SOLUTIONS

Chem 1200b Midterm Test

Sunday, March 4, 2012   2.00 hours

Section code 000   Exam code 111

Message from the Dean: All answer sheets are subject to a common data analysis which identifies anomalies of statistical significance in the selection of right and wrong answers by pairs of students. The course instructor is required to report all statistically significant results which suggest that cheating may have occurred. All such incidents will be subject to further investigation. All proven cases of cheating will be subject to severe academic penalties. If you are currently seated near someone with whom you studied, and you think you may choose many of the same answers as that person, please raise your hand now and ask a proctor to reseat you.

This test contains 34 multiple-choice questions of equal value. Be sure you have a complete test paper. Scrap paper, a periodic table, and other data are attached, and they may be removed for use.

Place your student ID card on your desk.

The only permitted calculator is a Sharp EL-510R(B). No other electronic devices may be in your possession, even for timekeeping purposes. A molecular model kit is permitted, but the sharing of models or model pieces is strictly forbidden.

You are welcome to keep the question booklet and any scrap sheets.

Proctors and instructors will not interpret, translate, clarify, or explain questions, nor will they confirm, verify, or assist you with your answers or your thinking. Therefore, you are not permitted to ask any questions related to the content of the test.

The Scantron must be completed within the duration of the test. At the end of the test, everyone must put their pencils down. The Scantron, whether completed or not, will be collected. Under no circumstances will this booklet be used for marking purposes. The answers on your Scantron are considered to be your official answers, so please ensure that you complete your Scantron accurately. Answers that cannot be read by the Scantron computer will be marked as being incorrect.

If you do not leave before the last 15 minutes, you must stay in your seat until you are dismissed. NO TALKING IS ALLOWED AT ANY TIME.

Before starting the test, be sure that you have correctly entered:

- Your student number
- Section code 000
- Exam code 111

1. Which of the following equilibrium constant expressions is/are correct for the corresponding equilibrium reactions, as written?

✓ 1. $Ag^+(aq) + Cl^-(aq) \rightleftharpoons AgCl(s)$ $K = \frac{1}{[Ag^+][Cl^-]}$

✓ 2. $BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^-(aq)$ $K = [Ba^{2+}][SO_4^{2-}]$

X 3. $\frac{1}{2}N_2O_4(g) \rightleftharpoons NO_2(g)$ $K = \frac{[NO_2]^2}{[N_2O_4]}$

[✓ A) 1 and 2 only]
B) 3 only
C) 1 and 3 only
D) 2 and 3 only
E) None are correct

2. Water can react with methane to produce carbon monoxide and hydrogen:

$$H_2O(g) + CH_4(g) \rightleftharpoons CO(g) + 3H_2(g)$$

A sealed reaction vessel initially contains only 100 KPa of $H_2O(g)$ and 100 kPa of $CH_4(g)$. After equilibrium is established, the pressure of CO(g) is determined to be 40 kPa. What is K_p for the reaction?

A) 711
B) 1.15×10^6
C) 1.13
[✓ D) 1.92×10^4]
E) 80.0

	H_2O +	CH_4 ⇌	CO +	$3H_2$
In:	100	100	0	0
Eq:	60	60	40	120

$$K_p = \frac{(P_{CO})(P_{H_2})^3}{(P_{H_2O})(P_{CH_4})}$$

$$= \frac{(40)(120)^3}{(60)(60)}$$

$$= \underline{1.92 \times 10^4}$$

3. The folding and unfolding of proteins are important biological processes. Studies on the immunoglobulin protein G have shown that a reversible transition from a folded to an unfolded form occurs when the temperature is increased.

$$\text{Protein G (folded)} \rightleftharpoons \text{Protein G (unfolded)} \qquad \Delta H^\circ = 210.9\ \text{kJ mol}^{-1} \quad \Delta S^\circ = 616.7\ \text{J mol}^{-1}\ \text{K}^{-1}$$

What is the value of K at 25°C?

