Thermochemistry

When the temperature of a substance changes, the amount of _kinetic energy changes. During a phase change, the temperature of the substance does not change but the amount of _kinetic energy does change. When molecules separate, as in melting, energy is _absorbed by the substance. When molecules bond together, as in freezing, energy is _released by the substance.

q = _energy (Joules)

m = _mass (grams)

c = _specific heat (j/gC)

ΔT = _change in temp ©

n = _moles

ΔH = __heat of... (Joules/mole)

Heat/Cool a substance
$q = m \times Cp \times \Delta T$
Phase Change
$q = n \times \Delta H$

Problems:

For Silver: Melting Point = 961 °C | Boiling Point = 2193 °C | Specific Heat = 230. J/kg°C

Heat of Fusion = 9504 J/mole | Heat of Vaporization = 93,970 J/mol | Molar Mass = 107.9g

1. How much energy is needed to raise the temperature of 2.50 kg of silver from 30.0°C to 961°C?
   535, 000 Joules

2. How many g of silver can be heated 3.00°C by 2.00×10^3 joules of energy?

   2.90 Kg or 2,900 g

3. How much heat is needed to vaporize 54.0 g of silver at its boiling temperature?

   47.0 kilo Joules

4. How much energy is released when 216 grams of silver solidify from the liquid state?
   -1.90×10^3 Joules

5. How many moles of silver can be liquefied at 961°C by 4.75×10^5 J of heat?

   50.0 moles

Physical properties of H_2O:

Melting Point. = 0.0°C	Boiling Point 100.°C
Heat of fusion = 6.01	Heat of solidification = -6.01
Heat of vaporization = 40.7	Heat of condensation = - 40.7
Specific heat (C_p) = 4.18	

1. Using the physical properties of water listed above, calculate the heat involved in the conversion of 20.0g of $H_2O_{(s)}$ to $H_2O_{(l)}$. Indicate if the reaction is endothermic or exothermic.

6.69 Kj

(Circle one)

Endothermic or Exothermic

2. Using the physical properties of water listed above, calculate the heat involved in the conversion of 95.0g of $H_2O_{(g)}$ to $H_2O_{(l)}$. Indicate if the reaction is endothermic or exothermic.

-215 Kj

(Circle one)

Endothermic or Exothermic

Physical properties of Hg:

Melting Point. = -39°C	Boiling Point 357°C
Heat of fusion = 2.34	Heat of solidification = - 2.34
Heat of vaporization = 59.37	Heat of condensation = - 59.37
Specific heat (C_p) = 1.25	

3. Using the physical properties of mercury (Hg) listed above, calculate the heat involved in the conversion of 40.0 g of $Hg_{(l)}$ to $Hg_{(g)}$.

11.8 Kj

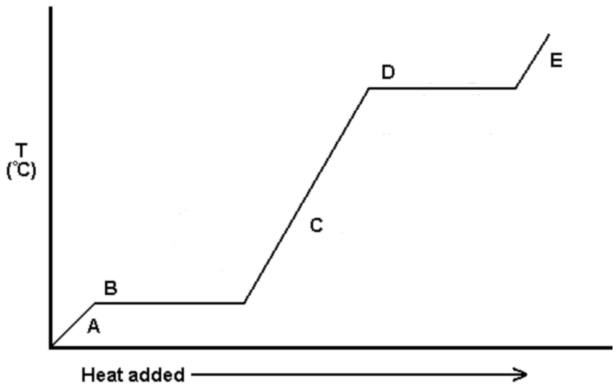
(Circle one)

Endothermic or Exothermic

4. Calculate the amount of heat involved to raise 50.0g of $Hg_{(l)}$ from 200°C to boiling AND to vaporize the entire 50.0g sample.

9812.5 J to heat to boiling point. Conver this to Kj before adding it to the 14.8 kj needed to vaporize the sample. Total Kj required is 24.6.

5. Label A-E with the following terms: Solid, liquid, gas, fusion, condensation, vaporization and solidification.



A =

B =

C =

D =

E =

Heats of Reaction and Stoichiometry

1. Describe what happens when two bodies of matter at different temperatures are brought together. Give an example.

The hot metal transers heat to the cold water in the calorimeter until the metal & the water reach the same temperature.

