Time: 10min | **Chemistry 203** Quiz 5 | **July 22, 2008** TTh sections

Name: Key | 15   7 1104859

Section: ______________

pb

1. Assuming that the reaction $BrO_3^- + 5Br^- + 6H^+ \rightarrow 3Br_2 + 3H_2O$ is a first order reaction,

   a. If at a particular time, $-\Delta[BrO_3^-]/\Delta t = 1.5 \times 10^{-2}$ M/s, what is $-\Delta[Br^-]/\Delta t$ at the same instant?

   b. What is the unit of its rate constant?

3pts

a- $\text{Rate} = -\frac{\Delta[BrO_3^-]}{\Delta t} = -\frac{1}{5}\frac{\Delta[Br^-]}{\Delta t} = 1.5\times10^{-2}\ M/s$

$\Rightarrow -\frac{\Delta[Br^-]}{\Delta t} = 5\times(1.5\times10^{-2}) \Rightarrow \boxed{-\frac{\Delta[Br^-]}{\Delta t} = 7.5\times10^{-2}\ M/s}$

3pts

b- $\text{Rate} = K[\ ]$ (since the reaction is a 1st order reaction).

$(M/s) : ?\ (M) \Rightarrow \boxed{K \text{ is in } s^{-1}}$

9 pts

2. The data below were determined for the reaction

$S_2O_8^{2-} + 3I^-(aq) \rightarrow 2SO_4^{2-} + I_3^-$.

Expt. #	$[S_2O_8^{2-}]$	$[I^-]$	Initial Rate
1	0.038	0.060	1.4×10^{-5} M/s
2	0.076	0.060	2.8×10^{-5} M/s
3	0.076	0.030	1.4×10^{-5} M/s

Determine the rate law and the rate constant for this reaction.

$\text{Rate} = K[S_2O_8^{2-}]^m[I^-]^n$

$\text{Rate}_1 = K[0.038]^m[0.060]^n = 1.4\times10^{-5}\ M/s$

$\text{Rate}_2 = K[0.076]^m[0.060]^n = 2.8\times10^{-5}\ M/s$

$\Rightarrow \frac{\text{Rate}_2}{\text{Rate}_1} = \left(\frac{0.076}{0.038}\right)^m = \frac{2.8\times10^{-5}}{1.4\times10^{-5}}$

$\text{Rate}_3 = K[0.076]^m[0.030]^n = 1.4\times10^{-5}\ M/s$

2pts

$\rightarrow 2.0^m = 2.0 \Rightarrow \boxed{m = 1}$

2pts

$\frac{\text{Rate}_2}{\text{Rate}_3} = \left(\frac{0.060}{0.030}\right)^n = \frac{2.8\times10^{-5}}{1.4\times10^{-5}} \Rightarrow 2.0^n = 2.0 \Rightarrow \boxed{n = 1}$

2pts

$\boxed{\text{Rate} = K[S_2O_8^{2-}][I^-]}$ ← Rate Law

3pts

1 pt on unit

K = ?

Expt #1 $\Rightarrow K = \frac{1.4\times10^{-5}\ M/s}{(0.038)(0.060)\ M} = 6.1\times10^{-3}\ M^{-1}s^{-1}$

Expt. #2 $\Rightarrow K = \frac{2.8\times10^{-5}}{0.076\times0.060} = 6.1\times10^{-3}\ M^{-1}s^{-1}$

Expt. #3 $\Rightarrow K = 6.1\times10^{-3}\ M^{-1}s^{-1}$

$\boxed{K_{average} = 6.1\times10}$

Time: 10min

Chemistry 203
Quiz 4

July 14, 2008
MW sections

Name: Key/15

Section: ______________

3 pts

1. Define a concentration cell.

Concentration cell is a cell that has a potential based only on differences in concentration.

In a concentration cell we use the same kind of electrodes and same solution (species) in the two compartments but different concentrations.

