Name: KEY

Chem. 115 Practice

Exam #4

Part 1: Multiple Choice

1. Calculate the solubility product constant for lead(II) iodide if 0.0024 mole of I ion is present in 2.0 L of a saturated lead(II) iodide solution

   a. 1.4×10^{-5}

   b. 8.6×10^{-10}

   c. 5.2×10^{-8}

   d. 3.5×10^{-6}

   e. 4.6×10^{-9}

2. Calculate the number of moles of Ag^+ ion present in 2.0 L of a saturated solution of silver chromate. For silver chromate, $K_{sp} = 1.1 \times 10^{-12}$.

   a. 2.6×10^{-4}

   b. 1.3×10^{-4}

   c. 2.1×10^{-4}

   d. 1.1×10^{-4}

   e. 4.1×10^{-4}

3. Calculate the molar solubility of silver carbonate in 1.0 M sodium carbonate solution. For silver carbonate, $K_{sp} = 8.1 \times 10^{-12}$.

   a. 8.1×10^{-12}

   b. 2.8×10^{-6}

   c. 1.4×10^{-6}

   d. 1.4×10^{-8}

   e. 2.0×10^{-4}

4. Calculate the pH of a solution necessary to just begin the precipitation of magnesium hydroxide when the concentration of magnesium ion = 0.001 M. For magnesium hydroxide $K_{sp} = 1.2 \times 10^{-11}$.

   a. 11

   b. 10

   c. 9

   d. 8

   e. 4

5. The line notation, $Mg(s) \mid Mg^{2+}(aq) \parallel Fe^{2+}(aq) \mid Fe(s)$, indicates that:

   a. iron metal is the reducing agent

   b. magnesium metal is the cathode

   c. Fe^{2+} ions are oxidized

   d. magnesium metal is the reducing agent

   e. Mg^{2+} ion is the reducing agent

6. Consider the following two electrode reactions and their standard electrode potentials:

$$Al^{+3}(aq) + 3e^- \rightarrow Al(s) \quad E° = -1.66 \text{ V}$$

$$Cd^{+2}(aq) + 2e^- \rightarrow Cd(s) \quad E° = -0.40 \text{ V}$$

   Write the cell reaction for a voltaic cell based on these two electrodes, and calculate the standard cell potential

   a. $2Al^{+3}(aq) + 3Cd^{+2}(aq) \rightarrow 2Al(s) + 3Cd(s) \quad E°_{cell} = 2.10 \text{ V}$

   b. $2Al(s) + 3Cd^{+2}(aq) \rightarrow 2Al^{+3}(aq) + 3Cd(s) \quad E°_{cell} = 1.26 \text{ V}$

   c. $2Al(s) + 3Cd^{+2}(aq) \rightarrow 2Al^{+3}(aq) + 3Cd(s) \quad E°_{cell} = 3.78 \text{ V}$

   d. $2Al^{+3}(aq) + 3Cd(s) \rightarrow 2Al(s) + 3Cd^{+2}(aq) \quad E°_{cell} = 1.26 \text{ V}$

   e. $2Al^{+3}(aq) + 3Cd(s) \rightarrow 2Al(s) + 3Cd^{+2}(aq) \quad E°_{cell} = 2.10 \text{ V}$

7. A voltaic cell consists of Mn/Mn^{2+} and Cd/Cd^{2+} half-cells with concentrations $[Mn^{2+}] = 0.75$ M and $[Cd^{2+}] = 0.15$ M. Calculate the cell potential at 25° C.

   $Cd^{+2}(aq) + 2e^- \rightarrow Cd(s)$ $E° = -0.40$ V

   $Mn^{+2}(aq) + 2e^- \rightarrow Mn(s)$ $E° = -1.18$ V

   a. 1.60 V

   b. 1.56 V

   c. 1.54 V

   d. 0.80 V

   e. 0.76 V

8. The standard reference electrode that is used to measure all other standard electrode potentials is called the "standard _____ electrode."

