

Day	Topic	Homework*
3.6	Quiz 3.1-3.2 Nuclear Introduction to Thermodynamics and heat flow.	Watch flip over calorimetry
3.7	Measuring Heat using Calorimetry	
3.8	How many calories are in a Cheeto?	
	Thanksgiving Break	Watch flip over heating/cooling curves
3.9	Enthalpy of phase changes Can you calculate the enthalpy of fusion for water?	Watch flip over Heats of Reaction
3.10	Heats of Reaction using heat of formation	Watch flip over Hess's Law
3.11	Quiz Hess's Law	
3.12	Which material would make the best slide?	
3.13	Exam 3	

* All homework is suggested. Solutions will be provided for all problems.

UNIT 3 Thermochemistry Objectives

- Student will understand energy and its forms, including kinetic, potential, chemical, and thermal energies
- Student will understand the law of conservation of energy and the processes of heat transfer
- Student will use thermochemical equations to calculate energy changes that occur in chemical reactions and classify reactions as exothermic or endothermic.
- Student will perform calculations involving heat, mass, temperature change, and specific heat
- Student will use calorimetry to calculate the heat of a chemical process.

Thermodynamics VOCABULARY

- Heat (q)
- System
- Surroundings
- Thermal equilibrium
- Calorimetry/Calorimeter
- Calorie (cal)
- (Dietary) Calorie (Cal)
- Kilocalorie
- Joule (J)
- Specific heat (capacity)
- Thermochemical equation
- Standard heat of formation (ΔH_f°)
- Hess's law
- Exothermic
- Endothermic
- Enthalpy (H)/Heat of reaction (ΔH)
- Entropy (S)/ Entropy of reaction (ΔS)
- Gibbs Free Energy (G)/Free energy change (ΔG)
- Latent heat
- Heating curve
- Molar enthalpy/heat of
 - fusion (ΔH_{fus})
 - solidification (ΔH_{solid})
 - vaporization (ΔH_{vap})
 - condensation (ΔH_{cond})
- Energy
- Work
- Change of state
- Solid

- Liquid

Lesson 3.6: Introduction to Thermodynamics and Heat Flow

Define Energy:

In the context of this chemistry course, this means that work is done when an object moves some distance in response to a force being applied. **Work = force x distance.** How can work be measured when using atoms, ions, or molecules?

Types of energy

- Mechanical Energy
 - Potential or stored energy
 - Wound up spring
 - Gasoline before burning
 - Kinetic energy or energy of motion
 - Moving objects
- Non-mechanical energy
 - Chemical
 - Electrical
 - Electromagnetic (radiant)
 - Sound
 - Heat
 - Magnetic

Chemical Potential Energy:

Energy is often converted from one form to another. In a car the chemical energy of the fuel is converted to heat energy, which is then converted to mechanical energy as gases expand to force the pistons to move.

Law of Conservation of Mass/Energy

As with the conservation of mass, energy cannot be either created or destroyed in a reaction. Einstein's research led to the discovery that mass can be converted into energy ($E=mc^2$)

Heat (q): energy that transfers from one object to another based on temperature differences. Heat always moves from hot to cold or high average kinetic energy to low average kinetic energy

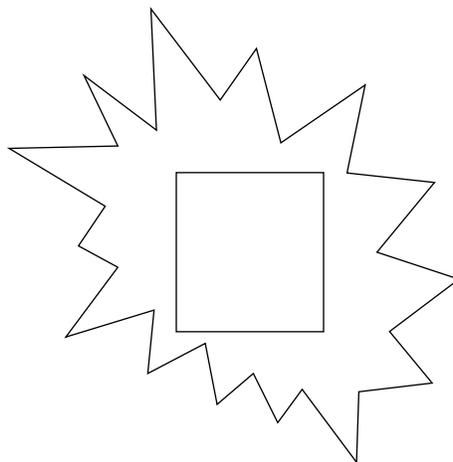
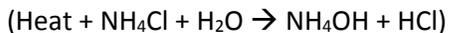
System:

Surroundings:

Universe:

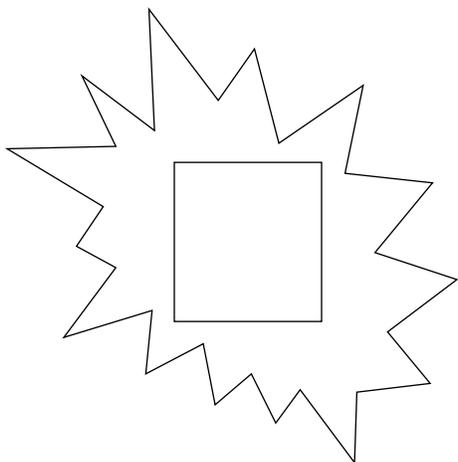
Law of Conservation of Energy: energy is never created nor destroyed it can all be accounted for in terms of work, stored energy, or heat

Endothermic: system absorbs energy from surroundings (+ q value). Energy (heat) is absorbed as the products of the reaction have more potential energy than reactants.