2. Consider the electrochemical cell whose reaction is given below

$3\,Zn\,(s) + 2Cr^{3+}\,(aq) \rightarrow 3\,Zn^{2+}\,(aq) + 2Cr\,(s)$

$[Cr^{3+}] = 0.010M$ $[Zn^{2+}] = 0.0085M$ $E^{\circ}_{Zn2+/Zn} = -0.76V$ $E^{\circ}_{Cr3+/Cr} = -0.74\ V$
$R = 8.314\ JK^{-1}mol^{-1}$ $F = 96500\ JV^{-1}$ $T = 25°C$

4 pts

a. Write the two half reactions and specify at which electrode each one occurs

2 pts Oxidation: $Zn_{(s)} \rightleftharpoons Zn^{2+}_{(aq)} + 2e^-$ occurs at anode

2 pts Reduction: $Cr^{3+}_{(aq)} + 3e^- \rightleftharpoons Cr_{(s)}$ occurs at cathode

5 pts

b. Calculate E° and E of the cell

2 pts

$$E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$$
$$= E^{\circ}_{Cr^{3+}/Cr} - E^{\circ}_{Zn^{2+}/Zn} = -0.74 - (-0.76) \Rightarrow \boxed{E^{\circ}_{cell} = +0.02V}$$

3 pts

$$E_{cell} = E^{\circ} - \frac{RT}{nF}\ln\frac{[Zn^{2+}]^3}{[Cr^{3+}]^2} = 0.02 - \frac{(8.314)(298)}{(6)(96500)}\ln\frac{(0.0085)^3}{(0.010)^2}$$
$$\Rightarrow E_{cell} = 0.02 - (-0.022) \Rightarrow \boxed{E_{cell} = 0.042V \text{ or } 0.04V}$$

3 pts
No partial credits!

c. Give the cell notation (cell representation) of the above cell.

$$Zn_{(s)} \,|\, Zn^{2+}_{(aq)}(0.0085M) \,||\, Cr^{3+}_{(aq)}(0.010M) \,|\, Cr_{(s)}$$

or $$Zn_{(s)} \,|\, Zn^{2+}_{(aq)} \,||\, Cr^{3+}_{(aq)} \,|\, Cr_{(s)}$$

Time: 10min | **Chemistry 203** Quiz 4 | **July 15, 2008** TTh sections

Name: Key/15

Section: ______________

3 pts

1. Define a galvanic cell.

A galvanic cell is an electrochemical cell where the redox reaction is a <u>spontaneous</u> reaction (or $E_{cell} > 0$)

2. Consider an electrochemical cell constructed from the following half cells, linked by an external circuit and by a KCl salt bridge.

- a Pb(s) cathode in 1.0 M $Pb(NO_3)_2$ solution
- an Al(s) anode in 1.0 M $Al(NO_3)_3$ solution

4 pts

a. Give the cell representation (cell notation)

$Al_{(s)} \mid Al(NO_3)_{3(aq)}\ (1.0M) \parallel Pb(NO_3)_{2(aq)}\ (1.0M) \mid Pb_{(s)}$

or $Al_{(s)} \mid Al(NO_3)_{3(aq)} \parallel Pb(NO_3)_{2(aq)} \mid Pb_{(s)}$

4 pts

b. Write the oxidation half reaction and the reduction half reaction

Oxidation: $Al_{(s)} \rightleftharpoons Al^{3+}_{(aq)} + 3e^-$

Reduction: $Pb^{2+}_{(aq)} + 2e^- \rightleftharpoons Pb_{(s)}$

4 pts

c. Give the balanced overall cell reaction

$(Al_{(s)} \rightleftharpoons Al^{3+} + 3e^-) \times 2$

$(Pb^{2+} + 2e^- \rightleftharpoons Pb_{(s)}) \times 3$

$2\,Al_{(s)} + 3Pb^{2+}_{(aq)} \rightleftharpoons 2\,Al^{3+} + 3Pb_{(s)}$

Time: 10min | **Chemistry 203** Quiz 3 | **July 10, 2008** TTh sections

Name: Key / 15

Section: ____________

7 pts

1. A solution prepared by dissolving 25.8mg of benzene (MW 78.11) in hexane and diluting to 250.0mL had an absorption peak at 256nm. The molar absorptiviy of benzene at 256nm is $201 M^{-1}cm^{-1}$. Determine:

a. the absorbance of the benzene solution in a spectrophotometer with a 1.00-cm cell

