CHM 152 Practice Final

1. Of the following, the one that would have the ***greatest entropy*** (if compared at the same temperature) is,

[a] H_2O (*s*) [b] H_2O (*l*) [c] H_2O (*g*)
[d] All would have the same entropy at the same temperature.
[e] More information would be needed to make a selection.

2. A process will be ***nonspontaneous*** at all temperatures if

[a] the enthalpy change is – and the entropy change is +.
[b] the enthalpy change is + and the entropy change is +.
[c] the enthalpy change is – and the entropy change is –.
[d] the enthalpy change is + and the entropy change is –.
[e] all reactions will be spontaneous at some temperature.

3. Given the following standard state free energies of formation,

	ΔG_f° (kJ/mol)
H_2O (*l*)	–237.2
HNO_3 (*l*)	–79.9
NO (*g*)	86.7
NO_2 (*g*)	51.8

Calculate $\boldsymbol{\Delta G^\circ}$ for the reaction: $3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(l) + NO(g)$

[a] $\Delta G^\circ = 8.7$ kJ [b] $\Delta G^\circ = 192$ kJ [c] $\Delta G^\circ = -414$ kJ [d] $\Delta G^\circ = -192$ kJ [e] $\Delta G^\circ = -155$ kJ

4. Consider a reaction for which $\Delta H = -13.2$ kJ and $\Delta S = -125$ J/K. Calculate $\boldsymbol{\Delta G_{rxn}}$ for this reaction at a temperature of 95.0 K.

[a] –25.1 kJ [b] –13.3 kJ [c] –13.1 kJ [d] –11.9 kJ [e] –1.33 kJ

5. Consider the reaction,

$$2SO_2(g) + O_2g) \rightarrow 2SO_3(g)$$

ΔG° for this reaction is –142 kJ. Calculate the $\boldsymbol{\Delta G}$ for this reaction at 25°C when the partial pressures of the gases are as follows: $P(SO_2) = 0.153$ atm, $P(O_2) = 0.164$ atm, and $P(SO_3) = 0.683$ atm.

[a] –130 kJ [b] 4.8 kJ [c] –142 kJ [d] –147 kJ [e] –154 kJ

6. The activation energy of a reaction may be *lowered* by,

[a] increasing the concentration of the reactants
[b] increasing the temperature
[c] adding a catalyst
[d] answers b and c are both correct
[e] answers a, b, and c are all correct.

Consider the following reaction and data for questions 7 and 8.

$$2NO + 2H_2 \rightarrow N_2 + 2H_2O$$

Exp#	[NO]	[H_2]	Rate (mol/L·s)
1	0.10 *M*	0.20 *M*	0.0150
2	0.10 *M*	0.30 *M*	0.0225
3	0.20 *M*	0.20 *M*	0.0600

7. If H_2 were *disappearing* at the rate of 0.080 mol/L·s, N_2 would be *appearing* at the rate of,

[a] 0.080 mol/L·s [b] 0.16 mol/L·s [c] 0.040 mol/L·s [d] 0.24 mol/L·s
[e] The rate of N_2 appearance can not be determined unless the temperature is known.

8. The ***rate law*** for the reaction is,

[a] rate = $k[NO]^2[H_2]^2$ [b] rate = $k[NO][H_2]^2$ [c] rate = $k[NO][H_2]$
[d] rate = $k[NO]^2[H_2]$ [e] none of these are correct.

9. The reaction

$$2NOCl \rightarrow 2NO + Cl_2$$

is found to obey the rate law, rate = $k[NOCl]^2$. Of the following three postulated mechanisms, the one(s) possible is/are,

1) $NOCl \rightarrow NO + Cl$ (slow)
$Cl + NOCl \rightarrow NOCl_2$ (fast)
$NOCl_2 + NO \rightarrow 2NO + Cl_2$ (fast)

2) $2NOCl \rightarrow NOCl_2 + NO$ (slow)
$NOCl_2 \rightarrow NO + Cl_2$ (fast)

3) $NOCl \rightarrow NO + Cl$ (slow)
$NOCl + Cl \rightarrow NO + Cl_2$ (fast)

[a] 1 [b] 2 [c] 3 [d] 1 and 3 [e] 2 and 3

10. At a certain temperature $K_c = 1 \times 10^9$ for the reaction:

$$Cl_2(g) + F_2(g) \rightleftharpoons 2ClF(g)$$

If 1.0 mol Cl_2 and 1.0 mol F_2 are placed in a reaction vessel and allowed to react, then at equilibrium...

