Here's some comments about a few questions I was asked about during the review or in office hours today concerning the "Practice Final" (1210 ACS PracFinal) and Chemistry Olympiad exams.

#6, 2012 ACS exam:

30.0 mL of 0.10 M Ca(NO₃)₂ and 15.0 mL of 0.20 M Na₃PO₄ solutions are mixed. After the reaction is complete, which of these ions has the **lowest** concentration in the final solution?

(A) Na^+ (B) NO_3^- (C) Ca^{2+} (D) PO_4^{3-}

Need to write the reaction (this is an exchange reaction), and balance the eqn and do a limiting reactant problem.

 $3 \operatorname{Ca}(\operatorname{NO}_3)_2(\operatorname{aq}) + 2 \operatorname{Na}_3\operatorname{PO}_4(\operatorname{aq}) \longrightarrow \operatorname{Ca}_3(\operatorname{PO}_4)_2(\operatorname{s}) + 6 \operatorname{NaNO}_3(\operatorname{aq})$

To determine the LR we have to calculate moles of each reactant (use M*V, V in liters).

? mol Ca(NO₃)₂ = $(0.10 \text{ M}) (0.0300 \text{ L}) = 0.0300 \text{ mol Ca}(NO_3)_2$

? mol Na₃PO₄ = $(0.20 \text{ M}) (0.0150 \text{ L}) = 0.0300 \text{ mol Na}_3PO_4$

The ratio from the balanced eqn is

 $3 \text{ mol } Ca(NO_3)_2$ ------ = 1.5 mol Ca(NO_3)_2/ 1 mol Na_3PO_4 2 mol Na_3PO_4

The ratio we have is:

 $\begin{array}{l} 0.0300 \ \text{mol} \ \text{Ca}(\text{NO}_3)_2 \\ \hline \\ 0.0300 \ \text{mol} \ \text{Na}_3\text{PO}_4 \end{array} = 1 \ \text{mol} \ \text{Ca}(\text{NO}_3)_2 / 1 \ \text{mol} \ \text{Na}_3\text{PO}_4 \end{array}$

Thus, $Ca(NO_3)_2$ is the LR and ALL the Ca^{2+} will react and wind up in the solid product. That means there won't be any Ca^{2+} remaining in the solution. There will be some Na_3PO_4 left over. This means there will be Na^+ , NO_3^- and PO_4^{3-} ions left in solution. How much Na_3PO_4 is left over? How much product is formed? Use an ICE table for this.

 $I = \begin{bmatrix} 3 & Ca(NO_3)_2 & (aq) + 2 & Na_3PO_4 & (aq) & \dots > Ca_3(PO_4)_2 & (s) + 6 & NaNO_3 & (aq) \\ 0.00300 & mol & 0.00300 & mol & 0 & 0 \end{bmatrix}$ $C = \begin{bmatrix} -0.00300 & mol & -0.00200 & mol & +0.00100 & mol & +0.00600 & mol \\ 0.00000 & mol & 0.00100 & mol & +0.00100 & mol & +0.00600 & mol \end{bmatrix}$

So no Ca^{2+} would be left in solution. There would be 0.00100 mol of PO_4^{3-} , 0.00600 mol of Na^+ , and 0.00600 mol of NO_3^- in solution.

Practice Final

#62)

 $\Delta E = q + w$

E is a state function and depends only on its present conditions and not how it got to its present state. So ΔE depends only on its initial and final states and not how it gets from initial to final. Heat, q, and work, w, are not state functions and depend on the path taken.

For pressure-volume work, $w = -P \Delta V$, at constant pressure so

$$E = q - P \Delta V$$

#41) Limiting reactant problem. Determine which is the LR and then figure how much of the excess reactant will react with the LR and therefore how much is left over. 6.00 g of CS₂ gas react with 10.0 g of Cl₂ gas according to the following reaction,

$$CS_{2}(g) + 3 Cl_{2}(g) ---> CCl_{4}(\ell) + S_{2}Cl_{2}(\ell)$$

Determine which is the LR

$$? \text{ mol } \text{CS}_2 = 6.00 \text{ g } \text{CS}_2 \text{ x} - \frac{1 \text{ mol } \text{CS}_2}{76.1407 \text{ g } \text{CS}_2} = 0.07880 \text{ mol } \text{CS}_2$$

 $\begin{array}{rl} 1 \ \text{mol} \ \text{Cl}_2 \\ ? \ \text{mol} \ \text{Cl}_2 \ = 10.00 \ \text{g} \ \text{Cl}_2 \ x \ ----- = \ 0.14103 \ \text{mol} \ \text{Cl}_2 \\ \hline 70.906 \ \text{g} \ \text{Cl}_2 \end{array}$

Which is the LR? Different ways to determine this.