2. The ΔH for the combustion of 1.00 mole of methane is -892 kJ.
   a. Write the energy term in the equation:

$$CH_{4(g)} + 2\,O_{2(g)} \rightarrow CO_{2(g)} + 2\,H_2O_{(l)} + \textbf{892kj}$$

   b. How much heat is given off when 1.00 gram of methane is burned?

-55.8 kj

3. How much heat would be released if 0.200 mol of nitrogen reacts in the following reaction: -36.9 Kj

$$N_{2(g)} + 3\,H_{2(g)} \rightarrow 2\,NH_{3(g)} \qquad \Delta H = -92.2 \text{ kJ per mole of } NH_3$$

   a. How much heat would be released from the same equation if 0.600 moles of hydrogen reacted?
   -36.9 Kj

2. Pentaborane reacts with oxygen gas to form B_2O_3 and water vapor.
   a. Balance the equation for the reaction and include the energy term.

$$2\,B_5H_9 + 12\,O_2 \rightarrow 5\,B_2O_3 + 9\,H_2O$$

   b. Calculate the change in enthalpy when 10.5 g of pentaborane burns. The ΔH for the balanced equation is –8686.6 kJ

-723 Kj

3. Welders use oxyacetylene torches which are fueled by the combustion of acetylene, C_2H_2, in the following reaction:

$$2\,C_2H_{2(g)} + 5\,O_{2(g)} \rightarrow 4\,CO_{2(g)} + 2\,H_2O_{(g)}$$

If the enthalpy change for the reaction above is –2511.14 kJ, how much heat can be produced by the reaction of:

a. 2 mol of C_2H_2

2511.4 kj

b. 2.00 g of C_2H_2

96.5 kj

4. Calculate the standard heat of reaction for the following reaction, if 1.00 g of magnesium gives off 46.22 kJ of heat when it reacts with excess fluorine? Write the energy term in the equation.

$$Mg_{(s)} + F_{2(g)} \rightarrow MgF_{2(s)} + 1123\ kj$$

5. Identify the following changes and classify as exothermic or endothermic:

a. $H_2O_{(s)} \rightarrow H_2O_{(l)}$ endo

b. $H_2O_{(l)} \rightarrow H_2O_{(g)}$ endo

c. $C_2H_5OH_{(g)}) \rightarrow C_2H_5OH_{(l)}$ exo

d. Combustion of Propane exo

Calorimetry

1. Explain what the specific heat of a substance is using your own words. Give an example of two substances which have relatively different significant heats.

TABLE 6.1 The Specific Heat Capacities of Some Common Substances

Substance	Specific Heat Capacity (J/°C · g)
$H_2O(l)$	4.18
$H_2O(s)$	2.03
$Al(s)$	0.89
$Fe(s)$	0.45
$Hg(l)$	0.14
$C(s)$	0.71

Specific Heat is how much energy is required to heat up one gram of a substance by one degree celsius.

2. If 5 moles of the metals; aluminum, silver, and iron, each absorb equivalent amounts of heat, which one of the metals will experience the largest increase in temperature? Why? HINT: Look up the CP for each substance.

Silver will experience the largest increase in temp. because it has the smallest specific heat value.

3. How much heat is required to raise the temperature of 500.0 g of iron from 25.0°C to 50.0°C?

563 Joules

4. The hot iron at 50.0°C is placed into 750.0 g of water at 25.0°C. If the iron loses $5.1x10^3$ J of heat, what will be the final temperature of the water?

Delta T = 1.63 so final temp is 27 C

5. What quantity of heat would have to be added to 5500 g of water to change its temperature from 25.0°C to 85.0°C?

1.38×10^6 Joules

6. If 500.0 g of water at 25°C loses $1.05x10^4$ J of heat, what will be the final temperature of the water?

Delta T = 5.02 C so the final temp will be 20 C

7. What is the specific heat of silver if a 93.9 g sample cools from 215.0°C to 196.0°C with the loss of 428 J of energy?

