

energy



international baccalaureate higher level chemistry

Name _____ Date _____

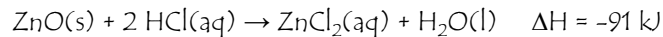
Thermochemistry Experiment

HEAT OF FORMATION (ΔH_f°) OF MgO

Background

The symbol used in thermodynamics for heat of reaction at constant pressure is ΔH . In an exothermic reaction, heat is liberated and the value of ΔH bears a negative sign, indicating that heat is released to the surroundings.

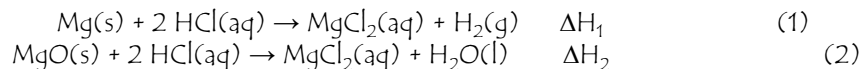
The results are shown as a thermochemical equation:



When the pressure is 1 atmosphere (atm), thermodynamic properties are termed standard, and the enthalpy change is symbolized ΔH° . Reference values of ΔH° and ΔH°_f (see the following discussion) are usually tabulated at 298 K.

Since ΔH varies little with temperature or pressure (especially for reactions in solution), you should expect close agreement between your experimental values and tabulated values of standard enthalpy changes, ΔH° and ΔH°_f .

In this experiment you will carry out the following reactions:



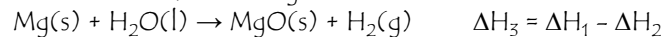
In both reactions (1) and (2) above, the HCl(aq) is in excess, thus the ΔH values are dependent upon the mass of Mg(s) or MgO(s) used. To find ΔH_1 and ΔH_2 above, first find the amount of heat associated with the reaction by

$$q_{\text{acid}} = C_{\text{acid}} m_{\text{acid}} \Delta t_{\text{acid}}$$

(assuming $C_{\text{acid}} = C_{\text{H}_2\text{O}}$ and the $d_{\text{acid}} = d_{\text{H}_2\text{O}}$) according to the solution. Dividing these values by the moles of Mg or MgO used will give the ΔH values in kJ/mol. Hence...

$$\Delta H_x = \frac{q_{\text{reaction}} \text{ (which equals } q_{\text{acid}})}{\text{moles of Mg or MgO used}}$$

According to Hess' law, when one adds two chemical reactions one must also add their ΔH 's to obtain ΔH of the combined reaction. A similar rule holds for subtraction. Thus, subtracting reaction (2) from reaction (1) gives:



For any reaction in which a single compound is formed from its elements in their standard states, the enthalpy change is symbolized ΔH_{of} , the heat of formation of the compound. Standard tables of ΔH_{of} are available for most common compounds, and from them one can obtain ΔH_{of} for any reaction by the following scheme, which can be demonstrated easily by Hess' law:

$$\Delta H_{\text{of}} = \sum \Delta H_{\text{of}}(\text{products}) - \sum \Delta H_{\text{of}}(\text{reactants})$$

Since $\Delta H_{\text{of}}(\text{H}_2\text{O})$ is readily measurable and is well known, you can apply the preceding method to deduce $\Delta H_{\text{of}}(\text{MgO})$ from your data,

$$\Delta H_3 = \Delta H_1 - \Delta H_2 = \Delta H_{\text{of}}(\text{MgO}) + \Delta H_{\text{of}}(\text{H}_2) - \Delta H_{\text{of}}(\text{Mg}) - \Delta H_{\text{of}}(\text{H}_2\text{O})$$

Note that by definition $\Delta H_{\text{of}}(\text{H}_2)$ and $\Delta H_{\text{of}}(\text{Mg})$ must be zero. Hence:

$$\Delta H_{\text{of}}(\text{MgO}) = \Delta H_1 - \Delta H_2 + \Delta H_{\text{of}}(\text{H}_2\text{O})$$

You can check your result against the literature value of $\Delta H_{\text{of}}(\text{MgO})$, which is well established. By using the tabulated ΔH_{of} values for HCl, MgO, and H_2O , you can, using Hess' law, deduce $\Delta H_{\text{of}}(\text{MgCl}_2)$ for each of your experiments, (1) and (2). and see how closely they agree.

Thus, in this experiment, we are determining the heat of the reaction of magnesium metal with hydrochloric acid and also of magnesium oxide with hydrochloric acid, using a simple styrofoam coffee cup calorimeter. The data obtained can then be combined with the known value of the heat of formation of H_2O to determine the heat of formation of MgO.

