CHEM 110 Final Exam - Practice Test 1 Solutions

1C
Given all the ions present in solution, we need to see which combinations would form an insoluble ionic compound. According to the solubility rules $CaCO_3$ is insoluble and will likely form a precipitate. Thus the remaining two ions, sodium and chloride, are the spectator ions that do not participate in the reaction.

2D
For this type of titration problem we can easily solve by using the equation :

$$M_A V_A N_A = M_B V_B N_B$$

Remember that often N is just equal to one, but in this case $N_B = 2$ because the base has two hydroxides ($N_A = 1$ and always will). Since we're looking to solve for the concentration of the acid, M_A, we write:

$$M_A = \frac{M_B V_B N_B}{V_A} = \frac{(0.2657)(35.65)(2)}{50.00} = 0.3789M$$

3A
For this problem we need to add together the heats of formation for the reactants and subtract the sum from the heats of formation for the products, making sure that we multiply each value by the corresponding stoichiometric coefficient, as the equation below shows:

$$\mathbf{\Delta H_{Rxn}} = \sum \mathbf{m} \Delta H_{f\,(products)} - \sum \mathbf{n} \Delta H_{f\,(reactants)}$$

$$[4(90.3) + 6(-241.8)] - [4(-45.9)] = -906kJ$$

Note that we do not include oxygen since the heat of formation is zero for any pure element.

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4A

This is a long question to solve, with lots of places to potentially make careless calculation errors, so be careful. First we need to draw lewis structures for the molecules so that we can assess which bonds are present. Since the bonds present on each side of the reaction are so different from each other we can't really use the short-cut of eliminating like bonds on each side. Instead it will be wiser to just add up all the bond dissociation energies for each bond on both sides of the reaction.

The bonds present in each molecule are as follows:

C_2H_5OH : 1 C-C bond, 1 C-O bond, 5 C-H bonds, 1 O-H bond = 3234 kJ
O_2 : 1 O=O double bond = 495 kJ
CO_2 : 2 C=O double bonds = 1598 kJ
H_2O : 2 O-H bonds = 926 kJ

We can then calculate heat of reaction using the equation below (taking into account the stoichiometry):

$$\mathbf{\Delta H_{Rxn}} = \sum \mathbf{D}_{\text{(reactants)}} - \sum \mathbf{D}_{\text{(products)}}$$

$$= [2(3234) + 7(495)] - [4(1598) + 6(926)] = -2015 \text{ kJ}$$

5B

In order to get our target reaction, we must modify the three reactions given as follows:

The first reaction is to be left alone, so its ΔH remains -622.2 kJ
The second reaction must be doubled, so its ΔH becomes (2 x -285.8) = -571.6 kJ
The third reaction must be flipped and doubled, so it becomes (-1 x 2 x -187.8) = 375.6kJ

When the modified reactions are added together they become our target equation, and so we can get the overall ΔH by adding together the modified ΔH's to get -818.2 kJ.

6A

This is an annoying way of testing your knowledge of solubility rules. If a reaction occurs it is because two ions have come together that are insoluble and so they drop out of solution as a precipitate (which just means a solid compound forming from a solution). The best way to look at this is to imagine you have all four ions present from the two solutions:

$$Ca^{2+}, Br^-, K^+, CO_3^{2-}$$

By looking at solubility rules we can see that $CaCO_3$ is insoluble and so it will form a precipitate.

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7C
The fastest way to complete these problems is to follow the metals on both sides. For part A, Ag becomes oxidized to Ag+. In b) Pb is oxidized to Pb +. In c) Fe stays 3+ on both sides. In part d) Sr is oxidized to Sr 2^+. Obviously C is the answer.

8D
The problem is best solved using the equation for diluting solutions: $M_1V_1 = M_2V_2$, where the initial concentration we're looking for is M_1. The solution is diluted from 5 mL to 27 mL, and so these are the initial and final volumes, respectively. The 10 mL sample is not needed to solve this problem and is just given to confuse you. Thus we can solve as follows:

$$M_1 = \frac{M_2V_2}{V_1} = \frac{(0.61)(0.027)}{(0.005)} = 3.294\,M$$

9C
Remember that "Like Dissolves Like". Since water mainly relies on H-bonding, we should expect our answer choice to do the same. The closest answers are (c) and (d) which both use H-bonding. However, since (d) is also used to a 4-carbon non-polar chain (which it will not get in water) it will not dissolve as well as (c) which only has a 1-carbon non-polar region. So (c) is the best choice.

