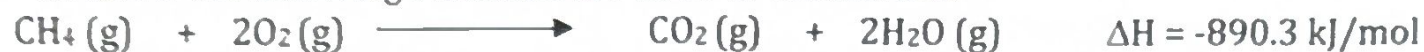


Practice/Thermochemistry-Thermochemical Equations and Stoichiometry

DO NOT SUBMIT

- 1) Indicate if the following reactions are endo or exothermic.



exo



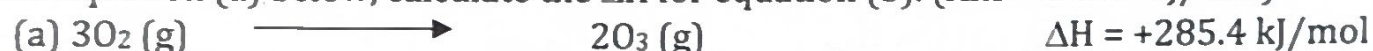
endo

- 2) Given equation (a) below, calculate the ΔH for equation (b). (Ans: -26.48 kJ/mol)



$\Delta H \text{ rev } \times \div 2$

- 3) Given equation (a) below, calculate the ΔH for equation (b). (Ans: +142.7 kJ/mol)



$\Delta H \div 2$

- 4) Write the thermochemical equation that expresses that at 0 °C ice melts by absorbing 334 J of heat per gram. (Ans: +6.02 kJ)



$$334 \frac{\text{J}}{\text{g}} \times \frac{18 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = \boxed{6.02 \text{ kJ}}$$

- 5) The complete combustion of liquid octane, C_8H_{18} , produces carbon dioxide and water at 25.0 °C and at constant pressure, it gives 47.9 kJ of heat per gram of octane. Write the thermochemical equation to show this information. (Ans: $-1.09 \times 10^4 \text{ kJ}$)

$$\text{C}_8\text{H}_{18} = 114.2 \text{ g/mol}$$

$$\frac{-47.9 \text{ kJ}}{1 \text{ g}} \text{C}_8\text{H}_{18} \times \frac{114.2 \text{ g}}{1 \text{ mol}} \text{C}_8\text{H}_{18} = -5.47 \times 10^3 \text{ kJ/mol}$$



6) Answer the following two questions using the equation given below.



a) What is the enthalpy change associated with the formation of 5.67 mol HCl gas in the following reaction? (Ans: -523 kJ)

$$5.67 \text{ mol HCl} \times \frac{-184.6 \text{ kJ}}{2 \text{ mol}} = \boxed{-523 \text{ kJ}}$$

b) What is the enthalpy change when 12.8 g of H_2 gas reacts with excess Cl_2 gas to form HCl? (Ans: $-1.17 \times 10^3 \text{ kJ}$)

$$12.8 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{-184.6 \text{ kJ}}{2 \text{ mol}} = \boxed{-1.17 \times 10^3 \text{ kJ}}$$

7) The reaction of sodium peroxide and water is a source of oxygen gas.



How many kilojoules of heat are evolved in a reaction of 226 g Na_2O_2 with 98.5 g H_2O ? (Ans: -416 kJ)

$$226 \text{ g Na}_2\text{O}_2 \times \frac{1 \text{ mol Na}_2\text{O}_2}{77.98 \text{ g}} \times \frac{1 \text{ mol O}_2}{2 \text{ mol Na}_2\text{O}_2} = 1.45 \text{ mol O}_2 \quad \boxed{\text{LR}}$$

$$98.5 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18 \text{ g}} \times \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2\text{O}} = 2.73 \text{ mol O}_2$$

$$\frac{-287 \text{ kJ}}{2 \text{ mol Na}_2\text{O}_2} \times 226 \text{ g Na}_2\text{O}_2 \times \frac{1 \text{ mol Na}_2\text{O}_2}{77.98 \text{ g}} = \boxed{-416 \text{ kJ}}$$

Practice/Thermochemistry - Calorimetry and Enthalpy of Reaction DO NOT SUBMIT

Show all the work for the calculations and give the answers in the correct significant figures.

