P C.2 (pg 1 of 7) **Enthalpy of Rx**

Name____________________Per____

You will need to use the $\Delta H_{formation}$ Tables, as well as the last page which has Bond Enthalpies and specific heat capacity values.

1. A student was burning propane to heat water just as you used butane to heat water in our LAD. She measured 325 ml of water at 3.5ºC and used 5.3 g of propane, C_3H_8 to heat the water to 88.3ºC. Using the data in this problem, calculate the *experimental* ΔH for the combustion of propane reporting the result in kJ/1 mole propane. (Look back at your lab calculations from today – this problem is nearly the same calculation.)

2. Write a balanced chemical equation for the combustion of propane, C_3H_8

3. Use the balanced equation you just wrote in the problem above, the equation below, and the heat of formation values on the orange sheet to calculate the theoretical value for the heat of combustion of propane.

$$\Delta H_{rx} = \sum n\Delta H_f \text{ products} - \sum n\Delta H_f \text{ reactants}$$

- $\sum$ means calculate the sum (add together)
- n stands for the stoichiometric coefficient of the substance in the balanced equation.
- ΔH_f stands for the enthalpy value looked up on the list for formation reactions for the particular substance

4. Compare your *theoretical* value (as calculated in #3) with your *experimental* value (as calculated in #1) as a percent yield which would tell you a % efficiency of heat transferred from the flame to the water.

5. Using the *experimental* value calculated in problem #1, calculate the temperature to which you could expect to heat 250 g of water that starts at 10.0ºC if you burned 1.5 g of propane.

6. Again, assuming the experimental ΔH calculated in problem #1 what mass of propane would you need to heat 100.0 g of liquid water from 5.0ºC to 55.0ºC?

Remember to put your work on paper before you even begin to touch a calculator.

7. Using the theoretical value for the ΔH_{comb} butane (you've got the theoretical value on your LAD C.3, or you can calculate it using your Thermodynamic Tables), how much heat would be released if you burned 2.50 g of butane?

   a. When burning 2.50 g of butane, (assume 100% transfer of energy) how much water could you heat from 20ºC to 100ºC?

   b. If you were heating 500.0 g of water that started at 25.0ºC, how much butane (assume 100% efficiency) would you need to burn in order to raise the temperature from 25.0ºC to 85.0ºC?

   c. If you were heating 350.0 g of water that was ended up at 55ºC and you burned 1.0 g of butane (assume 100% efficiency), to what temperature did the water start at?

   d. If you burned only 1.00 g of butane and you realized the transfer of energy is probably only 50% efficient, how hot could you heat 350.0 g of water that started at 24ºC?

8. Write a reaction for the combustion of sugar, $C_6H_{12}O_6$

   a. Determine the theoretical heat of combustion for burning sugar.

   b. Using the theoretical $\Delta H_{combustion}$ for sugar. Determine how much energy will be released when you burn 25.0 g of sugar.

   A twist in this next problem -

   c. When you burn sugar, assume that only 60% of the heat from the combustion of the sugar will be absorbed by the water you are heating. If you would like to heat 325 g of water from 20.0ºC up to boiling temp, how many grams of sugar should you burn?

9. When calcium chloride dissolves in water it dissociates. There is an energy exchange associated with this process. Write the equation that represents the dissolving process of calcium chloride.

   a. Calculate the theoretical $\Delta H_{dissolving}$ for calcium chloride.

   b. Use the theoretical value from part b to determine the final temperature of 200.0 g of 30.0ºC water when 5.60 g of calcium chloride is dissolved in it.

10. Calculate the theoretical $\Delta H_{dissolve}$ for the dissolving of copper(II) chloride.

11. Determine the ΔH^o_{Rx} for this reaction: $H_2O_{(L)} \rightarrow H_2O_{(g)}$

12. Determine the ΔH^o_{Rx} for this reaction: $H_2O_{(g)} \rightarrow H_2O_{(L)}$

13. Determine the theoretical heat of combustion for burning sugar. Assume water is vapor.

14. How much energy is produced when 4.0 g of sugar, $C_6H_{12}O_6$, is burned?

15. When you burn 4.0 g of sugar, only 60% of the energy is used to heat the water, the rest of the heat is lost. How hot can you get 50.0 g of water starting at 15.0ºC