10D
Lets find how many moles 73 grams represent. 73 grams/(12*6 + 12*1+ 6*16) = 0.41 moles. So for every 0.41 moles 732 kJ are released.
So 3 moles/0.41 moles = 7.39 *-732 = -5415 kJ.

11A
The balanced reaction looks as follows:

$$C_5H_{12}(g) + 8\ O_2(g) \rightarrow 5\ CO_2(g) + 6\ H_2O(g)$$

12E
In the first solution there are (0.05L)(0.1M) = 0.005 moles of HCl which means 0.005 moles Cl^-.
In the second solution there are (0.01L)(0.2M) = 0.002 moles of NaCl which means 0.002 moles of Cl^-.
Adding these two solutions gives us 0.007 moles of Cl^- in a 0.06 L solution. The new Cl- concentration is (0.007moles)/(0.06L) = 0.117 M.

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13A

Find the moles of $C_3H_5N_3O_9$ and convert it to moles of gas, then determine the volume.

$$642.41\ g\ C_3H_5N_3O_9 * \frac{1\ mole\ C_3H_5N_3O_9}{227\ g\ C_3H_5N_3O_9} * \frac{29\ moles\ gas}{4\ mole\ C_3H_5N_3O_9} = 20.51\ moles\ gas.$$

$$V = \frac{nRT}{P} = \frac{20.51\ moles * 0.08206 \frac{atm * L}{mole * K} * 273.15\ K}{1\ atm} = 459.7\ L$$

14B

Percent yield is (actual yield)/(theoretical yield) * 100. The actual yield is given as 3.6 grams. First find the limiting reagent.

7.73g H2O*(1 mole H_2O/18 grams H_2O)*(3 moles MgO/ 3 moles H_2O) = 0.43 mol MgO

3.82g Mg_3N_2*(1moleMg_3N_2/100.92gMg_3N_2)*(3 moles MgO/1moles Mg_3N_2) = 0.114 mol MgO

Mg_3N_2 is the limiting reagent, so we know 0.114 moles MgO is produced.
This gives the following mass:
(0.114moles MgO)*(40.31grams MgO/1 mole MgO) = 4.595 grams MgO.

(3.6grams/ 4.595grams) x 100% = 78.7% yield MgO

15E

We are looking for the mass the CFC. We know that $q_{sys}= m\Delta H_{vap.}$
ΔH_{vap} is given, and we know $q_{sys} = - q_{surr}$. We can calculate the q_{surr} by examining the water freezing. The water starts at 18°C. So it must first be cooled, then frozen. There are 2 steps in this case.

q1= mcΔT = (100g)(4.184J/gK)(-18°C) = -7531.2 J
q2= $-m\Delta H_{fus}$=(100g)(334J/g) = -33400 J
q_{surr} = - 40931.2 J

So $q_{sys} = - q_{surr}$ q_{sys}= 40931.2 J. Finally we can calculate the mass using $q_{sys}= m\Delta H_{vap.}$
m= $q_{sys}/\Delta H_{vap}$= 40931.2 J / 289J/g = 142 g CFC!!

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16D

(54g CH_4)*(1moleCH_4/16g CH_4)*(1mole CO_2/ 1 mole CH_4) = 3.375 moles CO_2

3.375 moles of product means the total heat released is 3.375 * (-1386.4 kJ) = -4679.1 kJ released.

17A

ppm is simply the mass (or volume, or moles) of the solute divided by the entire mass and then multiplied by a million:

$$\frac{0.0095\ g}{5\ \text{g}} = 1900\ \text{ppm} = 1.9 \times 10^3\ \text{ppm}$$

Remember to make sure that all values are in the same units before calculating.

18) 18378.3 kJ

Three steps in this endothermic process.

$q1 = mC\Delta T = (340.2g)(1.07J/gK)(-1 - -4) = 1.092$ kJ

$q2 = m\Delta H_{fus} = (340.2\ g)(54\ kJ/g) = 18370.8$ kJ

$q3 = mC\Delta T = (340.2\ g)(3.77\ J/gK)(4 - -1) = 6412.77$ J

q = (1.092 kJ + 18370.8 kJ + 6.41277 kJ) = 18378.3 kJ.

