

Ch. Thermochemistry (Thermodynamics)

-is the study of energy.

I. Energy is the capacity to do work or produce heat.

Law of Conservation of Energy

Energy cannot be created or destroyed, just transferred.

1st Law of Thermodynamics

The energy in the universe is constant.

State Functions

-depend only on the current state of the system. P, V, T, energy

Heat

Heat flows from hot to cold: high KE, high velocity particles collide with cold/slower particles, making them speed up, increasing their KE.

Exothermic

Heat energy flows out of a system into its surroundings.
(hot pack) **burning paper

Endothermic

Heat energy flows into a system from its surroundings.
(cold pack) **ice cube melting

$\Delta E = q + w$ ΔE = change in internal energy of a system (joules)

q = heat, w = work

+q = heat flows into the system (endothermic)

-q = heat flows out of the system (exothermic)

-w = the system does work (loses energy)

+w = the surroundings do work on the system

$w = -P\Delta V$ P = pressure

$\Delta V = V_f - V_i$ (change in volume)

**** 1 L·atm = 101.3 J**

Ex. 1) A system gives off 196 kJ of heat to its surroundings and the surroundings do 4.20×10^5 J of work on the system. Find ΔE of the system.

$$\frac{4.20 \times 10^5 \text{ J} \left| \frac{1 \text{ kJ}}{1 \times 10^3 \text{ J}} \right.}{1 \times 10^3 \text{ J}} = 420 \text{ kJ}$$

$$\Delta E = q + w = -196 \text{ kJ} + 420 \text{ kJ} = \mathbf{224 \text{ kJ}}$$

(-) q, since heat is lost,

(+) w, since work gained by system (work done on it)

Ex. 2) A balloon is being inflated. The volume of the balloon changes from 400 L to 750 L by the addition of $2.40 \times 10^6 \text{ J}$ of heat. Assume the balloon expands against a constant pressure of 1.00 atm. Find ΔE of the balloon.

$$\Delta E = q + w \quad w = -P\Delta V \quad \Delta V = V_f - V_i$$

$$w = -(1.00 \text{ atm})(750 \text{ L} - 400 \text{ L})$$

$$w = -350 \text{ L}\cdot\text{atm}$$

$$w = \frac{-350 \text{ L}\cdot\text{atm}}{1 \text{ L}\cdot\text{atm}} \left| \frac{101.3 \text{ J}}{1 \text{ L}\cdot\text{atm}} \right. = -3.54 \times 10^4 \text{ J}$$

$$\Delta E = q + w = 2.40 \times 10^6 \text{ J} - 3.54 \times 10^4 \text{ J}$$

$$\Delta E = \mathbf{2.36 \times 10^6 \text{ J}}$$

#48 Notes III. Calorimetry

-is the science of measuring heat.

$$\Delta E = q + w \quad w = -P\Delta V$$
$$q = sm\Delta T \quad \text{where } s = \text{specific heat capacity,}$$
$$m = \text{mass, and } \Delta T = T_f - T_i$$

Specific heat capacity (s)

-is the energy required to raise the temperature one degree Celsius for one gram of the substance.

$$s \text{ for H}_2\text{O} = 4.184 \text{ J/(g}\cdot\text{°C)} = 1 \text{ calorie, } s \text{ for Fe} = 0.45 \text{ J/(g}\cdot\text{°C)}$$

Molar heat capacity

-is the same except for 1 mol of the substance. $\text{J}/(\text{mol}\cdot\text{°C})$

**so the “q” equation would have mols for “m”, instead of mass

Ex. 1a) How much heat is required to raise 8.77 g CCl_4 from 37.1 °C to 56.4 °C ?

$$s \text{ for CCl}_4 = 0.856 \text{ J/(g}\cdot\text{°C)}$$

$$q = sm\Delta T = (0.856 \text{ J/(g}\cdot\text{°C)}) (8.77 \text{ g}) (56.4 \text{ °C} - 37.1 \text{ °C})$$
$$= (0.856 \text{ J/(g}\cdot\text{°C)}) (8.77 \text{ g}) (19.3 \text{ °C}) = \mathbf{1.45 \times 10^2 \text{ J}}$$

