Chemistry 102 2018 Discussion #9, Chapter 16

Student name____________________ TA name_________________ Section________________

Things you should know when you leave Discussion today:

1. Reduction potential $E°_{red}$ at standard conditions use Chemistry: Human Activity, Chemical reactivity, Mahaffy et al., 2e Appendix F for information.
2. Cell potential $E°_{cell}$ =[J/c=J/charge] is a measure of electrical potential difference.

$$\boldsymbol{E°_{cell} = E°_{red}(cathode) - E°_{red}(anode)} = \frac{0.05912V}{n_e} \log K \text{ (at 25°C)}$$

3. http://quantum.bu.edu/courses/ch102-spring-2016/notes/BalancingRedoxEquations.pdf
4. Gibbs Free Energy and Potential : $\boldsymbol{\Delta G° = -n_e \cdot \mathcal{F} \cdot E°}$; $\mathcal{F} = 96485.3 \left(\frac{charge}{mol(electrons)}\right)$

1. At 20 °C, an ionic solid, $A_2X_3(s)$, has $K_{sp} = 5.0 \times 10^{-14}$. A 0.50L solution of 1.00 M $A^{3+}(aq)$ and a 0.50 L solution of 1.0 M $X^{2-}(aq)$ are combined, and $A_2X_3(s)$ precipitates. Calculate the molarity of $X^{2-}(aq)$ after the combine solution has come to equilibrium.

2. Choose all that must be true:
   a. Anode is an electrode where oxidation takes place
   b. Cathode is an electrode where reduction takes place

   c. Most positive reduction potential represents the strongest Oxidizing Agent (an oxidizing agent is an electron acceptor and is most likely to get reduced)

   d. Most negative reduction potential represents strongest Reducing Agent (most likely to get oxidized)

   e. ***E*°** is an Intensive quantity as long as concentration does not change it does not depend on the size of the cell

   f. ***E*°** does not depend on the amount of charge transferred.

   g. ***E*°** does not depend on stoichiometry. (doubling the stoichiometry of all the species in the reaction has no effect on the ***E*°**)

   h. Sign of the cell potential tells the direction of electron flow.

   i. Only reduction potentials are listed in the reference tables. (oxidation potentials can be calculated from reduction potentials $\boldsymbol{E°_{ox} = -E°_{red}}$)

3. Write the balanced equation for the reduction of Fe^{2+}(aq) to Fe (s) using H_2 (g). **Assign for each half reaction if it is a cathode half reaction or an anode half reaction.** **(**E^o_{red} (Fe^{2+} / Fe) = −0.44)

Oxidation half RXN:

Reduction half RXN:

Net Redox RXN in Acidic Media:
Draw the galvanic cell corresponding to the redox reaction

Write the line notation for the galvanic cell corresponding to the redox reaction, using platinum electrodes, Pt(*s*) if needed.

E^o_{cell} =

4. Write the balanced equation for the oxidation of NO (g) to NO_3^- (aq) by reduction of dichromate, $Cr_2O_7^{2-}$ (aq) to Cr^{3+} (aq). **Assign for each half reaction if it is a cathode half reaction or an anode half reaction.**

Oxidation half RXN:

Reduction half RXN:

Net RXN in Acidic Media:

Write the line notation for the galvanic cell corresponding to the redox reaction, using platinum electrodes, Pt(*s*) if needed.

5. When zinc metal is burned in air, ZnO (s) is formed. Assuming the same chemical reaction, formation of zinc oxide from zinc metal, takes place instead in galvanic cell write the balanced reactions.

Oxidation half RXN:

Reduction half RXN:

Net RXN :

6. Answer the questions for the following redox reaction: Cu*(s)* | Cu^{2+}*(aq)* || Ag^{+}*(aq)* | Ag*(s)*

$E^o_{red}(Cu^{2+} | Cu) = 0.34V$ $E^o_{red}(Ag^{+} | Ag) = 0.80$ V

a. Cathode RXN:

b. Anode RXN:

c. Net RXN:

d. E^o_{cell} =

7. Write the balanced net reaction for aluminum solid reacting with lead (II) to yield lead solid and aluminum ion.

i. Give the line notation for the corresponding electrochemical cell.

ii. Given E^o_{red} (Pb^{2+}/ Pb)= −0.13 V and E^o_{red} (Al^{3+}/ Al) = −1.66 V, calculate the standard cell potential.

8. Answer the questions below using the list of standard reduction potentials.

	E°$_{red}$ (V)		***E°$_{red}$ (V)***
Pb^{4+}/Pb^{2+}	+1.67	In^{+}/In	−0.14
Ce^{3+}/Ce	+1.61	V^{3+}/V^{2+}	−0.26
Mn^{3+}/Mn^{2+}	+1.51	Cr^{3+}/Cr^{2+}	−0.41
O_3/O_2	+1.24	Fe^{2+}/Fe	−0.44
Br_2/Br^{-}	+1.09	Zn^{2+}/Zn	−0.76
Ag^{+}/Ag	+0.80	Al^{3+}/Al	−1.66

a. Which one is the best oxidizing agent?

b. Which one is the best reducing agent?

c. Circle the species that will reduce Br_2 but not V^{3+}:

In^{+} Al Ag Ce

d. Circle the species that will oxidize Cr^{2+} but not Mn^{2+}:

Pb^{4+} Zn^{2+} Fe^{2+} O_3

9. For each of the reactions below, circle all those that will have positive voltages. Assume all molecules are in their standard state concentrations unless otherwise given. The table below has standard reduction potentials and may be useful.