Examples

- Photosynthesis
- Melting and evaporation



Exothermic: system releases heat to surroundings (- q value). Energy (heat) is released as the products of the reaction have less potential energy than reactants. ($\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + \text{Heat}$).

Examples

- Burning
- Rusting of iron
- Condensation
- Freezing

Units for measuring energy

- Joule (J) - SI unit for heat and energy named after James Prescott Joule
- calorie (cal) - quantity of heat needed to raise the temperature of 1 g (1 ml or 1 cc or 1 cm³) of pure H₂O by 1.0 C°.

$$1 \text{ J} = 0.2390 \text{ cal or } 4.184 \text{ J} = 1 \text{ cal}$$

When reading nutrition labels, the number of calories stated are given as dietary calories, Cal (big C calories).

$$1 \text{ Calorie (dietary calorie)} = 1 \text{ kilocalorie} = 1000 \text{ calories}$$

The amount of heat that can be absorbed or transferred by a substance is dependent on the make-up of the substance. The amount of heat able to be absorbed or transferred can be calculated using:

- **Heat capacity (C):** amount of energy needed to raise or lower a substance 1°C. Heat capacity depends on makeup of substance as well as mass of substance
 - 1 mL of water takes less heat than 100 ml of water to raise 1 degree C
- **Specific Heat Capacity (c):** the amount of heat energy needed to raise or lower the temperature of 1 g of a substance by 1 °C.
 - Each substance has its own specific heat capacity which is defined as the amount of heat required to raise one mass unit 1 °C (J/g°C or cal/g°C)

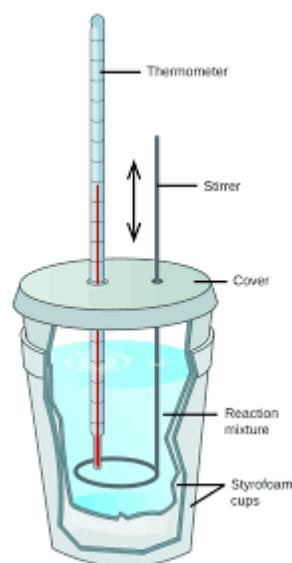
We will be using specific heat capacity in class to solve our heat related problems.

3.7 Measuring Heat using Calorimetry

If heat is transferred a temperature change or a phase change will be observed. The amount of energy that is transferred as heat is dependent on the materials present, the amount (grams) of materials, and the temperature of the material and the size of the temperature change. The amount of heat transferred can be calculated as follows:

$$q = m c \Delta T$$

Where q is heat (J or calories), m is mass (grams), c is specific heat, and ΔT is change in temperature (°C). Heat will flow from an object with more energy (higher temperature) to an object with less energy (lower temperature) until both objects reach the same temperature. Heat will always flow from high to low. It is possible to determine the heat transferred using **CALORIMETRY**. A calorimeter is a device used to measure the amount of heat transferred in a chemical or physical change. There are two types of calorimeters that can be used. A constant pressure calorimeter (coffee cup calorimeter image to the right) or a constant volume calorimeter (bomb calorimeter). Typically in both instances a substance is submerged in a known quantity of water and the temperature change of the water is measured and the heat gained or lost by the object is calculated based on the principle that energy is conserved. Whatever energy was gained or lost by the water is the same energy gained or lost by the object ($q_{\text{wat}} = q_{\text{obj}}$).



Specific Heat and Calorimetry

1. How much heat is required to raise the temp of 654 g of water from 34.5°C to 89.7°C?

Ans _____

2. How much heat is required to raise the temp of 654 g of silver from 34.5°C to 89.7°C?

Ans _____

3. If 7350 J were added to 152 g of ethanol, its temp would go up by how much?

Ans _____

4. 16.25 g of water at 54.0°C releases 402.7 J. What will be its final temp?

Ans _____

5. 697 J are added to a 36.8 g of kerosene and the temp increases from 22.5°C to 34.7°C. Determine kerosene's specific heat.

Ans _____

6. 25 copper pennies (each weighing 3.12 g) are placed in 36.0 g of ethanol at room temp (22.1°C). How much heat will it take to raise the temperature up to 65.8°C?

Ans _____

7. What mass of 54.0°C water must be added to 468 g of 21.0°C water to make the final temp of both come out to be 29.0°C?

Ans _____

8. What mass of 54.0°C gold must be added to 468 g of 21.0°C water to make the final temp of both come out to be 29.0°C?

substance	c (J/g°C)
water	4.184
ethanol	2.452
graphite	0.720
diamond	0.502
iron	0.444
copper	0.385
silver	0.237
gold	0.129
ice	2.092

Ans _____

9. A 325 g brass rod at 100.0°C is placed in a cup containing 162 g of 24.3°C water. The final temp comes out to be 37.4°C. Determine brass's specific heat.

