Practice Problems for Final - New Material Only

This practice exam covers only the new material from chapters 15, 16, 19 and 20. The actual **final** is cumulative and **covers ALL material from the semester**.

1. The balanced homogeneous vapor-phase reaction A + B ⇄ X + Y has $K_c = 9.5 \times 10^{-33}$ at 472 K. At equilibrium (pick the BEST answer)

a) products predominate (mostly products but measurable amounts of reactants)
b) reactants predominate (mostly reactants but measurable amounts of products)
c) roughly equal molar amounts of products and reactants are present
d) essentially only products exist
e) essentially only reactants exist

2. Given the following two equilibrium reactions,

(1) $NOBr\ (g) \rightleftarrows NO\ (g) + \frac{1}{2} Br_2\ (g)$

(2) $N_2\ (g) + O_2\ (g) \rightleftarrows 2\ NO\ (g)$

What is the equilibrium constant, K, for the reaction below,

$$2\ NOBr\ (g) \rightleftarrows N_2\ (g) + O_2\ (g) + Br_2\ (g)$$

a) $K = K_1^2 K_2^{-1}$ b) $K = K_1^{-1} K_2^2$ c) $K = K_1^2 K_2$
d) $K = \frac{1}{2} K_1 K_2^{-1}$ e) $K = (K_2 - 2K_1)$

3. The equilibrium constant K_c for the following reaction at 800.0 °C is 7.14×10^{-2}. What is K_p at this temperature?

$$2\ CO(g) \rightleftarrows C(\mathbf{s}) + CO_2(g)$$

a) 7.14×10^{-2} b) 8.11×10^{-4} c) 5.92×10^{2}
d) 6.29 e) 9.22×10^{-6}

4. Consider the reactions given below. In which case(s) will the reaction proceed more to the **RIGHT** by **inreasing** the **pressure**?

1) $2\ CO(g) + O_2(g) \rightleftarrows 2\ CO_2(g)$ $\Delta H = -566$ kJ

2) $2\ NO(g) \rightleftarrows N_2(g) + O_2(g)$ $\Delta H = -181$ kJ

3) $N_2O_4(g) \rightleftarrows 2\ NO_2(g)$ $\Delta H = 58.0$ kJ

4) $Ni(\mathbf{s}) + 4\ CO(g) \rightleftarrows Ni(CO)_4(g)$ $\Delta H = -160$ kJ

5) $N_2(g) + 3\ H_2(g) \rightleftarrows 2\ NH_3(g)$ $\Delta H = -91.8$ kJ

a) 3, 4 b) 1, 5 c) 1, 4, 5 d) 2, 3, 4 e) 1, 2, 3

5. For which of the following reactions is **REACTANT** formation favored by **low pressure AND high temperature**?

1) $N_2(g) + 3\ H_2(g) \rightleftarrows 2\ NH_3(g)$ $\Delta H = -91.8$ kJ

2) $2\ CO_2(g) \rightleftarrows 2\ CO(g) + O_2(g)$ $\Delta H = 566$ kJ

3) $2\ O_3(g) \rightleftarrows 3\ O_2(g)$ $\Delta H = -285$ kJ

4) $H_2(g) + F_2(g) \rightleftarrows 2\ HF(g)$ $\Delta H = -541$ kJ

5) $N_2(g) + 2\ O_2(g) \rightleftarrows 2\ NO_2(g)$ $\Delta H = 66.4$ kJ

a) 1 b) 2 c) 3 d) 4 e) 5

6. Consider the following reactions at equilibrium and determine which of the indicated changes will cause the reaction to proceed to the **right**. We are considering small changes in a substance (i.e. adding or removing small amounts)

1) $CO(g) + 3\ H_2(g) \rightleftarrows CH_4(g) + H_2O(g)$ (remove H_2O)

2) $N_2(g) + 3\ H_2(g) \rightleftarrows 2\ NH_3(g)$ (add NH_3)

3) $H_2(g) + F_2(g) \rightleftarrows 2\ HF(g)$ (add HF)

