

6 • Energy and Chemical Reactions

HESS'S LAW CALCULATIONS

The enthalpy of the reactants, $H_{\text{reactants}}$ and the enthalpy of the products, H_{products} depend on the bonding of the reactants and products... nothing else. So, the $\Delta H_{\text{reaction}}$ only depends on the initial and final state of the reaction, not how you got from one state to another state. It is called a "state function".

Practically speaking, if we can find several equations that "add up" to the equation we want, the $\Delta H_{\text{reactions}}$ will add up to the overall ΔH . This is called Hess's Law.

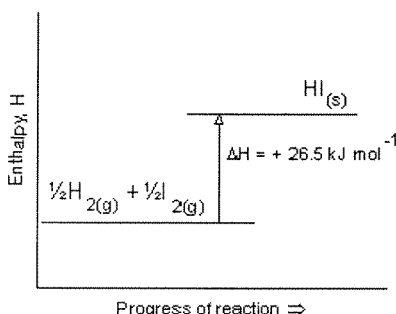
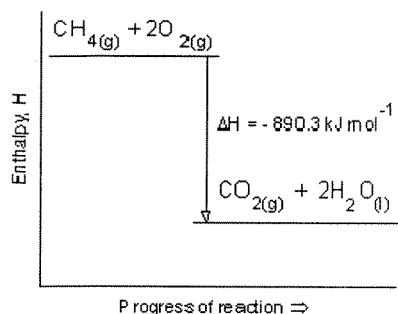
Heats of Formation: Write the formation equations for the following. [See Table 6.2 on page 270 of text.]

Compound	Formation Equation	ΔH_f (kJ·mol ⁻¹)
CH ₄ (g)	C(s) + 2H ₂ (g) → CH ₄ (g)	-74.8
H ₂ O(l)		-285.8
H ₂ O(g)		-241.8
CO ₂ (g)		-393.5
C ₂ H ₆ (g)		-84.7
C ₃ H ₈ (g)		-103.8
C ₄ H ₁₀ (g)		-125.6

- Calculate the heat of combustion, $\Delta H_{\text{combustion}}$, for each of the fuels in the chart.
- Develop a shortcut version of Hess's Law (when using H_f 's).
- Use the heat of combustion of C₃H₈ to calculate the ΔH_f of C₃H₈.

Thermodynamics practice WS #1

1. For each of the enthalpy diagrams shown below, state whether the reaction is endothermic or exothermic. Rewrite the equation for each reaction, including the energy term in the appropriate place (either on the products side or the reactants side)



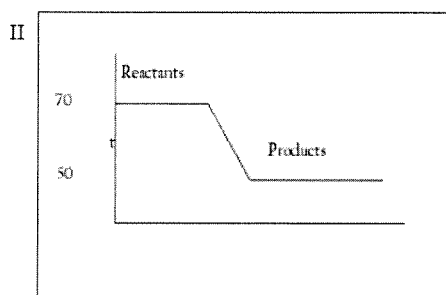
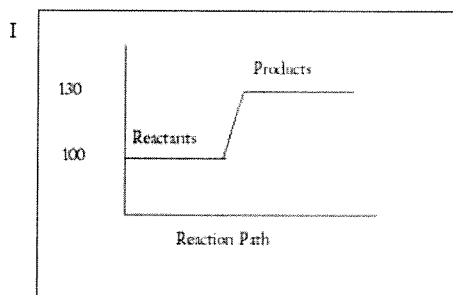
2. Define the following terms

- | | |
|----------------|-----------------------|
| a. Heat | d. Endothermic |
| b. Temperature | e. Enthalpy |
| c. Exothermic | f. Change in Enthalpy |

3. Identify the following processes as being endothermic or exothermic

- | | |
|---------------------|----------------------------------|
| a. melting lead | f. removing water from a hydrate |
| b. freezing water | g. decomposing a compound |
| c. boiling water | h. temperature increases |
| d. condensing steam | i. temperature decreases |
| e. burning wood | |

4. Use the following diagrams to answer the following questions



For diagram I

- a. What is the change in enthalpy in the reaction?
- b. Is this reaction endothermic or exothermic?

For diagram II

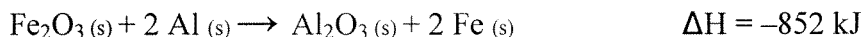
- a. What is the change in enthalpy in the reaction?
- b. Is this reaction endothermic or exothermic?