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19C

This is an acid base neutralization problem. Moles of H from the acid will neutralize moles of OH from the base. To determine the concentration of the resultant solution, you must determine the moles of H and OH present before the solution is mixed:

$$Moles\ Acid = M_a * V_a = .031\ mL * 0.9\ M = .0279\ moles\ H_2SO_4$$

$$Moles\ H = 0.0279\ moles\ H_2SO_4 * \frac{2\ moles\ H}{1\ moles\ H_2SO_4} = 0.0558\ moles\ H$$

$$Moles\ Base = M_B * V_B = 0.029mL * 1.69\ M = 0.0491\ moles\ RbOH$$

$$Moles\ OH = 0.0491moles\ RbOH * \frac{1\ mole\ OH}{1\ mole\ RbOH} = 0.0491\ moles\ OH$$

Since there are more moles of H, the resultant solution will be acidic. The concentration of H in the resultant solution will be the (moles of H after neutralization)/(Liters of final solution).

$$M_{a_2} = \frac{moles\ H\ after\ Neutrauliztion}{Total\ Volume} = \frac{0.0558\ moles\ H - 0.0491\ moles\ OH}{0.031 + 0.029\ L}$$

$$M_{a_2} = 0.11\ M\ H^+$$

20B

Let's set up and balance the combustion equation:

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

So 0.2 liters per mile means we burn 40 liters in this trip. Then we can calculate the mass by using the density, m = (40L)*(703g/L) = 28120g Octane.

(28120g C_8H_{18})*(1 mol C_8H_{18} / 114 g C_8H_{18})*(16 mol CO_2/ 2 mol C_8H_{18}) = 1.97×10^3 mol CO_2

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21B

Break the reaction into components showing the oxidation numbers on each element.

$$Cu^0 + 2Ag^+ + 2NO_3^- \rightarrow 2\,Ag^0 + Cu^2 + 2NO_3^-$$

Cancel out the components that are the same on both sides of the reaction and write down the redox reactions.

$$\text{Oxidation: } Cu^0 \rightarrow Cu^2 + 2e^-$$
$$Reduction: 2Ag^+ + 2e^- \rightarrow 2Ag^0$$

The oxidizing agent is the reactant that allows for oxidation to occur. Oxidation can only occur simultaneously with reduction. So the key point here is to realize that the species being reduced is IN the REACTANT that is the oxidizing agent. Since Ag is being reduced, $AgNO_3$ is the oxidizing agent.

22D

For this we want to use the dilution equation for calculating molarity with respect to volume:

$$M_1V_1 = M_2V_2$$

Please see the supplement for more info on using this equation.

To solve this question, we must find the volume of the resulting solution that has K^+ concentration of 0.5 M.

However, before we start plugging in values we must make sure that the concentrations for M_1 and M_2 are for the same molecule or ion. The problem we have is that the initial concentration of 2.5 M is for K_2SO_4 but the final concentration of 0.5 M is for K^+. To remedy this we can change the initial concentration of 2.5 M K_2SO_4 to **5.0 M K^+** since there are two K^+ ions for every one K_2SO_4 molecule.

Now we plug in the values:

$$V_2 = \frac{5.0 \times 50.0}{0.5} = 500\,mL$$

So the new solution must have volume 500 mL. However, this is not the answer since they ask how much must be added, not how much will the volume actually be. Instead, we will need **450 mL more water** to get from 50 mL to 500 mL.

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23C
Determine which ions are not present in the net ionic equation, these are the spectator ions. Start by determining the full molecular equation by using the solubility table.

$Full\ Molecular$: $(NH_4)_2CO_3\ (aq) + Sr(OH)_2\ (aq) \rightarrow SrCO_3\ (s) + 2\ NH_4OH(aq)$
$Full\ Ionic$: $2NH_4^+\ (aq) + CO_3^{-2}\ (aq) + Sr^{2+}(aq) + 2OH^-(aq) \rightarrow$
$SrCO_3(s) + 2NH_4^+\ (aq) + 2OH^-(aq)$
$Net\ Ionic$: $CO_3^{-2}\ (aq) + Sr^{2+}(aq) \rightarrow SrCO_3(s)$

Notice that NH4+ and OH- are the same on both sides. This means they did nothing in the reaction except watch (spectate). They are omitted from the net ionic and are the spectator ions.