4) $N_2(g) + 2\ O_2(g) \rightleftarrows 2\ NO_2(g)$ (remove NO_2)

5) $BaO(\mathbf{s}) + SO_3(g) \rightleftarrows BaSO_4(\mathbf{s})$ (add BaO)

a) 1, 4 b) 3, 4 c) 2, 3 d) 2, 3, 5 e) 1, 4, 5

7. If the **temperature** of the following endothermic reaction, already at equilibrium, is **raised** from 50°C to 100°C, in which direction will the reaction shift **AND** how will this affect the value of the equilibrium constant?

$$2\ H_2O(g) + Cl_2(g) \rightleftarrows 4\ HCl(g) + O_2(g)$$

a) shift in forward direction, K will decrease
b) shift in reverse direction, K will decrease
c) shift in forward direction, K will increase
d) shift in reverse direction, K will increase
e) shift in forward direction, K will not change

8. For the following system, 0.400 moles of I_2, 0.400 moles of Br_2 and 2.10 moles of IBr are placed in a **2.00**- L flask. The value of K_c for the reaction is 110.25. What are the concentrations (M) of **Br_2** and **IBr** (in this order) when equilibrium is reached?

$$I_2(g) + Br_2(g) \rightleftarrows 2\ IBr(g)$$

a) 0.116, 1.22 b) 0.109, 1.14 c) 0.100, 1.05

d) 0.013, 1.42 e) none of these

9. What is the conjugate **base** of methylamine, CH_3NH_2?

a) CH_3NH^+ b) $CH_3NH_2^-$ c) $CH_3NH_2^+$ d) $CH_3NH_3^+$ e) CH_3NH^-

10. Given that K_w for water is 2.40 x 10^{-14} (M^2) at 37°C, compute the pH of a neutral aqueous solution at 37°C (normal human body temperature). Answer the following **TWO** questions. What is the pH of a neutral solution at 37°C? **AND** If a solution has pH = 7.00 is it acidic, basic, or neutral at 37°C?

a) 7.19, acidic b) 7.19, basic c) 6.81, acidic

d) 6.81, basic e) 7.00, neutral

11. The pH of a 0.10 M solution of NH_4Cl containing 0.10 M NH_3 is 9.20. What is the **$[H_3O^+]$**?

a) 1.6 x 10^{-5} b) 1.0 x 10^{-1} c) 6.3 x 10^{-10}

d) 1.7 x 10^{-10} e) 2.0 x 10^{-9}

12. The K_a values for HS^- and HPO_4^{2-} are 1.2 x 10^{-13} and 4.8 x 10^{-13} respectively. Therefore it follows the HS^- is a __ acid than HPO_4^{2-} and S^{2-} is a ___ base than PO_4^{3-}.

a) stronger, stronger b) stronger, weaker

c) weaker, stronger d) weaker, weaker

13. What is the **ionization constant** of an acid if the hydronium ion concentration of a 0.500 M solution is 1.70 x 10^{-4} M?

a) 3.62 x 10^{-7} b) 2.89 x 10^{-8} c) 5.80 x 10^{-8}

d) 1.16 x 10^{-7} e) 1.70 x 10^{-3}

14. A 0.010 M solution of HNO_2 is 19% ionized. What is the K_a?

a) 4.4×10^{-4} b) 3.9×10^{-4} c) 3.6×10^{-4}

d) 5.0×10^{-4} e) 5.4×10^{-4}

15. What is the **pH** of a 0.20 M NH_4^+ solution (K_b: $NH_3 = 1.8 \times 10^{-5}$)?

a) 2.72 b) 3.11 c) 4.98 d) 5.12 e) 7.61

16. A 1.50 g sample of Vitamin C is dissolved in 100.0 mL of water and titrated with 0.250 M NaOH to the methyl orange equivalence point. The volume of the base used is 34.1 mL. What is the **molecular weight** of Vitamin C assuming one dissociable proton per molecule?

