Chemistry 152A — Tuesday Quiz (@ end of period)

Quiz # 6 (May 13, 2008) — Value: 10 points; Time: 10 minutes

Name ___________KEY______________________

Section __________________

Compare with problem Z11.52 This is an important warmup question for the next exam.

$$\Delta G^o_{rxn} = -RT\ln K = -nF\mathcal{E}^o \qquad \mathcal{E} = \mathcal{E}^o - \frac{RT}{nF}\ln Q$$

$$R = 8.3 \, {}^{J}\!/\!{}_{mol-K} \quad @T = 298 \quad RT = 2.48 \, {}^{kJ}\!/\!{}_{mole} \quad \frac{RT}{F} = .025\,V$$

For all of the short answer questions below, <u>explain</u> your conclusion.

Q6) Consider the Cu/Cu2+ battery ($@T = 25°C$). In each of the two beakers that will form the two halves of a battery there is a Cu metal strip, which acts as the terminal, and 100 ml of a solution containing $Cu(NO_3)_2$ (as Cu^{2+} and NO_3^- ions). In beaker A the $[Cu^{2+}]$ is 1.0 M, and in beaker B the $[Cu^{2+}]$ is 0.001 M. The two beakers are connected in standard battery configuration with a wire between the terminals and a voltmeter on the wire, and a salt bridge between the two solutions.

a) Explain why this battery will generate any voltage at all.
The concentration in the two cells is different; the voltage comes from the non-equilibrium nature of the concentration of the cells and the Q term. [This battery uses entropy entirely to generate the non-equilibrium Gibbs Energy.]

b) Which beaker A or B is the cathode and why?
The Red Cat: The Cathode is where reduction occurs. The Cu^{2+} goes to Cu (it is reduced) at the Cathode. In Beaker A the $[Cu^{2+}]$ is large and needs to be dropped, so in A the Cu^{2+} goes to Cu metal, so A is the cathode. The pole must be positive, and electrons are consumed at this pole.

c) Calculate the voltage of the battery:

$$\mathcal{E}_{Cell} = \mathcal{E}^o_{Cell} - \frac{RT}{nF}\ln Q \qquad \mathcal{E}^o_{Cell} = 0 \qquad n = 2 \quad Q = \frac{0.001}{1.0} = 1\cdot 10^{-3}$$

$$\mathcal{E}_{Cell} = 3\frac{RT}{nF}\ln 10 = 0.09V$$

d) What will the concentration of the $[Cu^{2+}]$ be in beaker A when the battery voltage is zero? (The battery runs down and goes to equilibrium.)

$$[Cu^{2+}]_{eq} \simeq \tfrac{1}{2}M = \tfrac{1}{2}(1.0 + 0.001)$$

e) If the battery is allowed to run until the concentration in A drops to 0.9 M, what was the maximum electrical work the battery could have performed?

Battery uses 0.1M Cu^{2+} in 100 ml, so 0.01 moles of Cu^{2+} were consumed (This is X below). The reversible work then:

$$w = X\Delta G_{rxn} = X\left(\Delta G^o_{rxn} + RT\ln Q\right) = 0.01(0 - 2.5\cdot 3\cdot 2.3)kJ = -170\ J$$