Ans _____

10. 100.0 g of water at 20.0°C are mixed with 200.0 g of copper at 40.0°C. What will the final temp come out to be?

Ans _____

Ans (IRO+1): 0.436 1.55 19.7 23.1 29.5 48.1 150 4860 5170 8560 151,000 units (IRO+1): J J J °C °C °C °C g g J/g°C J/g°C

3.9: Enthalpy of Phase Changes

All matter is made of uncountable numbers of particles. All particles are in constant motion. Even solid objects are in motion, even though we cannot detect motion. With two bodies in contact, energy will transfer from the body with the higher energy to the body with the lower amount of energy until the temperatures of each body are the same. Melting of ice is an example of this transfer. Heat enters melting ice, but as long as both solid and liquid are present the temperature does not change. The energy is being used to rearrange the molecules. The potential energy of the liquid water is higher than the ice, therefore the melting of ice is endothermic. Evaporation of sweat is another example.

More Thermodynamics:

Remember: If $q > 0$ we say that the object gained heat - heat flowed into our sample (endothermic), the system absorbed heat. Likewise if $q < 0$ we say that our object lost heat, heat flowed out of our sample (exothermic). The system lost heat.

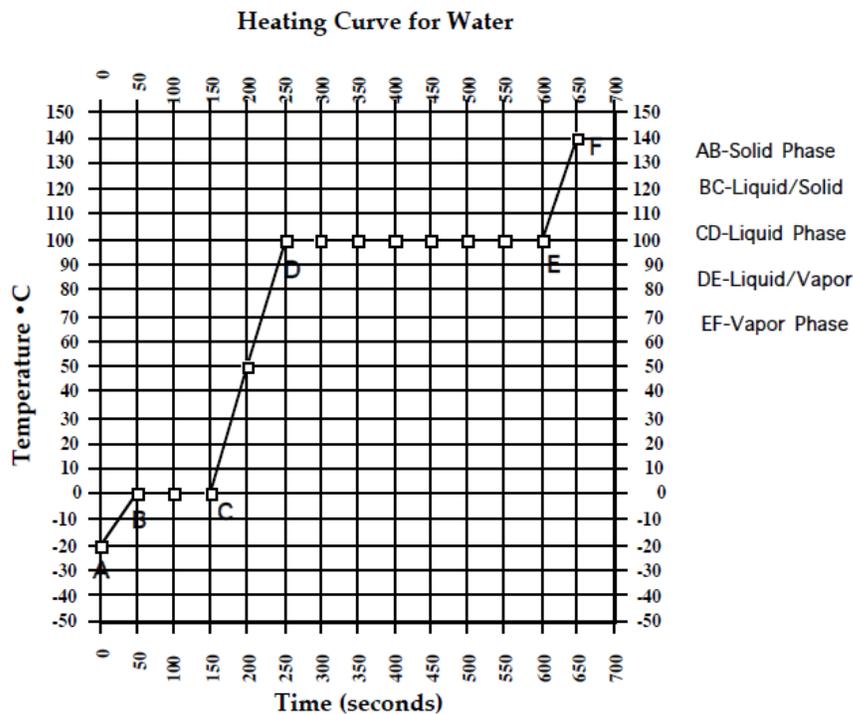
The temperature change, ΔT , is also related to heat. It can be positive or negative. It is obvious that if $\Delta T > 0$ then the temperature increased and if $\Delta T < 0$ then the temperature decreased.

Clearly whenever $\Delta T > 0$ then $q > 0$, and vice versa.

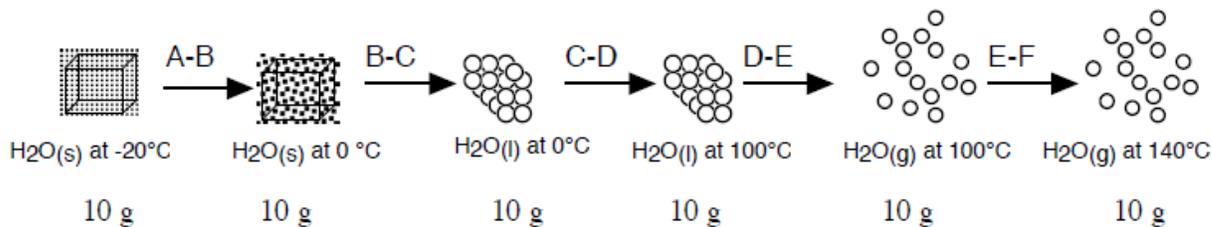
Changes of State: If you have a glass of ice water (the glass contains both ice and liquid water) and you add some heat to the ice water the temperature does not change. The heat melts some the ice rather than changing the temperature of the system. During a change of state even though energy is being

transferred all of the energy is being used to change the placement of the particles and **NO** change in temperature will be observed.