0.24 J/gC

Hess's Law

ΔH = Heat of Reaction = Enthalpy

Endothermic = ΔH is positive

Exothermic = ΔH is negative

1. Convert the following equation to the ΔH notation.

   a. $Cd(s) + \frac{1}{2} O_2(g) \rightarrow CdO(s) + 260\ kJ$ **Δ**H = -260 kj

   b. $Cl_2(g) + \frac{7}{2} O_2 + 272\ kJ \rightarrow Cl_2O_7(g)$ **Δ**H =+272 kj

   c. $2Au(s) + \frac{3}{2} O_2 \rightarrow Au_2O_3(s) + 8.36\ kJ$ **Δ**H =-8.36 kj

2. Which of the following equations are endothermic?

   a. $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g)$ $\Delta H = -242\ kJ$

   b. $\frac{1}{2} N_2(g) + \frac{1}{2} O_2(g) \rightarrow NO(g)$ $\Delta H = 88\ kJ$ endothermic

   c. $\frac{1}{2} N_2(g) + \frac{1}{2} O_2(g) + 33.9\ kJ \rightarrow NO_2(g)$ endothermic

   d. $\frac{1}{2} N_2(g) + \frac{3}{2} H_2(g) \rightarrow NH_3(g) + 46\ kJ$

   e. $NH_3(g) \rightarrow \frac{1}{2} N_2(g) + \frac{3}{2} H_2(g)$ $\Delta H = 46\ kJ$ endothermic

3. Identify the following reactions as exothermic or endothermic. Which has more energy the products or the reactants?

   a. $Hg(l) + \frac{1}{2} I_2(s) \rightarrow HgI$ $\Delta H = 138\ kJ$ endothermic   products

   b. $\frac{1}{2} N_2(g) + \frac{3}{2} F_2(g) \rightarrow NF_3$ $\Delta H = -113\ kJ$ exothermic   reactants

   c. $Si(s) + 2\ Br_2(l) \rightarrow SiBr_4(l)$ $\Delta H = -398\ kJ$ exothermic   reactants

d. $Na(s) \rightarrow Na(g)$ $\Delta H = 108.6$ kJ endothermic products

4. Given: $3C(s) + 2Fe_2O_3 + 463.1\ kJ \rightarrow 4Fe(s) + 3\ CO_2(g)$
   Rewrite the equation using one mole of carbon and the ΔH notation.

$C_{(s)} + \frac{2}{3} Fe_2O_3 \rightarrow 4/3\ Fe_{(s)} + CO_{2(g)}$ **Δ**H =154.4 Kj

5. Given: $\frac{1}{2} H_2(g) + \frac{1}{2} Br_2(g) \rightarrow HBr(g)$ $\Delta H = -35.9$ kJ
   Rewrite the equation for one mole of hydrogen gas and include the heat term as part of the equation.

$H_{2(g)} + Br_{2(g)} \rightarrow 2HBr_{(g)} + 71.8$ kj

6. Find the Heat of reaction for $H_2S(g) + \frac{3}{2} O_2(g) \rightarrow H_2O(g) + SO_2(g)$
   Given the following equations:

$H_2(g) + \frac{1}{8} S_8(s) \rightarrow H_2S(g)$ $\Delta H = -20$ kJ

$\frac{1}{8} S_8(s) + O_2(g) \rightarrow SO_2(g)$ $\Delta H = -297$ kJ

$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g)$ $\Delta H = -242$ kJ

ΔH = -519

7. Given the following equations:

$C_2H_2(g) + \frac{5}{2} O_2(g) \rightarrow 2\ CO_2(g) + H_2O(l)$ $\Delta H = -1290$ kJ

$CO_2(g) \rightarrow C(s) + O_2(g)$ $\Delta H = 390$ kJ

$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l)$ $\Delta H = -285$ kJ

Calculate the heat of formation for one mole of acetylene from carbon and hydrogen gas.

ΔH = 225 Kj

8. Calculate the heat of combustion of one mole of liquid methanol (CH_3OH) to carbon dioxide gas and water vapor. The methanol reacts with oxygen gas when combustion occurs. Use chart of "Thermodynamic Values" in your yellow packet to locate the heats of formation along with the following information:
   The formation of liquid methanol from its elements releases 238.6 kJ/mole.

ΔH = -638 kj

9. Given: $4\ C(s) + 5\ H_2(g) \rightarrow C_4H_{10}(g) + 125.4$ kJ
   Rewrite the following equation using one mole of carbon and the ΔH notation.