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

A) 2.70×10^{-4}
[✓ B) 1.83×10^{-5}]
C) 0.989
D) 10.9
E) Cannot be determined without knowing the concentrations of the folded and unfolded forms

$\Delta G^\circ = 210900\ \text{J mol}^{-1} - 298.15\ \text{K}\,(616.7\ \text{J mol}^{-1}\ \text{K}^{-1})$

$= 27030.9\ \text{J mol}^{-1}$

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

$27030.9 = (-8.314 \times 298.15) \ln K$

$-10.90 = \ln K$

$\therefore K = 1.837 \times 10^{-5}$

4. Calcium hydroxide, also known as hydrated lime, is used to neutralize excess acid in lakes. What is the pH of a saturated solution of $Ca(OH)_2$? The K_{sp} for $Ca(OH)_2$ is 1.4×10^{-6}.

A) 10.77
B) 11.84
C) 8.15
D) 11.07
[✓ E) 12.15]

$Ca(OH)_2 \rightleftharpoons Ca^{2+} + 2OH^-$

solid   0   0

solid − x   x   2x

$K_{sp} = [Ca^{2+}][OH^-]^2$

$1.4 \times 10^{-6} = (x)(2x)^2 = 4x^3$

$x = 7.047 \times 10^{-3}$

$[OH^-] = 2x = 1.41 \times 10^{-2}\ M$

pOH = 1.85

pH = 12.15

5. Consider the following silver salts:

AgCl $K_{sp} = 1.7 \times 10^{-10} = x^2$ $x = 1.3 \times 10^{-5}$
AgBr $K_{sp} = 5.0 \times 10^{-13} = x^2$ $x = 7.07 \times 10^{-7}$
Ag_2CO_3 $K_{sp} = 8.2 \times 10^{-12} = 4x^3$ $x = 1.27 \times 10^{-4}$

Which one of A – E correctly ranks the salts in the order of decreasing solubility?

x is the solubility

A) AgCl > AgBr > Ag_2CO_3
B) Ag_2CO_3 > AgBr > AgCl
C) AgCl > Ag_2CO_3 > AgBr
[✓ D) Ag_2CO_3 > AgCl > AgBr]
E) AgBr > Ag_2CO_3 > AgCl

6. NaOH is added to a 0.020 M solution of $Mg(NO_3)_2$ until the first trace of $Mg(OH)_2$ is seen. Assuming that the volume of the solution remains constant, what is the pH of the solution at this point? The K_{sp} of $Mg(OH)_2$ is 8.9×10^{-12}.

A) 4.65
B) 9.02
C) 9.63
[✓ D) 9.32]
E) 4.35

$Mg(OH)_2 \rightleftharpoons Mg^{2+} + 2OH^-$

$K_{sp} = [Mg^{2+}][OH^-]^2$

Set Q at first pptn = Ksp

$8.9 \times 10^{-12} = (0.020)[OH^-]^2$

$[OH^-] = 2.11 \times 10^{-5}$

pOH = 4.68

pH = $9.\underline{32}$

7. $CuCO_3$ has a K_{sp} of 1.4×10^{-10}. It does not dissolve appreciably in water, but it is very soluble in aqueous HCl. Which of the following reactions best explains why the solubility of $CuCO_3$ is higher in HCl than in water? $HCl \rightarrow H^+ + Cl^-$

A) $CuCO_3(s) \rightleftharpoons CuO(s) + CO_2(g)$
B) $Cu^{2+}(aq) + 2H_2O \rightleftharpoons Cu(OH)_2(s) + 2H^+(aq)$
C) $CuCO_3(s) + H_2O \rightleftharpoons CuO(s) + H_2CO_3(aq)$
D) $CO_3^{2-}(aq) + 2H_2O \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$
[✓ E)] $CO_3^{2-}(aq) + 2H^+(aq) \rightleftharpoons H_2CO_3(aq)$

