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Soccer – Thermodynamics Practice Problems

Hess's Law

1. Calculate the enthalpy change for the reaction:

$$P_4O_6\ (s) + 2\ O_2\ (g) \rightarrow P_4O_{10}\ (s)$$

Given the following enthalpies of reaction:

$P_4\ (s) + 3\ O_2\ (g) \rightarrow P_4O_6\ (s)$	$\Delta H = -1640.1$ kJ
$P_4\ (s) + 5\ O_2\ (g) \rightarrow P_4O_{10}\ (s)$	$\Delta H = -2940.1$ kJ

$P_4O_6 \rightarrow \cancel{P_4} + 3O_2$ $\quad +1640.1$

$\cancel{P_4} + 5O_2 \rightarrow P_4O_{10}$ $\quad + \ -2940.1$

$P_4O_6 + 2O_2 \rightarrow P_4O_{10}$ $\quad \boxed{-1300.0\ kJ}$

2. From the enthalpies of reaction:

$H_2\ (g) + F_2\ (g) \rightarrow 2\ HF\ (g)$	$\Delta H = -537$ kJ
$C\ (s) + 2\ F_2\ (g) \rightarrow CF_4\ (g)$	$\Delta H = -680$ kJ
$2\ C\ (s) + 2\ H_2\ (g) \rightarrow C_2H_4\ (g)$	$\Delta H = 52.3$ kJ

Calculate ΔH for the reaction:

$$C_2H_4\ (g) + 6\ F_2 \rightarrow 2\ CF_4\ (g) + 4\ HF\ (g)$$

$C_2H_4 \rightarrow 2C + 2H_2$ $\quad -52.3$ $\quad -\Delta H$

$2\ H_2 + \overset{2}{\ }F_2 \rightarrow 4\ HF$ $\quad -1074$ $\quad 2(\Delta H)$

$2C + 2/F_2 \rightarrow 2CF_4$ $\quad + \ -1360$

$\boxed{\Delta H = 2486\ kJ}$

3. How much energy is absorbed or released (identify which) when each of the following reactions takes place? The relevant bond energies are:

C – H	414 kJ/mol	C = O	730 kJ/mol	H – H	435 kJ/mol
O = O	502 kJ/mol	O – H	464 kJ/mol	Cl – Cl	243 kJ/mol
H – Cl	431 kJ/mol	C – Cl	331 kJ/mol		

a. $CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (g)$

ΔH = Bonds broken – bonds formed

$\Delta H = [4(414) + 2(502)] - [2(730) + 4(464)] = \boxed{-656\ kJ}$

b. $2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (g)$

$$\Delta H = [2(435) + 502] - [4(464)] = \boxed{-484\ kJ}$$

4. Using Appendix C of your textbook, calculate the ΔH° for the following reactions using the listed enthalpies of formation.

a. $2\ SO_2\ (g) + H_2\ (g) \rightarrow 2\ SO_3\ (g)$

$$[2(-395.2)] - [2(-296.9) + 1(0)] = \Delta H^\circ$$

$$\boxed{-196.6 \frac{kJ}{mol} = \Delta H^\circ}$$

b. $SiCl_4\ (l) + 2\ H_2O\ (l) \rightarrow SiO_2\ (s) + 4\ HCl\ (g)$

$$\Delta H = [4(-92.30) + 1(-910.9)] - [1(-640.1) + 2(-285.83)]$$

$$\Delta H = {}^{-}1280.1 - {}^{-}1211.76 = \boxed{-68.3 \frac{kJ}{mol}}$$

5. Hydrogen gas burns in air according to the equation below:

$$2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (l)$$

(a) Calculate the standard enthalpy change, ΔH°, for the reaction by the equation above. (The molar enthalpy of formation, ΔH_f°, for H_2O (l) is -285. kJ/mol^{-1}

$$\Delta H^\circ = [2(-285)] - [2(0) + 1(0)] = \boxed{-570 \frac{kJ}{mol}}$$

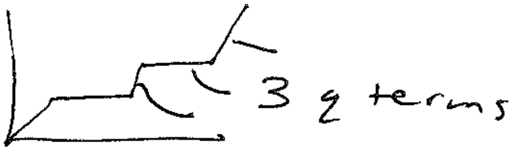
(b) Calculate the amount of heat, in kJ, that is released when 10.0 g of H_2 (g) is burned in air.

$$10.0\,g \left(\frac{1}{2.016\,g}\right) = 4.96\ mol \times \frac{-570\ kJ}{mol} = \boxed{-2827.2\ kJ}$$