The amount of heat required to change 10.0 g of ice at -20.0°C to steam at $140.^{\circ}\text{C}$ can be determined by breaking up the **heating curve**, below, into 5 different segments. A heating curve is used to show the temperature change over time for a specific substance. If the curve is read from right to left it is called a cooling curve and the energy would be negative, because it is being given off.



The diagram below illustrates the steps involved to convert 10 g of solid ice at -20.0°C to 10.0 g of gaseous steam at $140.^{\circ}\text{C}$.



Based on the diagram above the 10.0 g of water is ice from A-B, at B the ice begins to melt until time C. From time B to C the temperature remains constant even though energy is being added. The additional energy is being used to break apart the attractions between the molecules. At time C the now all melted ice (water) has a temperature of 0.0°C and with the addition of energy the temperature will begin to rise until time D. At time D, the temperature is now $100.^{\circ}\text{C}$, the boiling point of water, and steam will begin to form. While both water and steam are present the temperature will remain at $100.^{\circ}\text{C}$ until complete conversion of all water to steam. Just like the ice melting, energy is still being added but all of the energy is required to break the attractions between the molecules. At time E all of the water has been converted to steam and the temperature can begin to rise again. Looking at each of the individual segments we can begin to calculate the total energy required to raise the temperature from -20.0°C to

140.°C. Whenever there is a temperature change (only one phase present) the following can be used to calculate the amount of heat required for the given temperature change. The specific heat is specific to the phase of the substance present.

$$Q=mc\Delta T$$

Because during a phase change the temperature is constant the above equation cannot be used and the heat is calculated using the **latent heat** (ΔH , J/g) for the particular phase change. The latent heat is the amount of energy required to convert 1.0 gram of a substance from one phase to another phase.

$$Q=m\Delta H_x$$

For water:

Specific heats (c), J/(g°C):

Ice = 2.092

Water = 4.186

Steam = 2.006

Latent Heats (ΔH), J/g:*

Heat of fusion: 334

Heat of Vaporization: 2260

*If this was a cooling curve, all of the heats would be negative because this amount of heat would be released.

To solve the above equation we need to solve for the amount of heat required in each of the segments.

A-B

B-C

C-D

D-E

E-F

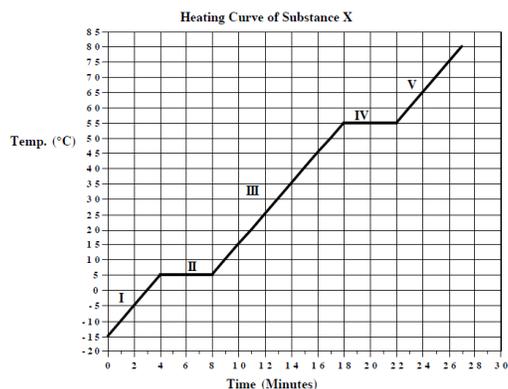
Check to make sure the heats from the individual regions are the same unit and sum.

Sample calculations:

1) How much heat (in J) does it take to vaporize 1.000 L of H₂O at 100°C? (1 mL of H₂O = 1 g)

2) How much heat (in J) is released when 10.87 mL of liquid H₂O freezes at 0°C?

3) The heating curve shown above is a plot of temperature vs time. It represents the heating of substance X at a constant rate of heat transfer. Answer the following questions using this heating curve:



- _____ a. In what part of the curve would substance have a definite shape and definite volume?
- _____ b. In what part of the curve would substance X have a definite volume but no definite shape?
- _____ c. In what part of the curve would substance X have no definite shape or volume?
- _____ d. What part of the curve represents a mixed solid/liquid phase of substance X?
- _____ e. What part of the curve represents a mixed liquid/vapor phase of substance X?
- _____ f. What is the melting temperature of substance X?
- _____ g. What is the boiling temperature of substance X?
- _____ h. In what part(s) of the curve would increasing kinetic energy be displayed?
- _____ i. In what part(s) of the curve would increasing potential energy be displayed?
- _____ j. In what part of the curve would the molecules of substance X be farthest apart?
- _____ k. In what part of the curve would the molecules of X have the lowest kinetic energy?
- _____ l. In what part of the curve would the molecules of X have the greatest kinetic energy?

Given the following information for substance X:

Specific Heat (c) of solid phase:	2.000 J/g°C
Heat of Fusion (ΔH_{fus}):	100.0 J/g
Specific Heat (c) of liquid phase:	5.000 J/g°C
Heat of Vaporization (ΔH_{vap}):	1000. J/g
Specific Heat (c) of vapor phase:	1.000 J/g°C

Heat Equations:

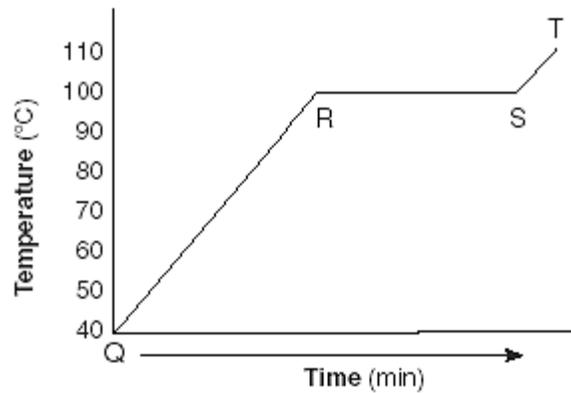
$Q = mc\Delta T$	$Q = m\Delta H_{fus}$	$Q = m\Delta H_{vap}$
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4) Use the information above to answer the following questions.