$\mathcal{E} = 201\ M^{-1}cm^{-1}$

$m = 25.8mg \quad M_m = 78.11 g/mol \quad V = 250.0 ml$

3 pts $\quad A = \mathcal{E} b C \qquad C = \frac{n}{V} = \frac{m}{M_m V} = \frac{25.8\times10^{-3} g}{78.11 g/mol \times 250.0\times10^{-3}}$

$C = 1.32\times10^{-3} M$

2 pts $\quad A = 201 \times 1.00 \times 1.32\times10^{-3} = 0.265 \Rightarrow \boxed{A = 0.265}$

b. the percent transmittance of the benzene solution.

$\%T = ?$

2 pts $\quad A = -\log T \Rightarrow T = 10^{-A} = 10^{-0.265} = 0.543$

$\boxed{\%T = 54.3\%}$

8 pts

2. 5.00mL of 2.00×10^{-3}M of a metal (M) solution are mixed with 2.00mL of 2.00×10^{-3}M of a ligand (L) solution, and 3.00mL of distilled water are added. The concentration of the formed complex (ML) is determined spectrophotometrically and found to be 5.07×10^{-5}M. Determine the equilibrium constant Kc of the above reaction.

1 pt $\quad$ 5.00 ml (M) $2.00\times10^{-3} M \Rightarrow n_{0M} = 2.00\times10^{-3} \times 5.00\times10^{-3} = 10.0\times10^{-6}$ mol

1 pt $\quad$ 2.00 ml (L) $2.00\times10^{-3} M \Rightarrow n_{0L} = 2.00\times10^{-3} \times 2.00\times10^{-3} = 4.00\times10^{-6}$ mol

3.00 ml H_2O

$[ML] = 5.07\times10^{-5} M$

$M + L \rightleftharpoons ML \qquad K_c = \frac{[ML]}{[M][L]}$

1 pt $\quad [M]_0 = \frac{n_0}{V_t} = \frac{10.0\times10^{-6}}{10.00\times10^{-3}} = 1.00\times10^{-3} M$

1 pt $\quad [L]_0 = \frac{4.00\times10^{-6}}{10.00\times10^{-3}} = 0.400\times10^{-3} M$

1 pt $\quad [M]_{eq} = [M]_0 - [ML] = 1.00\times10^{-3} - 5.07\times10^{-5} = 0.95..\times10^{-3} M$

1 pt $\quad [L]_{eq} = [L]_0 - [ML] = 0.400\times10^{-3} - 5.07\times10^{-5} = 0.349\times10^{-3} M$

2 pt $\quad K_c = \frac{5.07\times10^{-5}}{(0.95\times10^{-3})(0.349\times10^{-3})} = 152.9 \Rightarrow \boxed{K_c = 1.}$

Time: 10min **Chemistry 203** **July 8, 2008**

Quiz 2 TTh sections

Name: Key/15

Section: ______________

5pts

1. The solubility product for chromium(III) fluoride CrF_3 is $K_{sp} = 6.6 \times 10^{-11}$. How many grams of chromium(III) fluoride can be dissolved in one liter of water?
Cr=52.00 F=19.00

1pt $CrF_{3(s)} \rightleftharpoons Cr^{3+} + 3F^-$ $\quad K_{sp} = [Cr^{3+}][F^-]^3$

$\qquad\qquad\qquad\qquad s \qquad 3s$

$\Rightarrow K_{sp} = (s)(3s)^3 = 6.6 \times 10^{-11} \Rightarrow 27s^4 = 6.6 \times 10^{-11} \Rightarrow s = \sqrt[4]{\frac{6.6 \times 10^{-11}}{27}}$

2pts $\Rightarrow s = 1.3 \times 10^{-3} M$

2pts $s(g/l) = s(mol/l) \times M_m(g/mol) = 1.3 \times 10^{-3} mol/l \times 109.00\, g/mol$

$\Rightarrow \boxed{s = 1.4 \times 10^{-1}\, g/l}$ i.e. we can dissolve 0.14g of CrF_3 in 1l of water.

