EXAM #3 **May 3, 2005**
CHEMISTRY 212, Lect. Sect. 002
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I, _____, understand the ramifications of willful misconduct during examinations and am not guilty of receiving assistance completing this exam. I further agree that I did not observe misconduct and not report it, making me an accessory to the wrongful act, which is a violation of the honor code at George Mason University.

CLOSED BOOK EXAM—No notes or books allowed. Calculators may be used. Atomic masses of interest are included. Periodic tables are not allowed for this exam.

1. What is the solubility product expression for mercury(I) cyanide, $Hg_2(CN)_2$?

   (a) $[Hg^+]^2[CN^-]^2$ (b) $[Hg^+][CN^-]$ (c)$[Hg_2^{2+}][CN^-]^2$ (d) $[Hg_2^{2+}][2CN^-]^2$ (e) $[Hg_2][CN]^2$

2. Which of the following metal sulfides is the next to the most soluble (mol/L) in water?

   (a) CoS ($K_{sp} = 4 \times 10^{-21}$) (b) CuS ($K_{sp} = 8 \times 10^{-36}$) (c) FeS ($K_{sp} = 5 \times 10^{-18}$)
   (d) HgS ($K_{sp} = 4 \times 10^{-50}$) (e) MnS ($K_{sp} = 6 \times 10^{-16}$)

3. Silver oxalate, $Ag_2C_2O_4$, is slightly soluble in water. The silver ion concentration in a saturated solution is 2.2×10^{-4}M. What is the K_{sp} of $Ag_2C_2O_4$?

   (a) 1.1×10^{-4} (b) 2.4×10^{-8} (c) 1.1×10^{-11} (d) 5.3×10^{-12} (e) 1.3×10^{-12}

4. K_{sp} for $Pb_3(PO_4)_2$ is 10^{-44}. Two solutions are mixed, one containing Pb^{2+} and the other PO_4^{3-}. If, at the instant of mixing, Pb^{2+} is 10^{-6}M and PO_4^{3-} is 10^{-8}M, which one of the following statements is true?

   (a) A precipitate forms because $Q_{sp} < K_{sp}$.
   (b) A precipitate forms because $Q_{sp} > K_{sp.}$
   (c) No precipitate forms because $Q_{sp} = K_{sp}$.
   (d) No precipitate forms because $Q_{sp} < K_{sp}$.
   (e) No precipitate forms because $Q_{sp} > K_{sp}$.

5. The addition of dilute hydrobromic acid would clearly distinguish between solutions of

   (a) barium nitrate and sodium sulfate
   (b) lead nitrate and silver nitrate
   (c) mercury(I) nitrate and silver nitrate
   (d) silver nitrate and calcium sulfate
   (e) calcium nitrate and barium nitrate

6. In a solution in which the fluoride-ion concentration is 0.40 M, what is the molar solubility of MgF_2? (K_{sp} for MgF_2 is 8.0×10^{-8})

   (a) 1.0×10^{-7} (b) 2.0×10^{-7} (c) 5.0×10^{-7} (d) 1.4×10^{-4} (e) 7.1×10^{-4}

7. The total entropy of a system and its surroundings always increases for a spontaneous process. This is a statement of

   (a) the law of constant composition
   (b) the first law of thermodynamics
   (c) the second law of thermodynamics
   (d) the third law of thermodynamics
   (e) the law of conservation of matter

8. Arrange the following in order of **INCREASING** entropy, S°: Hg(*l*), Hg(*s*), C_6H_6(*l*), CH_3OH(*l*)

   (a) Hg(*s*), CH_3OH(*l*), C_6H_6(*l*), Hg(*l*)
   (b) CH_3OH(*l*), Hg(*s*), Hg(*l*), C_6H_6(*l*)
   (c) Hg(*l*), Hg(*s*), C_6H_6(*l*), CH_3OH(*l*)
   (d) Hg(*s*), Hg(*l*), C_6H_6(*l*), CH_3OH(*l*)
   (e) Hg(*s*), Hg(*l*), CH_3OH(*l*), C_6H_6(*l*)

9. The Gibbs free energy is given by the equation

(a) $\Delta G = \Delta S - T\Delta H$ (b) $\Delta G = G_{initial} - G_{final}$
(c) $\Delta G = \Delta H - T\Delta S$ (d) $\Delta G = \Delta S - H\Delta T$
(e) $\Delta G = \Delta H - S\Delta T$

10. The best criterion for the spontaneity of a chemical reaction is the sign of

(a) ΔH (b) $\Delta H°$ (c) $T\Delta S$ (d) ΔG (e) $\Delta G°$

11. All the following have free energy of formation values of zero **EXCEPT**

(a) Ca(*s*) (b) H(*g*) (c) He(*g*) (d) Ni(*s*) (e) U(*s*)

12. Given the following

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$; $\Delta G° = -29.4$ kJ
$3Fe_2O_3(s) + CO(g) \rightarrow 2Fe_3O_4(s) + CO_2(g)$; $\Delta G° = -61.6$ kJ

calculate $\Delta G°$ for $Fe(s) + Fe_2O_3(s) + CO_2(g) \rightarrow Fe_3O_4(s) + CO(g)$

(a) –32.2 kJ (b) –16.1 kJ (c) +16.1 kJ (d) +32.2 kJ (e) +48.3 kJ

13. For a reaction that has an equilibrium constant of $3x10^5$, which of the following statements must be true?

(a) $\Delta G° < 0$ (b) $\Delta G° > 0$ (c) $\Delta H° < 0$ (d) $\Delta H° > 0$ (e) $\Delta S° > 0$

14. The driving force for the endothermic dissolution of an ionic compound is an increase in

(a) entropy (b) enthalpy (c) internal energy (d) Gibbs energy (e) work

15. In an electrochemical cell, which statement is **ALWAYS** true of the cathode?

(a) It is considered the "negative" electrode.
(b) It is considered the "positive" electrode.
(c) Reduction occurs here.
(d) Metal is plated out here.
(e) Negative ions flow toward the cathode.