24E
For neutralization problems, plug the values into the following equation.

$$N_aM_aV_a = N_bM_bV_b$$
$$(1)(0.5M_a)(10mL_a) = (2)(0.25M_b)(V_b)$$
$$V_b = 10mL$$

Take note that this equation suggests that the moles of H would be equal to the number of moles of OH at neutralization.

$$N_a = \frac{moles\ H}{1\ mole\ Acid}$$

$$N_aM_aV_a = (N_a) * moles\ of\ Acid = Moles\ of\ H$$

25E
Like Dissolves Like.
Octane is a hydrocarbon, non-polar and therefore used to dispersion forces. Only (d) and (e) are non-polar and of the two (e) is by far the closest in design to octane considering it has a 6-carbon non-polar chain.

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CHEM 110 Final Exam – Practice Test 2 Solutions

1C

The standard balanced equation for the combustion reaction (which always has CO_2 and H_2O as products) looks like:

$$2\ C_6H_{14} + 19\ O_2 \rightarrow 12\ CO_2 + 14\ H_2O$$

Since we have only one mole of hexane (instead of two) we will need half the amount of oxygen, so 9.5 moles of O_2.

2D

Since neither reactant is said to be in excess, we must first decide what the limited reactant is before determining the max yield. To do this we start by converting all mass values to moles:

$$12.52g\ \mathrm{SrH}_2 \times \frac{1mol}{89.636g} = 0.140mols$$

$$5.21g\ H_2O \times \frac{1mol}{18.015g} = 0.289mols$$

We could determine the limiting reactant the conventional way by calculating the maximum amount of product we can get from the perspective of each reactant, however it's possible to quickly recognize that according to the balanced reaction if we use all 0.140 moles of SrH_2 we will need double the number of moles of H_2O, which is 0.280, and we have 0.289 moles of H_2O. So, it turns out that SrH_2 is the limiting reagent, since there would be some H_2O left over if we used up all of the SrH_2.

The next step is to figure out how much H_2 we can get when we use up all of our limiting reagent, the 0.140 moles of SrH_2/:

$$0.140mols\ \mathrm{SrH}_2 \times \frac{2mols\ \mathrm{H}_2}{1mol\ \mathrm{SrH}_2} = 0.280mols\ \mathrm{H}_2 \times \frac{2.016g}{1mol\ \mathrm{H}_2} = 0.564g$$

3E

For this type of titration problem we can easily solve by using the equation $M_AV_AN_A = M_BV_BN_B$.

Remember that often N is just equal to one, as is the case for both the acid and the base in this problem. Since we're looking to solve for the concentration of the acid, M_A, we write:

$$M_A = \frac{M_BV_B}{V_A} = 0.4605M$$

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4D

This problem talks about increasing the temperature of water and so we are dealing with calorimetry, and the equation q = mcΔT. But before we can use that equation to solve for the final temperature we first need one more variable, q, the energy that is liberated by burning 2.5g of the compound:

$$2.5g \times \frac{42kJ}{7.5g} = 14kJ = 14{,}000kJ$$

Notice how for this problem I convert from kJ to J since the q = mcΔT uses that unit.
Next we solve for final temperature:

$$q = mc(T_f - T_i),$$

$$T_f = \frac{q}{mc} + T_i = 633.4$$

5B

This problem requires Hess's law to solve. We need to combine the two reactions given to end up with the target equation for which we need to find ΔH.
The target equation shows 2 moles of NO_2 in the reactants side. The only place we see NO_2 is in the second reaction, but on the products side and there's only one mole, not two. To fix this we need to flip the second reaction and also multiply it by 2 so that it looks like:

$$2\ NO_2 \rightarrow 2\ NO + O_2, \qquad \Delta H = 114.2\ kJ$$

Notice that the ΔH value has changed because the reaction has changed. Since the reaction was flipped, the sign was flipped, and since we multiplied the reaction by two, we do the same to the ΔH.
Next, notice that the target equation has one mole of N_2O_5 on the products side, and we can only find N_2O_5 in the first reaction just on the reactants side. So we simply flip the first reaction, and we flip the sign of the ΔH value too to get:

$$2\ NO + 3/2\ O_2 \rightarrow N_2O_5, \qquad \Delta H = -223.7\ kJ$$

Lastly, we add the two reactions together and end up with the target equation after we cancel out any molecules that are on both sides. We also can then add together the two new ΔH values to get the total ΔH for the target reaction: 114.2 + (-233.7) = -109 5 kJ.