- 8) Calculate the heat capacity of a piece of iron if a temperature rises from 18 to 69 °C requires 672 J of heat. (Ans: 13 J/°C)

$$\Delta T = 69 - 18 = 51^{\circ}\text{C}$$

$$q = c \Delta T$$

$$c = \frac{q}{\Delta T} = \frac{672}{51} = \boxed{13 \text{ J/}^{\circ}\text{C}}$$

- 9) How much heat in kJ is required to raise the temperature of

- a) A 320.0 g of water from 15.0 to 96.0 °C? (sp heat water = 4.18 J/g °C) (Ans: 108 kJ)

$$q = ms \Delta T = 320 \text{ g} \times 4.18 \text{ J/g}^{\circ}\text{C} \times (96 - 15)^{\circ}\text{C}$$
$$= \boxed{108 \text{ kJ}}$$

- b) And 74.3 g of ethanol from -36.4 to 44.5 °C? (sp heat ethanol = 2.46 J/g °C) (Ans: 14.8 kJ)

$$q = 74.3 \times 2.46 \text{ J/g}^{\circ}\text{C} \times (44.5 - (-36.4))^{\circ}\text{C}$$
$$= \boxed{14.8 \text{ kJ}}$$

- 10) A 638 g block of lead initially at 27.0 °C absorbs 2044 J of heat. What is the final temperature of the lead block? (sp heat lead = 0.128 J/g °C) (Ans: 52.0 °C)

$$\Delta T = \frac{q}{ms} = \frac{2044}{638 \times 0.128 \text{ J/g}^{\circ}\text{C}} = 25^{\circ}\text{C}$$

$$25^{\circ}\text{C} = T_f - 27^{\circ}\text{C}$$

$$\boxed{T_f = 52^{\circ}\text{C}}$$

11) A 9.13 g sample of vanadium is heated to 99.10 °C and is then dropped into 20.0 g water in a calorimeter. The water temperature rises from 20.51 °C to 24.46 °C. Calculate the specific heat of vanadium. (Ans: -0.485 J/g °C)

$$-q_v = +q_{H_2O} \quad \text{OR} \quad q_{H_2O} + q_v = 0$$

$$[9.13g \times (s) \times (24.46 - 99.10)] + [20g \times 4.185 J/g^\circ C \times (24.46 - 20.51)]$$

$$-681.46(s) + 330.2 = 0$$

$$s = \frac{330.2}{681.46} = \boxed{0.485 J/g^\circ C}$$

12) What is the final temperature of water when 20.2 g of water at 95.4 °C is mixed with 50.5 g water at 26.7 °C? (Ans: 46.3 °C)

$$-q_{H_2O} = +q_{H_2O} \quad \text{OR} \quad q_{H_2O} + q_{H_2O} = 0$$

$$\text{hot} \quad \text{cold}$$

$$[20.2g \times 4.184 \times (T_f - 95.4)] + [50.5g \times 4.184 \times (T_f - 26.7)]$$

$$(84.5168T_f - 8062.9) + (211.29T_f - 5641.5)$$

$$295.81T_f - 13704.4 = 0$$

$$13704.4 / 295.81 = T_f = \boxed{46.3^\circ C}$$

13) A 500.0 mL sample of 0.500 M NaOH at 20.00 °C is mixed with an equal volume of 0.500 M HCl at the same temperature in a foam cup calorimeter. The temperature rises to 23.21 °C. What is the ΔH of the reaction? (Ans: -53.5 kJ/mol)

$$\text{total vol} = 500 + 500 = 1000 \text{ mL} \equiv 1000 \text{ g}$$

$$a) q_{cal} = m s \Delta T = 1000g \times 4.18 J/g^\circ C \times (23.21 - 20)^\circ C = 13.42 \text{ kJ}$$

$$q_{rxn} = -q_{cal} = -13.42 \text{ kJ}$$

$$b) \text{Calc. mols. } 0.5 \text{ L} \times \frac{0.5 \text{ mol}}{\text{L}} = 0.25 \text{ mol NaOH} \times \frac{1 \text{ mol H}_2\text{O}}{1 \text{ mol NaOH}} = 0.25 \text{ mol}$$

$$c) \text{molar } \Delta H = \frac{-13.42 \text{ kJ}}{0.25 \text{ mol}} = \boxed{-53.7 \text{ kJ}}$$