(a) Draw a cooling curve for substance X from 70.0°C to -10.0°C .

(b) Calculate the total heat energy in Joules needed to convert 20.0 g of substance X from -10.0°C to 70.0°C ?

5) A sample of water is heated from a liquid at 40°C to a gas at 110°C . The graph of the heating curve is shown below.



(a) Label each of the following regions on the above heating curve.

- Liquid, only
- Gas, only
- Phase change

(b) For section QR of the graph, state what is happening to the water molecules as heat is added.

(c) For section RS of the graph, state what is happening to the water molecules as heat is added.

(d) Calculate the heat required to raise the temperature from 40.0°C to 110.0°C .

3.10 Heat of Reactions using Heats of Formation

Enthalpy (ΔH)

Enthalpy is defined as heat at constant pressure. The **enthalpy of a reaction** (ΔH_{rxn}) can be calculated knowing how much heat was required to form the compounds used within the reaction. Each compound or molecules **heat of formation**, kJ/mol, can be calculated knowing that when a substance is formed at standard state (1.0 atm and 25.0°C) from its pure elements, at the same condition, all of the heat accompanying the reaction is due to the formation of the compound. This process can be exothermic or endothermic. The heats of formation for select elements can be found at the end of the packet.

The heat of reaction for any reaction can be calculated using the following formula.

$$\Delta H_{\text{reaction}} = \sum nH_f \text{ products} - \sum nH_f \text{ reactants}$$

If the calculated heat is negative the process is exothermic and if it is positive it is endothermic. The coefficients in a thermochemical reaction indicate the number of moles involved. It is possible to use fractions as coefficients in this kind of equation because they represent moles. Always read a chemical equation from left to right.

Example:

Given a simple chemical equation with the variables A, B and C representing different compounds:



and the standard enthalpy of formation values:

- $\Delta H_f^\circ[A] = 433 \text{ KJ/mol}$
- $\Delta H_f^\circ[B] = -256 \text{ KJ/mol}$
- $\Delta H_f^\circ[C] = 523 \text{ KJ/mol}$

the equation for the standard enthalpy change of formation is as follows:

$$\Delta H_{\text{reaction}}^\circ = n_C \Delta H_f^\circ[C] - (n_A \Delta H_f^\circ[A] + n_B \Delta H_f^\circ[B])$$

$$\Delta H_{\text{reaction}}^\circ = (1 \text{ mol})(523 \text{ kJ/mol}) - ((1 \text{ mol})(433 \text{ kJ/mol}) + (1 \text{ mol})(-256 \text{ kJ/mol}))$$

Because there is one mole each of A, B and C, the standard enthalpy of formation of each reactant and product is multiplied by 1 mole, which eliminates the mol denominator:

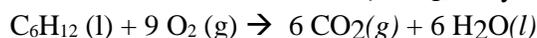
$$\Delta H_{\text{reaction}}^\circ = \mathbf{346 \text{ kJ}}$$

The result is 346 kJ, which is the standard enthalpy change of formation for the creation of variable "C" (LibreTexts Libraries, 2016).

If the heat of reaction is known the above equation can be used to calculate an unknown heat of formation for one of the compounds within the reaction.

Practice Problem (Anderson, 2016):

1. How many kJ are released when 1 mole of $\text{Pb}(\text{NO}_3)_2$ is formed from its elements? (see reference sheet):
2. Given the following reaction: $\text{Si}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SiO}_2(\text{s})$ $\Delta H_f = -910.7 \text{ kJ/mol}$ Calculate ΔH_f for 2 moles of SiO_2 : _____
3. Calculate ΔH_f for 0.5 moles of SiO_2 : _____
4. What is ΔH_f for this reaction? $2 \text{ P}(\text{s}) + 5/2 \text{ O}_2(\text{g}) \rightarrow \text{P}_2\text{O}_5(\text{s})$ _____
5. What is ΔH_f for this reaction? $4 \text{ P}(\text{s}) + 5 \text{ O}_2(\text{g}) \rightarrow 2 \text{ P}_2\text{O}_5(\text{s})$ _____
6. How many kJ are released when 1 mole of beer alcohol ($\text{C}_2\text{H}_5\text{OH}$) is burned? _____
(hint: write the combustion reaction for this reaction and see table of "heats of combustion")
7. Octane (in gasoline) is quite flammable. Change this sentence into a balanced equation: "When 1 mole of octane (C_8H_{16}) is burned, _____ kJ of heat energy is released." (hint: use fractions to balance the equation, and see reference sheet for values)
8. What is the value for ΔH_{comb} when 249.2 g of sugar ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) is burned? Write a balanced equation first.
9. The standard enthalpy of combustion of cyclohexane, $\text{C}_6\text{H}_{12}(\text{l})$, to $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{l})$ at 25°C is -3924 kJ/mol . Calculate the standard heat of formation, ΔH_f° , of cyclohexane.