8. Fe^{3+} that is released from mining operations results in the acidification of rivers and lakes. Two important reactions are:

1. $Fe^{3+}(aq) + 6H_2O \rightleftharpoons [Fe(H_2O)_6]^{3+}(aq)$
2. $[Fe(H_2O)_6]^{3+}(aq) + H_2O \rightleftharpoons [Fe(H_2O)_5(OH)]^{2+}(aq) + H_3O^+(aq)$

1) Fe^{3+} accepts e- prs from H_2O

Reactions 1 and 2 can be described as what types of acid-base reactions?

	Reaction 1	Reaction 2
A)	Arrhenius	Lewis
B)	Brønsted-Lowry	Brønsted-Lowry
C)	Brønsted-Lowry	Arrhenius
[✓ D)]	Lewis	Brønsted-Lowry
E)	Arrhenius	Brønsted-Lowry

2) H_2O accepts H^+ to form H_3O^+

→ Call it HA

9. What is the pH of a 0.135 M solution of formic acid, the acid excreted by fire ants? The K_a of formic acid is 1.9×10^{-4}.

	HA	$\rightleftharpoons$	H^+	+	A^-
In:	0.135		0		0
Eq:	$0.135 - x$		x		x

[✓ A) 2.30]
B) 3.72
C) 0.870
D) 4.59
E) Cannot be determined without knowing the volume of the acid

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$1.9 \times 10^{-4} = \frac{x^2}{0.135 - \cancel{x}}$$

$$x = 5.06 \times 10^{-3} = [H^+]$$

$$\therefore pH = \underline{2.30}$$

10. Chlorine is used in swimming pools to kill bacteria. When chlorine is dissolved in water, it forms HClO, the active form of chlorine that kills bacteria.

The pH of a pool is important because HClO is a weak acid ($K_a = 4.0 \times 10^{-8}$). The conjugate base ClO^- cannot pass through a cell membrane, so it is much less effective at killing bacteria than is HClO.

$$HClO \rightleftharpoons H^+ + ClO^-$$

At what pH is the concentration of HClO two times higher than that of ClO^-?

If $[ClO^-] = x$; then $[HClO] = 2x$

[✓ A) 7.10]
B) 7.40
C) 7.70
D) 7.00
E) Cannot be determined without knowing the concentration of HClO

$$K_a = \frac{[H^+][ClO^-]}{[HClO]}$$

$$4.0 \times 10^{-8} = \frac{[H^+](x)}{2x}$$

$$[H^+] = 8.0 \times 10^{-8} \qquad pH = \underline{7.10}$$

11. What is the pH of a solution prepared by mixing 10 mL of 0.030 M NH_3 and 20 mL of 0.015 M HCl? The K_b of NH_3 is 1.8×10^{-5}.

mol NH_3 = 0.030 M × 0.010 L = 0.0003 mol
mol H^+ = 0.015 M × 0.020 L = 0.0003 mol

A) 4.54
B) 3.37
C) 4.39
[✓ D) 5.63]
E) 3.28

Rxn: $NH_3 + H^+ \rightarrow NH_4^+$

$$[NH_4^+] = \frac{0.0003\ mol}{0.030\ L} = 0.010M$$

Eq:	NH_4^+	$\rightleftharpoons$	H^+	+	NH_3
In:	0.010		0		0
Eq:	$0.010 - x$		x		x

$$K_a = \frac{K_w}{K_b} = 5.56 \times 10^{-10} = \frac{x^2}{0.010 - \cancel{x}}$$

$$x = 2.36 \times 10^{-6} = [H^+] \qquad pH = \underline{5.63}$$

12. Hydrocyanic acid, HCN, is used in industrial processes to clean metals and has a K_a of 4.0 × 10^{-10}. Which one of the following statements describing the pH of a solution of KCN is correct? $KCN \rightarrow K^+ + CN^-$

CN^- is the conj base of HCN

A) The solution is neutral because KCN is a salt.

