**Quiz 3**
**CHEM 323**
**Fall 2008**

Name: ______________________________________________________________

1a. (10 Points) The enthalpy of dissolution, $\Delta_{diss}H^0$, for $CaCl_2$ in water -81.5 kJ/mole. If a 10.91-g sample of $CaCl_2$ is dissolved in 150.0 g of water, with both substances initially at 25.0 °C, what will be the final temperature of the solution? Assume that no heat was lost to the surroundings and that the specific heat capacity of the solution is 4.184 J·°C$^{-1}$·g$^{-1}$. Molar masses: $H_2O$ 18.0153 g/mole and $CaCl_2$ 110.9834 g/mole.

**Since no heat escapes, the heat generated by this exothermic reaction must go to warming the solution. Assuming that no work is done, we may write (by way of the First Law) that** $q_{rxn} = -q_{solution}$ **(where the reaction is the "system" and the solution is the "surroundings"). We may calculate** $q_{system}$ **from the given** $\Delta_{diss}H^0$ **and** $q_{solution}$ **from the specific heat capacity.** *Remember that* $C_p = \left(\frac{dH}{dT}\right)_p = \left(\frac{dq}{dT}\right)_p$, *which integrates to* $q = C_p \Delta T$, *assuming* $C_p$ *is independent of T.* $C_p$ *and C, the specific heat capacity are the same thing, only with different units.*

**We may, therefore, write** $q_{rxn} = n\Delta_{diss}H^0$ **and** $q_{solution} = mC\Delta T$**, which can be combined to give** $n\Delta_{diss}H^0 = -mC\Delta T$ **and** $\Delta T = -\frac{n\Delta_{diss}H^0}{mC}$**.**

**Substituting the given values in yields**

$$\Delta T = -\frac{n\Delta_{diss}H^0}{mC} = -\frac{(10.91\text{ g})\left(\frac{1\text{ mole}}{110.9834\text{ g}}\right)\left(-81.5x10^3\text{ J}\cdot\text{mole}^{-1}\right)}{(10.91+150.0\text{ g})\left(4.184\text{ J}\cdot^\circ\text{C}^{-1}\cdot\text{g}^{-1}\right)}$$

$$\Delta T = \frac{8.01_1 x10^3\text{ J}}{(160.9_1\text{ g})\left(4.184\text{ J}\cdot^\circ\text{C}^{-1}\cdot\text{g}^{-1}\right)} = \frac{8.01_1 x10^3\text{ J}}{673.2_4\text{ J}\cdot^\circ\text{C}^{-1}} = +11.9_0\ ^\circ\text{C}$$

**Therefore, the final temperature is ($11.9_0$ + 25.0 °C =) 36.9 °C.**

b. (3 Points) Write the balanced chemical equation that defines the standard enthalpy of formation for solid $CaCl_2$ (that is $\Delta_f H^0$ ($CaCl_2$, s)).

**Ca (s) + $Cl_2$ (g) → $CaCl_2$ (s)**

2. (8 Points) A certain liquid has an enthalpy of vaporization, $\Delta_{vap}H^0$, equal to 26.0 kJ·mole$^{-1}$ at its boiling point of 250.0 K. Calculate $q$, $w$, $\Delta H$ and $\Delta U$ when 0.500 mole of the liquid is vaporized at 250.0 K and 750.0 Torr. State all necessary assumptions.

**At constant pressure $q = \Delta H$ and so for this process $q = n\,\Delta_{vap}H^0$. If we assume that the volume of the gas formed is much larger than the volume of the liquid, $\Delta V \approx V_{gas}$, and so $w = -pV_{gas}$. Assuming that the vapor is ideal (perhaps not the best assumption, but it makes things easier), we can replace $pV_{gas}$ with $nRT$ to give $w = -nRT$. Since $\Delta U = q + w$, we can find $\Delta U$ once we've found $q$ and $w$.**

$$q = \Delta H = n\Delta H = (0.500\ \text{mole})\left(26.0\frac{\text{kJ}}{\text{mole}}\right) = +13.0\ \text{kJ}$$

$$w = -nRT = (0.500\ \text{mole})\left(8.31447\frac{\text{J}}{\text{K}\cdot\text{mole}}\right)(250.0\ \text{K}) = -1.03_9\ \text{kJ}$$

$$\Delta U = q + w = +13.0\ \text{kJ} + -1.03_9\ \text{kJ} = +12.0\ \text{kJ}$$

**In summary: $q = \Delta H = +13.0$ kJ, $w = -1.04$ kJ and $\Delta U = +12.0$ kJ.**

3. (5 Points) From the following, calculate $\Delta H$ of the reaction FeO (s) + CO (g) $\rightarrow$ Fe (s) + $CO_2$ (g).

$Fe_2O_3$ (s) + 3 CO (g) $\rightarrow$ 2 Fe (s) + 3 $CO_2$ (g) $\Delta H$ = -23. kJ
3 $Fe_2O_3$ (s) + CO (g) $\rightarrow$ 2 $Fe_3O_4$ (s) + $CO_2$ (g) $\Delta H$ = -39. kJ
$Fe_3O_4$ (s) + CO (g) $\rightarrow$ 3 FeO (s) + $CO_2$ (g) $\Delta H$ = +18. kJ

**Take the third equation, flip it and multiply it (and its $\Delta H$) by 2. Then take the first equation and multiply it and its $\Delta H$ by 3. The second equation just needs to be flipped (which also changes the sign on its $\Delta H$). This gives**

6 FeO (s) + 2 $CO_2$ (g) $\rightarrow$ 2 $Fe_3O_4$ (s) + 2 CO (g) $\Delta H$ = -36. kJ
3 $Fe_2O_3$ (s) + 9 CO (g) $\rightarrow$ 6 Fe (s) + 9 $CO_2$ (g) $\Delta H$ = -69. kJ
2 $Fe_3O_4$ (s) + $CO_2$ (g) $\rightarrow$ 3 $Fe_2O_3$ (s) + CO (g) $\Delta H$ = +39. kJ

**Adding these up gives**

6 FeO (s) + 6 CO (g) $\rightarrow$ 6 Fe (s) + 6 $CO_2$ (g) $\Delta H$ = -66. kJ

**which is six times the reaction we want, so divide the equation and its $\Delta H$ by 6 to give**

FeO (s) + CO (g) $\rightarrow$ Fe (s) + $CO_2$ (g) $\Delta H$ = -11. kJ