16. Using the reactions below,
$P_4 + 5\,O_2 \rightarrow P_4O_{10}$ $\Delta H_f = -2967$ kJ/mole of P_4O_{10}
$PCl_3 + Cl_2 \rightarrow PCl_5$ $\Delta H_{rx} = -84$ kJ/mole of PCl_5
$PCl_3 + \frac{1}{2}\,O_2 \rightarrow Cl_3PO$ $\Delta H_{rx} = -286$ kJ/mole Cl_3PO
$\frac{1}{4}\,P_4 + \frac{3}{2}\,Cl_2 \rightarrow PCl_3$ $\Delta H_f = -76.5$ kJ/ mole of PCl_3
Apply Hess's Law to determine the enthalpy for the reaction below.

$$P_4O_{10} + 6\,PCl_5 \rightarrow 10\,Cl_3PO$$

17. Determine the final temperature of 145 ml of water at 27.0ºC is mixed with 98.0 ml of 75.0º water.

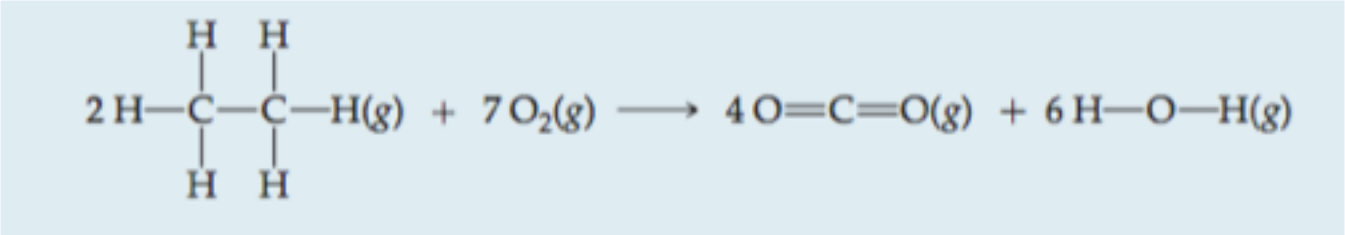
18. Determine the equilibrium temperature of 115 ml of acetone, C_3H_6O at 10.0ºC mixed with 115 ml of water, H_2O at 65.0ºC

19. Calculate how much heat is required to heat a 58.0 g piece of copper from −35.0ºC to 50ºC.

20. Use Bond Dissociation Energy values (on the last page of your Thermodynamic Table) to calculate ΔH_{rxn} for
$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$
Compute the ΔH_{rxn} using ΔH^o_f values. How do the values compare? Why might the values be slightly different.

21. Use Bond Dissociation Energy values (on the last page of your Thermodynamic Table) to calculate ΔH_{rxn} for
$CH_4 + 2\,H_2O \rightarrow CO_2 + 4\,H_2$
Compute the ΔH_{rxn} using ΔH^o_f values. How do the values compare? Why might the values be slightly different.

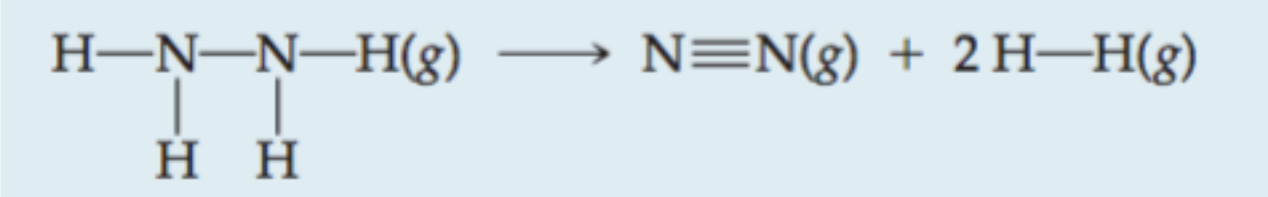
22. Use Bond Dissociation Energy values (on the last page of your Thermodynamic Table) to calculate ΔH_{rxn} for
$C_2H_6 + 7\,O_2 \rightarrow 4\,CO_2 + 6\,H_2O$



$$2\,H{-}CH_2{-}CH_2{-}H(g) + 7\,O_2(g) \longrightarrow 4\,O{=}C{=}O(g) + 6\,H{-}O{-}H(g)$$

Compute the ΔH_{rxn} using ΔH^o_f values. How do the values compare? Why might the values be slightly different.