B) The solution is acidic due to the reaction $HCN + H_2O \rightleftharpoons H_3O^+ + CN^-$.

C) The solution is acidic due to the reaction $K^+ + 2H_2O \rightleftharpoons KOH + H_3O^+$.

[✓ D)] The solution is basic due to the reaction $CN^- + H_2O \rightleftharpoons HCN + OH^-$.

E) The solution is basic due to the reaction
$K + H_2O \rightleftharpoons K^+ + OH^- + \frac{1}{2}H_2$.

13. Consider two weak acids, HX and HY. If HX is a stronger acid than HY, which one of the following statements is incorrect?

$K_a(HX) > K_a(HY) \quad \therefore pK_a(HX) < pK_b(HY)$

✓ A) The pK_a of HX is lower than that of HY.

✓ B) The $\Delta G°$ for the ionization of HX is more negative (less positive) than that of HY.

✓ C) A 0.100 M solution of HX is more ionized than that of HY.

✓ D) The conjugate base of HX is more stable than that of HY.

✗ [✓ E)] The pK_b of X^- is lower than that of Y^-.

$K_b(X^-) < K_b(Y^-) \quad \therefore pK_b(X^-) > pK_b(Y^-)$

14. The K_a of HF is greater than the K_a of CH_3COOH. Which of the following statements is/are correct? $K_a(HF) > K_a(CH_3COOH)$

✓ 1. A 1.0 M solution of CH_3COOH has a higher pH than a 1.0 M solution of HF.

✓ 2. A 1.0 M solution of CH_3COONa has a higher pH than a 1.0 M solution of NaF.

✗ 3. A 0.10 M solution of HF has a lower pH than a 1.0 M solution of HF.

$K_b(F^-) < K_b(CH_3COO^-)$

A) 1 and 3 only

[✓ B)] 1 and 2 only

C) 2 and 3 only

D) All statements are correct

E) None of the statements are correct

15. What is the pH of a solution formed by mixing 9.00 mL of 0.10 M CH_3COOH with 1.00 mL of 0.10 M HCl? The K_a of CH_3COOH is 1.8 × 10^{-5}.

A) 2.79

B) 3.40

[✓ C)] 2.00

D) 1.00

E) 1.48

$$[CH_3COOH] = \frac{0.10 \times 9}{10} = 0.090M$$

$$[HCl] = \frac{0.10 \times 1}{10} = 0.010M = [H^+]$$

$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$

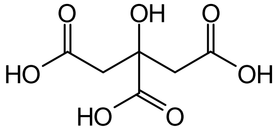
In: 0.090 0.010 0

EQ: 0.090 − x 0.010 + x x

$$K_a = 1.8 \times 10^{-5} = \frac{(0.010 + \cancel{x})(x)}{(0.090 - \cancel{x})} \qquad x = 1.62 \times 10^{-4}$$

$$\therefore [H^+] = 0.010M \qquad pH = \underline{2.00}$$

16. Citrus fruits contain a relatively high concentration of citric acid, the structure of which is shown below.

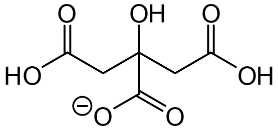


$K_{a1} = 8.1 \times 10^{-4}$
$K_{a2} = 1.8 \times 10^{-5}$ ∴ Acidic
$K_{a3} = 4.0 \times 10^{-7}$

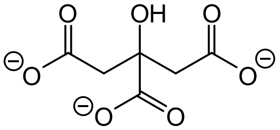
If 0.100 mol of monosodium citrate is dissolved in 1.00 L of water, which one of the following statements is/are correct about the solution?

✗ 1. The solution is basic.

✓ 2. The predominant form of citric acid is:



✓ 3. The concentration of the species shown on the right can be approximated to be equal to K_{a3}.