5pts

2. A solution contains 1.0×10^{-4} M Cu^- and 2.0×10^{-3} M Pb^{2+}. If a source of I^- is added gradually to this solution, will PbI_2 ($K_{sp}=1.4 \times 10^{-8}$) or CuI ($K_{sp}=5.3 \times 10^{-12}$) precipitate first? Justify your answer.

$[Cu^+] = 1.0 \times 10^{-4} M$ $\qquad CuI \rightleftharpoons Cu^+ + I^-$

$[Pb^{2+}] = 2.0 \times 10^{-3} M$ $\qquad PbI_2 \rightleftharpoons Pb^{2+} + 2I^-$

2pts $K_{sp}(CuI) = 5.3 \times 10^{-12} = [Cu^+][I^-] \Rightarrow [I^-] = \frac{5.3 \times 10^{-12}}{1.0 \times 10^{-4}} = 5.3 \times 10^{-8} M$

$\hookrightarrow$ $[I^-]$ needed to precipitate CuI is $5.3 \times 10^{-8} M$

2pts $K_{sp}(PbI_2) = 1.4 \times 10^{-8} = [Pb^{2+}][I^-]^2 \Rightarrow [I^-] = \sqrt{\frac{1.4 \times 10^{-8}}{2.0 \times 10^{-3}}} = 2.6 \times 10^{-3} M$

$\hookrightarrow$ $[I^-]$ needed to precipitate PbI_2 is $2.6 \times 10^{-3} M$

1pt since $5.3 \times 10^{-8} < 2.6 \times 10^{-3} \Rightarrow$ upon addition of I^- $\boxed{CuI_{(s)} \text{ will precipitate first}}$

5pts

3. Describe, using a flow chart, the separation procedure of Lead ion and Mercurous ion when present in a mixture.

2pts Pb^{2+}, Hg_2^{2+}

1pt $\downarrow$ HCl

$PbCl_2, Hg_2Cl_2$

1pt $\downarrow$ Hot water

Pb^{2+} $\qquad\qquad$ Hg_2Cl_2

1pt

Time: 10min

Chemistry 203
Quiz 1

July 3, 2008
TTh sections

Name: Key | 15

Section: ____________

6 pts

1. Answer the following short questions concerning today's experiment:

1 pt
a. What is the ligand?
Ethylene diamine tetraacetic acid (EDTA)

1 pt
b. What is the indicator?
Eriochrome Black T (Erio T)

1 pt
c. What is the buffer?
NH_3 / NH_4Cl

1 pt
d. What is the pH?
pH = 10.0

2 pts
e. What are the predominant forms of the indicator and the ligand at this pH?
HE^{2-} ; Y^{4-} (or HY^{3-})

4 pts

2. Write the equation of the complexation reaction taking place in this experiment and give the expression of the formation constant.

2 pts
$$Zn^{2+} + Y^{4-} \rightleftharpoons ZnY^{2-}$$

2 pts
$$K_f = \frac{[ZnY^{2-}]}{[Zn^{2+}][Y^{4-}]}$$

5 pts

3. The Zn (Mm= 65.38g/mol) in a 0.7556-g sample of foot powder was titrated with 21.27 mL of 0.01645M EDTA. Calculate the percentage of Zn in this sample.