6B

To answer this problem we need to track the oxidation numbers for every element in each compound as they go from reactants to products:

$$\underset{+1\ +5\ -2}{KClO_3\ (aq)} + \underset{+1\ -1}{HBr\ (aq)} \rightarrow \underset{0}{Br_2\ (l)} + \underset{+1\ -2}{H_2O\ (l)} + \underset{+1\ -1}{KCl\ (aq)}$$

Remember to check the rules for calculating oxidation numbers.
Notice that Cl goes from +5 to -1, which means it has gained 6 electrons and was reduced.
Notice, more importantly for this problem, that Br goes from -1 to 0, which means that it lost 1 electron, and thus HBr was oxidized.

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7B

To set this up we need to use calorimetry's energy conservation equation: $q_{sys} = -q_{surr}$. Let the unknown compound be the system, and the water the surroundings:

$$q_{sys} = q_{surr}$$

$$10.0 \times 0.803 \times (T_f - 108) = -[40.0 \times 4.184 \times (T_f - 25.0)]$$

$$T_f = 28.8$$

8E

When dealing with percent yield you usually need to find the theoretical yield first, as is the case for this problem. To do this, we first need a balanced equation:

$$H_2 + F_2 \rightarrow 2HF$$

Next we convert from mass to moles and then we determine the amount of moles of HF we could have gotten, which we then convert to mass:

$$16.1g\ H_2 \times \frac{1mol}{2.016g} \times \frac{2mol\ HF}{1mol\ H_2} \times \frac{20.006g}{1mol\ HF} = 319.5g\ HF$$

Now we just plug in the theoretical yield of 319.5 g into the percent yield equation:

$$\%Yield = \frac{300}{319.5} \times 100\% = 94\%$$

9E

The products of the reaction can be predicted using the solubility rules:

$$Ba(NO_3)_2\ (aq) + Na_2CO_3(aq) \rightarrow BaCO_3\ (s) + 2\ NaNO_3\ (aq)$$

This shows that Na^+ and Cl^- are always in aqueous form and thus are spectator ions.

10B

The best way to predict solubility is to follow the rule "like dissolves like".
With that in mind, we should recognize that water is dominated by hydrogen bonding, and so our best candidates will also have H-bonds. Of the options, choices 1 and 2 both have H-bonds, but option 2 also has a long 4-carbon nonpolar chain. This means that option 1 likes having H-bonds as well as a 4-carbon nonpolar region interaction, which it will not get when dissolved in water. So, of the two options, option 1 is less suited to be dissolved in water since it will not get the long nonpolar chain interaction that it's use to getting. This makes option 2 the most soluble, followed by option 1. After that, we have option 3 which despite not having H-bonds, is still polar and therefore closer to water's intermolecular forces than butane, which is clearly a nonpolar hydrocarbon that has completely different intermolecular forces than water.

11B

Based on the solubility rules provided on your data sheet, $CaSO_4$ is insoluble and will not dissolve in water.

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12C

When you're given the different ΔH_f values for each reactant and product we need to use the equation:

$$\mathbf{\Delta H_{Rxn}} = \sum \mathbf{m} \Delta H_{f\,(products)} - \sum \mathbf{n} \Delta H_{f\,(reactants)}$$

$$= [(-2346) + (-285) \times 2] - [(-986.6) + (-900.4) \times 2] = -130.4 kJ$$

13A

We are first gonna need a balanced reaction before we can determine exactly how many of each bond is present:

$$CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O$$

If we draw the correct Lewis structures we'll find the following bonds present:

Reactants: 4 C-H, 2 O=O, which when we add the values from the table gets 2642 kJ

