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Name: \_\_\_\_\_

SCH OAL  
Thermodynamics Test

1. For each of the following, label with a + for an increase in potential energy or a - for a decrease in potential energy:

- a) - the combustion of methane gas
- b) + the formation of diamond deep within the earth's crust
- c) - formation of ice on a window through a sublimation process
- d) + formation of toluene (l)
- e) + a reaction in which the reactants have greater forces or attraction than the products
- f) - heat energy is absorbed by the water jacket in a bomb calorimeter (consider reactant and product compounds only)
- g) + formation of an onion from all necessary raw materials
- h) + the conversion of ethyl alcohol to dimethyl ether
- i) - deflation of a balloon
- f) + build up of electrostatic charge in a thunderhead

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2. What is the difference between Q and ΔH? Why are these two quantities equal in magnitude but opposite of sign for a given chemical or physical change? (1 mark per main point)

✓ { Q → E<sub>k</sub> (chemical kinetic energy)

✓ { ΔH → E<sub>p</sub> (chemical potential energy)

✓ ΔH = -Q law of conservation of energy ✓

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✓ E<sub>T</sub> = E<sub>k</sub> + E<sub>p</sub>

ΔE<sub>T</sub> = ΔE<sub>k</sub> + ΔE<sub>p</sub>

ΔE<sub>T</sub> = 0 (law of conservation of energy)

∴ 0 = ΔE<sub>k</sub> + ΔE<sub>p</sub>

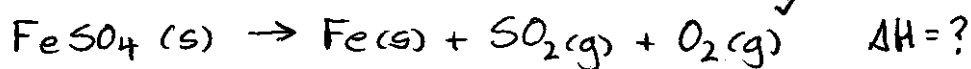
∴ ΔE<sub>p</sub> = -ΔE<sub>k</sub>

↓  
ΔU = -Q

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3. Using the summation of heats of formation technique, calculate the heat of reaction in kJ, if  $\text{FeSO}_4$  (s) is decomposed to form Fe (s) plus sulphur dioxide gas plus oxygen gas. Use the values in the text for help. Now determine the mass of natural gas ( $\text{CH}_4$  (g)) in kg that must be combusted to provide enough heat to decompose 750 kg of ferrous sulphate.

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$$\Delta H = [\Delta H^\circ_{\text{Fe} (\text{s})} + \Delta H^\circ_{\text{SO}_2 (\text{g})} + \Delta H^\circ_{\text{O}_2 (\text{g})}] - [\Delta H^\circ_{\text{FeSO}_4 (\text{s})}]$$

$$\Delta H = [0 + (-70.8 \text{ kcal}) + 0] - [-220.5 \text{ kcal}]$$

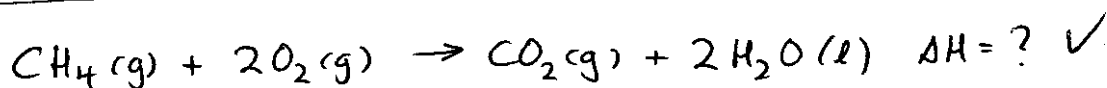
$$\Delta H = 149.7 \text{ kcal} \checkmark$$

$$149.7 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} = 626.35 \text{ kJ}$$

$$Q = -\Delta H \checkmark$$

$$Q = -626.35 \text{ kJ/mol FeSO}_4$$

heat absorbed



$$\Delta H = [\Delta H^\circ_{\text{CO}_2 (\text{g})} + 2\Delta H^\circ_{\text{H}_2\text{O} (\text{l})}] - [\Delta H^\circ_{\text{CH}_4 (\text{g})} + 2\Delta H^\circ_{\text{O}_2 (\text{g})}] \checkmark$$

$$\Delta H = [(-94.0 \text{ kcal}) + 2(-68.3 \text{ kcal})] - [(-17.9 \text{ kcal}) + 2(0)]$$

$$\Delta H = -212.7 \text{ kcal} \checkmark$$

$$Q = -\Delta H$$

$$Q = 212.7 \text{ kcal/mol CH}_4 \checkmark$$

$$750 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol FeSO}_4}{151.92 \text{ g FeSO}_4} \times \frac{149.7 \text{ kcal}}{1 \text{ mol FeSO}_4} \times \frac{1 \text{ mol CH}_4}{212.70 \text{ kcal}}$$

$$\rightarrow \times \frac{16.05 \text{ g CH}_4}{1 \text{ mol CH}_4} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 55.767 \text{ kg}$$

