Exam 4A
Chem 1142
Spring 2019

Name: KEY

MULTIPLE CHOICE. [2.5 pts ea.] Record the best response on the scantron sheet. [**50 pts total.**]

Assume all solutions are aqueous and at a temperature of 25 °C, unless stated otherwise.

Q1. The second law of thermodynamics states that:
A) The entropy of the reaction always increases
B) The entropy of the universe always increases
C) The entropy of the surroundings always increases
D) The entropy of the system always increases

Q2. Which substance would we expect to have the greatest molar entropy at 25 °C?
A) $NaF(s)$
B) $N_2(g)$
C) $H_2O(l)$
D) C(s, graphite)

Q3. Which chemical equation would most likely have $\Delta S°_{rxn} < 0$?
A) $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$
B) $Na(s) + ½\ Br_2(l) \longrightarrow NaBr(s)$
C) $H_2(g) + ½\ O_2(g) \longrightarrow H_2O(g)$
D) $C(s, graphite) + O_2(g) \longrightarrow CO_2(g)$

Q4. A chemical reaction has $\Delta H_{rxn} = -5.00$ kJ at 250 K. What will the entropy change of the **surroundings** be?
A) + 20 J/K
B) – 20 J/K
C) + 1300 kJ·K
D) – 1300 kJ·K

$\Delta S_{surr} = \frac{q}{T} = \frac{+5{,}000 J}{250 K} = +20$ J/K

$\Delta H = +ve$ $\Delta S = +ve$.

$\Delta G = \underset{+ve}{\Delta H} - \underbrace{T\overset{+ve}{\Delta S}}_{-ve}$

Q5. An endothermic reaction has $\Delta S_{rxn} > 0$. What can you say about its spontaneity?
A) It will always be spontaneous
B) It will always be non-spontaneous
C) It will be spontaneous at low temperatures, but non-spontaneous at high temperatures
D) It will be non-spontaneous at low temperatures, but spontaneous at high temperatures

Q6. When will the entropy of a substance be zero?
A) When it is an ion at a molar concentration of 1 M
B) When it is an element in its standard state at 25 °C
C) When it is a substance under 1 atm pressure at its normal melting point
D) When it is a perfect crystal at 0 K

$S = k \cdot \ln W$

#microstates.

non-spont std. cond. spont under

Q7. A chemical reaction has $\Delta G^o > 0$ and $\Delta G < 0$. This means that:
A) It is spontaneous under standard conditions, but non-spontaneous under current conditions
B) It is non-spontaneous under standard conditions, but spontaneous under current conditions
C) It will always be non-spontaneous under any condition
D) It will always be spontaneous under any condition

Q8. A chemical reaction with $\Delta G^o \ll 0$ will likely have an equilibrium constant, K, such that:
A) $K \gg 1$ very spont $A \gtrless B$ $K \sim \frac{P}{R}$.
B) $K \ll 1$
C) $K = 1$
D) $K = 0$

Q9. Which of the following chemical equations corresponds to the standard Gibbs energy of formation of $NH_3(g)$?
A) $NH_3(g) + O_2(g) \longrightarrow NO_2(g) + {}^3/_2\, H_2(g)$
B) $2NH_3(g) \longrightarrow N_2(g) + 3\, H_2(g)$
C) $\frac{1}{2}\, N_2(g) + {}^3/_2\, H_2(g) \longrightarrow NH_3(g)$ elements $\longrightarrow NH_3(g)$
D) $NH_2^-(aq) + H^+(aq) \longrightarrow NH_3(g)$

Q10. Which pair of chemical equations, when coupled (added), give rise to a spontaneous reaction?