PROCEDURE

1. Clean and prepare the calorimeter for the experiment.
2. Measure out 100.0 mL of 1.0 M HCl, using a graduated cylinder, into the calorimeter
3. Put the lid on
4. Suspend a digital thermometer in the HCl through a hole in the cover.
5. Cut about 22 cm of Mg ribbon, cleaning it using sandpaper before weighing to the nearest 0.001 g.
6. Record the initial temperature of the HCl.
7. Cut the Mg into pieces no more than 2 cm in length; then add it to the HCl through the blue lid. Close the lid.
8. Stir occasionally with the thermometer and observe the temperature.
9. Record the maximum temperature
10. Make a second determination following Steps 1-7. Compare the data between the two trials. If the data looks consistent, move on. If not, complete a third trial by following steps 1-7 again.
11. Repeat Steps 1-8, using approximately 0.5 g of powdered MgO instead of Mg metal (In the MgO experiment, be careful in handling the powder to ensure that it all gets into the HCl. Stir it right away to disperse it before it cakes on the bottom and overheats the solution locally).

CALCULATIONS/QUESTIONS. Provide your answers on a separate sheet of paper and hand it in.
The calculations for this lab will consist of the following :

1. As was described in the THEORY section, obtain ΔH_1 and ΔH_2 ; then combine these with ΔH_f (H_2O) from a table to find ΔH_f (MgO). Show your work and results below
2. Comparing your value of ΔH of (MgO) to the literature value, calculate the percentage error.
3. Another way to evaluate the consistency of your measurements is to use your values of ΔH_1 and ΔH_2 in conjunction with the literature value of ΔH_f for HCl , H_2O , and MgO to calculate ΔH_f of $MgCl_2(aq)$. Do this first for reaction (1), then reaction (2). Average your experimental values for ΔH_f of $MgCl_2(aq)$ and compare it to the literature value in terms of a percentage error (from the Handbook of Chemistry and Physics: ΔH_f of $MgCl_2(aq)$ = - 801.15 kJ/mol).
4. In this section of the discussion, you will examine the probable error in a number of factors in your determination to see their relative importance and to appreciate the refinements necessary to make precise thermochemical measurements. For each case, comment on the magnitude and direction of this error, as compared to other errors in the method.
5. What is the uncertainty in reading your thermometer? Express this as a percentage of ΔT in a typical experimental run
6. Do the same procedure as in (a) for your balance.
7. Do the procedure as in (a) for your graduated cylinder.
8. List at least two other sources of error that played a role in this experiment. Describe the effect the error had on the data collected.

CONCLUSION

Please summarize the results of your lab with a brief conclusion.

energy:





applied to classical physics involving movement:



energy applied to weight loss



sources

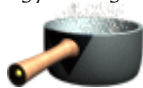








Energy changes in chemical reactions



Endothermic:



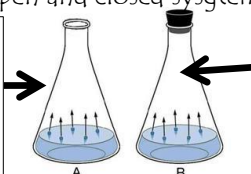
Exothermic:



open and closed systems

Thermodynamics

Thermochemistry



Enthalpy:

Calorimetry:



Standard Enthalpy of formation and Reaction



an average diet is about nutritional calories per day

People burn roughly nutritional calories per day without exercise

one pound of body fat contains nutritional calories

one hour of intense cardio exercise burns nutritional calories

if I eat normal, and run hard one hour per day I should lose nutritional calories per day

and I can expect to lose pound(s) per week.

how can we measure energy? 

by temperature change. For 1 L of water

each °C

= joule (j)

= calories (c)

= British Thermal Units (btu)

= Nutritional Calories (C)

= kilowatt hours (kwh)

Substance	Specific heat J/(g·°C)
Water(l) (liquid)	4.184
Water(s) (ice)	2.03
Water(g) (steam)	2.01
Ethanol(l) (grain alcohol)	2.44
Aluminum(s)	0.897
Granite(s)	0.803
Iron(s)	0.449
Lead(s)	0.129
Silver(s)	0.235
Gold(s)	0.129

forensic usefulness of specific heat
identify the unknown metals

mass of metal (g)	temp. increase (°C)	heat absorbed (J)	what metal is it?
4.68	182	256	
72	91	1336.6	
450	10	4603	



specific heat:



Periodic Trend in Specific Heats

Alkaline earth elements	Specific heat (J/g·°C)
Beryllium	1.825
Magnesium	1.023
Calcium	0.647
Strontium	0.301
Barium	0.204

how many joules of heat are needed to heat one cup (237 mL) of water from room temp. (25 °C) to boiling (100 °C)?






$$Q = mc\Delta T$$

energy and water

standard enthalpies of vaporization and fusion for water:
(energy for phase changes)

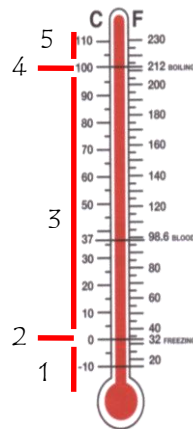
how many Nutritional Calories are required to heat 18 g of water from -20 to 110°C?

consider what happens as it heats up:

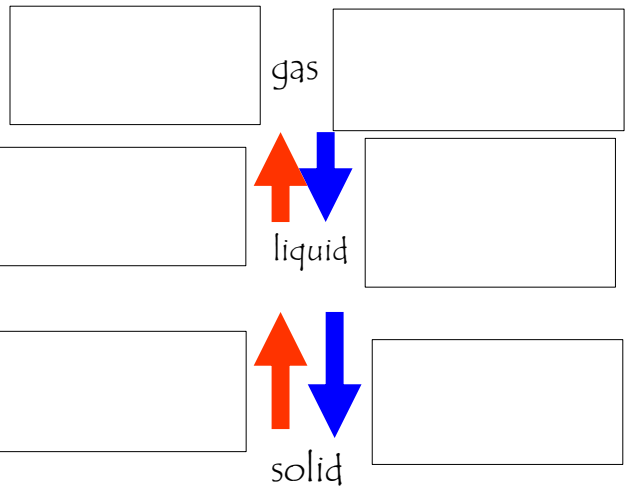
1.  First, it heats up from -20 to 0°C
2.  Then, it melts at 0°C:
3.  Then it heats up from 0 to 100 °C:
4.  Then, it boils at 100°C:
5.  Finally, the steam heats up from 100 to 110°C
(assuming it can pressurize)

Total

compare to Nutritional Calories (C):



water energy
 $C_{\text{ice}} = 2.03 \text{ J/g}^\circ\text{C}$
 $C_{\text{water(l)}} = 4.18 \text{ J/g}^\circ\text{C}$
 $C_{\text{steam}} = 2.01 \text{ J/g}^\circ\text{C}$
 melting: 334 J/g
 boiling: 2260 J/g
 4184 Joules = 1 Nut. Cal.

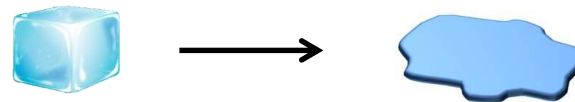




enthalpy

Heat at constant pressure (open system)

Observation:



ΔH

This requires 6010 joules of energy for an 18 gram (1 mole) cube of ice at exactly zero degrees Celsius

We can now write a thermochemical equation:

Q

Problem 1: The combustion of 16 g (1 mol) of methane (CH_4) releases 8,904 joules of heat. Write a thermochemical equation for this reaction.

How much energy is released during the formation of 12 grams of water by the combustion of methane?

= heat change at constant pressure

= ΔH or ΔH_{rxn}

ΔH vs. q ?

$\Delta H = q$ when pressure is constant



calorimetry:

50 g of a metal at 115.0 °C is placed in a fully insulated cup containing 125 g of water at 25.6 °C.

The water heats up to 29.3 °C. Use the table to identify the metal.

Strategy: 1. find heat gained by water (Q_{water}) This is equal to the heat lost by the metal (Q_{metal}). You can now solve for C_{metal} , and look it up.

Substance	Specific heat J/(g·°C)
Water(l) (liquid)	4.184
Water(s) (ice)	2.03
Water(g) (steam)	2.01
Ethanol(l) (grain alcohol)	2.44
Aluminum(s)	0.897
Granite(s)	0.803
Iron(s)	0.449
Lead(s)	0.129
Silver(s)	0.235
Gold(s)	0.129

Standard enthalpy of formation:

(ΔH°) is the heat change that results when **one mole** of a compound is formed from its **elements** at a pressure of 1 atm.

