Elite
Media

CHS 103 and 109
Exam past questions

RIVERS STATE UNIVERSITY

FIRST SEMESTER EXAM 2016-2019 SESSIONS (Compiled)

CHS 103 (Likely)

INSTRUCTION: ANSWER ALL QUESTIONS IN BOTH SECTIONS

SECTION A

1. Consider a molecular reaction involving two reactants A and B. Write down the reaction and show the product formed ____________

2. In the course of the reaction in Number one above reactants A & B are consumed hence their concentrations ______

3. Write down the rate of the reaction in question number one above ________

4. ____ Is the number of chemical species whose concentrations affect the rate of the reaction

5. Which of the following represents second order reaction (a) $2A \rightarrow$ Products (b) $A + B \rightarrow$ Products (c) $A \rightarrow$ Products (d) A and B (e) $2A + B \rightarrow$ Products

6. Which of the following reactions is an example of first order reaction (a) $C_2H_4 + H_2 \rightarrow C_2H_6$ (b) $2SO_2 + O_2 \xrightarrow{NO} 2SO_3$

(c) $CH_3CH_2Br \rightarrow C_2H_4 + HBr$

7. What happens when Nitrogen dioxide dissolves in water.

8. The equation for first order kinetics is given as

9. The equation for half – life of first order reactions is ____________.

10. What is the major concern about thermodynamics ____________________

11. What is the entropy of the reaction below (write down the standard entropy change of the reaction) ____________ $aA + bB \rightarrow cC + dD$

12. At constant temperature, the change in free energy (ΔG) of a system is given by ____________

13. Arrange the following ions in order of increasing strength as oxidizing agents: Ag^+ (aq), $Cr_2O_7^{2-}$ (aq), NO_3^- (aq)

(ii) (1) concentration
(2) Temperature
(3) pressure or volume

(iii) write the equation of the reaction and balance it first when even when the equation is not given to you

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

calculate their masses (i.e of both the reacting species and its products

Fe_2O_3 is $56\times2 + 16\times3$
$112 + 48 = 160g$ } Reacting species
$3CO$ is $(12+16)\times3 = 84g$

$2Fe$ is $56\times2 = 112$ } product
$3CO_2$ is $(12+16\times2)\times3$
$44\times3 = 132g$

From the equation

160g of Fe_2O_3 reacted with 84g of CO to give 112g of Fe and 132g of CO_2

$\therefore$ 84g CO = 160g Fe_2O_3
1g CO = x

$x \times 84g = 1\times160g$

$x = \frac{1\times160}{84}$ of Fe_2O_3

$\therefore$ 146g of CO will give

$x = \frac{1\times160}{84}\times146$ of Fe_2O_3

$\frac{23360}{84}g = \boxed{278.09g\ Fe_2O_3}$

(3) (a) According to Henry's Law

$C = K_HP \Rightarrow K_H = \frac{C}{P}$

where C = solubility of oxygen in water 1.38×10^{-3}

P = atmospheric pressure 1.00 atm

Thus $K_H = \frac{1.38\times10^{-3}\ mol\ L^{-1}}{1.00\ atm}$

$= 1.38\times10^{-3}\ mol\ L^{-1}\ atm^{-1}$

concentration of oxygen at 0.21 atm =

$1.38\times10^{-3}\ mol\ L^{-1}\ atm^{-1} \times 0.21\ atm = 2.9\times10^{-4}\ mol\ L^{-1}$

(b) using $\Delta S_v = \frac{\Delta H_v}{T_v} = \frac{9717\ cal\ mol^{-1}}{373.2k}$

$= 26.037\ cal\ mol^{-1}\ K^{-1}$

$\Rightarrow 26.037$ e.u mol^{-1} =

$26.037\times4.184\ J\ K^{-1}\ mol^{-1}$

$= 108.939\ J\ K^{-1}\ mol^{-1}$

(c) Using $\ln\frac{P_1}{P_2} = \frac{\Delta H_{vap}}{R}\left(\frac{T_1 - T_2}{T_1T_2}\right)$

where $P_1 = 401$ mmHg ; $T_1 = 18°C$ $(18+273) = 291k$

$P_2 = ?$; $T_2 = 32°C$ $(32+273) = 305k$

$\ln\frac{401}{P_2} = \frac{26{,}000\ J\ mol^{-1}}{8.314\ J\ K^{-1}\ mol^{-1}}\left(\frac{291 - 305k}{291k \times 305k}\right)$