[a] [ClF] will be much larger than [Cl_2] and [F_2].
[b] [ClF] will be much less than [Cl_2] and [F_2].
[c] [ClF] will be nearly equal to [Cl_2] and [F_2].
[d] the system will contain only Cl_2 and F_2.
[e] the system will contain only ClF.

11. The following reaction comes to equilibrium:

$$2H_2(g) + S_2(g) \rightleftharpoons 2H_2S(g)$$

If the volume of the reaction vessel is *reduced*, at constant temperature, the concentration of

[a] H_2 would increase. [b] H_2 would decrease. [c] S_2 would increase.
[d] H_2S would decrease. [e] H_2, S_2 and H_2S would not change.

12. For the equilibrium:

$$2N_2O(g) + O_2(g) \rightleftharpoons 4NO(g),\ \Delta H = +198 \text{ kJ}.$$

Which of the following changes will ***increase*** the amount of $\mathbf{N_2O}$ when the system reaches its new equilibrium state?

[a] The temperature of the system is lowered.
[b] The volume of the container is increased.
[c] Some O_2 is added.
[d] He(g) is added without changing the volume.
[e] Some NO is removed.

13. For the reaction: $H_2(g) + CO_2(g) \rightleftharpoons H_2O(g) + CO(g)$, the equilibrium constant, K_p, is 4.40 at 2000 K. Calculate $\Delta G°$ for this reaction.

[a] −73.2 kJ [b] −24.6 kJ [c] −243 J [d] 243 J [e] 24.6 kJ

14. What is the $\boldsymbol{K_b}$ value for the carbonate ion, CO_3^{2-} ?

[a] $K_b = 4.8 \times 10^{-11}$ [b] $K_b = \dfrac{1.0 \times 10^{-14}}{4.8 \times 10^{-11}}$ [c] $K_b = \dfrac{1.0 \times 10^{-14}}{4.2 \times 10^{-7}}$

[d] $K_b = \dfrac{1}{(4.2 \times 10^{-7})(4.8 \times 10^{-11})}$ [e] None of these are correct

15. Which of the following would *increase* the K_a for acetic acid?

[a] Decrease the pH of the solution. [b] Add some sodium acetate.
[c] Add some sodium hydroxide. [d] Add some water.
[e] None of the above, the K_a is temperature dependent only.

16. Which one of the following salts will form a ***basic*** solution on dissolving in water?

[a] NaCl [b] $KClO_4$ [c] NH_4NO_3 [d] NaCN
[e] None of these solutions are basic

17. Which is the ***conjugate base*** of $H_2PO_4^-$?

[a] H_3PO_4 [b] HPO_4^- [c] HPO_4^{2-} [d] PO_4^{3-} [e] $H_2PO_4^{2-}$

18. The pH of a solution is 8.50. The $\mathbf{[OH^-]}$ is:

[a] $[OH^-] = 3.2 \times 10^{-6}\ M$ [b] $[OH^-] = 3.2 \times 10^{-9}\ M$ [c] $[OH^-] = 5.50\ M$
[d] $[OH^-] = 1.0 \times 10^{-14}\ M$ [e] None of these is correct.

19. What is the ***pH*** of a 0.015 M HNO_3 solution?

[a] pH = 0.015 [b] pH = 1.82 [c] pH = 2.59 [d] pH = 2.6×10^{-3}
[e] This cannot be answered without more information.