   a. sulfur

   b. oxygen

   c. hydrogen

   d. iron

   e. platinum

9. A constant current was passed through a solution of $KAuCl_4$ between gold electrodes. Over a period of 20.00 min, the cathode increased in mass by 2.664 g. What was the current in amperes?
   (F = 96500 C/mol) Cathode half-reaction: $AuCl_4^-(aq) + 3e^- \rightarrow Au(s) + 4Cl^-(aq)$

   a. 1.08 A

   b. 3.26 A

   c. 2.17 A

   d. 6.52 A

   e. 3.48 A

10. Calculate the equilibrium constant K_c for this reaction at 25°C:
$2Cr(s) + 3Pb^{2+}(aq) \rightarrow 3Pb(s) + 2Cr^{3+}(aq)$ $E°_{cell} = 0.61$ V

a. 6.7×10^{61}

b. 8.1×10^{30}

c. 9.2×10^{45}

d. 3.2×10^{51}

e. 4.6×10^{22}

11. Consider the following half-reactions and select the strongest oxidizing agent present:
$Sr^{2+}(aq) + 2e^- \rightarrow Sr(s)$ $E° = -2.89$ V
$Cr^{2+}(aq) + 2e^- \rightarrow Cr(s)$ $E° = -0.913$ V
$Co^{2+}(aq) + 2e^- \rightarrow Co(s)$ $E° = -0.28$ V

a. $Cr^{2+}(aq)$

b. $Sr^{2+}(aq)$

c. $Co^{2+}(aq)$

d. $Sr(s)$

e. $Co(s)$

12. The standard free energy change for the following reaction is -210 kJ. What is the cell potential?

$$2H_2O_2(aq) \leftrightarrow 2H_2O(l) + O_2(g)$$

a. +0.640 V

b. +1.09 V

c. +0.420 V

d. +0.547 V

e. +0.752 V

13. In the electrolysis of aqueous sodium sulfate, which one of the following species is oxidized?

$S_2O_8^{2-}(aq) + 2e^- \rightarrow 2SO_4^{2-}(aq)$ $E° = 2.10$ V

$Na^+(aq) + e^- \rightarrow Na(s)$ $E° = -2.71$ V

$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ $E° = -0.38$V

$O_2(g) + 4H^+(aq) + 4e^- \rightarrow H_2O(l)$ $E° = 1.23$V

a. sodium ion

b. oxygen

c. sulfate ion

d. water

e. hydronium ion

14. Calculate $\Delta S°$ for the following reaction:

Standard molar entropies, $S°$ (J/mol·K): Cr(s), 23.8; $O_2(g)$, 205.1; $Cr_2O_3(s)$, 81.2

$$4Cr(s) + 3O_2(g) \rightarrow 2Cr_2O_3(s)$$

a. 548.1 J/K

b. 147.7 J/K

c. −147.7 J/K

d. −548.1 J/K

e. −66.5 J/K

15. When crystalline solid barium hydroxide octahydrate and crystalline solid ammonium nitrate are mixed in a beaker at room temperature, a spontaneous reaction occurs. The temperature of the beaker contents rapidly falls to below 0°C. Use this information to decide whether the reaction is exothermic or endothermic and what the signs of ΔH and ΔS are.

a. endothermic; $\Delta H > 0; \Delta S > 0$

b. exothermic; $\Delta H < 0; \Delta S > 0$

c. endothermic; $\Delta H < 0; \Delta S < 0$

d. endothermic; $\Delta H < 0; \Delta S > 0$

e. exothermic; $\Delta H > 0; \Delta S < 0$

16. A certain reaction has $\Delta H° = -14.2$ kJ and $\Delta S° = +87.9$ J/K. What is the value of $\Delta G°$ for this reaction? (Temperature is 25° C)

a. +40.4 kJ

b. -16.4 kJ

c. -26200 kJ

d. -40.4 kJ

e. -7820 kJ

17. A reaction has an equilibrium constant $K_c = 7.0$ at 35 °C. Calculate the value of $\Delta G°$ for the reaction

a. -4.98 kJ

b. -2.46 kJ

c. -5.66 kJ

d. -2.16 kJ

e. none of the above

18. The compound 1-pentanol has an enthalpy of vaporization of 55.5 kJ/mol and an entropy of vaporization of 148 J/K·mol. Calculate its approximate boiling point.

a. 45°C

b. 102°C

c. 93°C

d. 210°C

e. 375°C

19. Which of the following statements about entropy and enthalpy of a system is correct?

a. The absolute entropy of pure oxygen at 25°C and 1 atm is zero.

b. When ice melts, ΔS is positive and ΔH is negative.

c. When a candle burns, ΔS is positive and ΔH is negative.

d. The entropy of a system must increase for the reaction to be spontaneous.

e. None of the above statements are correct.

20. Sodium carbonate can be made by heating sodium bicarbonate carbonate:

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

For this reaction, $\Delta H° = 128.9$ kJ and $\Delta S° = 321$ J/K. At approximately what temperature will K = 1?

a. 401.6° C

b. 401.6 K

c. 33.1° C

d. 33.1 K

e. None of the above

Part 2: Free Response. Please show all work.