Useful values:

Specific Heat Capacity of H_2O (s): 2.06 J/(g x °C)
$\Delta H_{vaporization}$ of H_2O (at 100.0 °C) = 2256 J/g
Specific Heat Capacity of H_2O (l): 4.184 J/(g x °C)
ΔH_{fusion} of H_2O (at 0.0 °C) = 333 J/g
Density of H_2O = 1.00 g/mL

Name:
AP Chemistry
Date:
Period:



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Calorimetry

6. You have 15.0 g of water steam at 105 °C. You cool this steam down to 37 °C. The $\Delta H_{vaporization}$ is 2256 J/g. How much heat is required to complete this process. Express your final answer in kJ.

15.0 g m
$T_i = 105$ °C
$T_f = 37$°C

$q = q_{steam} + q_{condensation} + q_{liquid}$

$q = (15)(2.06)(100-105) + (15)(-2556) + (15)(4.184)(37-100)$

$q = -154.5 + -38340 + -3953.88$

$\boxed{q = 42{,}000\ J}$

7. An unknown metal with a mass of 75.8 g is heated to a temperature of 80.0 °C. It is then placed in 79.4 g of water that is at a temperature of 18.5 °C. The temperature of the water and metal than finishes at a temperature of 23.5 °C. What is the specific heat of the metal. Note – the specific heat of water is 4.184 J/(g x °C).

$0 = q_{metal} + q_{H_2O}$

$0 = mC_s\Delta T + mC_s\Delta T$

$0 = (75.8)C_s(23.5-80) + (79.4)(4.184)(23.5-18.5)$

$0 = -4282.7C_s + 1661.0$

$\frac{-1661.048}{-4282.7} = \frac{-4282.7C_s}{-4282.7}$

$\boxed{C_s = 0.388\ J/g\cdot°C}$

8. A 130.2 g sample of nickel heated to 95.0 °C is placed on 15.0 g of ice at -10.0 °C in an insulated container. The final temperature of the system is 1.55 °C. Assuming no heat escapes from the insulated container, what is the specific heat capacity of nickel?

$0 = q_{ice} + q_{melt} + q_{H_2O} + q_{metal}$

$0 = (15)(2.06)(10) + (15)(333) + (15)(4.184)(1.55) + (130.2)C_s(1.55-95)$

$0 = 309 + 4995 + 97.278 + -12167.19C_s$

$\boxed{0.44\ \frac{J}{g\cdot°C} = C_s}$

9. When 50.0 mL of 0.910 M H_2SO_3 at 23.1 °C is mixed with 150.0 mL of 0.582 M KOH at 23.1 °C in a coffee cup calorimeter, the temperature of the final solution is determined to be 32.8 °C. The density of the final solution is 1.00 g/mL; the specific heat capacity of the final solution is 4.184 J/(g x °C). Assuming no los of heat to the calorimeter, what is the value of ΔH_{rxn} for the reaction written below? HINT – Don't forget to find the limiting reactant!

$H_2SO_3\,(aq) + 2KOH\,(aq) \rightarrow 2\,H_2O + K_2SO_3$

① q_{rxn}

$0 = q_{solution} + q_{rxn} \Rightarrow -q_{rxn} = q_{solution}$

50+150 = 200 ml = 200 g $-q_{rxn} = (200\,g)(4.184)(32.8-23.1)$

$q_{rxn} = -8116.96\ J$

② Limiting Reactant

$0.05\,L\left(\frac{0.91\,mol}{L}\right)\left(\frac{2\,H_2O}{1\,H_2SO_3}\right) = 0.091\,mol$

$0.15\left(\frac{0.582\,mol}{L}\right)\left(\frac{2\,H_2O}{2\,KOH}\right) = 0.0873$

limiting reactant

0873

③ Calculate ΔH_{rxn}

$$\Delta H = -\frac{8116.96}{0.0873} \times \frac{2}{1} = -185955$$

$$\approx$$

$$\boxed{-185 \ \text{kJ/mol}}$$

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ΔS^o

10. In each of the following pairs, predict the substance that will have the higher standard molar entropy. Explain your reasoning!

a. C_2H_2 (g) or C_2H_6 (g)

C_2H_6 weighs more meaning it can take more orientations and, thus, have a higher ΔS

b. H_2O (l) or H_2O (g)

H_2O (g); gases have more disorder than liquids.