5. Draw Enthalpy Diagrams for the following conditions

- a. Reactants = 110 $\Delta H = +25$ Products = ?
- b. Reactants = 75 $\Delta H = -30$ Products = ?

c. Reactants = ?	$\Delta H = +40$	Products = 120
d. Reactants = ?	$\Delta H = -20$	Products = 50
e. Reactants = 90	$\Delta H = ?$	Products = 115
f. Reactants = 40	$\Delta H = ?$	Products = 20

6. Thermite mixtures are used for certain types of welding. The thermite reaction is highly exothermic.



1.00 mol Fe_2O_3 and 2.00 mol Al are mixed at room temperature (25°C), and a reaction is initiated. The liberated heat is retained within the products, whose combined specific heat over a broad temperature range is about $0.8 \text{ J g}^{-1} \text{ K}^{-1}$. The melting point of iron is 1530°C . Show that the quantity of heat liberated is more than sufficient to raise the temperature of the products to the melting point of iron.

7. A block of metal of mass 1 kg is heated to 400 K and then dropped into 0.3 kg of water. The water rises in temperature from 294 K to 310 K. The heat capacity of water is $4200 \text{ J/K}\cdot\text{kg}$. You can assume that all of the heat in the metal passed into the final system with no loss to the surroundings. Draw a cartoon showing the initial and final states of this system. Calculate the heatcapacity per kg of the metal.

8. A 6.22 kg piece of copper metal is heated from 20.5°C to 324.3°C . Given that the specific heat of Cu is $0.385 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}$, calculate the heat absorbed (in kJ) by the metal.

Table 3.2 The Specific Heat Capacities of Some Common Substances

Substance	Specific Heat Capacity ($\text{J/g } ^\circ\text{C}$)
water (l)* (liquid)	4.184
water (s) (ice)	2.03
water (g) (steam)	2.0
aluminum (s)	0.89
iron (s)	0.45
mercury (l)	0.14
carbon (s)	0.71
silver (s)	0.24
gold (s)	0.13

*The symbols (s), (l), and (g) indicate the solid, liquid, and gaseous states, respectively.

9. The molar heat of vaporization of water is $44.01 \text{ kJ mol}^{-1}$ at 298 K and $40.79 \text{ kJ mol}^{-1}$ at 373K. Give a qualitative explanation of the difference in these two values.

10. Calculate the total amount of heat that would be required to raise the temperature of a 24 g sample of iron 52° .

11. Calculate the amount of heat that it would take to raise the temperature of a 1.2kg block of Aluminum by 35° .

12. Calculate the amount of heat that it would take to raise 321 g of water from just above its freezing point to just below its boiling point.

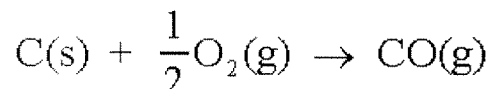
13. What is the mass of a sample of iron to which was applied 343 J which caused its temperature to change by 420° ?

14. When a pot of molten iron was dropped in a container of liquid water it delivered a total heat of 12,342 J which raised the water temperature by 12° . What is the mass of the water in that container.

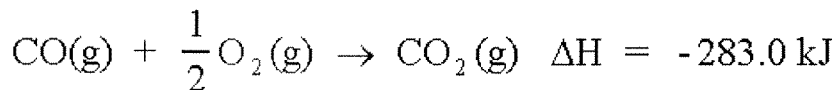
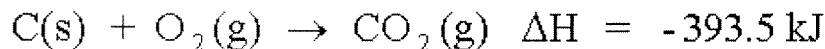
15. A 1.23 Kg block of Tofu was heated by 378 J and its temperature raised by 0.045° . What is the Specific Heat Capacity of that Tofu?

ENTHALPY PROBLEMS

1. Determine ΔH for the reaction

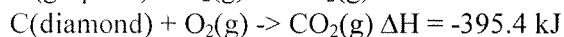
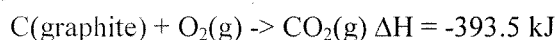


Use Hess's Law of Heat Summation and the following reactions to solve for the enthalpy of the above reaction



(-110.5 kJ)

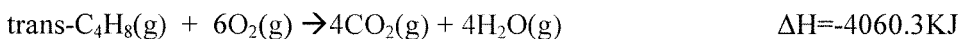
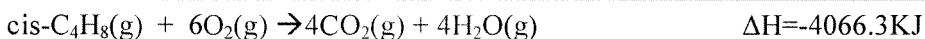
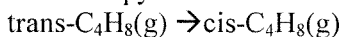
2. Carbon occurs in two forms: graphite and diamond. The enthalpy of combustion of graphite is -393.5 kJ, and that of diamond is -395.4 kJ



Calculate ΔH for the conversion of graphite to diamond

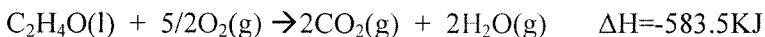
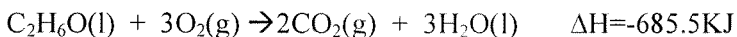
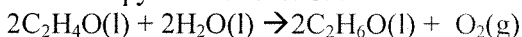
(1.9 kJ)

