)$

[ ] V

B. $Pt(s)|Ce^{4+}(aq), Ce^{3+}(aq)||I^{-}(aq)|I_2(s)|C(s)$

[ ] V

C. $Al(s)|Al^{3+}(aq)||Sn^{2+}(aq)|Sn(s)$

[ ] V

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

[ ]

A. $E°_{cell}$ is positive for spontaneous reactions.
B. Electrons will flow from the positive electrode to the negative electrode in a galvanic cell.
C. $E°_{cell}$ is the difference in voltage between the anode and the cathode, $E°_{cell} = E°_{cell}(\text{anode}) - E°_{cell}(\text{cathode})$.
D. The electrode potential of the standard hydrogen electrode is exactly zero.
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 $E°$.

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

$$Mn(s) \mid Mn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$$

A. $Cu(s) + Mn^{2+}(aq) \rightarrow Mn(s) + Cu^{2+}(aq)$
B. $Mn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Mn^{2+}(aq)$
C. $2\ Mn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + 2\ Mn^{2+}(aq)$
D. $2\ Cu(s) + Mn^{2+}(aq) \rightarrow Mn(s) + 2\ Cu^{2+}(aq)$
E. $3\ Mn(s) + 2\ Cu^{2+}(aq) \rightarrow 2\ Cu(s) + 3\ Mn^{2+}(aq)$

6. Which of the choices represents the correct cell notation for $Sn(s) + 2\ H^+(aq) \longrightarrow Sn^{2+}(aq) + H_2(g)$?

A. $H^+(aq) \mid H_2(g) \mid Pt \parallel Sn(s) \mid Sn^{2+}(aq)$
B. $H_2(g) \mid H^+(aq) \mid Pt \parallel Sn^{2+}(aq) \mid Sn(s)$
C. $Sn^{2+}(aq) \mid Sn(s) \parallel H_2(g) \mid H^+(aq) \mid Pt$
D. $Sn(s) \mid Sn^{2+}(aq) \parallel H^+(aq) \mid H_2(g) \mid Pt$
E. $Sn(s) \mid H_2(g) \parallel Sn^{2+}(aq) \mid H^+(aq) \mid Pt$

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

$Ag^+(aq) + e^- \rightarrow Ag(s)$ $E° = +0.80\ V$
$Pb^{2+}(aq) + 2\ e^- \rightarrow Pb(s)$ $E° = -0.13\ V$

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 Pb^{2+}(aq) is formed at the anode.
C. Pb(s) is formed at the cathode, and Ag^+(aq) is formed at the anode.
D. Pb^{2+}(aq) is formed at the cathode, and Cu(s) is formed at the anode.

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

$$Zn\ (s) + CuBr_2\ (aq) \longrightarrow Cu\ (s) + ZnBr_2(aq)$$
$$Cu\ (s) + ZnBr_2(aq) \longrightarrow \text{No Reaction}$$

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 Cu doesn't react with $ZnBr_2$ proves that copper attracts electrons more than does Zn.
E. None of the above.

9. 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.

$$Cu^{2+}(aq) + 2\ Fe^{2+}(aq) \rightarrow 2\ Fe^{3+}(aq) + Cu(s)$$

$$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)\ E° = +0.77\ V$$
$$Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s)\ E° = +0.34\ V$$

A. $E°$ = - 0.43 V, nonspontaneous.
B. $E°$ = - 0.43 V, spontaneous.
C. $E°$ = + 0.43 V, nonspontaneous.
D. $E°$ = + 0.43 V, spontaneous.

10. 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:

$$Cu(s) + 2\ Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2\ Ag(s)$$

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.
b. The silver electrode increases in mass as the cell operates.
c. There is a net movement of silver ions through the salt bridge from the silver half-cell to the copper half-cell.
d. There is a net movement of copper ions through the salt bridge from the copper half-cell to the silver half-cell.
e. The copper electrode decreases in mass as the cell operates.

11. Determine which of the following pairs of reactants will result in a spontaneous reaction at 25°C. Please refer to the table in question 1 for additional information.

A. $Sn^{2+(}aq) + Zn(s)$

B. $Cr^{3+(}aq) + Sn(s)$

C. $Zn(s) + K^{+}(aq)$

D. $Fe(s) + Ba^{2+}(aq)$

E. None of the above pairs will react.

12. Given the standard reduction potentials, which is the strongest oxidizing agent?

F_2 (g) + 2e⁻ → 2 F^- (aq)	+2.87
Cl_2 (g) + 2e⁻ → 2Cl^- (aq)	+1.359
Br_2 (l) + 2e⁻ → 2Br^- (aq)	+1.065
I_2 (s) + 2e⁻ → 2I^- (aq)	+0.536

A. F_2
B. Cl_2
C. Br_2
D. I_2
E. All have the same oxidizing strengths

13. Which of the following statements is *false* in reference to the two half-reactions in a voltaic cell? Select all that apply.