23. Use Bond Dissociation Energy values (on the last page of your Thermodynamic Table) to calculate ΔH_{rxn} for
$N_2H_4 \rightarrow N_2 + 2\,H_2$



$$H{-}NH{-}NH{-}H(g) \longrightarrow N{\equiv}N(g) + 2\,H{-}H(g)$$

Compute the ΔH_{rxn} using ΔH^o_f values. How do the values compare? Why might the values be slightly different.

Enthalpy of a Combustion Rx

Name____________________Per____

1. In this question we will assume that the heat gained by water = heat lost by propane as it is burning

$$c_{water} \times mass_{water} \times \Delta T_{water} = -\Delta H_{permolepropaneburned} \times moles_{propaneburned}$$

$$\left(\frac{4.18J}{1g1^{\circ}}\right)(325g)(88.3^{\circ}-3.5^{\circ}) = -\Delta H_{permolepropaneburned}\left(5.3g\times\frac{1mol}{44g}\right)$$

$\Delta H_{per\ mole\ of\ propane\ burned}$ = −956,384 J/mole thus −956 kJ/mole

2. Write out the balanced combustion equation: $C_3H_{8(g)} + 5\ O_{2(g)} \rightarrow 3\ CO_{2(g)} + 4\ H_2O_{(g)}$

3. Look up the ΔH_f for each of the compounds in the equation above.
$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ $\Delta H_f = -394$ kJ/mole of CO_2
$H_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(g)}$ $\Delta H_f = -242$ kJ/mole of $H_2O_{(g)}$
$3\ C_{(s)} + 4\ H_{2(g)} \rightarrow C_3H_{8(g)}$ $\Delta H_f = -104$ kJ/mole of $C_3H_{8(g)}$

$$\Delta H_{combustion} = \sum n\Delta H_{fusionproducts} - \sum n\Delta H_{fusionreac\tan ts}$$

(n refers to the coefficient for each product or reactant in the balanced equation)

$\Delta H_{combustion}$ = [3(−394) + 4(−242)] – [1(−104) + 5(0)] = −2046 kJ/1 mole $C_3H_{8(g)}$

4. $\left(\frac{956kJ}{2046kJ}\right)\times 100 = 47\% efficiency$

5. $\left(\frac{4.18J}{1g1^{\circ}}\right)(250g)(x^{\circ}-10^{\circ}) = 956kJ\left(\frac{1000J}{1kJ}\right)\left(1.5g\times\frac{1mol}{44g}\right)$ Solve for x, the water can be heated up to 41.2ºC

6. $\left(\frac{4.18J}{1g1^{\circ}}\right)(100g)(55^{\circ}-5^{\circ}) = 956kJ\left(\frac{1000J}{1kJ}\right)\left(xg\times\frac{1mol}{44g}\right)$ 0.96 g of propane need to be burned

7. Remember the calculation for the theoretical ΔH_{comb} for butane that was done in LAD 10.3:
$C_4H_{10} + 6.5\ O_2 \rightarrow 4CO_4 + 5H_2O$

$\Delta H_{rx} = \Sigma n\Delta H_{f\ products} - \Sigma n\Delta H_{f\ reactants}$ [4(−394) + 5(−242)] – [1(−125) + 6.5(0)] = −2661 kJ/1mole C_4H_{10}

$$\left(\frac{2661kJ}{1mol}\right)(2.50g)\left(\frac{1mol}{58g}\right) = 115kJ$$

(You can leave off the negative sign since it simply tells us that is energy coming out.)

a. Use the heat energy from the last problem - change to Joules to match the units on c_{water}
$q_{butane} = c_{water} \times m_{water} \times \Delta T_{water}$ (remember to make the energy units match on both sides of the equation.)

$\left(\frac{1000J}{1kJ}\right)\left(\frac{2661kJ}{1mol}\right)(2.50g)\left(\frac{1mol}{58g}\right) = \left(\frac{4.18J}{1g1^{\circ}C}\right)(m_w)(100^{\circ}-20^{\circ})$ $mass_{water}$ = **343 g of water**

b. Heat lost by butane = heat gained by water
$\Delta H_{combustion\ butane} \times \mathbf{mass_{butane}} \times 1/MM_{butane} = c_{water} \times m_{water} \times \Delta T_{water}$