1. Will a precipitate form when 125 ml of 0.0250 M aluminum nitrate and 25.0 ml of 0.000100 M calcium hydroxide are mixed together? Why? K_{sp} of aluminum hydroxide = 3.7×10^{-15}

$$Q_{sp} = [Al^{3+}][OH^-]^3$$

$$[Al^{3+}]_i = [Al(NO_3)_3]_i = 0.0250\ M$$

$$[OH^-]_i = 2 \cdot [Ca(OH)_2] = 2(0.000100\ M) = 0.000200\ M$$

$$Q_{sp} = \left[\frac{(0.0250M)(125mL)}{(150\ mL)}\right]\left[\frac{(0.000200M)(25.0\ mL)}{(150\ mL)}\right]^3$$

* remember to dilute!

$$Q_{sp} = 7.7 \times 10^{-16}$$

$Q_{sp} < K_{sp}$ ∴ **No precipitate.**

2. In the electrolysis of a molten mixture of potassium iodide and magnesium fluoride, identify which product forms at the positive electrode, and what product forms at the negative electrode.

$K^+(aq) + e^- \rightarrow K(s)$ $E° = -2.93$ V

$I_2(s) + 2e^- \rightarrow 2I^-$ $E° = -0.53$ V

$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$ $E° = -2.37$ V

$F_2(g) + 2e^- \rightarrow 2F^-$ $E° = 2.87$ V

In our mixture we have: K^+, I^-, Mg^{2+}, and F^-

Let's use the given $E°_{red}$ values to determine which half-reactions occur most easily (require the least amount of Energy).

Possible reductions:

$K^+ + e^- \rightarrow K(s)$ $E°_{red} = -2.93$ V

* $Mg^{2+} + 2e^- \rightarrow Mg(s)$ $E°_{red} = -2.37$ V

Possible oxidations:

* $2I^- \rightarrow 2e^- + I_2(s)$ $E°_{ox} = +0.53$ V

$2F^- \rightarrow 2e^- + F_2(g)$ $E°_{ox} = -2.87$ V

The 2 reactions with a * next to them will occur (higher Emf's).

∴ **$Mg(s)$ and $I_2(s)$ will form.**

3. In a $Cd^{2+}/Cd(s)$ and $Au^{3+}/Au(s)$ voltaic cell the electronic voltmeter measures to be 1.92-Volts. What concentration of cadmium (II) ion must be present in the cell if the gold (III) ion concentration is known to be 0.10 M.

$Au^{3+}(aq) + 3e^- \rightarrow Au(s)$ $E° = 1.50$

flip! → $Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$ $E° = -0.40$

$$E°_{cell} = 1.50 + 0.40\ V$$

$$E°_{cell} = 1.90\ V$$

$$E_{cell} = E°_{cell} - \frac{RT}{nF} \ln Q$$

$$(\ln Q)\frac{RT}{nF} = E°_{cell} - E_{cell}$$

$$\ln Q = [E°_{cell} - E_{cell}] \cdot \frac{nF}{RT}$$

$$Q = e^{[E°_{cell} - E_{cell}] \cdot \frac{nF}{RT}} \quad \text{——— (1)}$$

The overall rxn is:

$$2 \times \left[Au^{3+} + 3e^- \rightarrow Au(s) \right]$$

$$+ \; 3 \times \left[Cd(s) \rightarrow 2e^- + Cd^{2+} \right]$$

$$2Au^{3+} + 3Cd(s) \rightarrow 2Au(s) + 3Cd^{2+}$$

note: number of e^- transferred is $n = 6$. ——— (2)

and, $Q = \dfrac{[Cd^{2+}]^3}{[Au^{3+}]^2}$ ——— (3)

(2) & (3) plug into (1)

$$\frac{[Cd^{2+}]^3}{[Au^{3+}]^2} = e^{\left[E^\circ_{cell} - E_{cell}\right]\frac{nF}{RT}}$$

$$[Cd^{2+}]^3 = [Au^{3+}]^2 \cdot e^{\left[E^\circ_{cell} - E_{cell}\right]\frac{nF}{RT}}$$

$$[Cd^{2+}] = \left[[Au^{3+}]^2 \cdot e^{\left[E^\circ_{cell} - E_{cell}\right]\frac{nF}{RT}} \right]^{1/3}$$

$$[Cd^{2+}] = \left[(0.10\,M)^2 \cdot e^{\left[1.90 - 1.92\,V\right]\frac{(6\ mol\ e^-)\cdot(96500\ \frac{C}{mol\ e^-})}{(8.314\ \frac{J}{mol\,K})\cdot(298\,K)}} \right]^{1/3}$$

$[Cd^{2+}] = 0.18\,M$

$[Cd^{2+}] = 0.045\,M$

Always double check your calculator's answers!