$\ln\frac{401}{P_2} = 0.493$ $\therefore$ ~~$\frac{401}{P_2} = e^{-0.494}$~~

$\frac{401}{P_2} = e^{-0.493} = 0.611$

$P_2 = 656.3$ mmHg

(4)

NOTE: the one with the larger $E°$ is the reduction half reaction, to make 2nd reaction

(8) (a) it is because they largely or completely ionize in water

(b) $H_2O < H_2S < H_2Se < H_2Te$ and $HF < HCl < HBr < HI$

(c) (i) $K = \frac{[H_3O^+][A^-]}{[H_2O][HA]}$

$Ka = K[H_2O] = \frac{[H_3O^+][A^-]}{[HA]}$

(d) From the definition of PH and POH

$PH + POH = 14.0$

The POH of a solution is negative logarithm of the POH concentration i.e $POH = -\log[OH^-]$

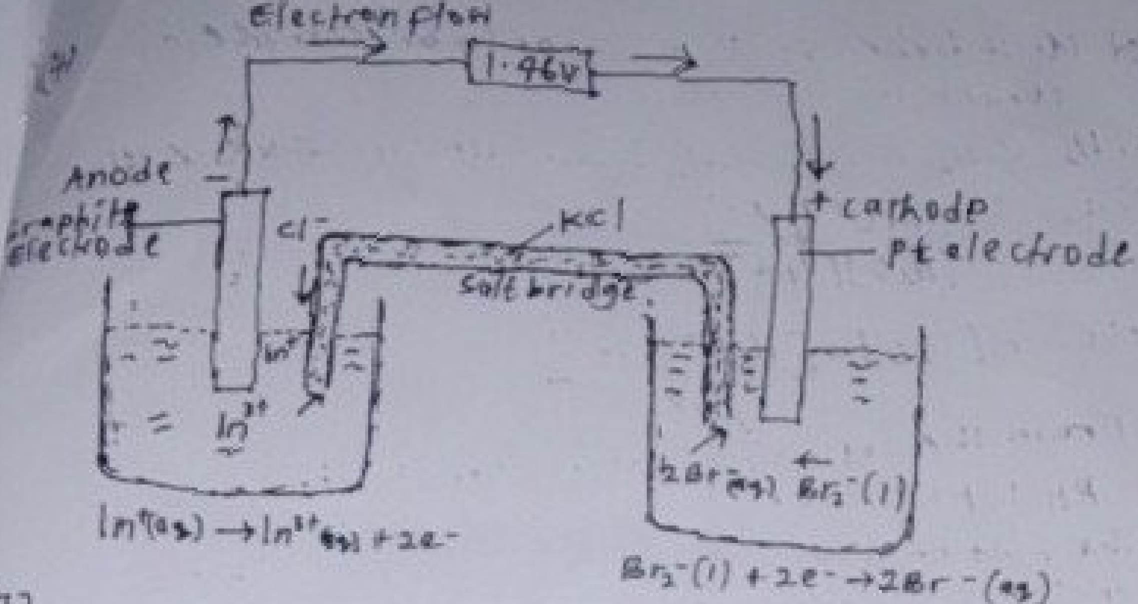
But $[H^+][OH^-] = K_w = 1.0 \times 10^{-14}$

Taking negative log of both sides, we have

$-(\log[H^+] + \log[OH^-]) = -\log(1 \times 10^{-14})$

$-\log[H^+] - \log[OH^-] = 14.00$

$\therefore PH + POH = 14.00$



(7)

(7)

(a)

salt bridge

Anodic half-cell | cathodic half-cell

$Zn(s) \,|\, Zn^{2+}(aq) \;||\; Cu^{2+}(aq) \,|\, Cu(s)$

Phase boundary | Direction of electron flow | Phase boundary

(b) * Oxidation half reaction

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

* Reduction half reaction

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

* overall cell Reaction

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

(c) It used to join the two solutions

(d) Anode is the negative electrode where oxidation occurs or where electrons are released while cathode is the positive electrode where reduction occurs