(i)	$\mathbf{A \longrightarrow B}$	$\mathbf{\Delta G^o = -20\ kJ}$
(ii)	$\mathbf{B \longrightarrow C}$	$\mathbf{\Delta G^o = +25\ kJ}$
(iii)	$\mathbf{A \longrightarrow C}$	$\mathbf{\Delta G^o = +10\ kJ}$
(iv)	$\mathbf{B \longrightarrow D}$	$\mathbf{\Delta G^o = -15\ kJ}$

A) (ii) & (iii) +35
B) (ii) & (iv) +10
C) (i) & (ii) +5
D) (i) & (iii) -10

Q11. Which of the following half-reactions is properly balanced?
A) $Br_2(l) \longrightarrow 2Br^-(aq)$ ×
B) $2H^+(aq) + e^- + IO^+(aq) \longrightarrow H_2O(l) + \frac{1}{2}\, I_2(s)$ ×
C) $2e^- + 2Cl^-(aq) \longrightarrow Cl_2(g)$ ×
D) $2e^- + H^+(aq) + Pb(OH)_3^+ \longrightarrow Pb(OH)_2 + H_2O(l)$ ✓

Q12. Where does reduction take place in a voltaic (galvanic) cell?
A) The salt bridge
B) The voltmeter red cat
C) The cathode an ox
D) The anode

Q13. Given the following two electrode potentials:
$\mathbf{Cu^{2+} + 2e^- \longrightarrow Cu}$ $\quad E^o = +0.34\ V$
$\mathbf{Zn^{2+} + 2e^- \longrightarrow Zn}$ $\quad E^o = -0.76\ V$ reduced: gain e^-s.
The best reducing agent would be:
A) Cu^{2+} makes something
B) Cu gain e^-
C) Zn^{2+}
D) Zn

anode, ox cat, red.

$Cr \rightarrow Cr^{3+} + 3e^-)\times 2$

$(2e^- + Mg^{2+} \rightarrow Mg\quad)\times 3$

$2Cr + 3Mg^{2+} + 6e^- \rightarrow 2Cr^{3+} + 3Mg$

Q14. What is the cell reaction for the voltaic cell:
$Cr(s)|Cr^{3+}(aq) || Mg^{2+}(aq)|Mg(s)$
A) $2Cr(s) + 2Cr^{3+}(aq) \longrightarrow 3Mg^{2+}(aq) + 3Mg(s)$
B) $2Cr(s) + 3Mg^{2+}(aq) \longrightarrow 2Cr^{3+}(aq) + 3Mg(s)$
C) $2Cr^{3+}(aq) + 3Mg(s) \longrightarrow 2Cr(s) + 3Mg^{2+}(aq)$
D) $3Mg^{2+}(aq) + 2Cr^{3+}(aq) \longrightarrow 3Mg(s) + 2Cr(s)$

Q15. Which of the following statements is true about a voltaic (galvanic) cell that has an $E°$ value of -0.50 V?

non-spont.

(i) The reaction is spontaneous ×
(ii) $K > 1$ ×
(iii) $\Delta G° > 0$ ✓

A) (i) & (ii)
B) (iii) only
C) (ii) & (iii)
D) (i) only

Q16. How would E_{cell} change if the concentration of $Ag^+(aq)$ was **increased** in the following voltaic (galvanic) cell?

$Ag(s)|Ag^+(aq) || Au^{3+}(aq)|Au(s)$

$(Ag \rightarrow Ag^+ + e^-)\,3$

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

$3Ag(s) + Au^{3+}(aq) \xrightarrow{3e^-} 3Ag^+(aq) + Au(s)$

$\downarrow E = E° - \frac{RT}{nF}\ln Q \uparrow$

$Q = \frac{[Ag^+]^3 \uparrow}{[Au^{3+}]}$

A) It would increase
B) It would not change
C) It would decrease
D) There is not enough information to decide

Q17. The reaction $MgCl_2 \longrightarrow Mg + Cl_2$ has a cell potential (E_{cell}) of -3.5 V. This means that:
A) The reaction is spontaneous and will produce 3.5 V in a cell
B) The reaction is non-spontaneous, and will produce 3.5 V in a cell
C) The reaction can be driven by applying a voltage of > 3.5 V ✓
D) The reaction is spontaneous, but can be stopped by applying a voltage of < -3.5 V

Q18. How many moles of electrons are produced when 12.0 A flows for 24.0 min?
A) 288 mol
B) 5.58 mol
C) 0.500 mol
D) 0.179 mol