Can compare the ratio of the reactants to that in the balanced eqn.

Balanced Eqn. ratio $3 \mod \text{Cl}_2 / 1 \mod \text{CS}_2 = 3/1$ Ratio from moles $0.14103 \mod \text{Cl}_2 / 0.07880 \mod \text{CS}_2 = 1.79/1$ Since this ratio is less than 3 the Cl₂ is the LR

So determine how many grams of CS_2 will react with this number of moles of Cl_2 .

 $? g CS_2 = 0.14103 \text{ mol } Cl_2 x \frac{1 \text{ mol } CS_2}{3 \text{ mol } Cl_2} x \frac{76.1407 \text{ g } CS_2}{1 \text{ mol } CS_2} = 3.579 \text{ g } CS_2$

So 3.579 g CS_2 react so how many grams of CS2 are in excess.

? g CS₂ excess = 6.00 g - 3.579 g = 2.42 g excess CS₂

What is the net ionic eqn for the reaction of HCN (aq) with NaOH (aq)?

This is a reaction of a weak acid with a strong base.

The net ionic eqn for a strong acid with a strong hydroxide base, such as HCl (aq) with NaOH (aq) is,

HCl (aq) + NaOH (aq) ---> NaCl (aq) + H₂O (
$$\ell$$
) molecular eqn
H⁺ (aq) + Cl⁻ (aq) + Na⁺ (aq) + OH⁻ (aq) ---> Na⁺ (aq) + Cl⁻ (aq) + H₂O (ℓ) ionic eqn
H⁺ + OH⁻ ---> H₂O (ℓ) net ionic eqn

For a weak acid, HCN (aq) with a strong hydroxide base NaOH (aq) get,

HCN (aq) + NaOH (aq) ---> NaCN (aq) + H₂O (
$$\ell$$
) molecular eqn
HCN (aq) + Na⁺ (aq) + OH⁻ (aq) ---> Na⁺ (aq) + CN⁻ (aq) + H₂O (ℓ) ionic eqn
HCN (aq) + OH⁻ (aq) ---> CN⁻ (aq) + H₂O (ℓ) net ionic eqn

The HCN (aq) is still written in molecular form in the ionic and net ionic eqns because it is a weak electrolyte and most of it remains in the molecular form in solution so we write it in molecular form.

Calorimetry problem.

A 37.5 g piece of gold at 83.0 °C is added to 100.0 g H_2O at 22.0 °C in a well-insulated cup. What is the temperature after the system comes to equilibrium? (The specific heat capacity of Au is 0.129 $J \cdot g^{-1} \cdot K^{-1}$ and that for H_2O is 4.184 $J \cdot g^{-1} \cdot K^{-1}$)

For calorimetry you need to remember a few of things,

 $q = C * \Delta T$ C = heat capacity

 $q = m * C_s * \Delta T$ $C_s = specific heat capacity$

$$q_{lost} = -q_{gained}$$

The Au at the higher temperature loses heat to the H_2O at the lower temperature.

$$\begin{split} m_{Au} * Cs_{Au} * (T_{f} - T_{i, Au}) &= -m_{H20} * Cs_{H20} * (T_{f} - T_{i, H20}) \\ (37.5 g) * (0.129 J \bullet g^{-1} \bullet ^{\circ} C^{-1}) * (T_{f} - 83.0 \ ^{\circ} C) &= -(100.0 g) * (4.184 J \bullet g^{-1} \bullet ^{\circ} C^{-1}) * (T_{f} - 22.0 \ ^{\circ} C) \\ (4.8375 J \bullet ^{\circ} C^{-1}) * (T_{f} - 83.0 \ ^{\circ} C) &= -(418.4 J \bullet ^{\circ} C^{-1}) * (T_{f} - 22.0 \ ^{\circ} C) \\ (T_{f} - 83.0 \ ^{\circ} C) &= -(86.49) * (T_{f} - 22.0 \ ^{\circ} C) \\ T_{f} - 83.0 \ ^{\circ} C &= -(86.49) * T_{f} + 1902.80 \ ^{\circ} C \\ - 83.0 \ ^{\circ} C - 1902.80 \ ^{\circ} C &= -(86.49) * T_{f} - T_{f} \\ - 1985.8 \ ^{\circ} C &= -(87.49) * T_{f} \\ T_{f} &= 22.697 \ ^{\circ} C &= 22.7 \ ^{\circ} C \end{split}$$