Half-Reaction	E°$_{red}$ (volts)
Ag^{+} / Ag	0.80
Fe^{3+} / Fe^{2+}	0.77
Fe^{2+} / Fe	−0.44

a) Fe^{3+} is able to oxidize Ag.

b) Ag^{+} is able to oxidize Fe^{2+}.

c) Fe^{2+} is able to oxidize Ag.

d) Write the line notation for the cell with most positive standard cell potential:

10. You have an electrochemical reaction that takes place in acidic solution of lead sulfate. Coming out of solution is a lead electrode and a platinum electrode.

$2H^{+}(aq) + 2e^{-} \leftrightharpoons H_2(g)$ E°$_{red}$ = 0 V

$PbSO_4(s) + 2e^{-} \leftrightharpoons Pb(s) + SO_4^{2-}(aq)$ E°$_{red}$ = −0.360 V

Pt(s) | H_2(g) | H_3O^{+}(aq) || $PbSO_4$(s) | SO_4^{2-}(aq) | Pb(s)

Anode RXN:

Cathode RXN:

NET RXN:

E^{o}_{cell} =

Things you should know and review:

- Oxidation Reduction Reactions
    i. Oxidation is <u>loss</u> of electrons (acts as a reducing agent)
    ii. Reduction is <u>gain</u> of electrons (acts as a oxidizing agent)
- Assigning Oxidation numbers
    iii. Oxidation number is **0** for atoms in an element.
    iv. The sum of all oxidation numbers in a molecule or ion must add up to the total charge.
    v. In compounds, alkalis (group 1) have oxidation number **+1**; alkaline earths (group 2) have oxidation number **+2**.
    vi. In compounds, fluorine (F) always has oxidation number **-1**. Other halogens (Cl ext.) have oxidation number **-1**, *except when bonded to elements that are more electronegative* such as fluorine or oxygen, where they can have positive oxidation numbers.
    vii. In compounds, hydrogen has oxidation number **+1** (*when bonded to elements with higher electronegativity*). When bonded to metals hydrogen has oxidation number **-1** *(when bonded to elements with lower elecronegativity)*.
    viii. In compounds, oxygen has oxidation number **-2**, except when bonded to fluorine, where fluorine's rule takes precedence (in OF_2 oxygen has oxidation number **+2**), and if there are O-O (peroxide) bonds. (in H_2O_2 or Na_2O_2 oxygen has oxidation number **-1).**
- **Steps for Balancing Redox reactions in Acidic or Basic media (always start with acidic).**
    Step 1. Write the two half reactions.
    Step 2. Balance all the elements except for O and H.
    Step 3. Balance O by adding H_2O.
    Step 4. Balance H by adding H^+.
    Step 5. Balance charge by adding electrons to the side of the half reaction that is most positive.
    Step 6. Balance electrons lost in one half reactions with electrons gained in the other half reaction.
    Step 7. Add half reactions together canceling what can be canceled.
    Step 8. Determine if your answer should be in acidic media or basic media
    If in acidic media: Add water to both sides in order to convert all H^+ to H_3O^+
        a) Combine H^+ and H_2O to form H_3O^+
        b) Simplify the reaction by canceling the H_2O that appear on both sides
    If in basic media: Add enough OH^- to neutralize H^+ to both sides of the reaction.
        a) Combine OH^- and H^+ to form H_2O.
        b) Simplify the reaction by canceling H_2O that appear on both sides.

- http://quantum.bu.edu/courses/ch102-spring-2016/notes/BalancingRedoxEquations.pdf

1. Balance the following reaction: $MnO_4^-(aq) + C_2O_4^{2-}(aq) \leftrightarrows MnO_2(aq) + CO_2(g)$.

   Oxidation half RXN:

   Reduction half RXN:

   Net RXN in Acidic Media:

   Net RXN in Basic Media:

2. Balance the following reaction: $IO^-_3(aq) + I^-(aq) \leftrightarrows I_2(aq)$.

   Oxidation half RXN:

   Reduction half RXN:

   Net RXN in Acidic Media:

   Net RXN in Basic Media:

3. Balance the redox reaction of H_2 (*g*) reducing Au^{3+} (*aq*) to produce solid gold in an acidic solution.

4. Balance the following reaction in basic media: **Fe^{2+} (*aq*) + MnO_4^- (*aq*) → Fe^{3+} (*aq*) + Mn^{2+} (*aq*)**

5. Use the steps provided to balance following oxidation reduction reactions in acidic and basic media.

   a. $Cu(s) + NO_3^-(aq) \leftrightarrows Cu^{2+}(aq) + NO_2(g)$

   b. $S^{2-}(aq) + I_2(aq) \leftrightarrows SO_4^{2-}(aq) + I^-(aq)$

6. Write down the balanced *oxidation half-reaction* for the complete combustion of glucose, $C_6H_{12}O_6$, in aqueous *acidic* solution.

Exam 2 Answers :

1. $K_a = 4;\ pH = -0.3$
2. 1.44M
3.
   a. $1.8 \cdot 10^{-4}$M
   b. $4.7 \cdot 10^{-5}$M - $5.0 \cdot 10^{-5}$M
   c. $1.67 \cdot 10^{-2}$M
   d. $1.2 \cdot 10^{-3}$M

4. $9.2 \cdot 10^{-9}$ - $9.6 \cdot 10^{-9}$
5.
   a. stay the same
   b. decrease
   c. decrease
   d. stay the same
   e. decrease
   f. stay the same
6. 3.0
7.
   a. $8.33 \cdot 10^{-4}$M
   b. $1.50 \cdot 10^{4}$M
8. 0.8420atm
9. −29.9 or −30 ºC
10. −15.8 ºC or −16ºC or −15.78 ºC