A) 1 and 3 only
[✓ B) 2 and 3 only]
C) 3 only
D) 1 and 2 only
E) 1 only

17. An acid buffer contains a weak acid and its conjugate base. Both the weak acid and its conjugate base must be present, ideally in comparable quantities, for the buffer to be as effective as possible. Which one of A – E best explains why both the weak acid and its conjugate base are necessary?

A) Acid-base indicators do not work unless both the weak acid and its conjugate base are present.

B) The weak acid removes any strong acid that may be added to the buffer, while the conjugate base removes any strong base that may be added to the buffer.

C) Having both the weak acid and its conjugate base is necessary to prevent the value of K_a from changing when a strong acid or base is added to the buffer.

D) The weak acid decreases the solubility of any strong base that may be added to the buffer, while the conjugate base decreases the solubility of any strong acid that may be added to the buffer.

[✓ E)] The weak acid removes any strong base that may be added to the buffer, while the conjugate base removes any strong acid that may be added to the buffer.

18. How many of the following mixtures are buffer solutions?

1) • 10 mL of 0.010M NaOH + 10 mL of 0.015 M CH_3COOH ✓
2) • 10 mL of 0.010M NaOH + 10 mL of 0.0050 M CH_3COOH X
3) • 10 mL of 0.010M CH_3COONa + 10 mL of 0.0050 M HNO_3 ✓
4) • 10 mL of H_2O + 10 mL of 0.015 M CH_3COOH X

A) 1 ✓1) Rxn forms CH_3COO^- with xs CH_3COOH
B) 3 X2) Rxn forms CH_3COO^- with xs NaOH
C) 4 ✓3) Rxn forms CH_3COOH with xs CH_3COO^-
[✓ D) 2] X4) Only dilution occurs
E) Cannot be determined without knowing the pH values of the solutions

19. What is the pH of a solution prepared by mixing 10 mL of 0.005 M NH_4Cl and 10 mL of 0.010 M NH_3. The K_b of NH_3 is 1.8×10^{-5}.

NH_3 - wk base, NH_4^+ - conj acid.

A) 4.44
[✓ B) 9.56] EQ: $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$
C) 5.05
D) 8.95
E) 9.08

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$1.8\times10^{-5} = \frac{[OH^-](.005)}{(0.010)}$$

$[OH^-] = 3.6\times10^{-5}$

pOH = 4.44, pH = 9.56

20. If a solution contains 0.40 mol of a weak acid (HA) and 0.25 mol of its conjugate base, (A^-) and the pH of the solution is 5.12, what is the K_a of the acid?

[✓ A) 4.74×10^{-6}] $[H^+] = 7.59\times10^{-6}$
B) 1.22×10^{-5} EQ: $HA \rightleftharpoons H^+ + A^-$
C) 8.22×10^{-10}
D) 2.45×10^{-5}
E) Cannot be determined without knowing the volume of the solution.

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$K_a = \frac{(7.59\times10^{-6})(0.25)}{0.40}$$

$$= \underline{4.74\times10^{-6}}$$

21. The pH inside a cell is maintained at about 7.40 by a phosphate buffer system.

$$H_2PO_4^-(aq) + H_2O \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq) \quad K_a = 6.2 \times 10^{-8}$$

If a cell has a pH of 7.40 and a total phosphate concentration ($H_2PO_4^- + HPO_4^{2-}$) of 0.020 M, what are the concentrations of $H_2PO_4^-$ and HPO_4^{2-} in the cell?

