name:__________________________

Short Answer [2 marks each]

1. List two of the four ways we have of predicting whether a reaction will be spontaneous or not?

$\Delta S_{univ} > 0$ $E^{\circ}_{cell} > 0$

$\Delta G < 0$ $K > 1$ 2 of these 4

2. What is the standard state for gases?

$P = 1\,atm$

$T = 298\,K$

3. Define enthalpy of formation.

ΔH for the reaction forming a compound from its constituent elements in their standard states.

4. For each pair listed below, select which is the stronger reducing agent.

[✓ Li] or K

[✓ Fe] or Cu

5. Define a Bronsted acid.

proton donor

6. For the following reaction, identify the base and its conjugate acid.

$$\underbrace{NH_3}_{\text{base}} + H_2O \rightarrow \underbrace{NH_4^+}_{\text{conj acid}} + OH^-$$

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name:__________________________

7. Write out the cell notation for an electrochemical cell containing electrodes of metals Fe and Ag immersed in a single electroytic solution of $Fe(OH)_2$ and AgOH.

$Fe(s)\,|\,Fe^{2+}, Ag^{+}, OH^{-}\,|\,Ag(s)$

8. What is the electron configuration for Cu?

$[Ar]\,3d^{10}4s^{1}$

9. Which of the following processes involves a phase change (put an X beside your choices)

a) dissolving sugar into a glass of water X

b) condensation of methanol X

c) mixing liquid silver and lead to form an alloy X

d) mixing oil and water _____

10. What are the two most significant differences between an ideal gas and a real gas?

real gases take up space & they interact

11. Why can a process that is spontaneous result in a decrease in the entropy of the system undergoing the change?

spontaneous means $\Delta S_{univ} > 0$

So, as long as ΔS_{surr} is +ve, ΔS_{sys} can be -ve

12. True or false: A strong oxidizing agent is less likely to be reduced in solution.

False

name:________________________

13. The data below show the temperature dependence of the equilibrium constant for the reaction

$Ag_2CO_3 (s) \rightleftharpoons Ag_2O (s) + CO_2 (g)$

$K = 3.98 \times 10^{-4}$ (350 K) and K = 1.48 (500 K)

a) Calculate the standard enthalpy change for the reaction, assuming it is independent of temperature [4].

$$\ln\left(\frac{K_2}{K_1}\right) = -\frac{\Delta H^\circ}{R}\left(\frac{1}{T_0} - \frac{1}{T_1}\right)$$

$$\ln\left(\frac{1.48}{3.98 \times 10^{-4}}\right) = \frac{-\Delta H^\circ}{8.314\,J K^{-1} mol^{-1}}\left(\frac{1}{500K} - \frac{1}{350K}\right)$$

$$\Delta H^\circ = 79.7\ kJ/mol$$

b) Calculate the standard entropy change for the reaction at 350 K [4].

need ΔG°

$$\Delta G^\circ = -RT \ln K_{350K}$$
$$= 22.8\ kJ/mol$$

$$\Delta S^\circ = \frac{\Delta H^\circ - \Delta G^\circ}{T}$$
$$= 163\ J K^{-1} mol^{-1}$$

name:________________________

14. For the following Daniell Cell at 298 K

$Zn(s) \mid ZnCl_2\ (m=0.01\ M) \parallel CuCl_2\ (m=0.005\ M) \mid Cu(s)$

The overall reaction is:

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

Calculate the following: E°, E, ΔG, ΔG° and K [12].

a) $E^\circ = 0.340 + 0.763 = 1.10V$

b) $E = E^\circ - \frac{RT}{nF} \ln\left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right)$

n = 2 when written as a 2e⁻ process

$= 1.10V - 0.091V =$ ~~1.09V~~ $1.09V$

c) $\Delta G = -nFE$

$= -210\ kJ/mol$

d) $\Delta G^\circ = -nFE^\circ$

$= -212\ kJ/mol$

e) $K = e^{-\Delta G^\circ/RT}$

$= 1.45 \times 10^{37}$

name:____________________

15. The C—C homolytic bond strength in ethane is the ΔH for the following reaction:

$$CH_3CH_3 \rightarrow 2CH_3^{\bullet}$$

a) To determine ΔH, we usually use the Δ_fH^o for the various species in the reaction. In this case, what are they [2]?

$\Delta_f H(CH_3CH_3)$, CH_3

b) Using the information provided, determine $\Delta_fH^o(CH_3^{\bullet})$ [10].

forward	$CH_3F + F^{\bullet} \rightarrow CH_3^{\bullet} + F_2$	k (300 K) = 1.6×10^{-3} $M^{-1}s^{-1}$; k (400 K) = 2.75×10^{12} $M^{-1}s^{-1}$
reverse	$CH_3^{\bullet} + F_2 \rightarrow CH_3F + F^{\bullet}$	k (300 K) = 33 $M^{-1}s^{-1}$; k (400 K) = 4950 $M^{-1}s^{-1}$

$\overline{\text{rate}}$ constants !