C/s

$I = \frac{Q}{t}$ $\quad Q = I \cdot t$

$= 12.0\frac{C}{s} \times 24.0\,\text{min} \times \frac{60s}{1\,\text{min}}$

$= 17{,}280C$

Q19. Copper can be plated out of a solution containing Cu^{2+} according to the half-reaction:
$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$
What mass of copper is formed when 15.0 A of current flows for 10.0 min?
A) 2.96 g
B) 5.92 g
C) 8.15 g
D) 12.6 g

$1F = 96{,}500\frac{C}{mol}$

$0.093\,mol\,e^- \times \frac{1\,mol\,Cu}{2\,mol\,e^-} \times \frac{63.55g\,Cu}{1\,mol\,Cu} = 2.96g\,Cu$

Q20. Which electrode is assigned a standard electrode potential of 0.00 V?
A) The standard hydrogen electrode
B) The standard pH electrode
C) The standard acidic electrode
D) The standard platinum electrode

Short Response.

Show ALL work to receive credit.

Q21. [12 pts.] Using the half-reaction method, balance the following redox reaction that occurs in **basic** solution:

(0) (+5)(−2) (+3) (−3)(+1)

$$Al + NO_3^- \longrightarrow Al(OH)_4^- + NH_3$$

−2 +1

$+\frac{1}{2}$ ea 3

Be sure to identify oxidation states of each atom as part of your answer. Clearly indicate whether each half-reaction is an oxidation or reduction process.

ox: $\left(4H_2O + \overset{(0)}{Al} \longrightarrow \overset{(+3)}{Al(OH)_4^-} + 4H^+ + 3e^- \right) \times 8$

(1. coeff) 2. H_2O 3. H^+ } +3 3 4. e^-

+1

cancel e^- 1

red: $\left(8e^- + 9H^+ + \overset{(+5)}{NO_3^-} \longrightarrow \overset{(-3)}{NH_3} + 3H_2O \right) \times 3$

(1. coeff) 2. H_2O 3. H^+ } +3 3 4. e^-

18 $\cancel{23}$ 5

$\cancel{32}H_2O + 8Al + \cancel{27}H^+ + 3NO_3^- \xrightarrow{24e^-} 8Al(OH)_4^- + \cancel{32}H^+ + 3NH_3 + \cancel{9H_2O}$

$+5OH^-$ $\cancel{+5OH^-}$

+1 OH^- 1

$\cancel{5H_2O}$

+1 balance 1

$18H_2O + 8Al + 3NO_3^- + 5OH^- \longrightarrow 8Al(OH)_4^- + 3NH_3$ ✓

ox → red →

Q22. [13 pts.] Consider the voltaic cell

Zn(s)|Zn^{2+}(aq)||Cr^{3+}(aq)|Cr(s)

(i) Write the half-cell reactions and the overall cell reaction.

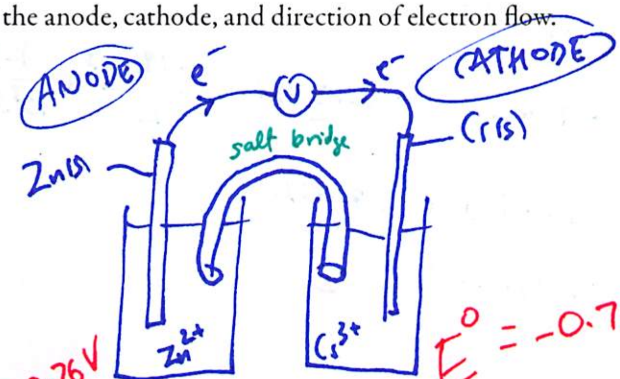
$$(Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-) \times 3$$

$$(3e^- + Cr^{3+}(aq) \rightarrow Cr(s)) \times 2$$

$$3Zn(s) + 2Cr^{3+}(aq) \xrightarrow{6e^-} 3Zn^{2+}(aq) + 2Cr(s).$$

3

(ii) Make a sketch of this voltaic cell and label it. At a minimum, be sure to include labels showing the anode, cathode, and direction of electron flow.