$m_{sample} = 0.7556\ g$

21.27 ml EDTA 0.01645 M

3 pts
$$n_{EDTA} = n_{Zn} \Rightarrow (M \times V)_{EDTA} = \frac{m_{Zn}}{M_{m\,Zn}}$$

$$\Rightarrow \underset{(mol/l)}{0.01645} \times \underset{(l)}{21.27 \times 10^{-3}} = \frac{m_{Zn}}{65.38\,(g/mol)} \Rightarrow m_{Zn} = 0.02288$$

2 pts
$$\% Zn = \frac{m_{Zn}}{m_{sample}} \times 100 = \frac{0.02288}{0.7556} \times 100 = \boxed{3.028\%\ Zn}$$

Total out of 12 points

Time: 11 min

Chem. 203
Drop Quiz 6

August 5, 2003
TT sections

NAME: KEY

Complete the following information:

+1 each

1. The Arrhenius equation gives the dependence of the rate constant of a chemical reaction on temperature.
2. Scale formed in kettles is mainly composed of $CaCO_3$ and $MgCO_3$, which can be tested for with HCl.
3. In a household-type resin, Na_2R becomes mainly CaR after ion exchange takes place.
4. In a de-ionizing type column, cations in the impure water are replaced by H^+, and anions are replaced by OH^-.
5. In a zero-order reaction with rate constant 0.0012 M s^{-1}, a solution of initial concentration 0.500 M in the reactant becomes 0.36 M after 2.0 minutes are elapsed; and the reaction terminates after 417 seconds are elapsed. [OK if 0.356]

see below

6. The minimum energy that a system of colliding molecules must possess to lead to products is called the activation energy.
7. The efficiency of an ion-exchange resin depends on pH and temperature

For zero-order: $[A] = -kt + [A]_0$

$$= -0.0012 \times 120 + 0.500$$

$$= -0.144 + 0.500 = 0.356\,M \longrightarrow \underline{\underline{0.36\,M}}$$

Termination when $[A]_t = 0 \Rightarrow -kt + [A]_0 = 0$

$$\hookrightarrow t = \frac{[A]_0}{k} = \frac{0.500}{0.0012} = \underline{\underline{417\,s}}$$

TOTAL out of 15 points

Time: 12 min

Chem. 203
Drop Quiz 4

July 14, 2004
TT sections

NAME : KEY

Two stock solutions provided to you are as follows:

Solution **A**: 2.00×10^{-3} M Fe^{3+}
Solution **B**: 2.00×10^{-3} M SCN^-

You placed 3.00 mL of solution **A** and 7.00 mL of solution **B** into a clean test tube, then 10.00 mL of distilled water are added to the mixture. After equilibrium was established (the solution has a red brown color), you measured the Absorbance of the solution in a spectrophotometer. After conversion, you found that the concentration of the complex $FeSCN^{2+}$ was 4.93×10^{-5} M.
Calculate the equilibrium constant for the formation of the $FeSCN^{2+}$ complex. Show your reasoning.

+3

$$n_{0\,Fe^{3+}} = 3.00 \times 2.00 \times 10^{-3} = 6.00 \times 10^{-3} \text{ mmol}$$

$$n_{0\,SCN^-} = 7.00 \times 2.00 \times 10^{-3} = 0.0140 \text{ mmol}$$

+1 At equilibrium: $n_{FeSCN^{2+}} \equiv x = 4.93 \times 10^{-5} \times 20.00 = 9.86 \times 10^{-4}$ mmol

+6

	Fe^{3+}	+ SCN^-	$\rightleftharpoons$ $FeSCN^{2+}$
Initially	0.00600	0.0140	–
Equilibrium	$5.01 \times 10^{-3} \equiv 0.00600 - x$	$0.0130 \equiv 0.0140 - x$	0.000986

+3 At equilibrium:

$$[Fe^{3+}] = \frac{5.01 \times 10^{-3}}{20.00} = 2.51 \times 10^{-4}\ M$$

$$[SCN^-] = \frac{0.0130}{20.00} = 6.50 \times 10^{-4}\ M$$

Hence

$$K = \frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]} = \frac{4.93 \times 10^{-5}}{(2.51 \times 10^{-4})(6.50 \times 10^{-4})} = \boxed{302}$$

+2

Total out of 15 points

Time: 16 min | **Chem. 203 Drop Quiz 7** | **July 28, 2005** TT sections

NAME : KEY

1. Write the rate law for a chemical reaction involving a single reactant A ($A \longrightarrow P$), given that the reaction is zero order in A. Call the rate constant k and the initial concentration of the reactant $[A]_0$.