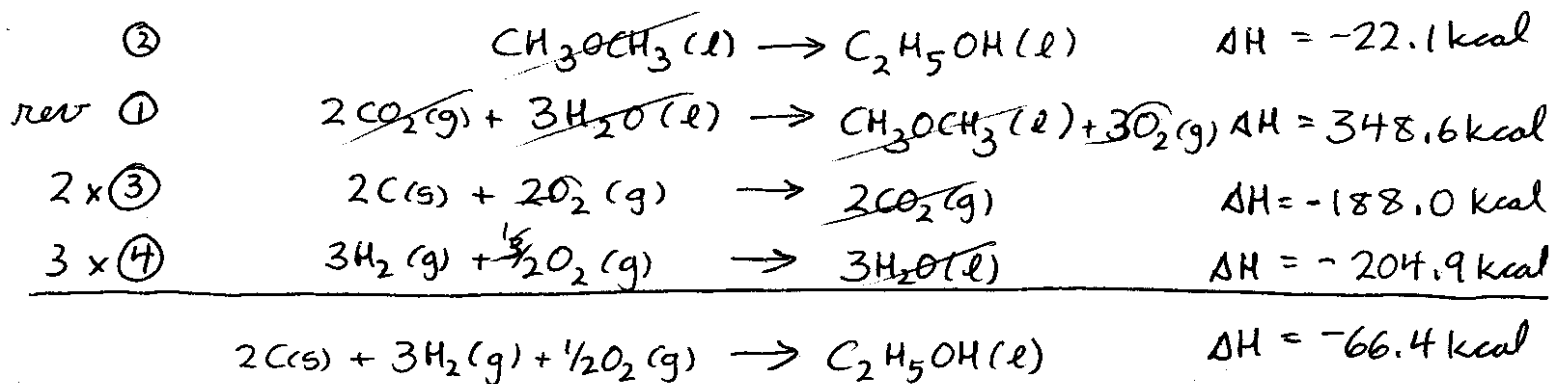
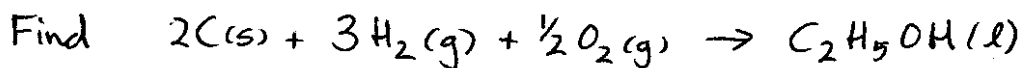
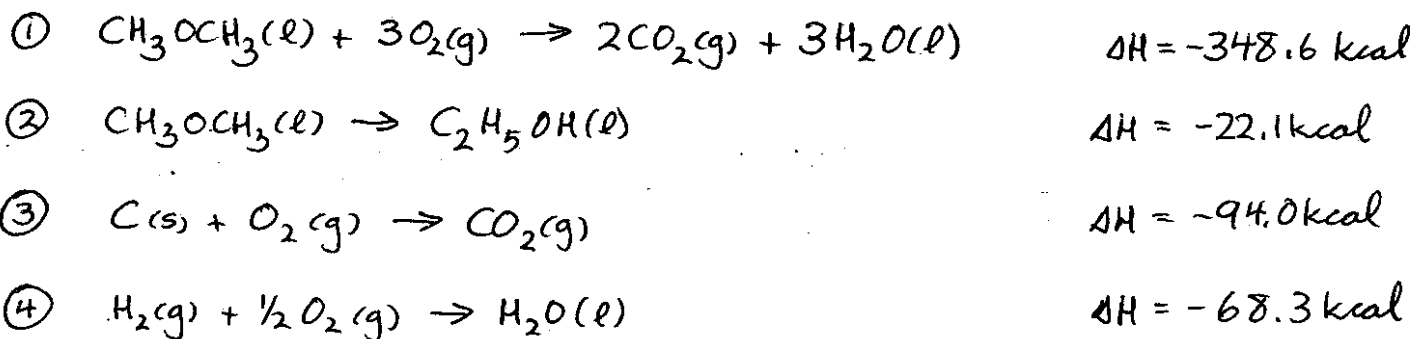
$\therefore$  55.767 kg of  $\text{CH}_4$  is required to decompose 750 kg of  $\text{FeSO}_4$

\* - sign is dropped for convenience, heat required by  $\text{FeSO}_4$  is provided by  $\text{CH}_4$

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4. Use Hess' Law (The Law of Hess) to prove that the heat of formation of ethyl alcohol ( $C_2H_5OH(l)$ ) is  $-66.4$  kcal/mol, using:
- The combustion of dimethyl ether ( $CH_3OCH_3(l)$ ) has a heat of reaction of  $\Delta H = -348.6$  kcal
  - The straight conversion (i.e. no other compounds present in the reaction) of dimethyl ether ( $CH_3OCH_3(l)$ ) to ethyl alcohol ( $C_2H_5OH(l)$ ) has a heat of reaction of  $-22.1$  kcal
  - the formation reaction for  $CO_2(g)$
  - the formation reaction for  $H_2O(l)$
- Please write out all four equations and the equation for the formation of ethyl alcohol before you attempt the Hessian Law!

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$$\frac{7}{2} - \frac{6}{2} = \frac{1}{2}$$

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5. When ammonium nitrate is mixed with water, the resulting solution quickly becomes very cold. Why? (1 mark per major point)

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- heat energy is absorbed,  $\Delta H = +$

- heat energy is converted to chemical potential energy

- particles are on average further apart and less attracted

- entropy is the driving force

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6. What are the two fundamental driving forces in thermodynamics?