$E^\circ = -0.76V$ $E^\circ = -0.73V$

½ ea
anode
cathode
wire w/ e^-
salt bridge
soln
sketch (gen)

3

(iii) Calculate E°_{cell}. $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = (-0.73V) - (-0.76V) = +0.03V$

3

(iv) What is E_{cell} if $[Zn^{2+}] = [Cr^{3+}] = 0.10$ M? Assume a temperature of 298 K.

$$E = E^\circ - \frac{RT}{nF} \cdot \ln Q \quad +1$$

$$Q = \frac{[Zn^{2+}]^3}{[Cr^{3+}]^2} = \frac{0.10^3}{0.10^2} = 0.10 \quad +1$$

$$= 0.03V - \frac{8.3145 J/mol \cdot K \times 298K}{6 \times 96{,}500 C/mol} \cdot \ln(0.10)$$

+1

4

$$= 0.03V - -0.01V = 0.04V$$

+1

Q23. [12 pts.] For the chemical reaction:

$$\mathbf{H_2O(g) + Cl_2(g) \longrightarrow 2HCl(aq) + \tfrac{1}{2}\,O_2(g)}$$

(i) Calculate ΔH^o, ΔS^o, and ΔG^o at 25 °C.

(Hint: be sure to look at the data table at the bottom of this question!)

$\Delta H^\circ = -92.6\,kJ$ 2

$\Delta S^\circ = -196.3\,J/K$ 2

$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$= -92.6\,kJ - 298.15\,K \times \frac{-196.3\,J}{K} \times \frac{kJ}{10^3\,J}$

$= -92.6\,kJ + 58.5\,kJ$

$= -34.1\,kJ$. 2

6

(ii) Use this information to calculate the equilibrium constant at 25 °C.

$\Delta G^\circ = -RT\ln K$

$\Rightarrow \ln K = \frac{-\Delta G^\circ}{RT} = \frac{+34{,}100\,J/mol}{8.3145\,J/mol\cdot K \times 298.15\,K} = 13.7$ 2

$\Rightarrow K = e^{13.7} = 9.3 \times 10^5$ 1

3

(iii) Predict how the spontaneity of this reaction will change (if at all!) at low vs. high temperature. Be sure to explain your answer!

Will be spont @ low T 1

but non-spont @ high T 1

because. $\underbrace{\Delta H}_{-} \underbrace{- T \underbrace{\Delta S}_{-}}_{+}$ 1

as T ↑, $-T\Delta S$ dominates, making ΔG +ve

3

Substance	ΔH_f^o (kJ/mol)	S^o (J/mol·K)
$H_2O(g)$	−241.8	188.8
$Cl_2(g)$	0	223.1
$HCl(aq)$	−167.2	56.5
$O_2(g)$	0	205.2

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Q24. [13 pts.] Consider the reaction:

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$$

Calculate ΔG_{rxn} at 25 °C under each of the following conditions:
(Hint: be sure to look at the data table at the bottom of this question!)

i) standard conditions

$$\Delta G^\circ = \sum n_p \cdot \Delta G_f^\circ(P) - \sum n_r \cdot \Delta G_f^\circ(R) = \left[1\,mol \times -162.3\,\frac{kJ}{mol}\right] - \left[1\,mol \times -137.2\,\frac{kJ}{mol} + 0\right]$$

$$= \boxed{-25.1\,kJ}$$

4 (4

ii) at equilibrium

$$\boxed{=0}$$

4 (4

iii) under conditions of $p_{CH_3OH} = 1.5$ atm, and $p_{CO} = p_{H_2} = 0.020$ atm.

$$\Delta G = \Delta G^\circ + RT\ln Q$$ 1

$$= -25.1\,\frac{kJ}{mol} + 8.3145\,\frac{J}{mol \cdot K} \times \frac{1\,kJ}{1000\,J} \times 298.15\,K \times \ln \frac{P_{CH_3OH}}{P_{CO} \times P_{H_2}^2}$$ 1