NB: Electrons released at the anode must be accepted at the cathode meaning that electron flow is from the anode to the cathode but in electrolytic cell (battery) electrons are generated at the cathode (negative) and must be accepted at the anode (positive)

(6)

(a) First of all seperate the reactions into oxidation half reaction and reduction half reaction before combining them to give a redox reaction

* Oxidation half Reaction

$In^{+}(aq) \rightarrow In^{3+}(aq) + 2e^{-}$ $E^{o}_{oxi} = ?$

* Reduction half reaction

$Br_2^{-}(l) + 2e^{-} \rightarrow 2Br^{-}(aq)$ $E^{o}_{red} = +1.06V$

combine the two equations

$In^{+}(aq) \rightarrow In^{3+}(aq) + 2e^{-}$ $E_{ox} = ?$

$Br_2^{-}(l) + 2e^{-} \rightarrow 2Br^{-}(aq)$ $E^{o} = +1.06V$

Net Redox Reaction $In^{+}(aq) + Br_2^{-}(l) \rightarrow In^{3+} + 2Br^{-}(aq)$

$E^{o}_{cell} = E^{o}_{ox} + E^{o}_{red}$

$E^{o}_{cell} = +1.06V$

$E^{o}_{ox} = E^{o}_{cell} - E^{o}_{red}$

$E^{o}_{ox} = [+1.46 - (+1.06)]V$

$E^{o}_{ox} = (+1.46 - 1.06)V$

$E^{o}_{ox} = +0.40V$

(b) The E^{o}_{red} for the reduction of In^{3+} to In^{+} is the reverse of oxidation process

$In^{+}(aq) \rightarrow In^{3+}(aq) + 2e^{-}$ $E^{o}_{ox} = +0.40$

* The Reverse process is

$In^{3+}(aq) + 2e^{-} \rightarrow In^{+}(aq)$ $E^{o}_{red} = -0.40V$

(c) In a cell diagram the oxidation (anodic) reaction are written on the left while reduction (cathodic) reaction on the right & double vertical lines (||) representing the salt bridge and single vertical line seperate phase boundary

$C(s) | In^{+}(aq), In^{3+}(aq) || Br_2^{-}(l) | Br^{-}(aq) | Pt(s)$

In oxid. graphite (carbon) is used as electrode

Platinum is used as electrode on the reduction half.

(iii) Freezing point
(iv) osmotic pressure

(22) (a) Lemon (citric acid)
(b) Human blood (amino acid)
(c) Digestive juice (HCl)
(d) Car battery electrolyte (H_2SO_4)
(e) vinegar (acetic acid)

(23) c

(24) (a) Lewis theory of acids and base
(b) Bronsted Lowry theory of acid and base

(25) D (26) D (27) spontaneous in the forward direction

(28) No net change so the system is at equilibrium

(29) second Law of thermodynamics

(30) Hydrogen carbonate ion (HCO_3^-)

(31) The concentration reaction quotient Q_c will be given by $Q_c = \frac{[HI]^2}{[I_2][H_2]} = \frac{[1.60]}{(0.15)(0.13)} = 131$

NOTE: If $Q_c < K_c$ the reaction goes from left to right i.e some of the I_2 and H_2 will combine to form more HI to attain a new equilibrium

(32) The equation for the dissociation of Ag_2CrO_4 at equilibrium with its ions is

$$Ag_2CrO_4(s) \rightleftharpoons 2Ag^-(aq) + CrO_4^{2-}(aq)$$

$1.3 \times 10^{-4}M$ $2 \times 1.3 \times 10^{-4}M$ 1.3×10^{-4}

$K_{sp} = [Ag]^2[CrO_4^{2-}]$

$= (2.6 \times 10^{-4}M)^2(1.3 \times 10^{-4}M) = 8.8 \times 10^{-12}$

(33) (a) $K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$

(b) (i) $aA(aq) + bB(aq) \underset{K_2}{\overset{K_1}{\rightleftharpoons}} cC(aq) + dD$

Thus the rate of reaction R will be given by

$R_1 = K_1[A]^a[B]^b$ for forward reaction

$R_2 = K_2[C]^c[D]^d$ for backward reaction

(ii) $R_1 = R_2 \therefore K_1[A]^a[B]^b = K_2[C]^c[D]^d$

Thus $\frac{K_1}{K_2} = \frac{[C]^c[D]^d}{[A]^a[B]^b} = K_c$

6. The standard cell potential is 1.46 v for a galvanic cell based on the following half-cell reactions.

$In^{+(aq)} \rightarrow In^{3+(aq)} + 2e^{-}$ E^{0}ox = ?

