

KEY

Recitation Week 11&12

(Heat Transfer, Calorimetry, Hess's Law)

PART A

A 500.0 gram sample of aluminum is initially at 25.0 °C. It absorbs 41.40 kJ of heat from its surroundings.

1. What type of change is this for the metal? Exothermic or Endothermic? Why?
2. What is the final temperature of the metal, in °C? (specific heat = 0.9930 J g⁻¹ °C⁻¹ for aluminum)
3. What is the specific heat capacity of a material? How is it linked to heat and temperature change?
4. What is the Heat Capacity of an object? How is it linked to heat and temperature change?
5. How is Heat Capacity of an object different to the specific heat capacity of an element or compound?
6. Can we obtain a Heat Capacity of an object made up of multiple compounds?

Aluminum at 25.0 °C 500g It absorbs 41.40 kJ of heat.

1. The metal takes up heat from its surroundings. Therefore this is an endothermic change.

$$2. \quad q = MC \Delta T \quad M = 500g \quad C = 0.9930 \text{ J g}^{-1} \text{ } ^\circ\text{C}^{-1}$$

$$q = +41.40 \text{ kJ} = 41400 \text{ J}$$

$$\Delta T = \frac{q}{MC} = \frac{41400 \text{ J}}{500g \cdot 0.9930 \text{ J g}^{-1} \text{ } ^\circ\text{C}^{-1}}$$

$$\Delta T = 83.38 ^\circ\text{C} = T_{\text{final}} - T_{\text{initial}}$$

$$T_{\text{initial}} = 25.0 ^\circ\text{C} \Rightarrow T_{\text{final}} = 83.38 ^\circ\text{C} + 25.0$$

$$= \underline{\underline{108.4 ^\circ\text{C}}}$$

3. The specific heat capacity is the heat energy required to raise the temperature of 1 g of material by 1°C or 1K .
4. The heat capacity of an object is the heat energy required to raise the temperature OF THAT OBJECT by 1°C or 1K .

C = heat capacity

$$q = C\Delta T.$$

5. The Heat capacity, C , has no mass associated with it, unlike the specific heat capacity, c .
6. Yes, we can calculate the heat capacity of ANY object independent of what it is comprised, as long as we know the heat is must absorb in order to have its temperature raised by 1°C . Its units are $\text{J}/^{\circ}\text{C}$.

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Recitation Week 9 11, 12

(Heat transfer, Calorimetry, Hess' law)

PART B

When pure sodium hydroxide is dissolved in water, heat is evolved. In a laboratory experiment to measure the **molar heat of solution of sodium hydroxide**, the following procedure was followed.

To a calorimeter containing 400. g of water at 20.00 °C, 12.65 g of NaOH, also at 20.00 °C was added. The temperature change, which was monitored by a digital probe with negligible heat capacity, stopped when it reached a temperature of 29.35 °C. If the specific heat of the mixture is 4.184 J g⁻¹ °C⁻¹, and the small heat capacity of the calorimeter is ignored, what is the heat evolved, per mole of sodium hydroxide?

1. Write down all the variables that you are given.
2. How much heat has been taken up by the water (the surroundings in this example)?
3. Explain why q for the water is positive.
4. Write out a chemical equation for the solvation of NaOH.
5. On what side of the arrow do we find the heat?
6. Is the heat best described as a product or a reactant?
7. Has the NaOH lost chemical energy or gained chemical energy in the solvation process?
8. What is the ΔH for the 12.65 g of NaOH?
9. When given the mass of a reagent (here NaOH), what should you do with this mass 99% of the time?
10. From your calculations, write down the enthalpy change associated with 1 mole of NaOH. This is called the Enthalpy of solution.
11. Is this an exothermic change or an endothermic change for the NaOH?
12. What is the concentration of NaOH in this solution, assuming there is no volume change upon adding the solute?

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1. 400g water at 20°C (20.00°C)
12.65g NaOH at 20°C is added.

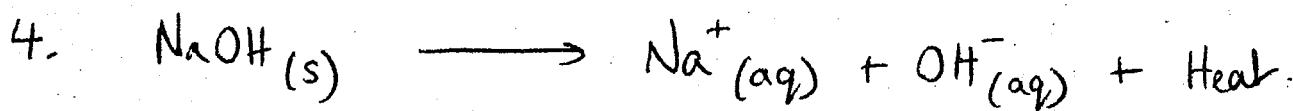
$$\Delta T = \underset{\text{final } \uparrow}{29.35^\circ\text{C}} - \underset{\text{initial } \uparrow}{20.00^\circ\text{C}} = 9.35^\circ\text{C}$$

$$2. \quad q = mc\Delta T = \left(\begin{array}{c} 400\text{g} \\ + 12.65\text{g} \end{array} \right) \times 4.184 \text{ Jg}^{-1} \text{ } ^\circ\text{C}^{-1} \times 9.35^\circ\text{C}$$

in this case we might want to account for
the additional mass of solute added

$$q = 16143 \text{ J}$$

3. q is positive because the temperature of the mixture increases
The thermal energy of the mixture increases.