+2
$$-\frac{d[A]}{dt} = k[A]^0 \equiv k$$

06 2. Integrate the rate law, and derive an expression of the concentration of A at time t $[A]_t$.

+2
$$-\frac{d[A]}{dt} = k \Rightarrow d[A] = -k\,dt$$

+3
$$\int_{[A]_0}^{[A]_t} d[A] = -k\int_0^t dt \Rightarrow [A]_t - [A]_0 = -kt$$

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

3. From the integrated equation, derive an expression for the half-life of the reaction (call it $t_{1/2}$, defined as the time at which the reactant A reaches half its initial concentration, $[A]_0/2$).

+3
At $t = t_{1/2}$ $\quad [A] = \frac{[A]_0}{2} \Rightarrow \frac{[A]_0}{2} = -k\,t_{1/2} + [A]_0$

$$\frac{[A]_0}{2} = kt_{1/2} \Rightarrow \boxed{t_{1/2} = \frac{[A]_0}{2k}}$$

4. Derive an expression for the time at which the reaction terminates (the concentration of reactant A becomes zero).

+3
At termination time t_f $\quad [A] = 0$

$$-kt_f + [A]_0 = 0 \Rightarrow \boxed{t_f = \frac{[A]_0}{k}}$$

5. Do chemical reactions with reactants following other rate laws than zero-order, reach termination?

+1
No. [A] would go asymptotically to zero.

Total out of 15 points

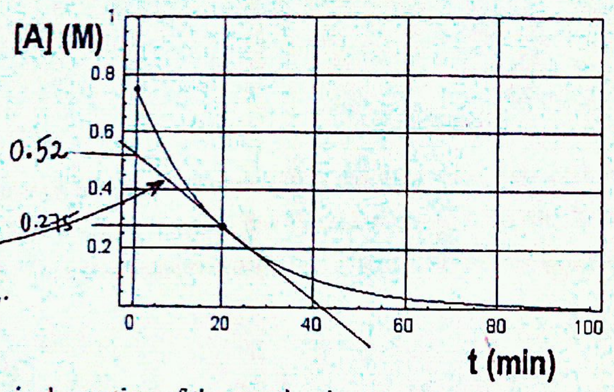
Time: 13 min | **Chem. 203** **Drop Quiz 7** | **July 27, 2005** MW sections

NAME : KEY

(11) 1. The concentration of a reactant A was monitored with time during the course of a chemical reaction.

grade the method! not the exact result.

+4 Draw tangent at $t = 20$ min.



$$\text{Slope} = \frac{y_2 - y_1}{x_2 - x_1} = \frac{0.275 - 0.52\ (M)}{20 - 0\ (min)}$$

+4 $$= \frac{-0.24}{20} = \boxed{-0.012\ M\,min^{-1}}$$

The chemical equation of the reaction is:

$$3\,A + 2\,B \longrightarrow C + 3\,D$$

Calculate the rate of the chemical reaction at $t = 20$ minutes.

+3 $$\text{Rate} = -\frac{1}{3}\frac{d[A]}{dt} = -\frac{1}{3} \times -0.012 = 0.0040\ M\,min^{-1}$$
$$= \boxed{4.0 \times 10^{-3}\ M\,min^{-1}}$$

2. Write the rate law of the reaction; given that it is first order in A, and second order in B.

+2 $$\text{Rate} = k[A][B]^2;\ \text{where } k \equiv \text{rate constant}$$

3. Express the rate of the reaction as a function of the rate of disappearance/formation of each species participating in the reaction.

+2 (+0.5 each) $$\text{Rate} = -\frac{1}{3}\frac{d[A]}{dt} = -\frac{1}{2}\frac{d[B]}{dt} = \frac{d[C]}{dt} = \frac{1}{3}\frac{d[D]}{dt}$$