$[H^+] = 3.98 \times 10^{-8}$

	$[H_2PO_4^-]$	$[HPO_4^{2-}]$
A)	0.012 M	0.0078 M
B)	0.0072 M	0.013 M
[✓ C)]	0.0078 M	0.012 M
D)	0.017 M	0.0030 M
E)	Cannot be determined without knowing the volume of the cell	

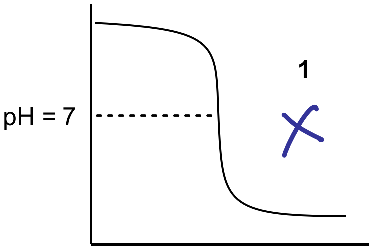
$$K_a = \frac{[H^+][HPO_4^{2-}]}{[H_2PO_4^-]}$$

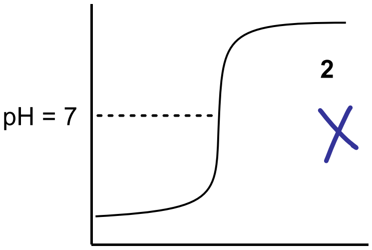
$$6.2 \times 10^{-8} = \frac{3.98 \times 10^{-8}\,[HPO_4^{2-}]}{[H_2PO_4^-]} \Rightarrow \frac{[HPO_4^{2-}]}{[H_2PO_4^-]} = 1.558$$

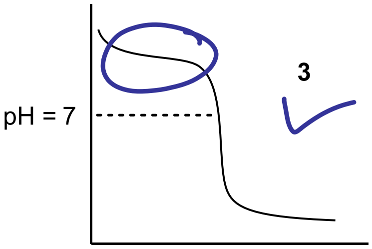
$$1.558 = \frac{x}{0.02 - x} \Rightarrow x = 0.012 = [HPO_4^{2-}]$$

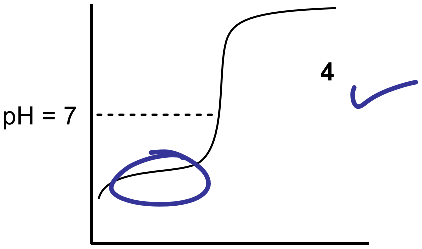
$$[H_2PO_4^-] = 0.078$$

22. Which of the following titration curves contain a buffering region?









A) 1 only
B) 1 and 2 only
C) 4 only
D) 3 only
[✓ E) 3 and 4 only]

23. Which one of A – E would be the best indicator for the titration of a weak acid, such as acetic acid ($K_a = 1.8 \times 10^{-5}$; $pK_a = 4.74$), with a strong base?

A) Methyl purple ($pK_{In} = 5.1$)
B) Thymol blue ($pK_{In} = 2.0$)
[✓ C) Phenol red ($pK_{In} = 7.6$)]
D) Bromophenol blue ($pK_{In} = 3.8$)
E) Bromocresol green ($pK_{In} = 4.8$)

At Eq. Pt, $pH > 7$ due to formation of conj. base

24. Which one of A – E is not a redox reaction?

A) $K + H_2O \rightarrow K^+ + \frac{1}{2}H_2 + OH^-$
B) $H_2O \rightarrow \frac{1}{2}O_2 + H_2$
C) $H_2SO_3 + \frac{1}{2}O_2 \rightarrow H_2SO_4$
D) $H_2 + C_2H_4 \rightarrow C_2H_6$
[✓ E)] $CH_3CH_2NH_2 + H_2O \rightarrow CH_3CH_2NH_3^+ + OH^-$

25. When the following reaction, which occurs in acidic solution, is balanced, what are the correct coefficients for Mn^{2+}, H_2O, and H^+?

$$MnO_2 \rightleftharpoons MnO_4^- + Mn^{2+}$$

Both A and B accepted

	MnO_4^-	H_2O	H^+
[✓ A)]	2	2	4
[✓ B)]	3 if Mn^{2+}	2	4
C)	2	2	2
D)	3	4	2
E)	2	4	2

$$MnO_2 + 2H_2O \rightarrow MnO_4^- + 4H^+ + 3e^-$$

$$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$$

$$2MnO_2 + \cancel{4H_2O} \rightarrow 2MnO_4^- + \cancel{8H^+} + \cancel{6e^-}$$

$$3MnO_2 + \overset{4}{\cancel{12}}H^+ + \cancel{6e^-} \rightarrow 3Mn^{2+} + \overset{2}{\cancel{6}}H_2O$$

