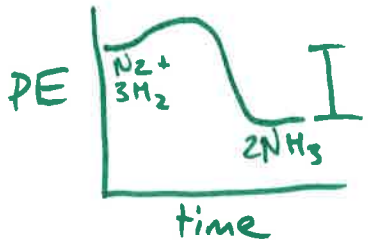
Heat Review Groupwork! Name: KEY

1) Sketch an **energy diagram** for the following reaction. Label the ΔH in the diagram and assign it a positive or negative sign:

$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)} + 92\ kJ$



ΔH = -92 kJ exothermic

2) If you reacted 1.00 grams of H_2 gas in the above reaction, how many kJ of energy could you theoretically produce?

$1.00\ g\ H_2 \left(\frac{1\ mol\ H_2}{2.02\ g}\right)\left(\frac{92\ kJ}{3\ mol\ H_2}\right) = 15.1815$ → 15.2 kJ

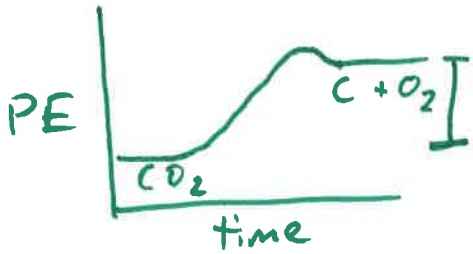
3) Imagine you performed the reaction described in (2) above and only produced 12.2 kJ of energy in the lab, calculate your percent yield and percent error.

% Yield = $\frac{12.2\ kJ}{15.2\ kJ} = 80.3\%$ Yield

% Error = $\frac{|15.2 - 12.2|}{15.2} \times 100 =$ 19.7% error

4) Sketch an energy diagram for the following reaction. Label the ΔH in the diagram and assign it a positive or negative sign:

$CO_{2(g)} + 394\ kJ \rightarrow C_{(s)} + O_{2(g)}$



ΔH = +394 kJ endo

5) If you needed to produce exactly 5.00 grams of carbon, how many kJ of energy would you need?

$5.00\ g\ C \left(\frac{1\ mol\ C}{12.01\ g}\right)\left(\frac{394\ kJ}{1\ mol\ C}\right) = 164.029$ → 164 kJ

6) How many grams of CO_2 could you decompose if you only had 100. kJ of energy available?

$100.\ kJ \left(\frac{1\ mol\ CO_2}{394\ kJ}\right)\left(\frac{44.01\ g}{1\ mol\ CO_2}\right) = 11.17$ → 11.2 g CO_2

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7) Write an equation to show the **formation** of one mole of $NH_4NO_{3(s)}$. Be sure to include the enthalpy (kJ) in the reaction.

$N_{2(g)} + 2H_{2(g)} + \frac{3}{2}O_{2(g)} \rightarrow NH_4NO_{3(s)} + 365.6\ kJ$

7b) If you were to form 1.00 gram of NH_4NO_3, how many kJ of energy would be released?

$1.00\,g\ NH_4NO_3 \left(\frac{1\ mol\ NH_4NO_3}{80.06\,g}\right)\left(\frac{365.6\ kJ}{1\ mol\ NH_4NO_3}\right) = 4.5665$ **4.57 kJ**

8) Write an equation to show the **combustion** of one mole of $C_8H_{18(l)}$. Be sure to include the enthalpy (kJ) in the reaction (look it up).

$C_8H_{18(l)} + \frac{25}{2}O_{2(g)} \rightarrow 8CO_{2(g)} + 9H_2O_{(g)} + 5{,}470\ kJ$

8b) If you were to combust 2.00 grams of octane ($C_8H_{18(l)}$), how many kJ of energy could you theoretically produce?

$2.00\,g\ C_8H_{18} \left(\frac{1\ mol\ C_8H_{18}}{114.26\,g}\right)\left(\frac{5470\ kJ}{1\ mol\ C_8H_{18}}\right) = 95.746$ **95.7 kJ**

9) Use your **bond enthalpy WS** to answer these questions. Calculate the ΔH of the following. Make sure to draw the correct Lewis structures first.

a) $I_{2(s)} + H_{2(g)} \rightarrow 2HI_{(g)}$

I-I H-H | H-I H-I

1(151 kJ) + 436 kJ | 2(299 kJ)

+587 kJ | −598 kJ

ΔH = −11 kJ

b) $2NCl_{3(s)} \rightarrow N_{2(g)} + 3Cl_{2(g)}$

2(Cl–N(Cl)–Cl) | N≡N Cl-Cl Cl-Cl Cl-Cl

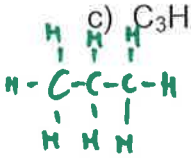
6(N-Cl) = | 941 kJ + 3(242 kJ)

6(200 kJ) = | −1667 kJ

+1200 kJ

ΔH = −467 kJ

c) $C_3H_{8(g)} + 5O_2 \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$



8(413 kJ) 2(348) + 5(495 kJ) | 3(O=C=O) + 4(H–O–H)

6475 kJ | 6(C=O) 6(799) + 8(O-H) 8(463)

| −8498 kJ

+ [redacted] kJ

ΔH = − [redacted] kJ

−2,023 kJ

d) $O_{2(g)} + 2F_{2(g)} \rightarrow 2OF_{2(g)}$

O=O F-F F-F | 

495 + 2(155) | 4(190)

+805 kJ | −760 kJ

ΔH = +45 kJ