16. What mass of chromium could be deposited by electrolysis of an aqueous solution of $Cr_2(SO_4)_3$ for 180.0 minutes using a constant current of 10.0 Amperes?

(a) 0.187 g (b) 0.373 g (c) 2.16 g (d) 6.47 g (e) 19.4 g

17. According to the following cell diagram, which species is undergoing oxidation?

$Sn \mid Sn^{2+} \mid\mid (Pt)MnO_2/Mn^{2+}$

(a) Sn (b) Sn^{2+} (c) Pt (d) MnO_2 (e) Mn^{2+}

18. Calculate E° for the cell reaction $2Cr + 3Sn^{4+} \rightarrow 3Sn^{2+} + 2Cr^{3+}$ given that

$Cr^{3+} + 3e^- \rightarrow Cr$ E° = –0.74V
$Sn^{4+} + 2e^- \rightarrow Sn^{2+}$ E° = 0.15V

(a) 1.93 V (b) 0.89 V (c) 0.59 V (d) 0.45 V (e) –0.59 V

KEY EQUATIONS

f.p. cyclohexane=6.55°C
k_f (cyclohexane) = 20.2°C/*m*
K.E. = $mv^2/2$
$m_e = 9.11 \times 10^{-31}$ kg
$m_p = 1.673 \times 10^{-27}$ kg
$m_n = 1.675 \times 10^{-27}$ kg
K_f (water) = 1.86°C/m
K_b(water) = 0.512°C/m
J = N•m
J = C•V
e = 2.718
F = 96,500 C
N = m•kg•s^{-2}
$N_0=N_A=6.02 \times 10^{23}$ units/mol
R= 0.0821 L•atm/mol•K
R= 8.314 J/mol • K
1 L•atm = 101 J = 0.101 kJ
1 Pa = 1 kg/(m• s^2) = 1 N/m^2
1 atm = 1.01×10^5 Pa
g = 9.807 m/s^2
M = n/V
w = -PΔV
ΔU = q + w
$\Delta H°_{rxn} = \Sigma \Delta H_f^\circ$(products) - $\Sigma \Delta H_f^\circ$(reactants)
$\Delta H_{rxn} = \Delta U_{rxn} + P\Delta V$
$\Delta H = nC_p\Delta T$
$\Delta H = c_p\Delta T$
$q = ms\Delta T$
1 Ci = 3.700×10^{10} dps
$t_{1/2}$ = 5730 years for ^{14}C

$$x_i = \left(\frac{n_i}{n_{total}}\right) = \left(\frac{P_i}{P_{total}}\right)$$

$P_i = x_i P_T$
$P_T = \Sigma P_i$

$$\log\left(\frac{P_2}{P_1}\right) = \left(\frac{\Delta H_{vap}}{2.303R}\right)\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\ln\left(\frac{K_2}{K_1}\right) = \left(\frac{\Delta H^0_{rxn}}{R}\right)\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$S = k_H P$

$$P_A = P^o_A X_A$$

$$\Delta P = P^o_A X_B$$

$\Delta T_b = K_b c_m$
$\Delta T_f = K_f c_m$
$\pi = MRT$

$$\ln\frac{[A]_t}{[A]_0} = -kt$$

$$\log\frac{[A]_t}{[A]_0} = -kt/2.303$$

$$\log\left(\frac{k_2}{k_1}\right) = \frac{E_a}{2.303R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$t_{1/2} = 0.693/k$
$k = pfZ$
$f = e^{-Ea/RT}$

$$\text{r.m.s.} = \sqrt{\frac{3RT}{M_m}}$$

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

$$Q_c = \frac{[C]_i^{\,c}[D]_i^{\,d}}{[A]_i^{\,a}[B]_i^{\,b}}$$

$K_p = K_c(RT)^{\Delta n}$

$K_{reverse} = 1/K_{forward}$

$$K = \frac{[H_3O^+]^2[S^{2-}]}{[H_2S]}$$

$w_{max} = -nFE_{cell}$
$E^0_{cell} = E^0_{cathode} - E^0_{anode}$
$\Delta G^0 = -nFE^0_{cell}$
$\Delta G_{rxn} = \Delta G^0_{rxn} + RT\ln Q$
$\Delta G^0_{rxn} = -RT \ln K$
$\Delta G_{rxn} = RT\ln(Q/K)$
$E^0_{cell} = (0.0592/n)\log K$
$E_{cell} = E^0_{cell} - (0.0592/n)\log Q$
$Q = It$
$\Delta S_{fus} = (\Delta H_{fus}/T_m)$

H 1.008	B 10.81	K 39.10	Ca 40.08
C 12.01	Br 79.90	Cl 35.45	F 19.00
N 14.00	Cr 52.00	P 30.97	Fe 55.85
O 16.00	S 32.00	Zn 65.39	Ni 58.70
Na 23.00	I 126.90		