ΔH = 31.35 kj

Hess's Law with Reactions

1. Calculate the standard enthalpy change, $\Delta H°$, for the formation of 1 mol of strontium carbonate (the material that gives the red color in fireworks) from its elements.

$$Sr\,(s) + C(graphite) + \tfrac{3}{2}O_2(g) \rightarrow SrCO_3(s)$$

The information available is

(1)	$Sr\,(s) + \tfrac{1}{2}O_2(g) \rightarrow SrO\,(s)$	$\Delta H° = -592$ kJ
(2)	$SrO\,(s) + CO_2\,(g) \rightarrow SrCO_3\,(s)$	$\Delta H° = -234$ kJ
(3)	$C(graphite) + O_2\,(g) \rightarrow CO_2\,(g)$	$\Delta H° = -394$ kJ

All equations are in the form needed, so just add up the delta H values -1220 kj

2. The combination of coke and steam produces a mixture called coal gas, which can be used as a fuel or as a starting material for other reactions. If we assume coke can be represented by graphite, the equation for the production of coal gas is

$$2\,C\,(s) + 2\,H_2O\,(g) \rightarrow CH_4\,(g) + CO_2(g)$$

Determine the standard enthalpy change for this reaction from the following standard enthalpies of reaction :

(1)	$C(s) + H_2O\,(g) \rightarrow CO\,(g) + H_2\,(g)$	$\Delta H° = 131.3$ kJ
(2)	$CO\,(g) + H_2O\,(g) \rightarrow CO_2\,(g) + H_2\,(g)$	$\Delta H° = -41.2$ kJ
(3)	$CH_4\,(g) + H_2O\,(g) \rightarrow 3\,H_2\,(g) + CO\,(g)$	$\Delta H° = 206.1$ kJ

Reaction #1 should be multiplied by 2 and reaction #3 should be flipped.
ΔH = 15.3 kj

Both Types of Hess's Law Problems

1. Given the following thermochemical reactions, determine the enthalpy change, ΔH, for the following:

$$NO_{(g)} + \frac{1}{2} O_{2(g)} \rightarrow NO_{2(g)}$$

Reaction	$\Delta H_f°$ (kJ/mol)
$\frac{1}{2} N_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow NO_{(g)}$	+90.4
$\frac{1}{2} N_{2(g)} + O_{2(g)} \rightarrow NO_{2(g)}$	+33.9

ΔH = -56.5

2. Determine ΔH for: $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$

Reaction	$\Delta H_f°$ (kJ/mol)
$C_{(s)} + 2H_{2(g)} \rightarrow CH_{4(g)}$	-74.8
$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$	-393.5
$H_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(l)}$	-285.8

ΔH = -890.3

3. Use chart of "Thermodynamic Values" in your yellow packet to locate the heats of formation along with the following information to determine how much energy is released when methanol (CH_3OH) is burned. $\Delta H_f°$ CH_3OH = -238.6 kJ/mol

$$CH_3OH_{(l)} + 3/2\ O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

ΔH = -638.4

4. Determine ΔH_{rxn} for: $4NH_{3(g)} + 3O_{2(g)} \rightarrow 2N_{2(g)} + 6H_2O_{(g)}$

$\Delta H = -638.4$

Entropy is the degree of randomness in a substance. The symbol for change in entropy is ΔS.

Solids are very ordered and have low entropy. Liquids and aqueous ions have more entropy because they move about more freely, and gases have an even larger amount of entropy. According to the Second Law of Thermodynamics, nature is always proceeding to a state of higher entropy.

Also, when there are more moles of products than moles of reactants, entropy is increasing. When total moles are reduced, entropy is decreasing.

Determine whether the following reactions show an **increase or decrease** in entropy and write the ΔS as (+) or (-).

1. $KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$ ___increases $+\Delta S$___
2. $H_2O(l) \rightarrow H_2O(s)$ __decrease $-\Delta S$
3. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ __ decrease $-\Delta S$
4. $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ __increase $+\Delta S$
5. $KCl(s) \rightarrow KCl(l)$ __increase $+\Delta S$
6. $CO_2(s) \rightarrow CO_2(g)$ __increase $+\Delta S$
7. $H^+(aq) + C_2H_3O_2^-(aq) \rightarrow HC_2H_3O_2(l)$ _decrease $-\Delta S$
8. $C(s) + O_2(g) \rightarrow CO_2(g)$ _decrease $-\Delta S$**
9. $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$ _decrease $-\Delta S$
10. $2N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$ _increase $+\Delta S$
11. $2Al(s) + 3I_2(s) \rightarrow 2AlI_3(s)$ _decrease $-\Delta S$
12. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ _decrease $-\Delta S$
13. $H_2O(g) \rightarrow H_2O(l)$ _decrease $-\Delta S$