$Br_2^{-}(1) + 2e^{-} \rightarrow 2Br^{-(aq)}$ $E^{0} = +1.06V$

(a)Calculate the E^{0}ox of the reaction

(b) Calculate E^{0}red for the reduction of In^{3+} to In^{+} (c) Draw a cell diagram for overall re

(d)Draw and label a voltaic cell IF Kcl gel was used as electrolyte the in the salt bridge and indicate the direction of electron flow as well as the direction of ions from the salt bridge

7 Using the cell diagram below answer the questions that follow:

$Zn_{(s)} \mid Zn^{2+(aq)} \| Cu^{2+(aq)} \mid Cu_{(s)}$

(a)Label the above cell Diagram

(b) Write down the oxidation half reaction, reduction half reaction and overall cell reaction in the spontaneous reaction in electro chemical cell based on the above cell diagram

(c)What is the function of salt bridge you have labeled in your cell diagram

(d) In a galvanic or Voltaic cell, what can you say about the anode and the cathode

8. (a) Why are all strong acids and bases strong electrolyte

(b) Arrange the following acids in increasing order of strength H_2Se, H_2S, H_2Te H_2O and HBr HF, HI, HCl

(C) Consider the ionization equilibrium of a weak acid H_A in water represented as H_A + H20 $\rightleftharpoons H_3O^{+} + A^{-}$ (i) The equilibrium constant is given by-------------- (ii) The acid dissociation constant or special

(d) Prove that PH +POH = 14.0

$\Delta G < O$ it means reaction is ________.

$\Delta G = O$ it means ________.

Which law states that "heatt cannot of itself pass from a colder to a warmer body unless work is provided by an external body"

30. If NH_3 is the conjugate base of the acid $NH4^+$, the conjugate base of carbonic acid (H_2CO_3) is ________.

31. A 2-liter flask contained a mixture of 0.12M hydrogen, 0.15M iodine and 1.60M hydrogen iodide at a temperature of 500K $I_2(g) + H_2(g) \rightleftharpoons 2HI(g)$ if the concentration equilibrium constant Kc is 160 determine, (a) the direction the equilibrium will shift

32. Analysis of a saturated solution of silver chromate at a certain temperature shows that the solution contains 1.3×10^{-4}M of dissolved Ag_2Cr O4. Calculate the solubility product of Ag_2Cr O4 at that temperature

33. (a) Consider the reversible reaction below and expressed the relationship between the product & reactant concentration at equilibrium aA (aq) + bB (aq) $\rightleftharpoons$ cC (aq) + dD (aq) (b) (i) from the above equation (33a) on the application of law of mass action to the rate of forward and backward reactions, the equation will appear as ________ while the rate of reaction R will be given by ________ (ii) where K_1 and K_2 are specific rate constants i.e. 33 (bi) for the forward & reverse reactions respectively. At equilibrium, the rates of forward and reverse reactions are equal. Write down the equation

SECTION B

1. The two half -reactions occur in a voltaic cell are $Zn^{(s)} \rightarrow Z_n^{2+}(aq) + 2e^-$ (Eleltrode $= z_n$) ClO_3^- (aq) + $6H^+$(aq) + $6e^- \rightarrow$ Cl (aq) + $3H2^0$ (Electrode = pt)
   (i) Indicate the reaction that occurs in the anode and the one that occurs at the cathode
   (ii) Does the zinc electrode gain, lose or retain its mass as the reaction proceeds?
   (iii) Does the platinum electrode gain, lose or retain its mass as the reaction proceeds

CHS 103 (Likely)

SOLUTIONS (2016-2019)

SECTION A

(1) $A + B \longrightarrow C + D$

(2) Decrease

(3) rate =
$$-\frac{d[A]}{dt} = -\frac{d[B]}{dt} = +\frac{d[C]}{dt} + \frac{d[D]}{dt}$$