The standard enthalpy of formation of any element in its most stable form is

$$\Delta H_f^\circ (\text{O}_2) =$$

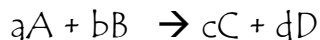
$$\Delta H_f^\circ (\text{O}_3) =$$

$$\Delta H_f^\circ (\text{C, graphite}) =$$

0

$$\Delta H_f^\circ (\text{C, diamond}) =$$

Substance	ΔH_f° (kJ/mol)	Substance	ΔH_f° (kJ/mol)
Ag(s)	0	H ₂ O ₂ (l)	-187.6
AgCl(s)	-127.04	Hg(l)	0
Al(s)	0	I ₂ (s)	0
Al ₂ O ₃ (s)	-1669.8	HI(g)	25.94
Br ₂ (l)	0	Mg(s)	0
HBr(g)	-36.2	MgO(s)	-601.8
C(graphite)	0	MgCO ₃ (s)	-1112.9
C(diamond)	1.90	N ₂ (g)	0
CO(g)	-110.5	NH ₃ (g)	-46.3
CO ₂ (g)	-393.5	NO(g)	90.4
Ca(s)	0	NO ₂ (g)	33.85
CaO(s)	-635.6	N ₂ O ₄ (g)	9.66
CaCO ₃ (s)	-1206.9	N ₂ O(g)	81.56
Cl ₂ (g)	0	O(g)	249.4
HCl(g)	-92.3	O ₂ (g)	0
Cu(s)	0	O ₃ (g)	142.2
CuO(s)	-155.2	S(rhombic)	0
F ₂ (g)	0	S(monoclinic)	0.30
HF(g)	-268.61	SO ₂ (g)	-296.1
H(g)	218.2	SO ₃ (g)	-395.2
H ₂ (g)	0	H ₂ S(g)	-20.15
H ₂ O(g)	-241.8	ZnO(s)	-347.98
H ₂ O(l)	-285.8	ZnS(s)	-202.9



$$\Delta H_{\text{rxn}}^\circ = \sum n \Delta H_f^\circ (\text{products}) - \sum m \Delta H_f^\circ (\text{reactants})$$

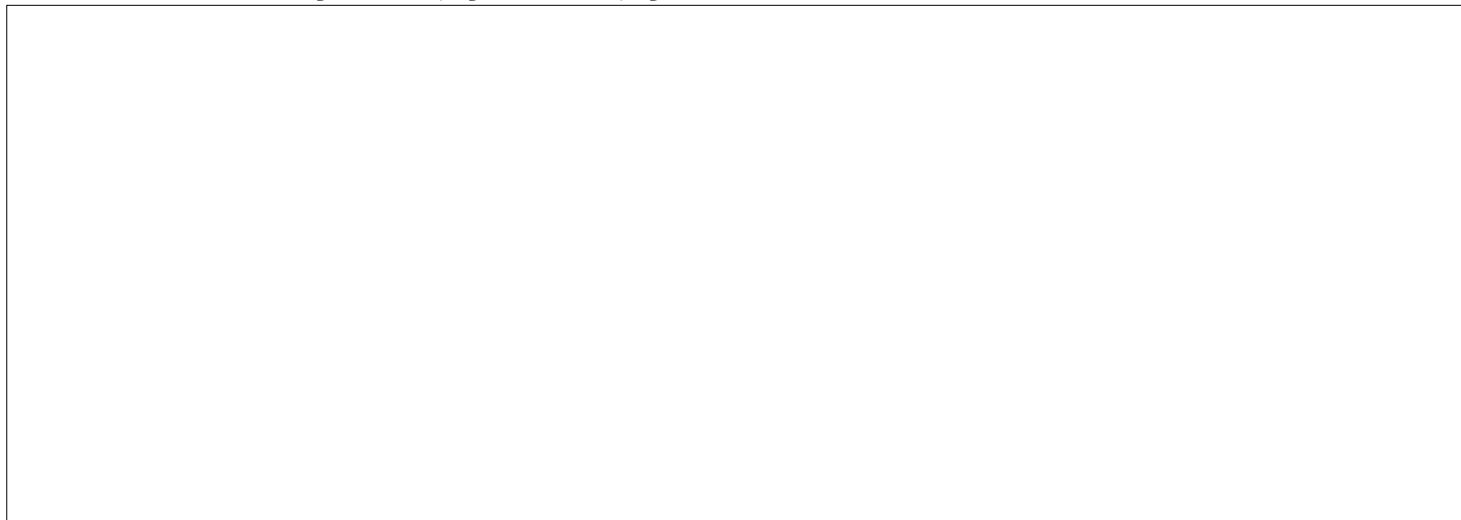
Hess's Law:

reaction energies can be divided into steps

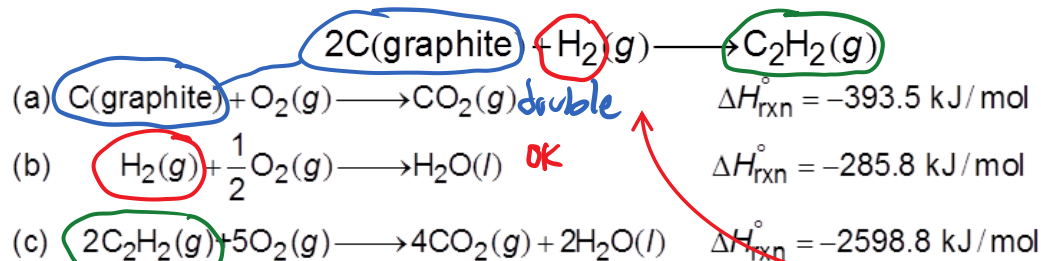
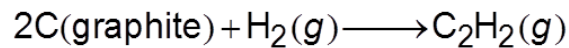
impress your friends by
saying it this way:

Enthalpy is a state function. (It doesn't
matter how you get there, only where you
start and end)

use the following information to determine the standard enthalpy of formation of carbon dioxide; draw an **energy diagram** that relates them.



Calculate the standard enthalpy of formation of acetylene (C_2H_2) from its elements given the data below:



Spontaneous Physical and Chemical Processes...yes or no?

A waterfall runs downhill

A lump of sugar dissolves in a cup of coffee

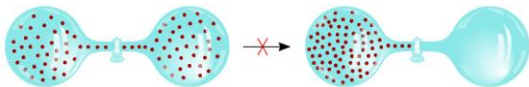
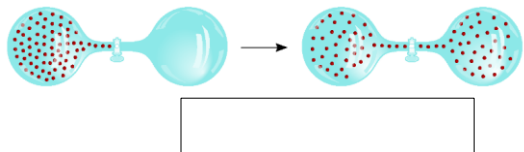
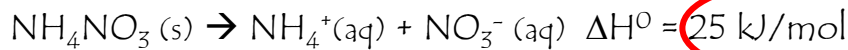
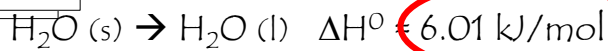
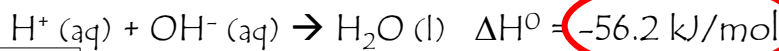
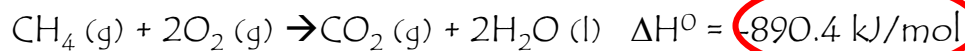
At 1 atm, water freezes below 0 °C and ice melts above 0 °C

Heat flows from a colder object to a warmer object

A gas expands in an evacuated bulb

Iron exposed to oxygen and water forms rust

These are all spontaneous reactions...what would you conclude?



...ΔG is
negative



Entropy (S) is

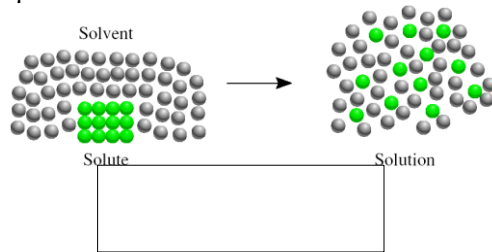
If the change from initial to final results in an **increase** in randomness the value of ΔS is

Predict the relative entropy values for solids, liquids and gases

$$S_{\text{solid}} < S_{\text{liquid}} \ll S_{\text{gas}}$$

For any substance, the solid state is more ordered than the liquid state and the liquid state is more ordered than gas state.