Gibbs Free Energy

For a reaction to be thermodynamically favorable, the sign of ΔG (Gibbs Free Energy) must be negative. The mathematical formula for this value is:

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

Where : ΔH = change in enthalpy or heat of reaction,
T = temperature in Kelvin,
ΔS = change in entropy

Complete the table for the sign of ΔG using
(+) OR (-) OR undetermined.

When conditions allow for an undetermined sign of ΔG, temperature will decide favorability.

ΔH	ΔS	ΔG
-	+	-
+	-	+
-	-	undetermined
+	+	undetermined

Answer the questions below:

1. The conditions in which ΔG is always negative is when ΔH is **-** and ΔS is **+**
2. The conditions in which ΔG is always positive is when ΔH is **+** and ΔS is -
3. When the situation is indeterminate, a low temperature favors the (entropy / enthalpy) factor, and a high temperature favors the (entropy / enthalpy) factor.

Answer Problems 4-6 with **"always, sometimes, or never"**.

4. The reaction: $NaOH(s) \rightarrow Na^{+}(aq) + OH^{-}(aq)$ + energy will always be favorable.
5. The reaction: energy + $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ will never be favorable.
6. The reaction: energy + $H_2O(s) \rightarrow H_2O(l)$ will sometimes be favorable.
7. What is the value of ΔG if ΔH = -32.0 kJ, ΔS = +0.250 J/K and T = 293 K? Tell whether this reaction is thermodynamically favorable.

   Delta G is negative this reaction is thermodynamically favorable (spontaneous)

8. What is the value of ΔG if ΔH = +12.0 kJ, ΔS = -0.510 J/K and T = 290 K? Tell whether this reaction is thermodynamically favorable.

Delta G is positive this reaction is thermodynamically not favorable (not spontaneous)

More ΔH, ΔS, ΔG Problems

Use chart of "Thermodynamic Values" in your yellow packet to locate the thermodynamic values that are needed to complete the following problems:

1. Calculate the standard heat of reaction (ΔH^0) for: $CoCO_{3(s)} \rightarrow CoO_{(s)} + CO_{2(g)}$

   90.5 kJ/mol

2. Calculate the change in entropy (ΔS^0) for: $CoCO_{3(s)} \rightarrow CoO_{(s)} + CO_{2(g)}$

   178 J/molK

3. Calculate the ΔG^0 for the reaction in the two problems above ($CoCO_{3(s)} \rightarrow CoO_{(s)} + CO_{2(g)}$) AND determine if this reaction is thermodynamically favorable or not. (Use $\Delta G^0 = \Delta H^0 - T\Delta S^0$ ... Hint: What is standard temperature in this unit?)

   41.6 kJ/mol

   (Circle one)

   Favorable or Non-Favorable

4. Use the table of "Thermodynamic Values" in your Yellow Packet to calculate ΔG^0 for the reaction in the problems above ($CoCO_{3(s)} \rightarrow CoO_{(s)} + CO_{2(g)}$). How does your value compare to you answer for ΔG^o in the problem above?

   42.6 kJ/mol

5. At what temperature is the above reaction thermodynamically favorable? (Assume $\Delta G^0 = 0$ is favorable)

   508K or 235℃

6. Calculate the standard heat of reaction (ΔH^0) for: $4\ NH_{3(g)} + 5\ O_{2(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$

-905.6 kJ/mol

7. Calculate the change in entropy (ΔS^0) for: $4\ NH_{3(g)} + 5\ O_{2(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$

205 J/molK

8. Calculate the ΔG^0 for the reaction in the two problems above ($4\ NH_{3(g)} + 5\ O_{2(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$) AND determine if this reaction is thermodynamically favorable or non-favorable.

-961.6 kJ/mol

(Circle one)

Favorable or *Non-Favorable*

9. What is the ΔG^0 for: $4\ Al_{(s)} + 3\ O_{2(g)} \rightarrow 2\ Al_2O_{3(s)}$

-3160 kJ/mol