(4) The order of a reaction

(5) D (6) C

(7) $2NO_2 + H_2O \rightarrow HNO_3 + HNO_2$ (Nitric acid and Nitrous acid are formed)

(8) $\ln[A]_t = \ln[A]_0 - kt$ or
$[A]_t = [A]_0 e^{-kt}$

(9) $t_{\frac{1}{2}} = \frac{0.692}{k}$

(10) To predict whether a reaction is feasible or not under specific conditions

(11) The entropy of reaction is the standard entropy change of the reaction under standard conditions of pressure and temperature, Thus

$$\Delta S^\circ_{rxn} = [cS^\circ(C) + dS^\circ(D)] - [aS^\circ(A) + bS^\circ(B)]$$

(12) $\Delta G = \Delta H - T\Delta S$

(13) $Ag^+(aq) < NO_3^-(aq) < Cr_2O_7^{2-}(aq)$

(14) a homogenous mixture of substances in which no settling occurs

(15) Because they contain portions (solutes) that are clearly distinguishable from other portions

(16) B (Blood cells are solutes whole blood is a solution)

(17) $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$ Rusting of iron (oxid. of Fe)

$4Fe_2O_3(s) + 3C(s) \rightarrow 4Fe(s) + 3CO_2(g)$ making of iron (reduction of Fe_2O_3)

(18) A

(19) Gas being compressed means work is done i.e W is positive. Heat was release to the environment (exothermic reaction) hence q is negative (amount of heat exchange)

using $\Delta E = q + W$

$\Delta E = -128 + 462 J = 334 J$

(20) (a) $H^\ominus_{mCO_2} =$
$-(\Delta H^\circ_C + \Delta H^\circ_{O_2} = -393.5 kJ mol^{-1})$
$C + O_2 \rightarrow CO_2, \Delta H^\circ_m = -393.5 kJ mol^{-1}$

(b) The mechanical equation (i.e a type of equation that relates a balanced equation of a reaction to the value of enthalpy of the reaction)

(21) (i) vapour pressure
(ii) boiling point

2. Which electrode is positive? (a) Define and illustrate the following terms

(i) Reversible reaction (ii) Chemical equilibrium

(ii) Equilibrium constant

(b) State Le chatelier's principle

(ii) Mention the three main factors that affect equilibrium

(iii) What mass of Fe_2O_3 in grams is required to react with 146g of co? The equilibrium reaction is $Fe_2O_3 + 3CO \rightarrow 2Fe + CO_2$

3. (a) The solubility of oxygen in water at 20°c and 1.00 atmospheric pressure is 1.38×10^{-3} molL^{-1}

Calculate the concentration of oxygen at 20^0C and partial pressure of 0.21 atmosphere.

(b) The molecular heat of vaporization of water is 9717 $cal\ m0l^{-1}$ at 373.2k. Determine the entropy change in Jk^{-1} $mo1^{-1}$ if 1 calk^{-1} $mo1^{-1}$ = 1 electron unit per m01 ($e.u\ m01^{-1}$) and $1\ e.u\ m0l^{-1}$ is = $4.184JK^{-1}m0l^{-1}$

(c) If the specific heat of vaporization of diethyl ether is 26,000J$mo1^{-1}$ and the vapor pressure is 401 mmH_g at 18°c Calculate the vapor pressure of the liquid at 32°c

4. Calculate the equilibrium constant for the reaction between the two half cells below which occurs at 25^0C in an alkaline solution.

$HO_2^-(aq) + H_2O(1) + 2e^- \rightarrow 3OH_{(aq)}\ E^0 = +0.878V$

$NiO_2(s) + 2H_2O(1) + 2e^- \rightarrow 2OH_{(aq)} + Ni(OH)2(S)\ E^0 = +0.490V$

5. (a) Based on the standard reduction potential decide the species expected to be the strongest oxidizing agent

(A) $C1^{-(aq)}$ (B) O_2(g) (c) $C1_2$ (D) $H^{+(aq)}$

$2Cl^{-(aq)} + 2e^- \rightarrow C1_2(g)\ E^o ox = -1.35v$

$O_{2(g)} + 4H^+ + 4e^- \rightarrow E^0 red = +1.23v$

$C1_{2(g)} + 2e^- \rightarrow 2c1_{(aq)}\ E^0 red = +1.35v$

$2H^{+(aq)} + 2e \rightarrow H_2(g)\ E^0 red = 0.000v$

(b). Calculate the entropy change for the reaction at 25^0C IF S^0 in $JK^{-1}m01^{-1}$ of $H_{2(g)}$ = 131.0 $CUO_{(s)}$ =43.5, $Cu_{(s)}$ = 33.3 $H2O_{(g)}$ = 188.7