20. What is the ***pH*** of a 0.10 *M* solution of CH_3COONa?

[a] pH = 4.74 [b] pH = 5.13 [c] pH = 8.87 [d] pH = 9.25 [e] None of these

21. Which of the following is the most ***acidic*** solution?

[a] 0.10 *M* CH_3COOH and 0.10 *M* CH_3COONa
[b] 0.10 *M* CH_3COOH
[c] 0.10 *M* HNO_2 and 0.10 *M* $NaNO_2$
[d] 0.10 *M* HNO_2
[e] 0.10 *M* CH_3COONa

22. Which of the following mixtures is suitable for making a buffer solution with an optimum pH of 9.2–9.3?

[a] CH_3COONa/CH_3COOH [b] NH_3/NH_4Cl [c] $NaOCl/HOCl$
[d] $NaNO_2/HNO_2$ [e] $NaCl/HCl$

23. Which of the following 0.10 *M* solutions ***cannot*** act as a good buffer?

[a] HF & NaF [b] CH_3COOH & CH_3COONa [c] $NaHCO_3$ & Na_2CO_3
[d] pure H_2O [e] All of these solutions are good buffers.

24. What is the ***pH*** of a solution that is 0.41 *M* HOCl and 0.050 *M* NaOCl?

[a] pH = 0.39 [b] pH = 3.94 [c] pH = 6.58 [d] pH = 7.49 [e] pH = 8.40

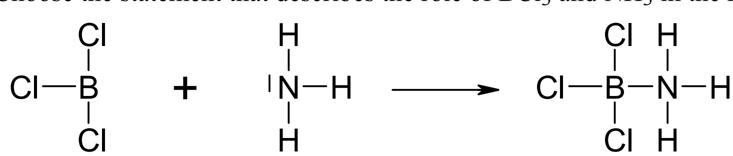
25. For which titration will the pH be ***basic*** at the equivalence point?

[a] HCl with NaOH [b] HCl with NH_3 [c] HOCl with NaOH
[d] All of these titrations will be neutral at the equivalence point.
[e] A prediction cannot be made without additional information.

26. The *pH* at the equivalence point of an acid-base titration may *differ* from 7.00 because of

[a] the indicator used. [b] the auto-ionization of water.
[c] the initial pH of the titrate. [d] the concentration of the titrant.
[e] the hydrolysis of the salt formed.

27. Choose the statement that describes the role of BCl_3 and NH_3 in the reaction:



[a] BCl_3 is a Lewis base and NH_3 is a Lewis acid.
[b] Both NH_3 and BCl_3 are Lewis acids.
[c] Both NH_3 and BCl_3 are Lewis bases.
[d] BCl_3 is a Lewis acid and NH_3 is a Lewis base.

28. What is the ***pH*** of the solution resulting from the addition of 10.0 mL of 0.10 M NaOH to 50.0 mL of 0.10 M HCl solution?

[a] pH = 1.00 [b] pH = 1.08 [c] pH = 1.18 [d] pH = 7.00
[e] Insufficient data to answer.

29. The molar solubility of $MgCO_3$ is 1.8×10^{-4} mol/L. What is $\boldsymbol{K_{sp}}$ for this compound?

[a] 1.8×10^{-4} [b] 3.6×10^{-4} [c] 1.3×10^{-7} [d] 3.2×10^{-8} [e] 2.8×10^{-14}

30. The K_{sp} of Ag_2CrO_4 is 1.9×10^{-12}. The ***molar solubility*** of Ag_2CrO_4 in 0.10 M $AgNO_3$ solution is:

[a] 7.8×10^{-5} M [b] 9.5×10^{-12} M [c] 1.4×10^{-6} M
[d] 1.3×10^{-4} M [e] 1.9×10^{-10} M

31. For $PbCl_2$, $K_{sp} = 2.4 \times 10^{-4}$. Will a *precipitate* of $PbCl_2$ *form* when 0.50 L of 3.0×10^{-2} M $Pb(NO_3)_2$ is added to 0.50 L of 9.0×10^{-2} M NaCl and *why*?

[a] Yes, $Q > K_{sp}$ [b] No, $Q < K_{sp}$ [c] No, $Q = K_{sp}$
[d] Yes, $Q < K_{sp}$ [e] More information is needed to solve the problem.