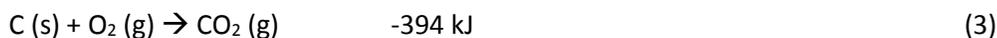
Ans (IRO+1): -6019.8, -5470.5, -4110 -3009.9, -2014.1, -1821.4, -1366.8, -455.4, -153, -125.4, 451.9 Units: kJ/mol (for all)
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3.11 Hess's Law

Hess's Law in chemistry is a practical application of the transitive law in math. According to the transitive law if $A = B$ and $B = C$, therefore $A = C$. In chemistry we use this law to simplify reactions. Usually a reaction does not take place in one step but in multiple steps. For example carbon monoxide can be converted to carbon dioxide by the addition of oxygen.



The overall heat of this reaction can be calculated by looking at the individual heats of reaction for the formation of 1 mole of carbon monoxide and carbon dioxide.



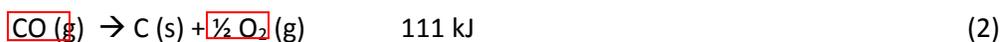
If the above equations (2 & 3) are manipulated, meaning multiplied by a coefficient or reversed, the overall reaction (1) can be obtained.

Follow the steps below to solve any Hess's Law problem.

1. Balance the individual equations
2. If necessary look up standard enthalpies
3. Put a square around the substances that should be reactants. Circle everything that should be products.
4. Flip equations around if necessary to cancel out terms on opposite sides and ensure reactants are reactants and products are products.
5. Changing the equation around requires a sign change of the H of that individual step
6. Sum up the individual steps



Checking to ensure all substances that need to be reactants are on the left hand side of the arrow and everything that needs to be products are on the right hand side of the arrow, it is observed that carbon monoxide (CO) needs to be a reactant, therefore reaction number 2 needs to be reversed. Once it is reversed the sign on the heat also needs to be reversed. Also oxygen is on both sides of the equation, but that is okay because only half of a mole is required.



Knowing the reactions are balanced the next step would be to cancel like terms on opposite sides of the arrow. Solid carbon is found on opposite sides of the equation and therefore can be cancelled as well as $\frac{1}{2}$ oxygen. Once all like terms have been cancelled the next step is to sum up the reactants and products and the heats of reactions to give the overall reactions.



One the reactions have been summed, as well as the heats, the overall heat required to convert carbon monoxide to carbon dioxide is -283 kJ.

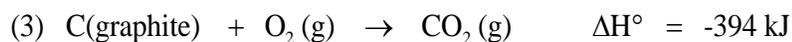
To sum it up, Hess's law says that if two or more equations for which the enthalpy is given or otherwise known can be added together to create a target equation, the enthalpy changes can also be summed to find the enthalpy change of the target equation.

HESS'S LAW PRACTICE PROBLEMS

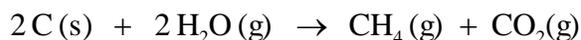
- 1) Calculate ΔH° for the formation of 1 mol of strontium carbonate (the material that gives the red color in fireworks) from its elements.



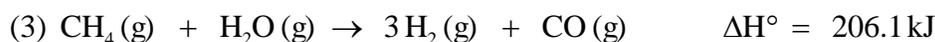
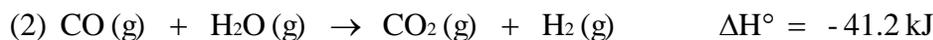
The information available is



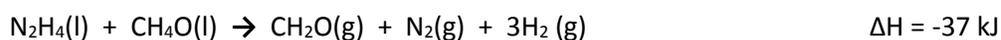
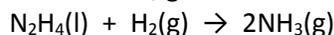
- 2) The combination of coal and steam produces a mixture called coal gas, which can be used as a fuel or as a starting material for other reactions. If we assume coal can be represented by graphite, the equation for the production of coal gas is:



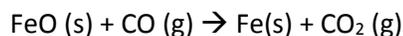
Determine the standard enthalpy change for this reaction from the following standard enthalpies of reaction :



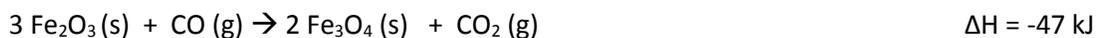
- 3) Find the ΔH for the reaction below, given the following steps and subsequent ΔH values:



- 4) One reaction involved in the conversion of iron ore to the metal is:



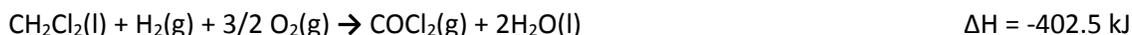
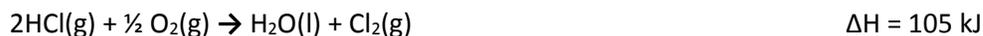
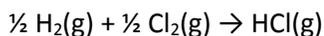
Calculate the standard enthalpy change for this reaction from the following reactions of iron oxides with carbon monoxide.