$H_{2(g)} + CuO(s) \rightarrow Cu(s) + H_2O(g)$

SECTION B
CHS 103 SOLUTIONS

(1a)(i) The anodic half reaction is where oxidation occurs $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$ while reduction occurs at the cathode
$ClO_3^-{}_{(aq)} + 6H^+{}_{(aq)} + 6e^- \rightarrow$
(ii) $Cl^-{}_{(aq)} + 3H_2O_{(l)}$
(iii) The platinum electrode retain its mass as the reaction proceeds (it is not involved in the reaction) hence maintains its mass)
(ii) The zinc electrodes lose some mass as the reaction proceeds.
(iv) The Pt is positive electrode (cathode) while the zinc is the negative electrode (Anode)

(2)

(a)(i) Reversible Reaction
This is a reaction that does not go to completion and occurs in either direction. That is both forward and backward reaction occurs simultaneously.
A typical example is the formation of ammonium hydroxide when ammonia gas dissolved in water
$NH_{3(g)} + H_2O_{(l)} \rightleftharpoons NH_4OH_{(l)}$

(ii) Chemical Equilibrium is a condition that exist when two opposing reactions occur simultaneously at the same rate. It can be generally represented as
$aA + bB \rightleftharpoons cC + dD$
where capital letters are the reacting species while small letters are the stoichiometric coefficients

(iii) Equilibrium constant (Kc) or mass action expression is the product of the equilibrium concentration of the products, each raised to the power that corresponds to its coefficient in the balanced equation.
$N_{2(g)} + O_2 \rightleftharpoons 2NO_{(g)}$

$$K_c = \frac{[NO]^2}{[N_2][O_2]}$$

NOTE: The Kc unit depends on the coefficients of the reacting species and product

(2b)
(i) Le Chatelier's principle
This states that if a change of conditions (stress) is applied to a system at equilibrium, the system shift in the direction that reduces the stress so as to move toward a new state of equilibrium

oxidation half cell reaction we will reverse the reaction

- oxidation half cell

$2OH^-(aq) + Ni(OH)_2(s) \rightarrow NiO_2(s) + 2H_2O(l) + 2e^-$ $E^\circ = +0.490v$

- Reduction half cell

$HO_2^-(aq) + H_2O(l) + 2e^- \rightarrow 3OH^-(aq)$ $E^\circ = +0.878v$

Now sum the two half cells reactions to have the overall redox reaction

Oxidation half cell

$2OH^-(aq) + Ni(OH)_2(s) \rightarrow NiO_2(s) + 2H_2O(l) + 2e^-$ $E^\circ = -0.490v$

Reduction half cell

$HO_2^-(aq) + H_2O(l) + 2e^- \rightarrow 3OH^-(aq)$ $E^\circ = +0.878$

OVERALL REACTION is

$Ni(OH)_2(s) + HO_2^-(aq) \rightarrow NiO_2(s) + OH^-(aq) + H_2O(l)$

$E^\circ_{cell} = E^\circ_{ox} + E^\circ_{red}$

$E^\circ_{cell} = (-0.490 + 0.878)v$

$E^\circ_{cell} = +0.388v$

Using $E^\circ_{cell} = \frac{\Delta G^\circ}{-nF} = \frac{-2.303RT \log K}{-nF} = \frac{2.303RT}{nF}\log K$

where $n=2$; $F = 96500\ J/v$; $E^\circ_{cell} = +0.388V$; $R = 8.314\ J/mol\,K$

$T = 278K$; $\therefore E^\circ_{cell} = \frac{2.303R}{nF}\log K$

$\log K = \frac{nF E^\circ_{cell}}{2.303RT} = \frac{2 \times 96500\ J/v\,mol \times 0.388V}{2.303 \times 8.314\ J/mol\,K \times 298K} = 13.124$

$K = 10^{13.124} = 1.33 \times 10^{13}$

(5b)