a) 176 b) 164 c) 152 d) 146 e) 139

17. Predict which of the following reactions has a positive entropy change.

I $2\ N_2(g) + O_2(g) \longrightarrow 2\ N_2O(g)$

II $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

III $Zn(s) + 2\ HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$

a) I b) II c) III d) I and II e) II and III

18. Calculate the ΔS (J/mol•K) of fusion for ethane which melts at -183°C. The heat of fusion is 2.86 kJ/mole.

a) 21.4 b) 31.8 c) 15.6 d) 28.1 e) 34.3

19. Calculate the entropy change (J/mole•K) of the reaction. The molar entropies are given below each substance.

$Br_2(g) + F_2(g) \longrightarrow 2\ BrF(g)$

S° (J/K·mole)	245.5	202.8	229.0

a) 9.7 b) -9.7 c) +15.3 d) -219.3 e) -126.3

20. The standard free energy of formation of $H_2O(\ell)$ is -237.1 kJ/mol. Using the data for the following reaction, calculate the free energy of formation, ΔG_f°, for $N_2H_4(\ell)$ (in kJ/mol).

$N_2H_4(\ell) + O_2(g) \longrightarrow N_2(g) + 2\ H_2O(\ell)$ $\Delta G^\circ = -623.4$ kJ

a) 623.4 b) 149.2 c) 386.3 d) 193.2 e) 860.5

21. Consider the ΔG°_f and ΔH°_f (kJ/mole) for the following oxides. Which oxide can be **most easily <u>decomposed</u>** to form the metal and oxygen gas.

	ΔG_f°	ΔH_f°
a) PbO	-187.9	-217.3
b) ZnO	-318.4	-348.3
c) Ag_2O	-11.2	-31.1
d) CdO	-228.4	-258.2

22. From the following ΔH° and ΔS° values predict which of reactions I, II, and III would be spontaneous at 25°C.

	ΔH°(kJ)	ΔS°(J/K)
I	-5.0	-20
II	-10.0	-10
III	-25.0	+75

a) I b) II c) III d) I and II e) II and III

23. The $K_p = 2.47 \times 10^{-29}$ at 25°C for the following reaction. Calculate ΔG° (kJ).

$3/2\ O_2(g) \longrightarrow O_3(g)$

a) 157 b) 163 c) 169 d) 172 e) 175

24. The $K_p = 0.113$ at 25°C and $\Delta H° = +57.2$ kJ for the following reaction. Calculate the **temperature** (°C) where $K_p = 1.00$

$$N_2O_4(g) \rightleftarrows 2\ NO_2(g)$$

a) 45 b) 56 c) 65 d) 70 e) 76

25. For the following reaction, calculate the value of ΔG (kJ) at 25°C in a solution when $[Co^{3+}] = 0.0050$ M, $[NH_3] = 0.10$ M, and $[Co(NH_3)_6^{3+}] = 1.00$ M.

$$Co^{3+}(aq) + 6\ NH_3(aq) \rightleftarrows Co(NH_3)_6^{3+}(aq) \qquad \Delta G° = -41.7\ kJ$$

a) −47.4 b) +20.5 c) +5.7 d) −20.5 e) −5.7

26. From the listed standard electrode potentials, what is E° for a voltaic cell made from the following electrodes?

$$Mg^{2+}(aq) + 2\ e^- \longrightarrow Mg(s) \qquad E° = -2.363\ V$$
$$Sn^{2+}(aq) + 2\ e^- \longrightarrow Sn(s) \qquad E° = -0.136\ V$$

a) +2.227 b) 2.499 c) -2.227 d) -2.499 e) +1.113

27. The E° = 1.29 V for the following reaction. The standard electrode potential for Ce^{4+} as written is +1.61 V. What is the standard electrode potential for the reduction of BiO^+?

$$3\ Ce^{4+}(aq) + Bi(s) + H_2O(\ell) \rightleftarrows 3\ Ce^{3+}(aq) + BiO^+(aq) + 2\ H^+(aq) \qquad E° = +1.29\ V$$

$$BiO^+(aq) + 2\ H^+(aq) + 3\ e^- \longrightarrow Bi(s) + H_2O(\ell) \qquad E° = ?$$
$$Ce^{4+}(aq) + 1\ e^- \longrightarrow Ce^{3+}(aq) \qquad E° = +1.61\ V$$

a) -0.32 b) +0.32 c) +2.90 d) -2.90 e) -1.45

28. Consider the following half-cell reactions and associated standard half-cell potentials and determine which species is the **best <u>reducing</u>** agent.