For questions 32 and 33 consider the following redox reaction,

$$Zn + NO_3^- \xrightarrow{\text{acidic}} Zn^{2+} + NO$$

32. In the balanced equation in acidic solution, the ***coefficient*** of H_2O is,

[a] 1 H_2O [b] 2 H_2O [c] 3 H_2O [d] 4 H_2O [e] 5 H_2O

33. The ***reducing agent*** is,

[a] Zn [b] NO_3^- [c] NO [d] H^+ [e] H_2O

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34. Referring to the Standard Reduction Potential Table on this exam, the strongest ***oxidizing agent*** listed below is

[a] Mn^{2+} [b] Mn [c] ClO_3^- [d] Cl_2 [e] Ag^+

35. Of the following reactions, the ***spontaneous*** one would be,

[a] $Mn^{2+} + Fe \rightarrow Mn + Fe^{2+}$
[b] $3Ag + NO_3^- + 4H^+ \rightarrow 3Ag^+ + NO + 2H_2O$
[c] $3Sn^{4+} + 2Cr^{3+} + 7H_2O \rightarrow 3Sn^{2+} + Cr_2O_7^{2-} + 14H^+$
[d] all are spontaneous.
[e] none are spontaneous.

36. In a galvanic cell, the reaction occurring at the ***cathode*** is

[a] reduction [b] oxidation [c] hydrolysis [d] isomerization [e] acid-base

37. Use data from the Table of Standard Reduction Potentials to calculate the ***equilibrium constant*** for the following reaction at 25°C.

$$Cl_2(g) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(l)$$

[a] $K = 1.5 \times 10^{-10}$ [b] $K = 6.3 \times 10^{9}$ [c] $K = 1.3 \times 10^{41}$
[d] $K = 8.1 \times 10^{4}$ [e] $K = 9.8$

38. The radioactive decay of ^{87}Kr *produces* a beta particle and,

[a] an alpha particle [b] a positron [c] H^+
[d] ^{87}Br [e] ^{87}Rb

39. A first order reaction has a rate constant of $k = 2.4 \times 10^{-3}\ s^{-1}$ at 25°C. The half-life of the reaction is:

[a] 1660 s [b] 576 s [c] 289 s [d] 1.70×10^{-3} s [e] 2.90×10^{-4} s

40. A certain radioactive element has a half-life of 10.0 minutes. If 2.00 grams of this element were present initially, the amount of the element remaining after 8.30 minutes has elapsed would be:

[a] 1.00 grams [b] 1.13 grams [c] 1.32 grams [d] 1.08 grams [e] 0.288 grams

Potentially Useful Information

$\Delta G° = \Sigma n\Delta G_f°(\text{products}) - \Sigma n\Delta G_f°(\text{reactants})$

$\Delta H° = \Sigma n\Delta H_f°(\text{products}) - \Sigma n\Delta H_f°(\text{reactants})$

$\Delta S° = \Sigma nS°(\text{products}) - \Sigma nS°(\text{reactants})$

$\Delta S_{univ} = \Delta S_{sys} + \Delta S_{surr}$

$\Delta G° = \Delta H° - T\Delta S°$

$\Delta G° = -RT\ln K$

$\Delta G = \Delta G° + RT\ln Q$

$R = 8.314$ J/mol·K = 0.0821 L·atm/mol·K

K = °C + 273

$PV = nRT$

$$\text{molarity}(M) = \frac{\text{moles of solute}}{\text{L of solution}}$$

$$\text{Rate} = \frac{\Delta[\text{conc}]}{\Delta t}$$

$\text{Rate} = k[A]^x[B]^y$

$\ln[A]_t = -kt + \ln[A]_0$

$$\ln\left(\frac{[A]_t}{[A]_0}\right) = -kt$$

$$\ln\left(\frac{[A]_0}{[A]_t}\right) = kt$$

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

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

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

$$\ln k = \left(\frac{-E_a}{R}\right)\left(\frac{1}{T}\right) + \ln A$$

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

A quadratic equation of the form $ax^2 + bx + c = 0$, has the solutions: $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