Products: 2 C=O, 4 O-H, which when we add the values from the table gets 3450 kJ

Lastly we combine the energy values for reactants and products using the following equation:

$$\mathbf{\Delta H_{Rxn}} = \sum \mathbf{D}_{(reactants)} - \sum \mathbf{D}_{(products)} = -808\ kJ$$

14B

Decomposition the following reaction:

$$C_6H_{12}O_6\ (g) \rightarrow 6C\ (s) + 6H_2(g) + 3O_2(g)\ \ \Delta H°_{rxn} = -1200\ kJ$$

The reaction will release 1200 kJ for every 1 mole of $C_6H_{12}H_6$ decomposed.

$$90g\ C_6H_{12}O_6 * \frac{1\ mole\ C_6H_{12}O_6}{180\ g\ C_6H_{12}O_6} * \frac{-1200\ kJ}{1\ mole\ C_6H_{12}O_6} = -600\ kJ$$

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15D

This problem can be very deceiving. Way too much information is given. The fact that we have 1 L of stock solution is irrelevant. Recall molarity is calculated by moles/L. Also remember that molarity is consistent throughout the solution. If I take a random sampling of the stock solution, it will have the exact same molarity. Do not confuse molarity with the number of moles though. This random sampling of the stock solution will also have a random amount of moles depending on the volume of the sample. So 0.05L of the stock of 0.346 M is how many moles?

n/V = M, n = VM = (.346) x (.05) = 0.0173 moles.

This number of moles would then be diluted into a new volume, but we are looking for moles not molarity. Finally, notice that for every mole of K_2O, there are 2 moles of K. So moles of K will be 2 x (0.0173) = 0.0346

16E

Break each of these into their ions:

Na^+ CO_3^{2-} Rb^+ OH^- Li^+ SO_4^{2-}

Na^+ Rb^+ Li^+ are always soluble.

SO_4^{2-} is usually soluble.

CO_3^{2-} OH^- are usually insoluble.

We must pick the answer choice that is an exception to SO_4^{2-} but not to CO_3^{2-} OH^-.

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17A

This is a much trickier version of the standard titration problems you have encountered. Once you balance the equation it shows that we need 2 moles of HCl for every 1 mole of K_2CO_3. So first we must determine how many moles of K_2CO_3 we have and then how many moles of HCl we need:

$$1.068g\ \mathrm{K}_2CO_3 \times \frac{1mol}{138.204g} \times \frac{2mol\ \mathrm{HCl}}{1\mathrm{mol}\ \mathrm{K}_2CO_3} = 0.0155mol\ \mathrm{HCl}$$

Next we calculate molarity by dividing moles by volume:

$$M_{HCl} = \frac{0.0155mol}{0.02355L} = 0.66M$$

18C

Since phosphorus doesn't have a consistent oxidation number, we have to analyze the oxidation number for all the elements and set their sum equal to the charge of the ion:

P O_4 $^{3-}$

x + -2(4) = -3, so we can solve for x = +5

Thus the oxidation number for P is +5.

19B

Write the reaction and balance it.

$$Mg + O_2 + H_2 \rightarrow Mg(OH)_2$$

This is a Hess's law problem. We must use the partial reactions to create the overall reaction. Start with the furthest left reactant and determine which partial reaction contains that specimen. If only one partial reaction contains the specimen, it must be used to obtain the overall reaction. If that reactant is in more than 1 partial reaction, skip it and go onto the next reactant. Once a partial reaction is used, cross is off and do not use it again. Remember that any manipulations done the reaction must also be done to the $\Delta H°_{rxn}$. Below is the solution:

Reactants:	Products:	$\Delta H°_{rxn}$
Mg + (1/2) O_2	MgO	(1/2) (-452)
(1/2) O_2 + (1/2)H_2	OH	(1/2)(-771)
MgO + OH + (1/2) H_2	$Mg(OH)_2$	(-1/2) (321)
Sum: Mg + O_2 + H_2	$Mg(OH)_2$	-722 kJ

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20E

We know:

$$M_{solution} = \frac{moles_{solute}}{Liters_{solution}}$$

Find the moles of the solute:

$$moles_{solute} = MV = (1.25 * .01) = 0.0125\ moles\ Li_2CO_3$$

Now convert to mass of Li^+

$$0.0125\ moles\ Li_2CO_3 * \frac{2\ moles\ Li^+}{1\ mole\ Li_2CO_3} * \frac{6.941\ grams\ Li^+}{1\ mole\ Li^+} = 0.174\ g\ Li^+$$