Ex. 1b) Calculate the molar heat capacity.

$$\frac{0.856 \text{ J}}{\text{g}\cdot\text{°C}} \left| \frac{154.13115 \text{ g CCl}_4}{1 \text{ mol CCl}_4} \right. = \mathbf{132 \text{ J}/(\text{mol}\cdot\text{°C})}$$

Ex. 1c) Calculate the heat of raising 24.1 mol CCl_4 from 25 °C to 325 °C .

$$q = sm\Delta T = (132 \text{ J/(mol}\cdot\text{°C)}) (24.1 \text{ mol}) (325 \text{ °C} - 25 \text{ °C})$$
$$= (132 \text{ J/(mol}\cdot\text{°C)}) (24.1 \text{ mol}) (300 \text{ °C}) = \mathbf{9.5 \times 10^5 \text{ J}}$$

Ex. 2) A 25.0 g sample of metal was heated to 90.0 °C and added to 50.0 cm³ of water at 20.0 °C. What is the specific heat of the metal, if the final temperature was 22.0 °C?

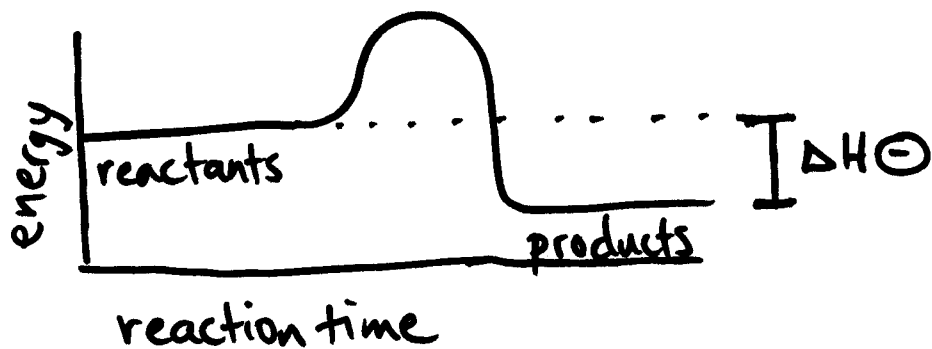
<u>Metal</u>	<u>Water</u>	D = m/v
Loses heat	Gains heat	$(1.0 \text{ g/cm}^3) = \frac{m}{(50.0 \text{ cm}^3)}$
-q =	+ q	50.0 g = m
- smΔT =	+ smΔT	

$$\begin{aligned}
 -s (25.0 \text{ g}) (22.0 \text{ }^\circ\text{C} - 90.0 \text{ }^\circ\text{C}) &= (4.184 \text{ J/(g}\cdot\text{ }^\circ\text{C)}) (50.0 \text{ g}) (22.0 \text{ }^\circ\text{C} - 20.0 \text{ }^\circ\text{C}) \\
 -s (25.0 \text{ g}) (-68 \text{ }^\circ\text{C}) &= (4.184 \text{ J/(g}\cdot\text{ }^\circ\text{C)}) (50.0 \text{ g}) (2.0 \text{ }^\circ\text{C}) \\
 s (1700) &= (418.4) \\
 \mathbf{s = 0.246 \text{ J/(g}\cdot\text{ }^\circ\text{C)}}
 \end{aligned}$$

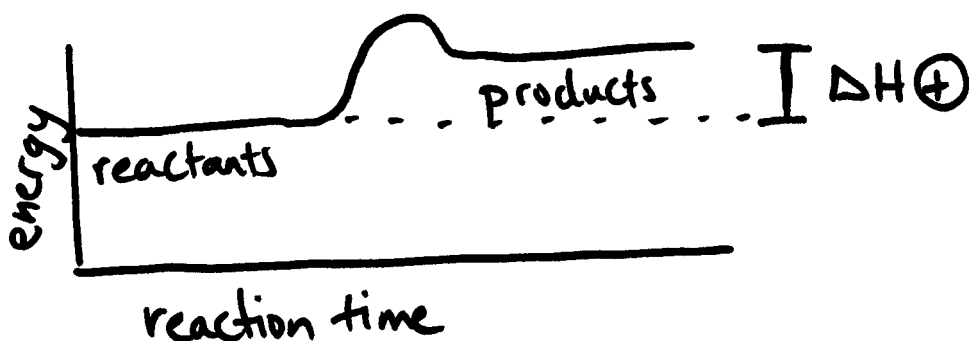
#49 Notes II. Enthalpy (H)