Half-reaction 1: Zn(s) ⟶ Zn^{2+}(aq) + 2 e^- (electrode = Zn)
Half-reaction 2: ClO_3^-(aq) + 6 H^+(aq) + 6 e^- ⟶ Cl^-(aq) + 3 H_2O(l) (electrode = Pt)

A. Half-reaction 2 is the anode and half-reaction 1 is the cathode
B. The Zn electrode in half-reaction 1 increases in mass due to oxidation of Zn(s) to Zn^{2+} ions
C. The Pt electrode in half-reaction 2 does not lose nor gain mass during the reaction
D. Pt in half-reaction 2 is the cathode
E. Pt in half-reaction 2 is the anode
F. All the above statements are true
G. All the above statements are false

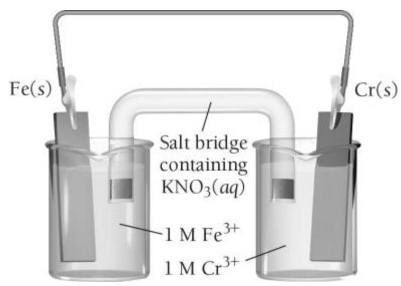
14. You are provided with a table of standard electrode potentials for a series of hypothetical reactions in aqueous solutions:

Reduction Half-Reaction	E° (V)
$A^{+}(aq) + e^{-} \longrightarrow A(s)$	1.33
$B^{2+}(aq) + 2e^{-} \longrightarrow B(s)$	0.87
$C^{3+}(aq) + e^{-} \longrightarrow C^{2+}(aq)$	−0.12
$D^{3+}(aq) + 3e^{-} \longrightarrow D(s)$	−1.59

Based on this information, which of the following statements is *true*?

A. A^{+}(aq) is the strongest oxidizing agent
B. D^{3+}(aq) is the weakest oxidizing agent
C. D(s) is the strongest reducing agent
D. A(s) is the weakest reducing agent
E. C^{2+} can be oxidized by either B^{2+}(aq) or A^{+}(aq)
F. All the above statements are true

15. Refer to the picture of the voltaic cell below to answer the next set of questions. You can reference the table in question 1 for more information.



A. Which electrode represents the anode?

B. Which electrode represents the cathode?

C. What is E°_{cell}?

V

D. Cations from the salt bridge flow towards the ______ and the anions flow towards the ______

Formula Sheet

Length
1 kilometer = 0.62137 mile
1 inch = 2.54 centimeters (exactly)
1 Ångstrom = 1×10^{-10} meter

Energy
1 joule = 1 $kg \cdot m^2/s^2$
1 calorie = 4.184 joules
1 Calorie = 1 kilocalorie = 1000 calories
1 L·atm = 101.325 joules

Pressure
1 pascal = 1 N/m^3 = 1 $kg/m \cdot s^2$
1 atmosphere = 101.325 kilopascals = 760 mm Hg = 760 torr = 14.70 lb/in^2
1 bar = 1×10^5 Pa (exactly)

Temperature
0 K = −273.15°C
K = °C + 273.15
°C = (5/9)(°F − 32)

Mass
1 kg = 2.205 lbs

Volume
1 mL = 1 cm^3 = 1 cc

Constants
c = 2.998×10^8 m/sec
h = 6.626×10^{-34} $J \cdot sec^{-1}$
R = 0.08206 L·atm/mol·K = 8.314 J/mol·K
Specific heat of water = 4.184 J/g·K
Mass of an electron: 9.109×10^{-31} kg
Mass of a proton: 1.673×10^{-27} kg
RH = 2.18×10^{-18} J
Specific heat of water = 4.184 J/g·K
STP = 273.15 K and 1 atm
Avogadro's number: 6.022×10^{23}