$\Delta_fH^o(F^{\bullet})$ = 79 kJ mol^{-1}
$\Delta_fH^o(CH_3F)$ = -234 kJ mol^{-1}

$E_{a,forward} - E_{a,reverse} = \Delta H$

forward $\ln\left(\frac{k_{400}}{k_{300}}\right) = \frac{-E_{a,forward}}{R}\left(\frac{1}{400K} - \frac{1}{300K}\right)$

$E_{a,forward} = 350$ kJ/mol

reverse $E_{a,reverse} = 50$ kJ/mole

$\Delta H = E_{a,forward} - E_{a,reverse} = 300$ kJ/mol

$\Delta H = \Delta_f H(CH_3^{\bullet}) + \Delta_f H(\overset{0}{F_2}) - \Delta_f H(CH_3F) - \Delta_f H(F)$
✓ ✓ ✓

$\Delta_f H(CH_3) = 145$ kJ/mol

c) Given that Δ_fH^o (ethane) = -84 kJ mol^{-1}, what is the homolytic bond strength in ethane [2]?

$\Delta H = 2\Delta_f H(CH_3^{\bullet}) - \Delta_f H(\text{ethane})$
$= 374$ kJ/mol

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name:____________________

16. Determine the percent ionization of acetic acid ($HC_2H_3O_2$) at the following concentrations: 1.0 M and 0.010M. $K_a = 1.8 \times 10^{-5}$ M [12]

a) $CH_3CO_2H \rightleftharpoons CH_3CO_2^- + H^+$

1.0	0	0
$-x$	x	x
$1.0-x$	x	x

$K = 1.8\times10^{-5} = \frac{x^2}{1-x}$ assume $x \ll 1$

$x = 0.0042M$

% ionization $\frac{0.0042}{1.0} \times 100 = 0.42\%$

b)

0.01	0	0
$-x$	x	x
$.01-x$	x	x

$1.8\times10^{-5} = \frac{x^2}{.01-x}$ assume $x \ll 0.01$

$x = 0.00043$

% ionization $\frac{0.00043}{.01} \times 100 = 4.3\%$

name:____________________________

17. What concentration of formate ion, CHO_2^-, must be present in 0.366 M $HCHO_2$ to produce a buffer solution with pH = 4.06? K_a $HCHO_2 = 1.8 \times 10^{-4}$ M [8]

$$HCHO_2 + H_2O \rightleftharpoons CHO_2^- + H_3O^+$$

a) using Henderson – Hasselbach

pH = 4.06

pKa = 3.74 $(-\log(1.8 \times 10^{-4}))$

$$4.06 = 3.74 + \log\left(\frac{[CHO_2^-]}{[HCHO_2]}\right)$$

$$\frac{[CHO_2^-]}{[HCHO_2]} = 2.1$$

~~&~~ since $[HCHO_2] = 0.366$, $[CHO_2^-] = 0.77M$

OR do an ICE table

I	0.366	x	0
eq	$0.366 - 8.71 \times 10^{-5}$	$x + 8.71 \times 10^{-5}$	8.71×10^{-5} (from pH)

$$1.84 \times 10^{-4} = \frac{(x + 8.71 \times 10^{-5})(8.71 \times 10^{-5})}{0.366 - 8.71 \times 10^{-5}}$$

x = 0.76M.

name:

18. Balance the redox reaction $MnO_2(s) + ClO_3^- \rightarrow MnO_4^- + Cl^-$ occuring in basic solution [12]

step 1 $\overset{+4}{Mn}\underset{-2}{O_2} + \overset{+5}{Cl}\underset{-2}{O_3^-} \rightarrow \overset{+7}{Mn}\underset{-2}{O_4^-} + \underset{-1}{Cl^-}$

ox $MnO_2 \rightarrow MnO_4^-$

red $ClO_3^- \rightarrow Cl^-$

Step 2 ox. ½ rxn

$4OH^- + 2H_2O + MnO_2 \rightarrow MnO_4^- + 4H^+ + 4OH^-$

$\rightarrow MnO_4^- + 4H_2O + 3e^-$

$4OH^- + MnO_2 \rightarrow MnO_4^- + 2H_2O + 3e^-$ (1)

Step 3 red ½ rxn

$6OH^- + 6H^+ + ClO_3^- \rightarrow Cl^- + 3H_2O + 6OH^-$

$6e^- + 6H_2O + ClO_3^- \rightarrow Cl^- + 3H_2O + 6OH^-$

$3H_2O + 6e^- + ClO_3^- \rightarrow Cl^- + 6OH^-$ (2)

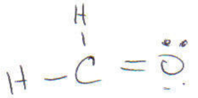
Step 4 (1) x2 + 2 & cancel common species

$\boxed{2OH^- + 2MnO_2 + ClO_3^- \rightarrow 2MnO_4^- + Cl^- + H_2O}$

name:

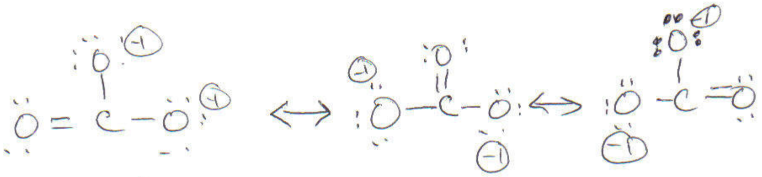
19. Write Lewis structures (including formal charges) for the following compounds [10].

a) CH_2O

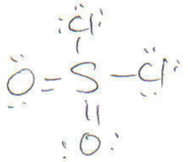


no formal charges

b) CO_3^{2-}



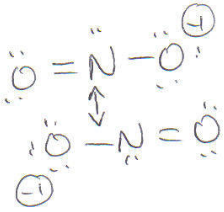
c) O_2SCl_2



no formal charge
max double bonds around S

d) NO_2^-

half marks | full marks



e) OCN^-