	E°
$S_2O_6^{2-}(aq) + 4\ H^+(aq) + 2\ e^- \rightarrow 2\ H_2SO_3(aq)$	+0.60 V
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.771 V
$VO_2^+(aq) + 2\ H^+(aq) + e^- \rightarrow VO^{2+}(aq) + H_2O(\ell)$	+1.00 V
$N_2O(aq) + 2\ H^+(aq) + 2\ e^- \rightarrow N_2(g) + H_2O(\ell)$	+1.77 V

a) Fe^{2+} b) H_2SO_3 c) N_2 d) VO^{2+} e) VO_2^+

29. Use the following E° for the electrode potentials, calculate ΔG° in kJ for the indicated reaction.

$$5\ Bi(s) + 3\ MnO_4^-(aq) + 14\ H^+(aq) \rightleftarrows 5\ BiO^+(aq) + 3\ Mn^{2+}(aq) + 7\ H_2O(\ell)$$

$BiO^+(aq) + 2\ H^+(aq) + 3\ e^- \longrightarrow Bi(s) + H_2O$	E° = +0.32 V
$MnO_4^-(aq) + 8\ H^+(aq) + 5\ e^- \longrightarrow Mn^{2+}(aq) + 4\ H_2O$	E° = +1.51 V

a) -1.72×10^2 b) -1.42×10^2 c) -1.20×10^4

d) -1.72×10^3 e) -1.42×10^3

30. Consider an electrochemical cell in which the following reaction occurs and predict which changes will increase the cell voltage.

$$2\ HCl(aq) + Ca(s) \longrightarrow CaCl_2(aq) + H_2(g)$$

I increase the amount of Ca(s)
II decrease the pressure of $H_2(g)$
III decrease the [HCl(aq)]

a) I b) II c) III d) I & II e) II & III

31. Which of the following is characteristic of the **anode** in an **<u>electrolysis</u>** cell?

a) It is where reduction occurs.
b) It attracts negative ions.
c) It receives electrons from the wire.
d) It may gain weight during electrolysis.
e) More than one of the above is correct.

32. The standard electrode potential of Ag^+ is 0.800. The measured voltage of the following cell is 0.900 V at 25°C. Calculate the **pH** of the solution.

$$Pt(s)|H_2(g, 1.00\ atm)|H^+(aq, ?\ M)||Ag^+(aq, 1.00\ M)|Ag(s)$$

a) 1.69 b) 3.38 c) 1.12 d) 2.15 e) 2.78

33. Using standard electrode potentials, calculate the ΔG (kJ) for the following electrochemical cell

$$Zn(s)|Zn^{2+}(aq, 0.010\ M)||Cd^{2+}(aq, 0.050\ M)|Cd(s)$$

$Zn^{2+}(aq) + 2\ e^- \rightarrow Zn(s)$ $E° = -0.763$ V
$Cd^{2+}(aq) + 2\ e^- \rightarrow Cd(s)$ $E° = -0.403$ V

a) 67.0 b) -69.5 c) -73.4 d) -65.0 e) 71.5

34. How many minutes will it take an electric current of 3.64 A to deposit all the copper from 740 mL of 0.250 M $CuSO_4(aq)$? (atomic weight: Cu = 53.55)

a) 182 b) 163 c) 144 d) 102 e) 98

Answers to Practice "Final" Problems

1) E	13) C	25) C
2) A	14) A	26) A
3) B	15) C	27) B
4) C	16) A	28) B
5) A	17) E	29) D
6) A	18) B	30) B
7) C	19) A	31) B
8) A	20) B	32) A
9) E	21) C	33) C
10) D	22) E	34) B
11) C	23) B	
12) C	24) B	