2

$$= -25.1\,\frac{kJ}{mol} + 2.478968\,\frac{kJ}{mol} \cdot \ln \frac{1.5}{0.020 \times 0.020^2}$$

(5

1

$$= -25.1\,\frac{kJ}{mol} + 30.1\,\frac{kJ}{mol} = \boxed{+5.0\,\frac{kJ}{mol}}$$

Substance	ΔG_f° (kJ/mol)
$CH_3OH(g)$	−162.3
$H_2(g)$	0
$CO(g)$	−137.2

/13

Useful Information

Periodic Table of the Elements

IA 1	IIA 2	3	4	5	6	7	8	9	10	11	12	IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	VIIIA 18
1 H 1.008																	2 He 4.003
3 Li 6.941	4 Be 9.012											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 32.07	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 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92160	34 Se 78.96	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 95.94	43 Tc [98]	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.60	53 I 126.9	54 Xe 131.3
55 Cs 132.9	56 Ba* 137.3	71 Lu 175.0	72 Hf 178.5	73 Ta 180.9	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po [210]	85 At [210]	86 Rn [222]
87 Fr [223]	88 Ra** [226]	103 Lr [262]	104 Rf [261]	105 Db [262]	106 Sg [266]	107 Bh [264]	108 Hs [265]	109 Mt [268]	110 [269]	111 [272]	112 [277]	113	114 [285]	115	116 [289]	117	118 [293]

*	57 La 138.9	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm [145]	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.50	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0
**	89 Ac [227]	90 Th 232.0	91 Pa 231.0	92 U 238.0	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]

$N_A = 6.022 \times 10^{23}\ \text{mol}^{-1}$

$$R = 8.3145 \frac{\text{J}}{\text{mol}\cdot\text{K}} = 0.08206 \frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$$