$$K_c = \frac{[\text{products}]^x}{[\text{reactants}]^y}$$

$K_p = K_c(0.0821\ T)^{\Delta n}$

$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25°C

$pH = -\log[H_3O^+]$

$K_w = K_a \times K_b$

pH + pOH = 14

$pOH = -\log[OH^-]$

$$pH = pK_a + \log\frac{[\text{base}]}{[\text{acid}]}$$

$\Delta G° = -nF\,E°_{cell}$

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

$$E°_{cell} = \frac{RT}{nF}\ln K$$

$$E = E° - \frac{RT}{nF}\ln Q$$

$F = 96{,}500$ J/V·mol

$E°_{cell} = E°_{oxi} + E°_{red}$

$$E°_{cell} = \frac{0.0257\,\text{V}}{n}\ln K \quad \text{at 25°C}$$

$$E = E° - \frac{0.0257\,\text{V}}{n}\ln Q \quad \text{at 25°}$$

Acid Ionization Constants (25°C)

Acid	Formula	K_a	pK_a
Hydrofluoric	HF	7.1×10^{-4}	3.15
Nitrous	HNO_2	4.5×10^{-4}	3.35
Benzoic	C_6H_5COOH	6.5×10^{-5}	4.19
Acetic	CH_3COOH	1.8×10^{-5}	4.74
Carbonic Bicarbonate ion	H_2CO_3 HCO_3^-	4.2×10^{-7} 4.8×10^{-11}	6.38 10.32
Hypochlorous	$HOCl$	3.2×10^{-8}	7.49
Ammonium ion	NH_4^+	5.6×10^{-10}	9.25
Hydrocyanic	HCN	4.9×10^{-10}	9.31

Standard Reduction Potentials

Half-Reaction	E° (Volts)
$Mg^{2+}\ (aq) + 2\ e^- \rightarrow Mg\ (s)$	–2.37
$Mn^{2+}\ (aq) + 2\ e^- \rightarrow Mn\ (s)$	–1.18
$Zn^{2+}\ (aq) + 2\ e^- \rightarrow Zn\ (s)$	–0.76
$Cr^{3+}\ (aq) + 3\ e^- \rightarrow Cr\ (s)$	–0.74
$Fe^{2+}\ (aq) + 2\ e^- \rightarrow Fe\ (s)$	–0.44
$2\ H^+\ (aq) + 2\ e^- \rightarrow H_2\ (g)$	0.00
$Sn^{4+}\ (aq) + 2\ e^- \rightarrow Sn^{2+}\ (aq)$	+0.13
$Fe^{3+}\ (aq) + e^- \rightarrow Fe^{2+}\ (aq)$	+0.77
$Ag^+\ (aq) + e^- \rightarrow Ag\ (s)$	+0.80
$NO_3^-\ (aq) + 4\ H^+\ (aq) + 3\ e^- \rightarrow NO\ (g) + 2\ H_2O\ (l)$	+0.96
$Br_2\ (l) + 2\ e^- \rightarrow 2\ Br^-\ (aq)$	+1.07
$Cr_2O_7^{2-}\ (aq) + 14\ H^+\ (aq) + 6\ e^- \rightarrow 2\ Cr^{3+}\ (aq) + 7\ H_2O\ (l)$	+1.33
$Cl_2\ (l) + 2\ e^- \rightarrow 2\ Cl^-\ (aq)$	+1.36
$ClO_3^-\ (aq) + 12\ H^+\ (aq) + 10\ e^- \rightarrow Cl_2\ (g) + 6\ H_2O\ (l)$	+1.47
$Ce^{4+}\ (aq) + e^- \rightarrow Ce^{3+}\ (aq)$	+1.70

Answer Key

1) c	2) d	3) a	4) e	5) a	6) c	7) c	8) d	9) b	10) a
11) b	12) a	13) b	14) b	15) e	16) d	17) c	18) a	19) b	20) c
21) d	22) b	23) d	24) c	25) c	26) e	27) d	28) c	29) d	30) e
31) b	32) d	33) a	34) c	35) b	36) a	37) b	38) e	39) c	40) b