- 14) A 250.0 mL sample of 0.250 M NaOH at 21.00 °C is mixed with an equal volume of 0.500 M H₂SO₄ at the same temperature in a foam cup calorimeter. The temperature rises to 26.52 °C. What is the ΔH of the reaction? (Ans: 92.4 kJ/mol)



a) $q_{\text{cal}} = m s \Delta T$

$$= 500 \times 4.184 \times (26.52 - 21)$$

$$= 1.155 \times 10^4 \text{ J} \approx 11.5 \text{ kJ}$$

$$\begin{aligned} \text{total vol} &= \frac{250}{250} \\ &= 500 \text{ mL} \\ &\approx 500 \text{ g} \end{aligned}$$

$$q_{\text{rxn}} = -q_{\text{cal}} = -11.5 \text{ kJ}$$

b) mol NaOH $0.25 \text{ L} \times 0.25 \text{ M} = 0.0625 \text{ mol}$ - LR
 H₂SO₄ $0.25 \text{ L} \times 0.5 \text{ M} = 0.125 \text{ mol}$

c) $\Delta H = 11.5 \text{ kJ} / 0.0625 \text{ mol} = \boxed{184 \text{ kJ/mol}}$

- 15) A 1.50 g sample of NH₄NO₃(s) is added to 35.0 g of water in a foam cup calorimeter and stirred until dissolved. The temperature of the solution drops from 22.7 to 19.4 °C. What is the enthalpy the NH₄NO₃ solution? (Ans: +25.8 kJ/mol)

a) $q_{\text{cal}} = m s \Delta T = 35 \text{ g H}_2\text{O} \times 4.184 \text{ J/g}^\circ\text{C} \times (19.4 - 22.7)^\circ\text{C}$
 $= -483 \text{ J}$

$$q_{\text{rxn}} = -q_{\text{cal}} = +483 \text{ J}$$

$$(\text{MW}) \text{NH}_4\text{NO}_3 = 80.03 \text{ g/mol}$$

b) $\frac{483 \text{ J}}{1.5 \text{ g NH}_4\text{NO}_3} \times \frac{80.03 \text{ g NH}_4\text{NO}_3}{1 \text{ mol}} = \boxed{26 \text{ kJ/mol}}$

- 16) Express the equation below as a thermochemical equation when 30.12 g of ethane is combusted in oxygen in a bomb calorimeter. The temperature in the calorimeter changes from 25.6 °C to 40.0 °C. The heat capacity of the calorimeter is 2.33 kJ/°C. (Ans: -33.4 kJ/mol)



$$-q = C \Delta T = 2.33 \text{ kJ/}^\circ\text{C} \times (40.0 - 25.6)^\circ\text{C}$$

$$\hookrightarrow 14.4^\circ\text{C}$$

$$= -33.552 \text{ kJ}$$

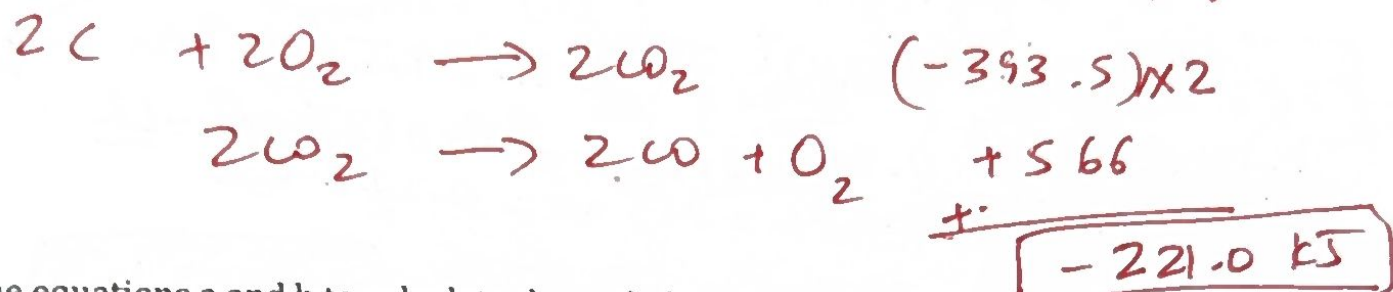
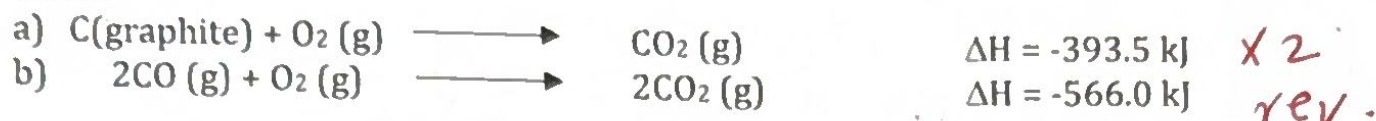
$$30.12 \text{ g C}_2\text{H}_6 \times \frac{1 \text{ mol}}{30 \text{ g}} = 1.004 \text{ mol}$$