$M_1V_1 = M_2V_2$

$\Delta G = -nFE_{cell}$

$\Delta G° = -nFE°_{cell}$

$$E^{o}_{cell} = \frac{RT}{nF}\ln K$$

$$E_{cell} = E^{\circ}_{cell} - \frac{RT}{nF}\ln Q$$

$E°_{cell} = E°_{cathode} - E°_{anode}$

$F = 96{,}500\ \text{C/mol e}^-$

$1\ \text{V} = 1\ \text{J/C}$

$R = 8.3145\ \text{J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$

$Q\,(\text{charge}) = I\cdot t$

$\Delta G = \Delta H - T\Delta S$

$\Delta S_{surr} = q_{surr}/T$

$\Delta G = \Delta G° + RT\ln Q$

$\Delta G° = -RT\ln K$

Standard reduction potentials

ion			$E°$(V)
$F_2(g) + 2e^-$	$\longrightarrow$	$2F^-(aq)$	2.87
$H_2O_2(aq) + 2H^+(aq) + 2e^-$	$\longrightarrow$	$2H_2O(l)$	1.78
$PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^-$	$\longrightarrow$	$PbSO_4(s) + 2H_2O(l)$	1.69
$MnO_4^-(aq) + 4H^+(aq) + 3e^-$	$\longrightarrow$	$MnO_2(s) + 2H_2O(l)$	1.68
$MnO_4^-(aq) + 8H^+(aq) + 5e^-$	$\longrightarrow$	$Mn^{2+}(aq) + 4H_2O(l)$	1.51
$Au^{3+}(aq) + 3e^-$	$\longrightarrow$	$Au(s)$	1.50
$PbO_2(s) + 4H^+(aq) + 2e^-$	$\longrightarrow$	$Pb^{2+}(aq) + 2H_2O(l)$	1.46
$Cl_2(g) + 2e^-$	$\longrightarrow$	$2Cl^-(aq)$	1.36
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^-$	$\longrightarrow$	$2Cr^{3+}(aq) + 7H_2O(l)$	1.33
$O_2(g) + 4H^+(aq) + 4e^-$	$\longrightarrow$	$2H_2O(l)$	1.23
$MnO_2(s) + 4H^+(aq) + 2e^-$	$\longrightarrow$	$Mn^{2+}(aq) + 2H_2O(l)$	1.21
$IO_3^-(aq) + 6H^+(aq) + 5e^-$	$\longrightarrow$	$\frac{1}{2}I_2(aq) + 3H_2O(l)$	1.20
$Br_2(l) + 2e^-$	$\longrightarrow$	$2Br^-(aq)$	1.09
$VO_2^+(aq) + 2H^+(aq) + e^-$	$\longrightarrow$	$VO^{2+}(aq) + H_2O(l)$	1.00
$NO_3^-(aq) + 4H^+(aq) + 3e^-$	$\longrightarrow$	$NO(g) + 2H_2O(l)$	0.96
$ClO_2(g) + e^-$	$\longrightarrow$	$ClO_2^-(aq)$	0.95
$Ag^+(aq) + e^-$	$\longrightarrow$	$Ag(s)$	0.80
$Fe^{3+}(aq) + e^-$	$\longrightarrow$	$Fe^{2+}(aq)$	0.77
$O_2(g) + 2H^+(aq) + 2e^-$	$\longrightarrow$	$H_2O_2(aq)$	0.70
$MnO_4^-(aq) + e^-$	$\longrightarrow$	$MnO_4^{2-}(aq)$	0.56
$I_2(s) + 2e^-$	$\longrightarrow$	$2I^-(aq)$	0.54
$Cu^+(aq) + e^-$	$\longrightarrow$	$Cu(s)$	0.52
$O_2(g) + 2H_2O(l) + 4e^-$	$\longrightarrow$	$4OH^-(aq)$	0.40
$Cu^{2+}(aq) + 2e^-$	$\longrightarrow$	$Cu(s)$	0.34
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$	$\longrightarrow$	$H_2SO_3(aq) + H_2O(l)$	0.20
$Cu^{2+}(aq) + e^-$	$\longrightarrow$	$Cu^+(aq)$	0.16
$Sn^{4+}(aq) + 2e^-$	$\longrightarrow$	$Sn^{2+}(aq)$	0.15
$2H^+(aq) + 2e^-$	$\longrightarrow$	$H_2(g)$	0
$Fe^{3+}(aq) + 3e^-$	$\longrightarrow$	$Fe(s)$	−0.036
$Pb^{2+}(aq) + 2e^-$	$\longrightarrow$	$Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^-$	$\longrightarrow$	$Sn(s)$	−0.14
$Ni^{2+}(aq) + 2e^-$	$\longrightarrow$	$Ni(s)$	−0.23
$Cd^{2+}(aq) + 2e^-$	$\longrightarrow$	$Cd(s)$	−0.40
$Fe^{2+}(aq) + 2e^-$	$\longrightarrow$	$Fe(s)$	−0.45
$Cr^{3+}(aq) + e^-$	$\longrightarrow$	$Cr^{2+}(aq)$	−0.50
$Cr^{3+}(aq) + 3e^-$	$\longrightarrow$	$Cr(s)$	−0.73
$Zn^{2+}(aq) + 2e^-$	$\longrightarrow$	$Zn(s)$	−0.76
$2H_2O(l) + 2e^-$	$\longrightarrow$	$H_2(g) + 2OH^-(aq)$	−0.83
$Mn^{2+}(aq) + 2e^-$	$\longrightarrow$	$Mn(s)$	−1.18
$Al^{3+}(aq) + 3e^-$	$\longrightarrow$	$Al(s)$	−1.66
$Mg^{2+}(aq) + 2e^-$	$\longrightarrow$	$Mg(s)$	−2.37
$Na^+(aq) + e^-$	$\longrightarrow$	$Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^-$	$\longrightarrow$	$Ca(s)$	−2.76
$Ba^{2+}(aq) + 2e^-$	$\longrightarrow$	$Ba(s)$	−2.90
$K^+(aq) + e^-$	$\longrightarrow$	$K(s)$	−2.92
$Li^+(aq) + e^-$	$\longrightarrow$	$Li(s)$	−3.04