$$\Rightarrow 5MnO_2 + 4H^+ \rightarrow 2MnO_4^- + 3Mn^{2+} + 2H_2O$$

26. How many of the following statements are correct?

1) - Under standard conditions, Cl_2 is a stronger oxidizing agent than H_2. ✓
2) - Under standard conditions, I^- is a weaker reducing agent than Cu. ✓
3) - Under standard conditions, H_2 gas would be formed if Ni were dissolved in HCl. ✓
4) - Under standard conditions, Zn metal can reduce Al^{3+} to Al. X

[✓ A) 3]
B) 2
C) 4
D) 1
E) None are correct

1) $Cl_2 + 2e^- \rightarrow 2Cl^-$ $E^0 = +1.36$
2) $2I^- \rightarrow I_2 + 2e^-$ $E^0 = -0.53$ $Cu \rightarrow Cu^{2+} + 2e^-$ $E^0 = -0.34$
3) $Ni \rightarrow Ni^{2+} + 2e^-$ $E^0 = +0.25$
4) $Zn \rightarrow Zn^{2+} + 2e^-$ $E^0 = +0.76$
$Al^{3+} + 3e^- \rightarrow Al$ $E^0 = -1.66$

27. What is the $E°_{cell}$ for the cell shown below?

$Co(s) \mid Co^{2+}(aq) \parallel MnO_4^-(aq), Mn^{2+}(aq) \mid Pt(s)$

A) 4.74 V
[✓ B) 1.79 V]
C) 1.23 V
D) −1.23 V
E) 1.62 V

$Co \rightarrow Co^{2+} + 2e^-$ $E^0_{ox} = +0.28$
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ $E^0_{red} = +1.51$
$E^0_{cell} = \underline{1.79\ V}$

28. What is the voltage for the cell shown below at 25°C?

$Al(s) \mid Al^{3+}(aq, 0.30\ M) \parallel Cd^{2+}(aq, 0.020\ M) \mid Cd(s)$

A) 2.06 V
B) 2.03 V
C) 1.53 V
[✓ D) 1.22 V]
E) 1.26 V

$Al \rightarrow Al^{3+} + 3e^-$ $E^0_{ox} = +1.66\ V$
$Cd^{2+} + 2e^- \rightarrow Cd$ $E^0_{red} = -0.40\ V$
RXN: $2Al + 3Cd^{2+} \rightarrow 2Al^{3+} + 3Cd$. $n = 6$
$E^0_{cell} = +1.26\ V$

$$E_{cell} = E^0_{cell} - \frac{0.0257}{n} \ln \frac{[Al^{3+}]^2}{[Cd^{2+}]^3}$$

$$= 1.26 - \frac{0.0257}{6} \ln \frac{(0.30)^2}{(0.020)^3}$$

$$= \underline{1.22\ V}$$

29. Which statements is/are correct about the cell based on the following reaction under standard conditions and basic pH?

$$NO_3^-(aq) + Mn(s) \rightarrow NO_2^-(aq) + MnOH_2(s)$$

✓ 1. The reaction, as written, proceeds spontaneously.
✓ 2. The cathode requires an inert electrode, such as Pt or graphite.
✓ 3. The Mn half-cell is the anode.