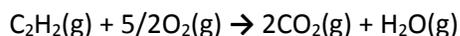
5) Find the ΔH for the reaction below, given the following steps and subsequent ΔH values:



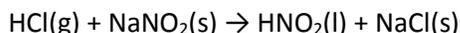
6) Find the ΔH for the reaction below, given the following steps and subsequent ΔH values:



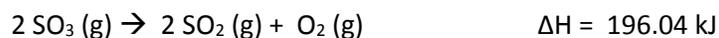
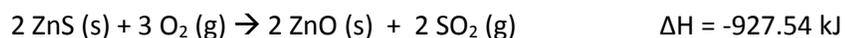
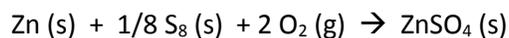
7) Find the ΔH for the reaction below, given the following steps and subsequent ΔH values:



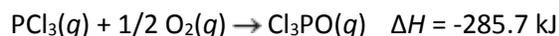
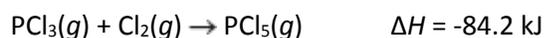
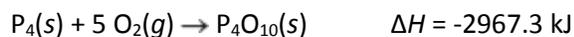
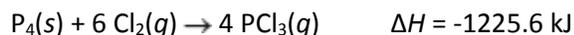
8) Find the ΔH for the reaction below, given the following steps and subsequent ΔH values:



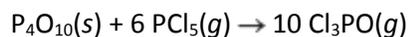
9) Calculate the standard enthalpy change for this reaction from the combustion of Zn (g) with oxygen:



10) You are given the following data:



Calculate ΔH for the following reaction.



Works Cited

Anderson, J. (2016, May 1). *Honors Chemistry*. Retrieved from Mr. A's Chemistry:

<http://www.pkwy.k12.mo.us/west/teachers/anderson/honors/honors.html>

LibreTexts Libraries. (2016, July 13). *Standard Enthalpy of Formation*. Retrieved from Chemistry

LibreTexts: Given a simple chemical equation with the variables A, B and C representing different compounds:

Heat of Formation

Substance	State	ΔH_f	Substance	State	ΔH_f
ammonia	<i>g</i>	-45.9	lead(IV) oxide	<i>s</i>	-274.5
ammonium chloride	<i>s</i>	-314.4	lead(II) nitrate	<i>s</i>	-451.9
ammonium sulfate	<i>s</i>	-1180.9	lead(I) sulfate	<i>s</i>	-919.94
barium chloride	<i>s</i>	-858.6	lithium chloride	<i>s</i>	-408.6
barium nitrate	<i>s</i>	-992.1	lithium nitrate	<i>s</i>	-483.1
barium sulfate	<i>s</i>	-1473.2	magnesium chloride	<i>s</i>	-641.5
benzene	<i>g</i>	+82.88	magnesium oxide	<i>s</i>	-601.6
benzene	<i>l</i>	+49.080	magnesium sulfate	<i>s</i>	-1284.9
calcium carbonate	<i>s</i>	-1207.6	manganese(IV) oxide	<i>s</i>	-520.0
calcium chloride	<i>s</i>	-795.4	manganese(II) sulfate	<i>s</i>	-1065.3
calcium hydroxide	<i>s</i>	-983.2	mercury(I) chloride	<i>s</i>	-264.2
calcium nitrate	<i>s</i>	-938.2	mercury(II) chloride	<i>s</i>	-230.0
calcium oxide	<i>s</i>	-634.9	mercury(II) oxide (red)	<i>s</i>	-90.8
calcium sulfate	<i>s</i>	-1434.5	methane	<i>g</i>	-74.9
carbon (diamond)	<i>s</i>	+1.9	nitrogen dioxide	<i>g</i>	+33.2
carbon (graphite)	<i>s</i>	0.00	nitrogen monoxide	<i>g</i>	+90.29
carbon dioxide	<i>g</i>	-393.5	dinitrogen monoxide	<i>g</i>	+82.1
carbon monoxide	<i>g</i>	-110.5	dinitrogen tetroxide	<i>g</i>	+9.2
copper(II) nitrate	<i>s</i>	-302.9	oxygen (O ₂)	<i>g</i>	0.00
copper(II) oxide	<i>s</i>	-157.3	ozone (O ₃)	<i>g</i>	+142.7
copper(II) sulfate	<i>s</i>	-771.4	diphosphorus pentoxide	<i>s</i>	-3009.9
ethane	<i>g</i>	-83.8	potassium bromide	<i>s</i>	-393.8
ethyne (acetylene)	<i>g</i>	+228.2	potassium chloride	<i>s</i>	-436.49
hydrogen (H ₂)	<i>g</i>	0.00	potassium hydroxide	<i>s</i>	-424.58
hydrogen bromide	<i>g</i>	-36.29	potassium nitrate	<i>s</i>	-494.6
hydrogen chloride	<i>g</i>	-92.3	potassium sulfate	<i>s</i>	-1437.8
hydrogen fluoride	<i>g</i>	-273.3	silicon dioxide (quartz)	<i>s</i>	-910.7
hydrogen iodide	<i>g</i>	+26.5	silver chloride	<i>s</i>	-127.01 = 0.5
hydrogen oxide (water)	<i>g</i>	-241.8	silver nitrate	<i>s</i>	-120.5
hydrogen oxide (water)	<i>l</i>	-285.8	silver sulfide	<i>s</i>	-32.59
hydrogen peroxide	<i>g</i>	-136.3	sodium bromide	<i>s</i>	-361.8
hydrogen peroxide	<i>l</i>	-187.8	sodium chloride	<i>s</i>	-385.9
hydrogen sulfide	<i>g</i>	-23.9	sodium hydroxide	<i>s</i>	-425.9
iodine (I ₂)	<i>s</i>	0.00	sodium nitrate	<i>s</i>	-467.9
iodine (I ₂)	<i>g</i>	+62.4	sodium sulfate	<i>l</i>	-1387.1
iron(II) chloride	<i>s</i>	-399.4	sulfur dioxide	<i>g</i>	-296.8
iron(III) oxide	<i>s</i>	-825.5	sulfur trioxide	<i>g</i>	-395.7
iron(III) oxide	<i>l</i>	-863.0	tin(IV) chloride	<i>l</i>	-511.3
iron(II) sulfate	<i>s</i>	-928.4	zinc nitrate	<i>s</i>	-483.7
iron(II) sulfide	<i>s</i>	-100.0	zinc oxide	<i>s</i>	-350.5
lead(II) oxide	<i>s</i>	-217.3	zinc sulfate	<i>s</i>	-980.14

ΔH_f is heat of formation of the given substance from its elements. All values of ΔH_f are expressed as kJ/mol at 25°C. Negative values of ΔH_f indicate exothermic reactions. *s* = solid, *l* = liquid, *g* = gas

TABLE A-5 HEAT OF COMBUSTION

Substance	Formula	State	ΔH_c	Substance	Formula	State	ΔH_c
hydrogen	H ₂	<i>g</i>	-285.8	benzene	C ₆ H ₆	<i>l</i>	-3267.6
graphite	C	<i>s</i>	-393.5	toluene	C ₇ H ₈	<i>l</i>	-3910.3
carbon monoxide	CO	<i>g</i>	-283.0	naphthalene	C ₁₀ H ₈	<i>s</i>	-5156.3
methane	CH ₄	<i>g</i>	-890.8	anthracene	C ₁₄ H ₁₀	<i>s</i>	-7076.5
ethane	C ₂ H ₆	<i>g</i>	-1560.7	methanol	CH ₃ OH	<i>l</i>	-726.1
propane	C ₃ H ₈	<i>g</i>	-2219.2	ethanol	C ₂ H ₅ OH	<i>l</i>	-1366.8
butane	C ₄ H ₁₀	<i>g</i>	-2877.6	ether	(C ₂ H ₅) ₂ O	<i>l</i>	-2751.1
pentane	C ₅ H ₁₂	<i>g</i>	-3535.6	formaldehyde	CH ₂ O	<i>g</i>	-570.7
hexane	C ₆ H ₁₄	<i>l</i>	-4163.2	glucose	C ₆ H ₁₂ O ₆	<i>s</i>	-2803.0
heptane	C ₇ H ₁₆	<i>l</i>	-4817.0	sucrose	C ₁₂ H ₂₂ O ₁₁	<i>s</i>	-5640.9
octane	C ₈ H ₁₈	<i>l</i>	-5470.5				
ethene (ethylene)	C ₂ H ₄	<i>g</i>	-1411.2				
propene (propylene)	C ₃ H ₆	<i>g</i>	-2058.0				
ethyne (acetylene)	C ₂ H ₂	<i>g</i>	-1301.1				

ΔH_c = heat of combustion of the given substance. All values of ΔH_c are expressed as kJ/mol of substance oxidized to H₂O(*l*) and/or CO₂(*g*) at constant pressure and 25°C. *s* = solid, *l* = liquid, *g* = gas