11. Using Appendix C in your textbook, calculate ΔS^o values for the following reactions:

a. $C_2H_4 (g) + H_2 (g) \rightarrow C_2H_6 (g)$

$$\Delta S = [1(229.5)] - [1(219.4) + 1(130.58)] = \boxed{-120.48\ J/mol \cdot K}$$

b. $Be(OH)_2 (s) \rightarrow BeO (s) + H_2O (g)$

$$\Delta S = [1(13.7) + 1(188.83)] - [1(50.21)] = \boxed{152.32\ J/mol \cdot K}$$

Gibbs Free Energy

12.

Substance	Standard Heat of Formation, DH_f^o, in kJ mol^{-1}	Absolute Entropy, S°, in J mol^{-1} K^{-1}
C*(s)*	0.00	5.69
CO_2*(g)*	-393.5	213.6
H_2*(g)*	0.00	130.6
H_2O*(l)*	-285.85	69.91
O_2*(g)*	0.00	205.0
C_3H_7COOH*(l)*	?	226.3

The enthalpy change for the combustion of butyric acid at 25°C, ΔH^o_{comb}, is -2,183.5 kilojoules per mole. The combustion reaction is

$$C_3H_7COOH(l) + 5\ O_2(g) \rightarrow 4\ CO_2(g) + 4\ H_2O(l)$$

(a) From the above data, calculate the standard heat of formation, ΔH_f^o, for butyric acid.

$$-2183.5 = [4(-393.5) + 4(-285.85)] - [1(\Delta H) + 5(0)]$$

$$-2183.5 = -2717.4 - \Delta H \Rightarrow \boxed{-533.9 = \Delta H}$$

(b) Write a correctly balanced equation for the formation of butyric acid from its elements.

$$4C + 4H_2 + O_2 \longrightarrow C_3H_7COOH$$

(c) Calculate the standard entropy change, $\Delta S_f°$, for the formation of butyric acid at 25°C. The entropy change, $\Delta S°$, for the combustion reaction above is -117.1 J K^{-1} at 25°C.

$$-117.1 = \Delta S - [4(5.69) + 4(130.6) + 4(205)]$$

$$-117.1 = \Delta S - 1365.16$$

$$1248.1\ J/K = \Delta S$$

(d) Calculate the standard free energy of formation, $\Delta G°_f$, for butyric acid at 25°C.

$$\Delta G = \Delta H - T\Delta S$$

$$\Delta G = -533.9 - (298)(1.248.1)$$

$$\boxed{\Delta G = -905\ kJ}$$

13. Propane, C_3H_8, is a hydrocarbon that is commonly used as fuel for cooking.

(a) Write a balanced equation for the complete combustion of propane gas, which yields $CO_2(g)$ and $H_2O(l)$.

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$$

(b) Calculate the volume of air at 30°C and 1.00 atmosphere that is needed to burn completely 10.0 grams of propane. Assume that air is 21.0 percent O_2 by volume.

$$10\ g\ C_3H_8 \left(\frac{1}{44.094}\right)\left(\frac{5\ O_2}{1\ C_3H_8}\right) = 1.13\ mol\ O_2$$

$$V_{O_2} = \frac{nRT}{P} = \frac{(1.13)(0.0821)(303)}{1} = 28.2\ L$$

$$.21 = \frac{28.2}{V_{tot}} \Rightarrow \boxed{134\ L}$$

(c) The heat of combustion of propane is -2,220.1 kJ/mol. Calculate the heat of formation, $\Delta H_f°$, of propane given that $\Delta H_f°$ of $H_2O(l)$ = -285.3 kJ/mol and $\Delta H_f°$ of $CO_2(g)$ = -393.5 kJ/mol.

$$-2220.1 = [4(-285.3) + 3(-393.5)] - [\Delta H + 5(0)]$$

$$-2220.1 = -2321.7 - \Delta H$$

$$\boxed{-101.6\ kJ/mol = \Delta H}$$

(d) Assuming that all of the heat evolved in burning 30.0 grams of propane is transferred to 8.00 kilograms of water (specific heat = 4.18 J/g·K), calculate the increase in temperature of water.

$$q = 30.0\ g\left(\frac{1}{44\ g}\right)\left(\frac{2220.1\ kJ}{mol}\right) = 1514\ kJ \longrightarrow q = mC_s\Delta T$$

$$1514 = (8000\ g)(4.18)\ \Delta T$$

$$\boxed{45.2°C = \Delta T}$$