$\left(\frac{1000J}{1kJ}\right)\left(\frac{2661kJ}{1mol}\right)(m)\left(\frac{1mol}{58g}\right) = \left(\frac{4.18J}{1g1^{\circ}C}\right)(500g)(85^{\circ}-25^{\circ})$ **$\mathbf{mass_{butane}}$ = 2.74 g**

c. Heat lost by butane = heat gained by water

$\Delta H_{\text{combustion butane}} \times \text{mass}_{\text{butane}} \times 1/\text{MM}_{\text{butane}} = c_{\text{water}} \times m_{\text{water}} \times \Delta T_{\text{water}}$

$\left(\frac{1000J}{1kJ}\right)\left(\frac{2661kJ}{1mol}\right)(1g)\left(\frac{1mol}{58g}\right)=\left(\frac{4.18J}{1g1^{\circ}C}\right)(350g)(\Delta T)$ $\Delta T_{\text{water}} = 31.3^{\circ}C$ thus $55^{\circ} - 31.3^{\circ} = 23.7^{\circ}C$

Of course you could put the $T_{final} - T_{initial}$ into the equation and solve for $T_{initial}$. I chose not to, so I could avoid distribution.

d. $\left(\frac{1000J}{1kJ}\right)\left(\frac{2661kJ}{1mol}\right)(1g)\left(\frac{1mol}{58g}\right)\times 0.05(efficiency)=\left(\frac{4.18J}{1g1^{\circ}C}\right)(350g)(\Delta T)$ $\Delta T = 15.7^{\circ}$ thus $15,7^{\circ} + 24^{\circ} = 39.7^{\circ}C$ will be the final temp, *notice that this is not nearly as hot as the 55ºC final temp in problem (c), because when the efficiency is accounted for, the temperature of course can not go up so high.*

8. First you need to write a balanced equation: $C_6H_{12}O_{6(s)} + 6\ O_{2(g)} \rightarrow 6\ CO_{2(g)} + 6\ H_2O_{(g)}$

a. Use the equation: $\Sigma\ \Delta H_{f\ \text{products}} - \Sigma\ \Delta H_{f\ \text{reactants}} =$ theoretical $\Delta H_{\text{combustion}}$

to calculate theoretical $\Delta H_{\text{combustion}} = [6(-394) + 6(-242)] - [(-1273) + 6(0)] = -2543$ kJ/mole of sugar

b. Use the equation: $\Delta H_{\text{combustion sugar}} \times \text{mass}_{\text{sugar}} \times 1/\text{MM}_{\text{sugar}} = q$

$$\left(-\frac{2543kJ}{1mol}\right)(25g)\left(\frac{1mol}{180g}\right)=-353\ kJ\ \textit{released}$$

(The negative sign of course does not mean negative energy, it simply means exothermic, so use the absolute value in part b)

c. In this problem it is important to realize that 60% of the energy lost by the butane burning will be equal to the energy that will be absorbed by the water.

$\Delta H_{\text{combustion sugar}} \times 60\% \times \text{mass}_{\text{sugar}} \times 1/\text{MM}_{\text{sugar}} = c_{\text{water}} \times m_{\text{water}} \times \Delta T_{\text{water}}$

$$-\left(-\frac{2543kJ}{1mol}\right)\left(\frac{1000J}{1kJ}\right)(0.60(efficiency))(m)\left(\frac{1mol}{180g}\right)=\left(\frac{4.18J}{1g1^{\circ}C}\right)(325g)(100^{\circ}-20^{\circ})$$

m = 12.8 g of sugar must be burned to provide enough heat.

9. First write an equation to represent the dissolving of calcium chloride: $CaCl_2 \rightarrow Ca^{2+}_{(aq)} + 2\ Cl^{-}_{(aq)}$

a. Then calculate the energy needed for this process using the equation: $\Delta H_{\text{dissolve}} = \Sigma\ \Delta H_{f\ \text{products}} - \Sigma\ \Delta H_{f\ \text{reactants}}$
theoretical $\Delta H_{\text{dissolve}} = [(-543) + 2(-167)] - [-796] = -81$ kJ/mole of $CaCl_2$ (the negative sign tells you that energy is given off. This tells you that the temp of the water is going to go up.)

When you use the −81 kJ/mole in the next part of the problem, use the absolute value in the equation because really there is no such thing as negative energy. The sign is just a convention to tell us energy in or out.

b. For this problem you need to realize that the heat lost by the dissolving calcium chloride will equal the heat gained by the water

$q_{dissolve} = q_{water}$

theoretical $\Delta H_{\text{dissolve}} \times \text{mass}_{\text{calcium chloride}} \times 1\text{mol}/\text{MM}_{\text{calcium chloride}} = c_{\text{water}} \times \text{mass}_{\text{water}} \times \Delta T_{\text{water}}$

$$-\left(-\frac{81kJ}{1mol}\right)\left(\frac{1000J}{1kJ}\right)(5.6g)\left(\frac{1mol}{111g}\right)=\left(\frac{4.18J}{1g1^{\circ}C}\right)(200g)(\Delta T)$$

(Note that the $\Delta H_{dissolve}$ was in KJ and thus converted to Joules to match the fact that c is in Joules.)

Solve: $\Delta T_{\text{water}} = 4.9^{\circ}$, and since it is exothermic, the water will heat $30^{\circ} + 4.9^{\circ} = 34.9^{\circ}C$

10. First write an equation: $CuCl_2 \rightarrow Cu^{+2} + 2\ Cl^{-}$

$\Delta H_{\text{comb}} = \Sigma\Delta H_{\text{prod}} - \Sigma\Delta H_{\text{react}}$

$\Delta H_{\text{rx}} = [2(-167) + 65] - (-206) = -63$kJ/mole

11. $\Delta H_{\text{comb}} = \Sigma\Delta H_{\text{prod}} - \Sigma\Delta H_{\text{react}}$

$(-242) - (-286) = +44$ kJ/1 mole H_2O

12. Since the equation is reversed, the sign changes = −44 kJ/1 mole H_2O

13. Remember, you need a balanced equation: $C_6H_{12}O_{6(s)} + 6\ O_{2(g)} \rightarrow 6CO_{2(g)} + 6\ H_2O_{(g)}$
$\Delta H_{comb} = \Sigma \Delta H_{prod} - \Sigma \Delta H_{react}$
$\Delta H_{comb} = [6(-394) + 6(-242)] - [-1273 + 6(0)] = -2543$ kJ/mole

14. Calculate the energy for 4 g >> 4 g × 1mole/180g × 2543 kJ/mole = 56.5 kJ produced

15. 4.0 g of sugar would in theory produce 56.5 kJ of energy, however, only 60% of the energy reaches the water.
56.5 kJ × 0.60 = 33.9 kJ >> 33,900 J = 4.18 × 50 × ΔT solve and ΔT = 162ºC
ΔT = 162ºC >> the water would reach 162ºC and begin to boil, never going above 100ºC

16. $\Delta H_{rx} = -613$ kJ/mole of P_4O_{10}

Reverse	$P_4O_{10(s)} \rightarrow P_{4(s)} + 5\ O_{2(g)} \quad \Delta H = +2967$
Multiply by 6 and reverse	$6\ PCl_{5(g)} \rightarrow 6\ PCl_{3(g)} + 6\ Cl_{2(g)} \quad \Delta H = +504$
Multiply by 10	$10\ PCl_{3(g)} + 5\ O_2 \rightarrow 10\ Cl_3PO_{(g)} \quad \Delta H = -2860$
multiply the ΔH by 4 since it's listed per mole	$P_{4(s)} + 6\ Cl_{2(g)} \rightarrow 4\ PCl_{3(g)} \quad \Delta H = -76.5 \times 4$
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	$P_4O_{10(s)} + 6\ PCl_{5(g)} \rightarrow 10\ Cl_3PO_{(g)} \quad \Delta H = +305$ kJ

17. $-\left(\frac{4.18J}{1g1^{o}}\right)(145g)(x^{o}-27^{o})=\left(\frac{4.18J}{1g1^{o}}\right)(98g)(x^{o}-75^{o})$ x = T_f = 46.4ºC

18. Remember that 115 ml of acetone does not weigh 115 g $115ml \times \frac{0.79g}{1ml} = 91g$

$$-\left(\frac{2.17J}{1g1^{o}}\right)(91g)(x^{o}-10^{o})=\left(\frac{4.18J}{1g1^{o}}\right)(115g)(x^{o}-65^{o})$$

x = T_f = 49ºC

19. $\left(\frac{0.386J}{1g1^{o}}\right)(58g)(85^{o}) = 1902J\ of\ heat\ Required$

20. −113 kJ/mol*rxn*

21. −1700 kJ/mol*rxn*

22. −2831 kJ/mol*rxn*

23. −86 kJ/mol*rxn*

24.