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- tendency toward minimum energy (1st Law of Thermody)

- tendency toward maximum randomness (2nd Law of Thermody)

7. Given that:

a)  $\Delta G = \Delta H - T\Delta S$  where:  $\Delta G =$  Gibb's free energy (J)  
 $\Delta H =$  Enthalpy (J)  
 $T =$  temperature (K)  
 $\Delta S =$  Entropy

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b)  $\Delta G < 0$  for a spontaneous reaction

c)  $\Delta H$  value is 10 times more significant than the  $\Delta S$  value

State what the sign (+ or -) is for each of the following situations:

Situation	$\Delta H$ sign	$\Delta S$ sign
evaporation of any liquid to a gas	+	+
combustion of a hydrocarbon	-	+
solidification of a liquid to a crystal solid	-	-
dissolving of a salt in water becomes cold	+	+
dissolving of a salt in water becomes hot	-	+

Note that this question tests your understanding of chemical potential energy in the  $\Delta H$  column and your understanding of entropy in the  $\Delta S$  column.

\* best condition for spontaneous reaction

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8. Given that the combustion of 22.0 g of acetic acid ( $\text{CH}_3\text{COOH}(\text{l})$ ) is able to warm 4 L of water from  $20.000^\circ\text{C}$  to  $39.066^\circ\text{C}$ , calculate the heat of formation of acetic acid. (Use kcal)

$$Q = mc\Delta T \quad \checkmark$$

$$Q = 4000\text{g} \times 1.000\frac{\text{cal}}{\text{g}^\circ\text{C}} \times 19.066^\circ\text{C}$$

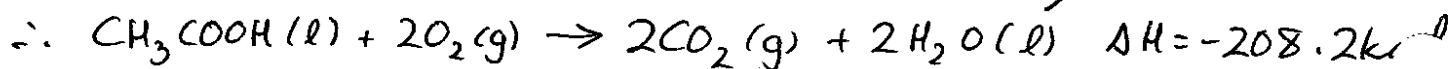
$$Q = 76264\text{cal}$$

$$Q = 76.264\text{kcal}$$

$$\frac{76.264\text{kcal}}{22.0\text{g CH}_3\text{COOH}} \times \frac{60.06\text{g CH}_3\text{COOH}}{1\text{mol CH}_3\text{COOH}} = 208.2\text{kcal/mol CH}_3\text{COOH}$$

$$\Delta H = -Q \quad \checkmark$$

$$\Delta H = -208.2\text{kcal}$$



$$\Delta H = [2\Delta H^\circ_{\text{CO}_2(\text{g})} + 2\Delta H^\circ_{\text{H}_2\text{O}(\text{l})}] - [\Delta H^\circ_{\text{CH}_3\text{COOH}(\text{l})} + 2\Delta H^\circ_{\text{O}_2(\text{g})}] \quad \checkmark$$

$$-208.2\text{kcal} = [2(-94.0\text{kcal}) + 2(-68.3\text{kcal})] - [\Delta H^\circ_{\text{CH}_3\text{COOH}} + 2(0)]$$

$$-208.2\text{kcal} = -324.6\text{kcal} - \Delta H^\circ_{\text{CH}_3\text{COOH}}$$

$$\Delta H^\circ_{\text{CH}_3\text{COOH}} = -116.40\text{kcal} \quad \checkmark$$

If the values in this question are correct the answer you should get is -116.4 kcal!!!

Bonus:

Given the following values at 25 °C and 1 atm pressure:

$$\Delta H_{\text{H}_2\text{O}(l)}^\circ = -68.32 \text{ kcal}$$

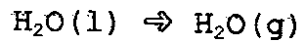
$$\Delta H_{\text{H}_2\text{O}(g)}^\circ = -57.80 \text{ kcal}$$

$$S_{\text{H}_2\text{O}(l)}^\circ = 16.72 \text{ cal}$$

$$S_{\text{H}_2\text{O}(g)}^\circ = 45.11 \text{ cal}$$

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Calculate the temperature above which the physical change of



becomes spontaneous!!!

$$\Delta H = \Delta H_{\text{H}_2\text{O}(g)}^\circ - \Delta H_{\text{H}_2\text{O}(l)}^\circ$$

$$\Delta H = -57.80 \text{ kcal} - (-68.32 \text{ kcal})$$

$$\Delta H = 10.52 \text{ kcal}$$

$$\Delta H = 10520 \text{ cal}$$

$$\Delta S = S_{\text{H}_2\text{O}(g)}^\circ - S_{\text{H}_2\text{O}(l)}^\circ$$

$$\Delta S = 45.11 \text{ cal} - 16.72 \text{ cal}$$

$$\Delta S = 28.39 \text{ cal}^*$$

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

$\Delta G < 0$  for a spontaneous event

$$\Delta H - T\Delta S < 0$$

$$\Delta H < T\Delta S$$

if  $\Delta S$  is positive \*

$$\frac{\Delta H}{\Delta S} < T$$

$$\therefore T > \frac{10520 \text{ cal}}{28.39 \text{ cal}}$$

$$T > 370.55 \text{ K}$$

close !!