21A

First determine the amount of energy required to heat the water, and then to boil it.

$$q = q_1 + q_2$$
$$q_1 = mC\Delta T = (216g)\left(4.184\frac{J}{g°C}\right)(100 - 25) = 67.78\ kJ$$
$$q_2 = m\Delta H_{vap} = (12\ mole)\left(40.66\frac{kJ}{mol}\right) = 487.92\ kJ$$
$$q_{sys} = 555.4\ kJ$$

This energy must have all come from the combustion of the propane, so we can say:

$$q_{sys} = -q_{surr} \quad and \quad q_{surr} = q_{comb} = -555.7\ kJ$$

$$q_{comb} = m\Delta H_{comb} \qquad (-555.7\ kJ) = m\left(-555.7\frac{kJ}{mole}\right)$$

So m=1 mole of propane which is 54g need to combust.

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22D
Write the reaction and determine which the limiting reagent. The maximum amount of Carbon dioxide produced is the amount produced by the limiting reagent.

Reaction:

$$C_9H_{20} + 14O_2 \rightarrow 9CO_2 + 10H_2O$$

$$320g\ C_9H_{20} * \frac{1mole\ C_9H_{20}}{128\ g\ C_9H_{20}} * \frac{9\ mole\ CO_2}{1\ mole\ C_9H_{20}} * \frac{44\ g\ CO_2}{1mole\ CO_2} = 990\ g\ CO_2$$

$$117g\ C_9H_{20} * \frac{1mole\ o_2}{32\ g\ O_2} * \frac{9\ mole\ CO_2}{14\ mole\ O_2} * \frac{44\ g\ CO_2}{1mole\ CO_2} = \ 103.4\ g\ CO_2$$

23B
Write the reaction:

$$C_6H_{12}O_6 + 12O_2 \rightarrow 6CO_2 + 6H_2O$$

Now plug the values into this equation.

$$\Delta H°_{rxn} = \Sigma H_{f\ products} - \Sigma H_{f\ reactants}$$

$$\Delta H°_{rxn} = [6(-400) + 6(-300)] - [-1200] = \ -3000\ kJ$$

The answer is positive because the question asked about the magnitude of energy, not the change of energy. Keep an eye out for this! Energy cannot have a negative magnitude.

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24C

First determine and balance the reaction. When dealing with the formation of a compound with no given reaction, assume the reactants as the constituent elements in their elemental form.

$$N_2 + 4H_2 \rightarrow 2NH_4$$

We want to determine the H_f of NH_3:

$$\Delta H°_{rxn} = \Sigma H_{f\ products} - \Sigma H_{f\ reactants}$$

The $\Sigma H_{f_{reac}} = 0$ because they are all in their elemental form. So $\Delta H°_{rxn} = \Sigma H_{f\ pro}$

We must determine $\Delta H°_{rxn}$ with the bond enthalpies.

$$\mathrm{H}_{\mathrm{f}_{\mathrm{NH}_3}} = \Delta H°_{rxn} = \Sigma(reactant\ bonds) - \Sigma(product\ bonds)$$

So draw the Lewis structures, determine how many and which types of bonds are present, then finally, plug the bond values into the equation above.

$\Sigma(reac\ bonds) = (1\ N \equiv N) + 4(1\ H - H) = 1(1 * 910) + 4(1 * 421) = 2594\ kJ$
$\Sigma(product\ bonds) = \ 2(4\ N - H) = 2(4 * 347) = 2776\ kJ$

Finally:

$\Delta H°_{rxn} = 2594 - 2776 = \ -182\ kJ$

The $\Delta H°_{rxn}$ pertains to 2 moles of NH_4 being produced (based on the balanced reaction.) To determine the H_f for just 1 mole of NH_4, divide the $\Delta H°_{rxn}$ by 2.

$$\mathrm{H_f} = \frac{\Delta H°_{rxn}}{2} = \ -\frac{182\ kJ}{2\ moles} = \ -91\frac{kJ}{mole}$$

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