Equations
d (density) = m/V
$P_1V_1=P_2V_2$
$V_1/T_1=V_2/T_2$
$P_1V_1/n_1T_1=P_2V_2/n_2T_2$
$PV=nRT$
$(P + a(n^2/V^2))\cdot(V - nb) = nRT$
molar mass (M) = mRT/PV
density (d) = MP/RT
$x_A = n_A/n_{tot} = P_A/P_{tot} = V_A/V_{tot}$
$P_{tot} = P_A + P_B + \ldots$
$n_{tot} = n_A + n_B + \ldots$

$$\mu_{rms} = \sqrt{\frac{3RT}{M}}$$

$$\frac{Rate\ of\ effusion\ A}{Rate\ of\ effusion\ B} = \sqrt{\frac{MW_B}{MW_A}}$$

$Q = C \times \Delta T = c_{specific} \times m \times \Delta T$
$Q = n \times \Delta H$ (kJ/mol) $= m \times \Delta H$ (kJ/g)
$w = -P\Delta V$
$\Delta E = q + w$
$\Delta H° = \Sigma n\Delta Hf°$(products) $- \Sigma n\Delta Hf°$(reactants)
$\Delta H° = \Sigma n\Delta H°$(bonds broken) $- \Sigma n\Delta H°$(bonds formed)
$E=h\nu$
$c=\lambda\nu$
$\lambda = h/mv$

$$\Delta E = -2.18 \times 10^{-18} J\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_{vap}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$C_g = kP_g$
$P_{solution}=P_{solvent}\, X_{solvent}$
$P_{solution}=\sum P_j=\sum P_jX_j$
$\Delta T_b = K_b m i$
$\Delta T_f = K_f m i$
$\pi = MRTi$

Thermodynamic and Electrochemistry

$S = k_b \times \ln(W)$
$k_b = 1.381 \times 10^{-23}$ J/K
$\Delta S = q_{rev}/T$
$\Delta S_{surr} = q_{surr}/T = -q_{rev}/T$
$\Delta S_{univ} = \Delta S_{sys} + \Delta S_{surr}$
$\Delta S°_{rxn} = \Sigma v S°_{products} - \Sigma v S°_{reactants}$
$\Delta H°_{rxn} = \Sigma v H°_{products} - \Sigma v H°_{reactants}$
$\Delta G°_{rxn} = \Sigma v G°_{products} - \Sigma v G°_{reactants}$
$\Delta G = \Delta H - T\Delta S$
$\Delta G = \Delta G° + RT\cdot\ln Q$
R=8.314 J/mol.K
$\Delta G° = -RT\cdot\ln K$
$\Delta G = -nFE_{cell}$
F = 96485 J/(V·mol e^-)
$E°_{cell} = RT/nF \ln K$
$E°_{cell} = (0.0257/n) \ln K = (0.0592/n) \log K$
$E_{cell} = E°_{cell} - (RT/nF) \ln Q$
$E_{cell} = E°_{cell} - (0.0257/n) \ln Q$
Electrolysis: Q (total charge) = I × t = n × F

Integrated Rate Laws & half-life

$$\ln\frac{[A]}{[A]_0} = -kt$$

$$\frac{1}{[A]} = kt + \frac{1}{[A]_0}$$

$$[A] = -kt + [A]_0$$

$$t_{1/2} = \frac{[A]_0}{2k}$$

$$t_{1/2} = \frac{\ln 2}{k} = \frac{0.693}{k}$$

$$t_{1/2} = \frac{1}{k[A]_0}$$

$$ln\frac{k_2}{k_1} = -\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

<u>Equilibrium and Acid / Base</u>

$K_P = K_c \times (RT)^{\Delta n}$
pH = $-\log[H_3O^+]$
$K_w = 1.0 \times 10^{-14}$ at 25 °C
$K_w = [H_3O^+] \times [OH^-]$
$K_w = K_a \times K_b$
$pK_a = -\log[K_a]$
Buffer: $pH = pK_a + \log\frac{[A^-]}{[HA]}$

$$ln\frac{K_2}{K_1} = \frac{\Delta H^\circ_{rxn}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

Periodic Table of the Elements

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.01																	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.97	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc [97]	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
37 Cs 132.91	56 Ba 137.33		72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po [209]	85 At [210]	86 Rn [222]
87 Fr [223]	88 Ra [226]		104 Rf [267]	105 Db [268]	106 Sg [269]	107 Bh [270]	108 Hs [269]	109 Mt [277]	110 Ds [281]	111 Rg [282]	112 Cn [285]	113 Nh [286]	114 Fl [290]	115 Mc [290]	116 Lv [293]	117 Ts [294]	118 Og [294]

57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm [145]	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97
89 Ac [227]	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np [237]	94 Pu [244]	95 Am [243]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]	103 Lr [262]