3. Use the thermochemical equations shown below to determine the enthalpy for the reaction:



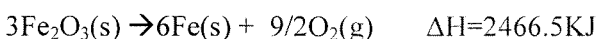
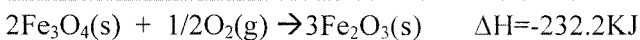
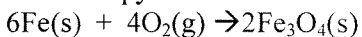
(6 kJ)

4. Use the thermochemical equations shown below to determine the enthalpy for the reaction:



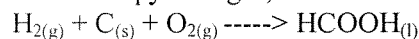
(204 kJ)

5. Use the thermochemical equations shown below to determine the enthalpy for the reaction:



(-2234.3 kJ)

6. Given the following reactions and their enthalpy changes, calculate the enthalpy change for

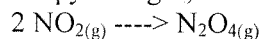


Equations	Change in Heat Energy

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$	-394 kJ
$H_{2(g)} + 1/2 O_{2(g)} \rightarrow H_2O_{(l)}$	-286kJ
$HCOOH_{(l)} + 1/2 O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$	-275kJ

(-405 kJ)

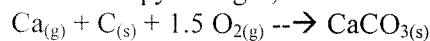
7. Given the following reactions and their enthalpy changes, calculate the enthalpy change for



Equations	Change in Heat Energy
$N_{2(g)} + 2 O_{2(g)} \rightarrow 2 NO_{2(g)}$	67.8 kJ
$N_{2(g)} + 2 O_{2(g)} \rightarrow N_2O_{4(g)}$	9.67 kJ

(-58.13 kJ)

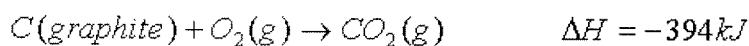
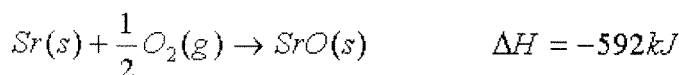
8. Given the following reactions and their enthalpy changes, calculate the enthalpy change for



Equations	Change in Heat Energy
$Ca_{(g)} + 2 C_{(s)} \rightarrow CaC_{2(s)}$	-62.8 kJ
$CO_{2(g)} \rightarrow C_{(s)} + O_{2(g)}$	394 kJ
$CaCO_{3(s)} + CO_{2(g)} \rightarrow CaC_{2(s)} + 5/2 O_{2(g)}$	1538 kJ

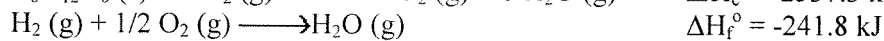
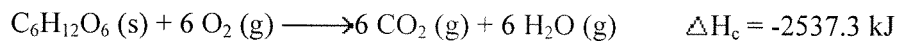
(-1994.8 kJ)

9. Calculate the heat change for the formation of 1.00mol of strontium carbonate from its elements. Given the following experimental information:



(-1220 kJ)

10. Calculate the ΔH_f° for $6 C(s) + 6 H_2(g) + 3 O_2(g) \rightarrow C_6H_{12}O_6(s)$ given that:



(-1274.5 kJ)

Big Questions – Thermodynamics

1. What is enthalpy? How are enthalpy changes measured? Why can we not measure absolute enthalpy?
2. What is the difference between an exothermic reaction and an endothermic reaction? What are some examples of exothermic reactions?
3. Is bond breaking an exothermic or endothermic process? What about making new bonds?
4. In an endothermic reaction, are the products or the reactants more stable? How can you tell?
5. What is specific heat capacity?
6. How can the enthalpy change of a reaction be calculated? What information is required?
7. What is the basic principle behind Hess's Law? Why is it useful?
8. What is average bond enthalpy? How does bond enthalpy relate to the endothermic or exothermic nature of a reaction? Will the ΔH of a reaction calculated using average bond enthalpies be the same as the value calculated using ΔH of formation? Why or why not?
9. What is meant by the term standard state? standard enthalpy change of formation? standard enthalpy change of combustion?
10. Define the term lattice enthalpy. How does lattice enthalpy change if the size or charge of the ions changes?
11. What is electron affinity? What is the difference between electron affinity and ionization energy?
12. What is the meaning of a large difference between the theoretical and experimental values for lattice enthalpy? What is the name of the cycle that is used to calculate experimental values for lattice enthalpy?
13. What is entropy? Can entropy ever be zero? State and explain four things that increase the entropy of a system.
14. What is the relationship between Gibbs Free energy and spontaneity of a reaction?
15. What is the equation that allows us to calculate ΔG ?
16. Will an exothermic reaction always be spontaneous? Explain.
17. If a reaction causes an increase in entropy, will it necessarily be spontaneous? Explain.