$NO_3^- + H_2O + 2e^- \rightarrow NO_2^- + 2OH^-$ $\mathcal{E}^\circ_{red} = 0$

$Mn + 2OH^- \rightarrow Mn(OH)_2 + 2e^-$ $\mathcal{E}^\circ_{ox} = +1.55$

$\mathcal{E}^\circ_{cell} = +1.55$ V

A) 2 and 3 only
B) 2 only
C) 3 only
D) 1 and 3 only
[✓ E) All statements are correct]

30. Given the following information, find the solubility product for AgBr at 298 K.

$AgBr(s) + e^- \rightarrow Ag(s) + Br^-(aq)$ $E° = +0.070$
$Ag^+(aq) + e^- \rightarrow Ag(s)$ $E° = +0.80$ Reverse, $\therefore \mathcal{E}^\circ = -0.80$

WANT: $AgBr(s) \rightleftharpoons Ag^+(aq) + Br^-(aq)$

$\mathcal{E}^\circ_{cell} = -0.73$ V

$\Delta G^\circ = -nF\mathcal{E}^\circ_{cell}$

$= [-(1)(96485)(-0.73)]$

$= 70434$

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

$70434 = (-8.314 \times 298) \ln K$

$K = 4.5 \times 10^{-13}$

A) 5.18×10^{14}
B) 1.9×10^{-15}
C) 2.2×10^{12}
[✓ D) 4.5×10^{-13}]
E) 7.0×10^{4}

31. How many of the following ions will form a precipitate when mixed with CO_3^{2-}?

Na^+ Ca^{2+} ✓ NH_4^+ Ba^{2+} ✓

A) 4
B) 1
[✓ C) 2]
D) 3
E) None

32. $Zn(OH)_2$ is insoluble in water. Which statement best explains why it is soluble in aqueous NH_3?

A) It is reduced by NH_3 into soluble Zn^{4+}.
B) It is oxidized by NH_3 into soluble Zn^{4+}.
C) It is reduced by NH_3 into soluble Zn.
D) The OH^- groups are neutralized by NH_3, leaving soluble Zn^{2+} behind.
[✓ E)] It reacts with NH_3 to form the soluble coordination complex $[Zn(NH_3)_4]^{2+}$.

33. When completing the preparation of a solution in a volumetric flask, water is added until…

A) All of the solid dissolves.
B) There is no air space left in the flask.
[✓ C)] The bottom of the meniscus on the line in the neck of the flask.
D) The bottom is filled but there is no water in the neck of the flask.
E) The meniscus turns upside-down.

34. In the equilibrium experiment, the following reaction was studied.

$$Fe^{3+} + SCN^- \rightleftharpoons Fe(SCN)^{2+}$$

Which statement best explains why $Fe(SCN)^{2+}$ concentrations were determined by taking absorbance measurements at 470 nm and not at another wavelength?

A) 470 nm is a wavelength that our eyes can see.
B) It is the only wavelength of light supported by the MicroLab unit.
C) Shorter wavelengths would be in the ultraviolet region and would have damaged the bonds present in $Fe(SCN)^{2+}$.
[✓ D)] It is the wavelength of light that $Fe(SCN)^{2+}$ absorbs the most.
E) Water does not absorb light at 470 nm.

End of test. Solutions and marks will be posted after the tests have been marked and statistical analyses have been performed.

R values

$0.0821\ \mathrm{L\ atm\ mol^{-1}\ K^{-1}}$

$0.0831\ \mathrm{L\ bar\ mol^{-1}\ K^{-1}}$

$8.314\ \mathrm{L\ kPa\ mol^{-1}\ K^{-1}}$

$8.314\ \mathrm{J\ mol^{-1}\ K^{-1}}$

$8.314\ \mathrm{kg\ m^{2}\ s^{-2}\ mol^{-1}\ K^{-1}}$

1 atm = 101.3 kPa

1 L kPa = 1 J

1 L atm = 101.3 J

K = °C + 273.15

$N_A = 6.02 \times 10^{23}$

$1\ \text{ampere} = 1\ \mathrm{C\ s^{-1}}$

$\Delta G° = -RT \ln K$

$K_w = 1.0 \times 10^{-14}$

$\text{Faraday} = 96485\ \mathrm{C\ mol^{-1}}$

$\Delta G° = -nFE°$

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

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \left(\frac{RT}{nF}\right)\ln Q$$

Scrap paper