$H_2(g) + CuO(s) \rightarrow Cu(s) + H_2O(g)$

$\Delta S^\circ = \underbrace{(33.3 + 188.7)}_{product} - \underbrace{(131.0 + 43.5)}_{reactant}$

$= 222 - 174.5 = +47.5\ JK^{-1}mol^{-1}$

Product − reactant $= \Delta S^\circ$

(5)

(a) The strongest oxidizing agent are the ones with the largest standard reduction potential i.e $Cl_2(g) = +1.35v$ has the largest SRP hence has the strongest oxidizing species (C)

14. A true solution is ________________.

15. Why are heterogeneous mixtures not regarded as true solution ________________.

16. Which of the following is not a solution (a) sea water (b) blood cells (c) urine (d) alloy

17. Write an equation to show rusting of iron (oxidation of Fe) and making of iron (reduction Fe_2O_3) via reduction of iron ore Fe_2O_3 with charcoal (C)

18. Which of the following is true of a galvanic cell under standard condition (A) 1M, 25°C, 1 atm (B) 1110 V, 1M, 25°K, 1atm (C) 1M, 25°K, 1atm (D) None of the above

19. The work done when a gas is compressed in a cylinder is 462J and the heat transferred from the gas to the surrounding is 128J. calculate the energy change for the process

20. (a) For the reaction C + O_2 → CO_2 $H\theta_m$ CO_2 is – 393,5KJ mol^{-1} what is the standard molar enthalpy for the formation of carbon dioxide (b) the type of equation above is called ________________.

21. Four colligative properties of solutions are i. ____________ ii. ____________ iii. ____________ iv. ____________.

22. Which type of acid is contained in the following (a) Lemon (b) Human blood (c) Digestive juice (d) Car battery electrolyte (e) Vinegar

23. ____________ define acid as a substance that contains hydrogen and produces H^+ in aqueous solution while a base is a substance that contains hydroxyl group (OH) and produces hydroxide ions (OH-) in aqueous solution

(a) Bronsted – Lowry theory (b) Lewis's theory of acids and base (c) Arrhenius definition

24. (a) $H^+ + \ddot{N}H_3 \rightarrow [NH_4]^+$ (b) $NH_3 + H^+ \rightleftharpoons NH4^+$. Equations (a) and (b) above represents which theories of acid and base.

25. Which of the following reactions does not occur through second order reaction (a) $C_2H_4 + H_2 \rightarrow C_2H_6$ (b) 2HI → $H_2 + I_2$ (c) $CH_3COOC_2H_5$ + NaoH → CH_3COONa + C_2H_5OH (d) $CH_3CHO \xrightarrow{I2} CH_4$ + CO

26. One of the following is not a third order rate of reaction (a) A + B + C → products (b) 2A + B → products (c) 3A → products (d) A + B → products

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Department of Chemistry
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CHM 109 (General Chemistry Practical II) Examination Time allowed: 45mins
Instruction: Answer All Questions

Surname:..............................Other Names:..............................

Matric No:..............................Faculty/Dept:..............................

Signature:..............................No. on Attendance List:....................

1(a) *The reaction between dil H_2SO_4 and dil NaOH is called* ....

(b) Write a balanced equation for the reaction.

....

2. Identify each of the following as a chemical or physical change.

(a) dissolving sugar in tea ....

(b) combustion of gasoline ....

(c) souring of milk ....

(d) decaying of garbage ....

3. List four Apparatus and two reagents you used during experiments.

(i) .... (ii) .... (iii) ....

(iv) .... (v) .... (vi) ....

4. Mention four laboratory safety rules.

(i) ....

(ii) ....

(iii) ....

(iv) ....

5. List four ways a chemical change can be identified.

(i) .... (ii) ....

(iiii) .... (iv) ....

6. What is the colour of methyl orange in

(a) A solution of strong Acid ....

(b) A solution of strong Base ....

You don't have to be great to start, but you have to start to be great. –Zig Ziglar

My dear, I just want to remind you that you have started your journey to greatness and success is sure for you if you diligently follow the path of excellence.

Do your best and leave the rest for God, because I am sure that God, who began the good work within you, will continue his work until it is finally finished on tha day when Christ Jesus comes back again. Phil.1.6 (NLT)

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