$$\frac{-33.552 \text{ kJ}}{1.004 \text{ mol}} = \boxed{-33.4 \text{ kJ/mol}}$$

Practice/Thermochemistry-Hess's Law and Heat of Formation DO NOT SUBMIT

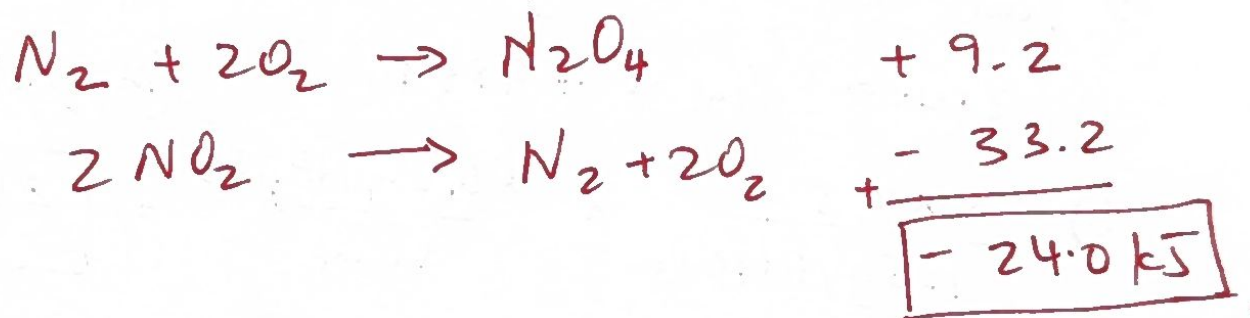
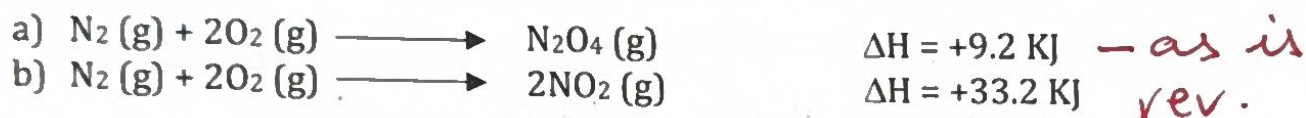
17) Use the equations a and b to calculate the enthalpy change for the equation given below.
 $2C(\text{graphite}) + O_2(g) \longrightarrow 2CO(g) \quad \Delta H = ? \text{ (Ans: -221.0 kJ)}$

Given:



18) Use the equations a and b to calculate the enthalpy change for the equation given below.
 $2NO_2(g) \longrightarrow N_2O_4(g) \quad \Delta H = ? \text{ (Ans: -24.0 kJ)}$