Predict the sign of ΔS for the processes below:



Entropy Changes and gases

When gases are produced (or consumed)

If a reaction produces more gas molecules than it consumes, $\Delta S^\circ > 0$.

If the total number of gas molecules diminishes, $\Delta S^\circ < 0$.

If there is no net change in the total number of gas molecules, then ΔS° may be positive or negative BUT ΔS° will be a small number.

Predict whether the entropy change of the system in each of the following reactions is positive or negative.

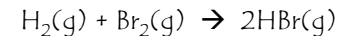
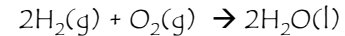


Table 18.1**Standard Entropy Values (S°) for Some Substances at 25°C**

Substance	S° (J/K · mol)
H ₂ O(l)	69.9
H ₂ O(g)	188.7
Br ₂ (l)	152.3
Br ₂ (g)	245.3
I ₂ (s)	116.7
I ₂ (g)	260.6
C (diamond)	2.4
C (graphite)	5.69
CH ₄ (methane)	186.2
C ₂ H ₆ (ethane)	229.5
He(g)	126.1
Ne(g)	146.2

what patterns do you observe for these standard entropy values?

predict ΔS = positive or negative

freezing ethanol

evaporating a beaker of liquid bromine at room temperature

dissolving glucose in water

cooling nitrogen gas from 80°C to 20°C

ΔS and microstates:

First Law of Thermodynamics

Second Law of Thermodynamics

Spontaneous process:

$$\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} > 0$$

Equilibrium process:

$$\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} = 0$$

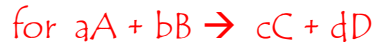
$\Delta G^\circ_{\text{rxn}}$ and ΔG°_f
define:

The standard free-energy of reaction (ΔG°) is the free-energy change for a reaction when it occurs under standard-state conditions.

Standard free energy of formation (ΔG°_f) is the free-energy change that occurs when **1 mole** of the compound is formed from its elements in their standard states.

how are they calculated?:

$$\Delta G^\circ_{\text{rxn}} = \sum n \Delta G^\circ_f(\text{products}) - \sum m \Delta G^\circ_f(\text{reactants})$$



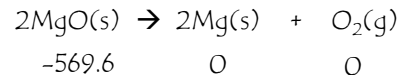
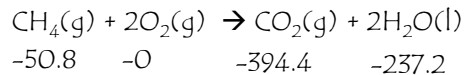
$$\Delta G^\circ_{\text{rxn}} = [d\Delta G^\circ_f(C) + d\Delta G^\circ_f(D)] - [a\Delta G^\circ_f(A) + b\Delta G^\circ_f(B)]$$

look them up!
 ΔG°_f of any element in its stable form is zero.

calculating standard free energy changes

Calculate the standard free-energy changes for the following reactions at 25°C

(given in kJ/mol).



Free energy predictions based on the sign of enthalpy, entropy, and temperature

Temperature and Spontaneity of Chemical Reactions



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

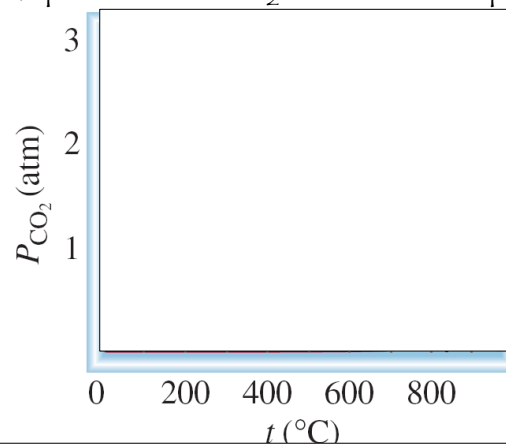
$$\Delta G^\circ = 130.0 \text{ kJ/mol at } 25^\circ\text{C}$$

$$\Delta G^\circ = 0 \text{ at } 835^\circ\text{C}$$

predict the partial pressure of CO_2 over this temperature range.

Free energy predictions:

ΔH	ΔS	ΔG
+	+	
+	-	
-	+	
-	-	



explanation:

Thermochemistry

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Miscellaneous

A different perspective on energy: chemteam

additional topics

- bond energy
- ozone depletion
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- born-haber cycle
- enthalpy of solution
- hydration enthalpy