5+6. Heat is a product here, as shown by experiment.

7. $\text{NaOH}_{(s)}$ has lost energy. $\text{Na}^+_{(aq)}$ and $\text{OH}^-_{(aq)}$ are lower in energy than the "Reactant", $\text{NaOH}_{(s)}$

$$8. \quad \Delta H = -q_{\text{surroundings}} = -16143 \text{ J}$$

$$9. \quad 12.65\text{g NaOH} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g}} = 0.31625 \text{ mol}$$

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10. 0.31625 mol yields 16143 J of heat.

\Rightarrow ~~0.31625 mol~~ 16143 J per 0.31625 mol.

$\Rightarrow \frac{16143 \text{ J}}{0.31625 \text{ mol}} = 51045 \frac{\text{J}}{\text{mol}}$ of heat.

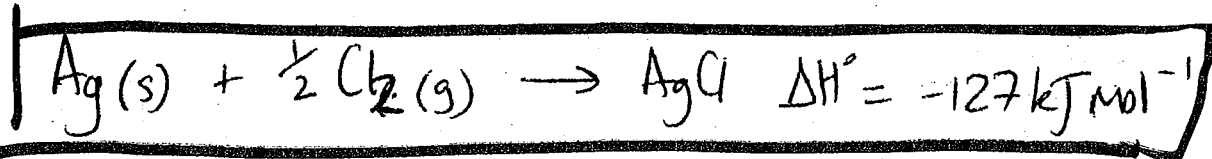
$$\Delta H = -q = -\frac{51045 \text{ J}}{\text{mol}} = -51.05 \text{ kJ/mol}$$

11. ΔH is negative \Rightarrow This is an exothermic change.

$$12. \frac{0.31625 \text{ mol}}{\frac{\text{mL}}{412.65 \text{ mL}}} = \frac{0.31625 \text{ mol}}{0.41265 \text{ L}} = \boxed{0.7664 \text{ mol L}^{-1}}$$

Assuming density of water is 1 g/mL.

$$\begin{aligned} & (400.00 \text{ g}_{\text{H}_2\text{O}} + 12.65 \text{ g}_{\text{NaOH}}) \times \frac{1 \text{ mL}}{1.000 \text{ g}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = \\ & = (412.65 \text{ mL}) = 0.41265 \text{ L} \end{aligned}$$

Part C

- Heat is produced. ΔH is negative. The reaction is EXOTHERMIC.
- $$50.5 \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.87 \text{ g Ag}} = 0.468 \text{ mol.}$$

$$-127 \text{ kJ mol}^{-1} \times 0.468 \text{ mol} = -59.5 \text{ kJ. heat}$$

$$\Rightarrow 59.5 \text{ kJ of heat is transferred.}$$
- If 25.8 kJ of heat were to be released from this reaction then clearly we must have started with less than one mole of Silver.

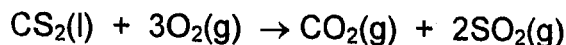
$$\Rightarrow 25.8 \text{ kJ} \times \frac{1 \text{ mol}}{-127 \text{ kJ}} = 0.203 \text{ mol Ag.}$$

$$0.203 \text{ mol Ag} \times \frac{1 \text{ mol AgCl}}{1 \text{ mol Ag}} = 0.203 \text{ mol AgCl.}$$
- To DECOMPOSE 102.5g AgCl

$$102.5 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.3 \text{ g AgCl}} = 0.7152 \text{ mol AgCl.}$$

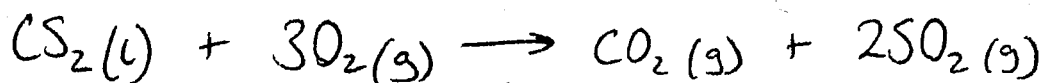
$$0.7152 \text{ mol AgCl} \times +127 \text{ kJ mol}^{-1} \text{ AgCl} = \underline{\underline{90.8 \text{ kJ}}}$$

PART 3



Species (state)	ΔH_f° in kJ/mole
$\text{CS}_2(\text{l})$	+ 87.9
$\text{CO}_2(\text{g})$	- 393.5
$\text{SO}_2(\text{g})$	-296.8

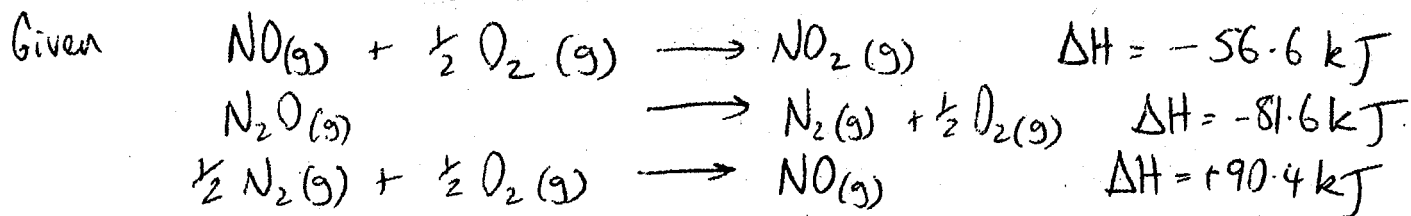
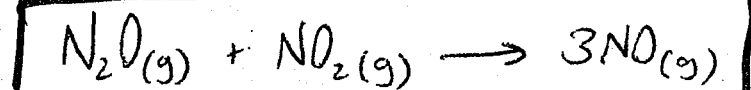
Determine the standard-state enthalpy of this reaction
Is this reaction exothermic or endothermic?



$$\begin{aligned} \Delta H_{\text{rxn}} &= \sum \Delta H_f^\circ (\text{products}) - \sum \Delta H_f^\circ (\text{reactants}) \\ &= \left(-393.5 \frac{\text{kJ}}{\text{mol}} \times 1 \text{ mol} \right) + \left(-296.8 \frac{\text{kJ}}{\text{mol}} \times 2 \text{ mol} \right) - \left(87.9 \frac{\text{kJ}}{\text{mol}} \times 1 \text{ mol} \right) \\ &= -1075.0 \text{ kJ/mol}^{-1} \text{ reaction) (EXOTHERMIC)} \end{aligned}$$

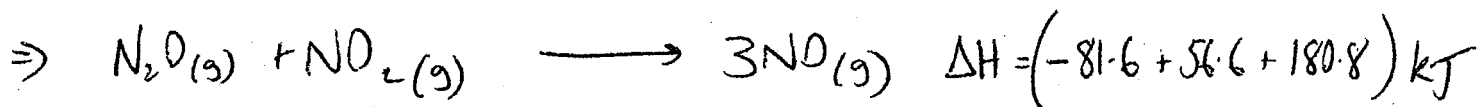
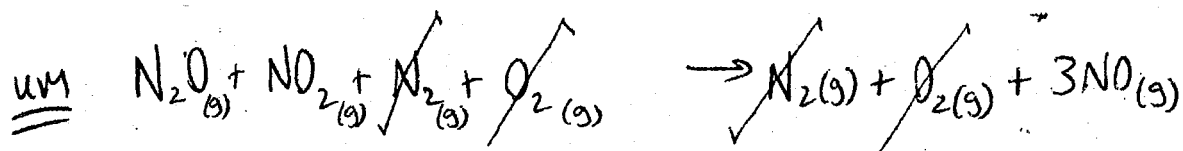
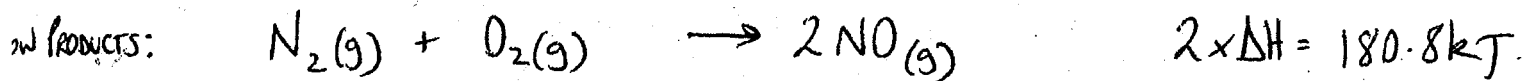
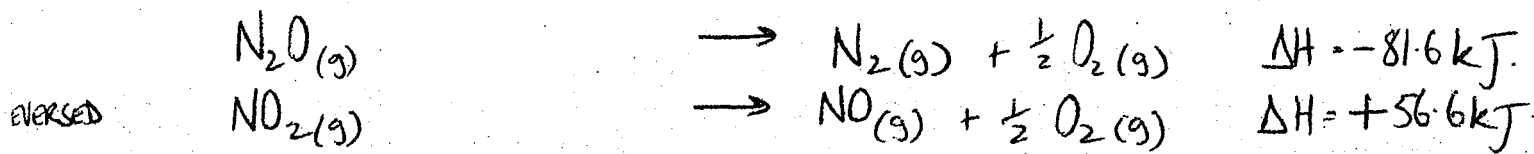
We say this is -1075.0 kJ per mole of reaction.

PART E

Calculate $\Delta H^\circ_{\text{rxn}}$ 

Try to sum these equations in such a way as to obtain the original.

Try to obtain the reactants first:



$$\Delta H = +155.8 \text{ kJ}$$

ENDOTHERMIC

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Recitation Week 11, 12

(Heat transfer, Calorimetry, Hess' law)

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PART F (an additional practice problem):

A chemical compound has a molecular weight of 89.05 g/mole. 1.400 grams of this compound underwent complete combustion under constant pressure conditions in a special calorimeter. This calorimeter had a heat capacity of 2980 J °C⁻¹ (Note that the calorimeter was made of a metal shell, a water "substitute" - a special oil, and a thermocouple). The temperature went up by 11.95 degrees. Calculate the molar heat of combustion of the compound.

1. How much heat was given off by the combustion reaction, based on the temperature rise of the calorimeter?
2. How many moles of compound were used up to provide this heat?
3. How much heat would one mole of compound provide?
4. Is the combustion of this compound exothermic or endothermic?
5. Would the molar enthalpy of combustion for this compound be a negative quantity or a positive quantity?

Chemical compound of molecular weight 89.05 g/mole
We have 1.400g
 $C_{\text{calorimeter}} = 2980 \text{ J/}^\circ\text{C}$

$$\Delta T = 11.95^\circ\text{C}$$

$$1. q_{\text{calorimeter}} = C \Delta T = 2980 \text{ J/}^\circ\text{C} \times 11.95^\circ\text{C} = 35611 \text{ J}$$

$$2. \text{ moles compound} = 1.400\text{g} \times \frac{1 \text{ mol}}{89.05\text{g}} = 0.01572 \text{ mol}$$

3. If 0.01572 mol provides 35611 J we can write this as:

$$35611 \text{ J per } 0.01572 \text{ mol or } \frac{35611 \text{ J}}{0.01572 \text{ mol}} = 2265114 \text{ J/mol} = \underline{\underline{2265 \text{ kJ/mol}}}$$

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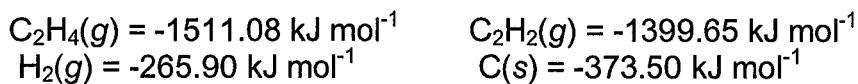
Recitation Week 11, 12

(Heat transfer, Calorimetry, Hess' law)

PART G (an additional practice problem):

Complete combustion of hydrocarbons, or compounds with C, H, and O as the only elements, gives CO_2 and H_2O as the only products. If carried out under standard conditions, the CO_2 is a gas while the H_2O is a liquid.

A professor suggests the following standard enthalpies of **combustion** (although a bright student later recognizes that these values were arbitrarily chosen and were not the true values recognized in the literature):

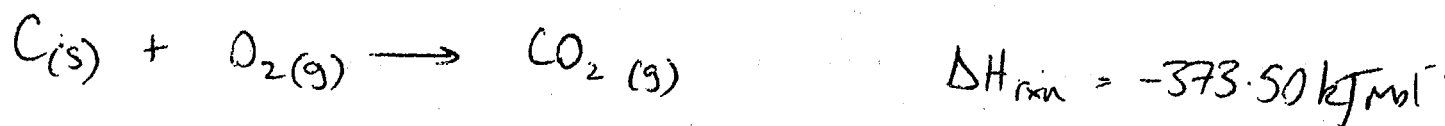
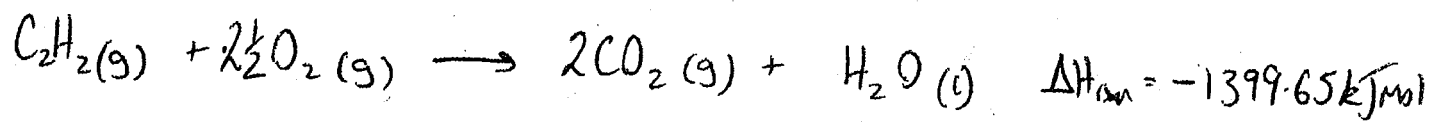
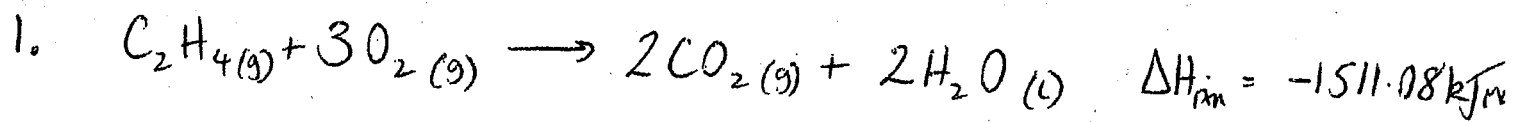
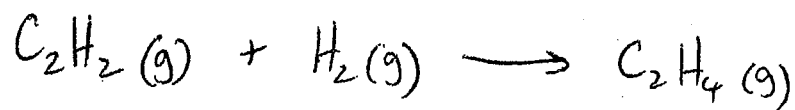


Calculate the standard enthalpy of reaction for the process, $\text{C}_2\text{H}_2(g) + \text{H}_2(g) \rightarrow \text{C}_2\text{H}_4(g)$

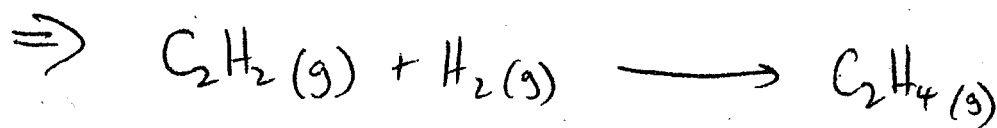
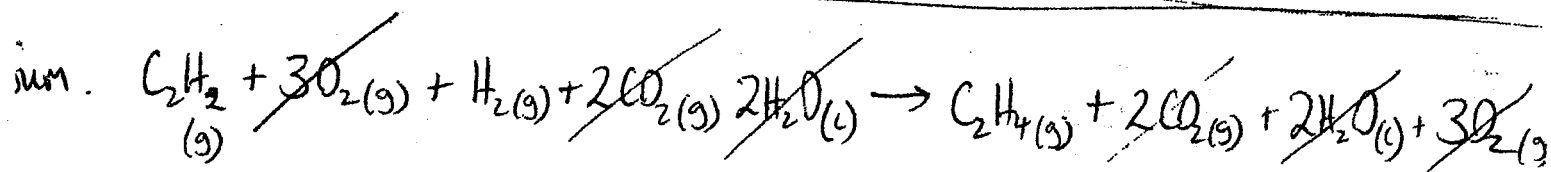
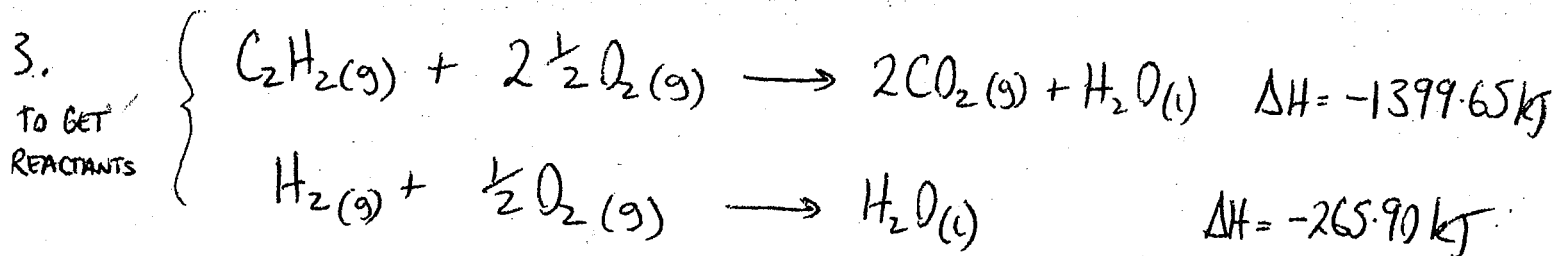
1. Write down the chemical equations for the combustions of each of the four chemicals given above.
2. Manipulate these equations so that by adding these reactions together you arrive at the reaction for the addition of hydrogen to Ethyne, $\text{C}_2\text{H}_2(g)$ to form Ethene, $\text{C}_2\text{H}_4(g)$.
 - a. Remember that you may reverse an equation so long as the sign of the associated enthalpy change is changed
 - b. Remember that you may multiply an equation by any number as long as you multiply the associated enthalpy change by that same number.
3. Hence calculate the enthalpy for the addition reaction described above.

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part 6

2. We are allowed to use fractions as coefficients because by definition these are thermochemical reactions - They always involve molar quantities and never involve individual molecules or atoms.



$$\Delta H = -1399.65 \text{ kJ} + (-285.90 \text{ kJ}) + 1511.08 \text{ kJ} \\ = -154.47 \text{ kJ/mol rxn.}$$