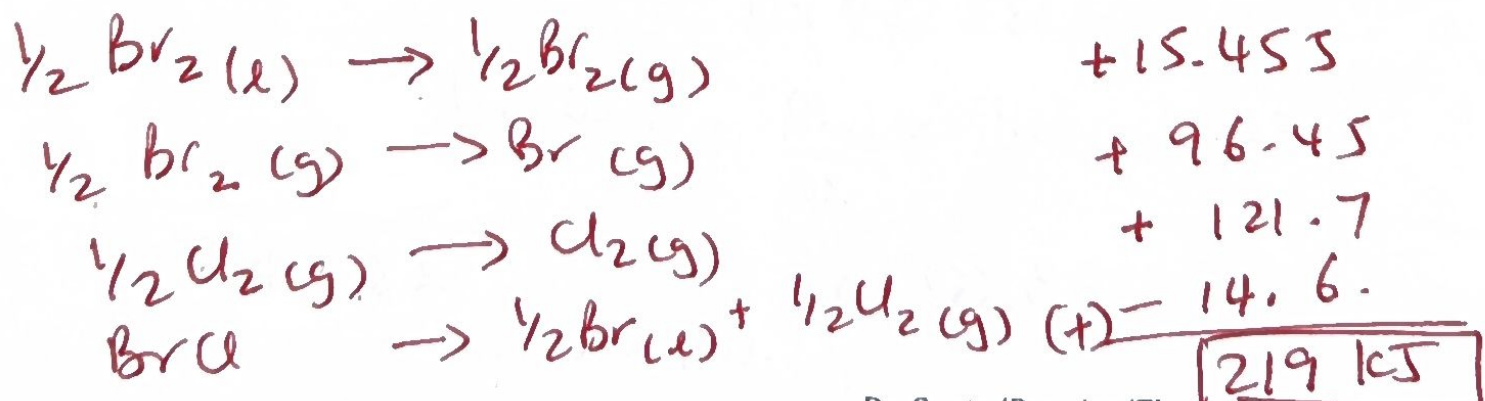
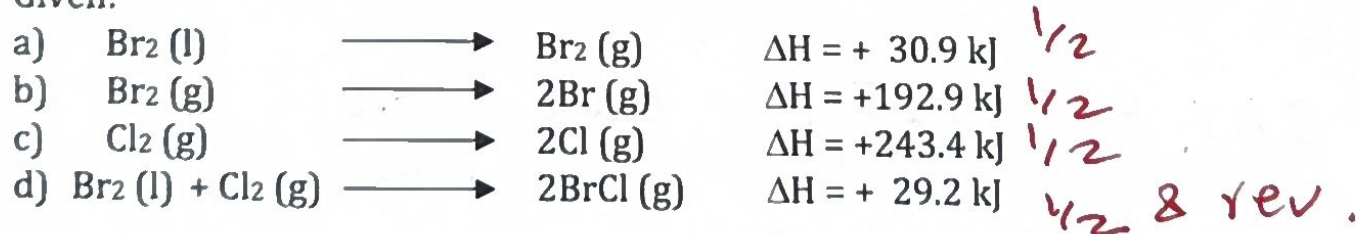
Given:



19) Use the equations given to calculate the enthalpy change for the equation given below. Show the process to get to your answer. (Ans: 219 kJ)



Given:



*20) Use the standard enthalpies of formation given below to calculate the standard enthalpy change for the following two reactions.

(ΔH_f° in kJ/mol: $\text{Al}_2\text{O}_3 = -1676$, $\text{Al}(\text{OH})_3 = -1276$, $\text{H}_2\text{O} = -285.8$; $\text{C}_2\text{N}_2 = 308.9$, $\text{CO}_2 = -393.5$)



$$\underbrace{-1676 + (-285.8) \times 3}_{-2533.4} \quad \underbrace{(-1276) \times 2}_{-2552}$$

$$\Delta H = \text{prod} - \text{react} = -2552 - (-2533.4) = \boxed{-18.6 \text{ kJ}}$$



$$\underbrace{(308.9) + 2(0)}_{308.9} \quad \underbrace{3(-393.5) + 0}_{-787}$$

$$\Delta H = \text{prod} - \text{react} = -787 - (308.9) = \boxed{-1095.9 \text{ kJ}}$$