

Exam 4A

Chem 1142

Spring 2017

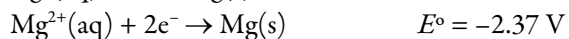
Name: _____

MULTIPLE CHOICE. [3 pts ea.] Record the best response on the scantron sheet. [45 pts total.]

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

- Q1. Which version of the exam do you have?
- 4A
 - 4B
- Q2. Which law of thermodynamics states that the entropy of the universe keeps increasing?
- First law
 - Second law
 - Third law
 - Fourth law
- Q3. Which of the following substances would be expected to have the highest entropy at a given temperature?
- H₂O(s)
 - Au(s)
 - Hg(l)
 - He(g)
- Q4. Which of the following chemical reactions would likely have a $\Delta S^\circ \approx 0$?
- $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$
 - $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$
 - $\text{CO}(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{CH}_2\text{O}(\text{g})$
 - $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$
- Q5. A chemical reaction loses 25 kJ of heat to its surroundings at a temperature of 25 °C. What will the entropy change of the surrounding be?
- 1 kJ/°C
 - +1000 J/°C
 - 84 J/K
 - +84 J/K
- Q6. An endothermic chemical reaction has $\Delta S^\circ_{\text{rxn}} < 0$. What can you say about the spontaneity of this reaction?
- The reaction is always spontaneous
 - The reaction is always non-spontaneous
 - The reaction is spontaneous at low temperatures, but non-spontaneous at high temperatures
 - The reaction is non-spontaneous at low temperatures, but spontaneous at high temperatures

- Q7. Which of the following substances will have a Gibbs free energy of formation of zero?
- $\text{H}_2(\text{l})$
 - $\text{CH}_4(\text{g})$
 - $\text{C}(\text{s, graphite})$
 - $\text{CO}_2(\text{s})$
- Q8. A reaction with a large and negative value of ΔG° will have an equilibrium constant, K , whereby which statement best applies:
- $K \gg 1$
 - $K \ll 1$
 - $K = 1$
 - $K = 0$
- Q9. At equilibrium, what can you say about the value of ΔG ?
- $\Delta G = 1$
 - $\Delta G > 0$
 - $\Delta G = 0$
 - $\Delta G \gg 1$
- Q10. The oxidation number of chromium in $\text{Cr}_2\text{O}_7^{2-}$ is:
- +8
 - +7
 - +6
 - +5
- Q11. Which of the following is **not** a redox reaction?
- $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$
 - $\text{Zn}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g})$
 - $\text{PCl}_3(\text{l}) + \text{Cl}_2(\text{g}) \rightarrow \text{PCl}_5(\text{l})$
 - $\text{KCl}(\text{aq}) + \text{AgNO}_3(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{KNO}_3(\text{aq})$
- Q12. Oxidation takes place at which part of a galvanic cell?
- Salt bridge
 - Voltmeter
 - Anode
 - Cathode
- Q13. Given the following two standard electrode potentials:



Which of the following species would be the best **oxidizing** agent?

- $\text{Ag}^+(\text{aq})$
- $\text{Ag}(\text{s})$
- $\text{Mg}^{2+}(\text{aq})$
- $\text{Mg}(\text{s})$

Q14. A spontaneous redox reaction would best be described as having:

- a) $E^\circ_{\text{cell}} > 0, \Delta G^\circ > 0$
- b) $E^\circ_{\text{cell}} > 0, \Delta G^\circ < 0$
- c) $E^\circ_{\text{cell}} < 0, \Delta G^\circ > 0$
- d) $E^\circ_{\text{cell}} < 0, \Delta G^\circ < 0$

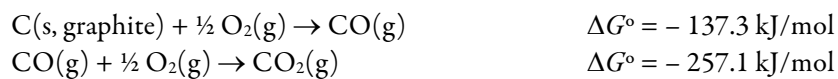
Q15. The charge on 2 moles of electrons is equal to:

- a) $-193,000 \text{ C}$
- b) $-2 \times 6.022 \times 10^{23}$
- c) $96,500 \text{ C}$
- d) $2-$

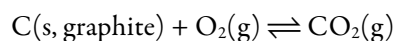
Short Response.

Show ALL work to receive credit.

Q16. [10 pts.] Given the following chemical equations:

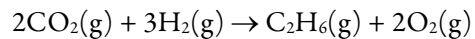


Calculate the value of the equilibrium constant, K , at 25°C for the reaction:



Be sure to show all work and explain clearly your solution.

Q17. [15 pts.] (a) Show how to, then calculate, ΔG° at 45 °C for the chemical reaction:

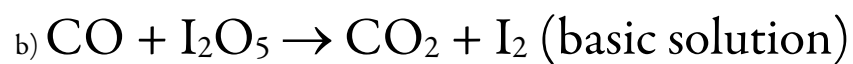


Substance	CO ₂ (g)	H ₂ (g)	C ₂ H ₆ (g)	O ₂ (g)
ΔH_f° (kJ/mol)	-393.5	0	-84.7	0
S° (J/mol·K)	213.6	131.0	229.5	205.0

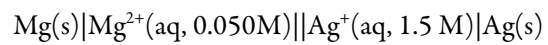
(b) Comment on the value you obtained from part (a), and its meaning.

(c) Explain what will happen to the reaction as the temperature is increased.

Q18. [15 pts.] Balance the following two redox equations using the half-reaction method. Be sure to clearly identify all oxidation numbers in the original skeleton equation.



Q19. [15 pts.] Calculate E_{cell} for the following cell diagram:



As part of your answer you should calculate E°_{cell} and also write the overall balanced chemical equation. Assume the cell temperature is 25 °C.

Periodic Table of the Elements

IA		IIA										IIIA										IVA										VA										VIA										VIIA										VIIIA									
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	57 Lu 175.0	58 Hf 178.5	59 Ta 180.9	60 W 183.8	61 Re 186.2	62 Os 190.2	63 Ir 192.2	64 Pt 195.1	65 Au 197.0	66 Hg 200.6	67 Tl 204.4	68 Pb 207.2	69 Bi 209.0	70 Po [210]	71 At [210]	72 Rn [222]																																																						
87 Fr [223]	88 Ra** [226]	89 Lr [262]	90 Rf [261]	91 Db [262]	92 Sg [266]	93 Bh [264]	94 Hs [265]	95 Mt [268]	96 [269]	97 [272]	98 [277]	99 [285]	100 [285]	101 [289]	102 [289]	103 [293]	104 [293]																																																						
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$$N_A = 6.022 \times 10^{23} \text{ mol}^{-1}$$

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

$$\Delta G = -nFE_{\text{cell}}$$

$$\Delta G^\circ = -nFE_{\text{cell}}^\circ$$

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

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

$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$$

$$F = 96,500 \text{ C/mol } e^- \quad 1 \text{ V} = 1 \text{ J/C}$$

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

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

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

Partial list of standard electrode potentials:

