

Key

CHM 152 Exam 3 100 Pts Fall 2006 Name: \_\_\_\_\_

1. Calculate the pH of a buffer solution that contains 0.25 M benzoic acid ( $C_6H_5CO_2H$ ) and 0.15 M sodium benzoate ( $C_6H_5COONa$ ). Given  $K_a = 6.5 \times 10^{-5}$  for benzoic acid.  $pH = pK_a + \log \frac{0.15}{0.25}$
- (A) 3.97      B. 4.83      C. 4.19      D. 3.40      E. 4.41
2. A solution is prepared by mixing 500 mL of 0.10 M NaOCl and 500 mL of 0.20 M HOCl. What is the pH of this solution?  $K_a(HOCl) = 3.2 \times 10^{-8}$   $pH = pK_a + \log \frac{0.05}{0.1}$
- A. 4.10      B. 7.00      C. 7.19      D. 7.49      E. 7.80
3. The pH at the equivalence point of a titration may differ from 7.0 because of
- A. the initial concentration of the standard solution.  
B. the indicator used.  
C. the self-ionization of  $H_2O$ .  
D. the initial pH of the unknown.  
(E) hydrolysis of the salt formed.
4. For which type of titration will the pH be basic at the equivalence point?
- A. Strong acid vs. strong base.  
B. Strong acid vs. weak base.  
(C) Weak acid vs. strong base.  
D. All of the above.  
E. None of the above.
5. What is the pH at the equivalence point in the titration of 100 mL of 0.10 M HCl with 0.10 M NaOH?
- A. 1.0      B. 6.0      C. 7.0      D. 8.0      E. 13.0
6. For  $PbCl_2$ ,  $K_{sp} = 2.4 \times 10^{-4}$ , will a precipitate of  $PbCl_2$  form when 0.10 L of  $3.0 \times 10^{-2}$  M  $Pb(NO_3)_2$  is added to 400 mL of  $9.0 \times 10^{-2}$  M NaCl? Choose one of the following.  $Q_{sp} = [0.006][0.072]^2 = 7.1 \times 10^{-5}$
- A. Yes,  $Q > K_{sp}$ .      (B) No,  $Q < K_{sp}$ .      C. No,  $Q = K_{sp}$ .      D. Yes,  $Q < K_{sp}$ .
7. The molar solubility of  $MgCO_3$  is  $1.8 \times 10^{-4}$  mol/L. What is  $K_{sp}$  for this compound?
- A.  $1.8 \times 10^{-4}$   
B.  $3.6 \times 10^{-4}$   
C.  $1.3 \times 10^{-7}$   
(D)  $3.2 \times 10^{-8}$   
E.  $2.8 \times 10^{-14}$
- $MgCO_3(s) \rightleftharpoons Mg^{2+} + CO_3^{2-}$   
 $K_{sp} = (Mg^{2+})(CO_3^{2-}) = (1.8 \times 10^{-4})^2$
8. The molar solubility of tin iodide ( $SnI_2$ ) is  $1.28 \times 10^{-2}$  mol/L. What is  $K_{sp}$  for this compound?
- (A)  $8.4 \times 10^{-6}$       B.  $1.28 \times 10^{-2}$       C.  $4.2 \times 10^{-6}$       D.  $1.6 \times 10^{-4}$       E.  $2.1 \times 10^{-6}$

$$K_{sp} = (1.28 \times 10^{-2}) [2 \times 1.28 \times 10^{-2}]^2$$

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9. The  $K_{sp}$  for  $Ag_3PO_4$  is  $1.8 \times 10^{-18}$ . Determine the  $Ag^+$  ion concentration in a saturated solution of  $Ag_3PO_4$ .

- A.  $1.6 \times 10^{-5} M$
- B.  $2.1 \times 10^{-5} M$
- C.  $3.7 \times 10^{-5} M$
- D.  $1.1 \times 10^{-13} M$
- E.  $4.8 \times 10^{-5} M$

$$Ag_3PO_4(s) \rightleftharpoons 3Ag^+ + PO_4^{3-}$$

$$K_{sp} = [3x]^3 [x]$$

$$K_{sp} = 27x^4$$

$$\sqrt[4]{\frac{K_{sp}}{27}} = x = 1.6 \times 10^{-5}$$

$$3x = 4.8 \times 10^{-5}$$

10. Which of the following would decrease the  $K_{sp}$  for  $PbI_2$ ?

- A. Lower the pH of the solution
- B. Add a solution of  $Pb(NO_3)_2$
- C. Add a solution of KI
- D. None of the above—the  $K_{sp}$  of a compound is constant at constant temperature.

11. Calculate the minimum concentration of  $Mg^{2+}$  that must be added to 0.10 M NaF solution in order to initiate a precipitate of magnesium fluoride. For  $MgF_2$ ,  $K_{sp} = 6.9 \times 10^{-9}$ .

- A.  $1.4 \times 10^7 M$
- B.  $6.9 \times 10^{-9} M$
- C.  $6.9 \times 10^{-8} M$
- D.  $1.7 \times 10^{-7} M$
- E.  $6.9 \times 10^{-7} M$

$$MgF_2(s) \rightleftharpoons Mg^{2+} + 2F^-$$

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

$$6.9 \times 10^{-9} = [Mg^{2+}][0.10]^2 =$$

$$[Mg^{2+}] = 6.9 \times 10^{-7}$$

12. Arrange the following compounds in order of increasing standard molar entropy at 25°C:

$C_3H_8(g)$ ,  $C_2H_4(g)$ ,  $ZnS(s)$ , and  $H_2O(l)$ .

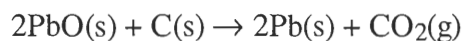
- A.  $ZnS(s) < H_2O(l) < C_3H_8(g) < C_2H_4(g)$
- B.  $C_2H_4(g) < H_2O(l) < C_3H_8(g) < NaCl(s)$
- C.  $ZnS(s) < C_3H_8(g) < C_2H_4(g) < H_2O(l)$
- D.  $C_3H_8(g) < C_2H_4(g) < H_2O(l) < ZnS(s)$
- E.  $ZnS(s) < H_2O(l) < C_2H_4(g) < C_3H_8(g)$

13. Which response includes all the following processes that are accompanied by an increase in entropy?

1.  $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_3(\text{g})$
2.  $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{s})$
3.  $\text{Br}_2(\text{l}) \rightarrow \text{Br}_2(\text{g})$
4.  $\text{H}_2\text{O}_2(\text{l}) \rightarrow \text{H}_2\text{O}(\text{l}) + 1/2\text{O}_2(\text{g})$

- A. 1, 2, 3, 4      B. 1, 2      C. 2, 3, 4      **D. 3, 4**      E. 1, 4

14. Given the equation for the reduction of  $\text{PbO}(\text{s})$ .



Calculate  $\Delta S^\circ$  for this reaction at  $25^\circ\text{C}$ , given the following absolute entropies:

	$S^\circ\text{J}/(\text{K}\cdot\text{mol})$
$\text{PbO}(\text{s})$	69.45
$\text{C}(\text{s})$	5.7
$\text{Pb}(\text{s})$	64.89
$\text{CO}_2(\text{g})$	213.6

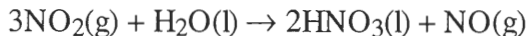
$$\Delta S = \sum S_{\text{prod}} - \sum S_{\text{react}}$$

- A.** +198.8 J/K      B. +488.0 J/K      C. +353.6 J/K      D. -203.3 J/K      E. +203.3 J/K

15. A negative sign for  $\Delta G$  indicates that, at constant T and P,

- A. the reaction is exothermic.
- B. the reaction is endothermic.
- C. the reaction is fast.
- D.** the reaction is spontaneous.
- E.  $\Delta S$  must be  $> 0$ .

16. Calculate  $\Delta G^\circ$  for the following reaction



Given the following free energies of formation

$$\Delta G = \sum G_{\text{prod}} - \sum G_{\text{react}}$$

	$\Delta G^\circ_f(\text{kJ/mol})$
$\text{H}_2\text{O}(\text{l})$	-237.2
$\text{HNO}_3(\text{l})$	-79.9
$\text{NO}(\text{g})$	86.7
$\text{NO}_2(\text{g})$	51.8

- A. 8.7 kJ      B. 192 kJ      C. -414 kJ      D. -192 kJ      E. -155 kJ

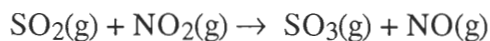
17. For the reaction  $\text{H}_2(\text{g}) + \text{S}(\text{s}) \rightarrow \text{H}_2\text{S}(\text{g})$

$\Delta H^\circ = -20.2 \text{ kJ}$  and  $\Delta S^\circ = +43.1 \text{ J/K}^{-1}$ . Which of the following statements is true?

- A. The reaction is only spontaneous at low temperatures.  
 B. The reaction is spontaneous at all temperatures.  
 C.  $\Delta G^\circ$  becomes less favorable as temperature increases.  
 D. The reaction is spontaneous only at high temperatures.  
 E. The reaction is at equilibrium at  $25^\circ\text{C}$  under standard conditions.

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

18. Calculate  $K_p$  for the reaction below at 298 K.



Given the following free energies of formation:

	$\Delta G^\circ_f$
$\text{SO}_2(\text{g})$	-300.4 kJ/mol
$\text{SO}_3(\text{g})$	-370.4 kJ/mol
$\text{NO}(\text{g})$	86.7 kJ/mol
$\text{NO}_2(\text{g})$	51.8 kJ/mol

- A.  $6.99 \times 10^{-7}$   
 B.  $5.71 \times 10^{-8}$   
 C. 14.2  
 D. 475  
 E.  $1.43 \times 10^6$

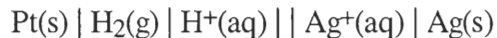
$$\Delta G^\circ = -35.1 \text{ kJ}$$

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

$$\frac{35.1 \times 10^3 \text{ J}}{(8.314 \text{ J})(298 \text{ K})} = \ln K_p$$

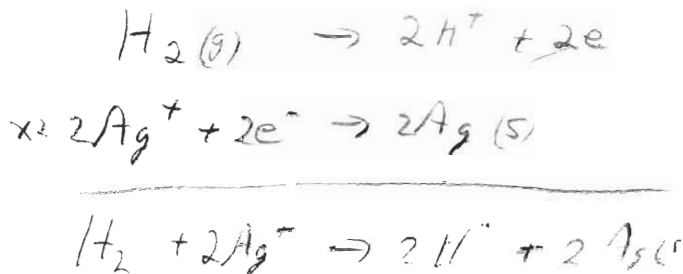
$$14.167 = \ln K_p$$

19. Given the following notation for an electrochemical cell

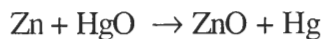


what is the balanced overall (net) cell reaction?

- A.  $2\text{H}^+(\text{aq}) + 2\text{Ag}^+(\text{aq}) \rightarrow \text{H}_2(\text{g}) + 2\text{Ag(s)}$
- B.  $\text{H}_2(\text{g}) + 2\text{Ag(s)} \rightarrow \text{H}^+(\text{aq}) + 2\text{Ag}^+(\text{aq})$
- C.  $2\text{H}^+(\text{aq}) + 2\text{Ag(s)} \rightarrow \text{H}_2(\text{g}) + 2\text{Ag}^+(\text{aq})$
- D.  $\text{H}_2(\text{g}) + \text{Ag}^+(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{Ag(s)}$
- E.  $\text{H}_2(\text{g}) + 2\text{Ag}^+(\text{aq}) \rightarrow 2\text{H}^+(\text{aq}) + 2\text{Ag(s)}$**



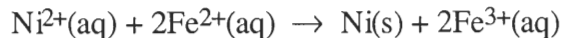
20. A certain electrochemical cell has for its cell reaction:



Which is the half-reaction occurring at the anode?

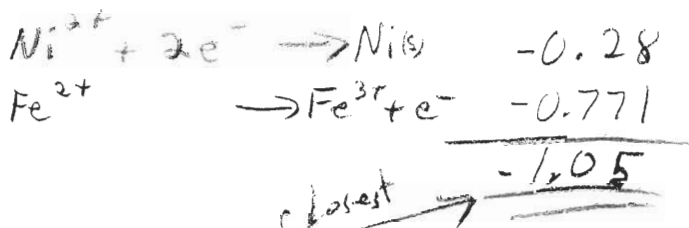
- A.  $\text{HgO} + 2\text{e}^- \rightarrow \text{Hg} + \text{O}^{2-}$
- B.  $\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$
- C.  $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$**
- D.  $\text{ZnO} + 2\text{e}^- \rightarrow \text{Zn}$

21. For the reaction

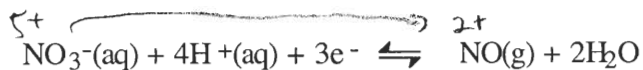


The standard cell potential  $E^\circ_{\text{cell}}$  is

- A. +2.81 V.
- B. +1.02 V.
- C. +0.52 V.
- D. -1.02 V.**
- E. -2.81 V.



22. In the following half equation, which is the oxidizing agent?



- A.  $\text{NO}_3^-$**
- B.  $\text{H}^+$
- C.  $\text{e}^-$
- D.  $\text{NO}$
- E.  $\text{H}_2\text{O}$

23. Given the following standard reduction potentials in acid solution

	$E^\circ(V)$	
$Al^{3+} + 3e^- \rightleftharpoons Al(s)$	-1.66	
$AgBr(s) + e^- \rightleftharpoons Ag(s) + Br^-$	+0.07	
$Sn^{4+} + 2e^- \rightleftharpoons Sn^{2+}$	+0.14	
$Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$	+0.77	← most positive for

*Ox agent gains e<sup>-</sup>*

The strongest oxidizing agent among those shown above is

- A. Fe<sup>3+</sup>      B. Fe<sup>2+</sup>      C. Br<sup>-</sup>      D. Al<sup>3+</sup>      E. Al

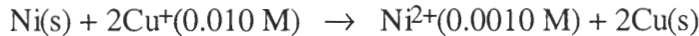
24. Which one of the following reagents is capable of transforming Cu<sup>2+</sup>(1 M) to Cu(s)? *Cu<sup>2+</sup> + 2e<sup>-</sup> → Cu +0.337V*

- A. I<sup>-</sup>(aq)      B. Ni(s)      C. Al<sup>3+</sup>(1 M)      D. F<sup>-</sup>(1 M)      E. Ag(s)

25. Which of the following species is the strongest oxidizing agent under standard state conditions?

- A. Ag<sup>+</sup>(aq)      B. H<sub>2</sub>(g)      C. H<sup>+</sup>(aq)      D. Cl<sub>2</sub>(g)      E. Al<sup>3+</sup>(aq)

26. Calculate the cell emf for the following reaction:



- A. 0.40 V      B. -0.43 V      C. 0.43 V      D. 0.34 V      E. 0.37 V

0.800V

27. A metal object is to be gold-plated by an electrolytic procedure using aqueous AuCl<sub>3</sub> electrolyte. Calculate the number of moles of gold deposited in 3.0 min by a constant current of 10 A.

- A. 6.2 × 10<sup>-3</sup> mol  
 B. 9 × 10<sup>-3</sup> mol  
 C. 1.8 × 10<sup>-2</sup> mol  
 D. 3.5 × 10<sup>-5</sup> mol  
 E. 160 mol

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

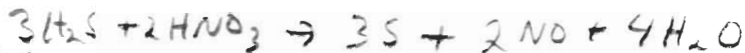
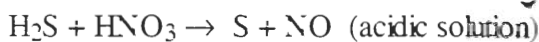
3.0 min	60s	10A	mol e <sup>-</sup>	1 mol Au	=
	1 min	8	9.65 × 10 <sup>4</sup> e <sup>-</sup>	3 mole	

28. Aluminum does not corrode as does iron, because

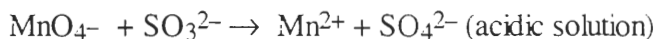
- A. Al does not react with O<sub>2</sub>.  
 B. A protective layer of Al<sub>2</sub>O<sub>3</sub> forms on the metal surface.  
 C. Al is harder to oxidize than is Fe.  
 D. Fe gives cathodic protection to Al.  
 E. The electrical circuit cannot be completed on an Al surface.

NO  
ANSWER

29. Complete and balance the following redox equation. When properly balanced with whole number coefficients, the coefficient of S is



- A. 1                      B. 2                      **C. 3**                      D. 5                      E. 6
30. Complete and balance the following redox equation. What is the coefficient of H<sub>2</sub>O when the equation is balanced with the set of smallest whole numbers?



**A. 3**

B. 4

C. 5

D. 8

E. None of the above.



31. Complete and balance the following redox equation. What is the coefficient of H<sub>2</sub>O when the equation is balanced with the set of smallest whole numbers?



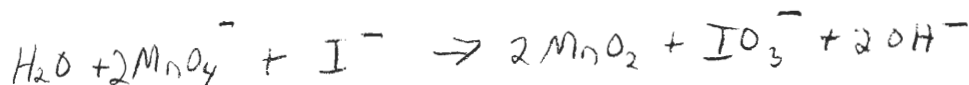
**A. 1**

B. 2

C. 4

D. 10

E. None of the above.



32. Which choice gives the correct oxidation numbers for all three elements in Rb<sub>2</sub>SO<sub>3</sub> in the order that the elements are shown in the formula?

A. -2, +6, -2

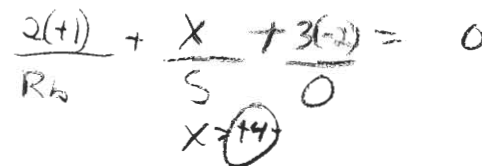
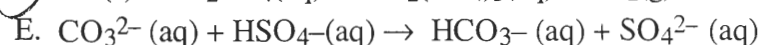
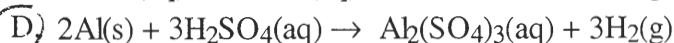
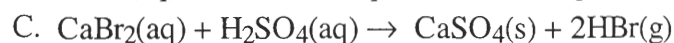
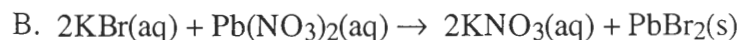
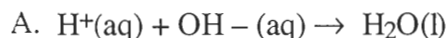
B. -1, +4, -3

C. +2, +4, -2

**D. +1, +4, -2**

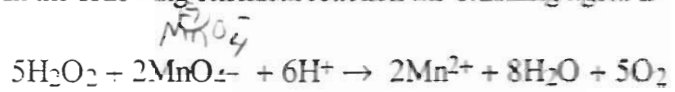
E. +1, +6, -6

33. Which one of the following is a redox reaction?



— Over —

34. In the following chemical reaction the oxidizing agent is



- A.  $\text{H}_2\text{O}_2$       B.  $\text{MnO}_4^-$       C.  $\text{H}^+$       D.  $\text{Mn}^{2+}$       E.  $\text{O}_2$