-is the heat energy.

q = H at constant P



Exothermic
Heat lost.



Endothermic
Heat gained.

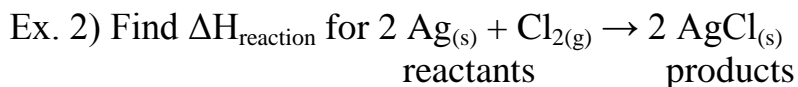
Ex. 1) Are the following endothermic or exothermic.

- a) boiling water: absorbing heat to boil, so **endothermic**
- b) burning paper: gives off more heat, than it absorbs to begin burning, so **exothermic**
- c) melting ice cube: absorbing heat to melt, so **endothermic**

We can measure the difference in heat for the formation of compounds from their elements.
See textbook for ΔH_f° values ($^\circ$ for standard state, $_f$ for formation) cr = crystalline solid

$$\Delta H_{\text{reaction}} = \sum n \Delta H_{\text{products}(\text{final})} - \sum n \Delta H_{\text{reactants}(\text{initial})}$$

↑
Gr. letter sigma (summation)



$$\Delta H_{\text{reaction}} = \sum n \Delta H_{\text{products}(\text{final})} - \sum n \Delta H_{\text{reactants}(\text{initial})}$$

$$\Delta H_{\text{reaction}} = [(2 \text{mol AgCl}) (-127.068 \text{ kJ/mol})] - [(2 \text{ mol Ag}) (0 \text{ kJ/mol}) + (1 \text{ mol Cl}_2)(0 \text{ kJ/mol})]$$

$$\Delta H_{\text{reaction}} = -254.136 \text{ kJ} - 0 - 0 = \mathbf{-254.136 \text{ kJ exothermic}}$$

Ex. 3) Find $\Delta H_{\text{reaction}}$ for $\text{Zn}_{(s)} + 2 \text{HCl}_{(aq)} \rightarrow \text{ZnCl}_{2(aq)} + \text{H}_{2(g)}$

$$\Delta H_{\text{reaction}} = \sum n \Delta H_{\text{products}(\text{final})} - \sum n \Delta H_{\text{reactants}(\text{initial})}$$

$$\Delta H_{\text{reaction}} = [(1 \text{ mol ZnCl}_2) (-488.19 \text{ kJ/mol}) + (1 \text{ mol H}_2) (0 \text{ kJ/mol})] \\ - [(1 \text{ mol Zn}) (0 \text{ kJ/mol}) + (2 \text{ mol HCl}) (-167.159 \text{ kJ/mol})]$$

$$\Delta H_{\text{reaction}} = -488.19 \text{ kJ} + 0 - 0 + 334.318 \text{ kJ} = \mathbf{-153.87 \text{ kJ exothermic}}$$

**Make sure subscripts match (_(cr) = _(s))

#50 Notes IV. Hess's Law

The enthalpy change (ΔH_r°) for a reaction is the sum of the enthalpy changes for a series of reactions, that add up to the overall reaction.

Steps:

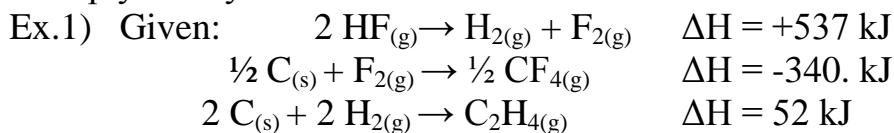
For each reaction:

1) Check to see, if the compounds are on the correct sides of the reaction.

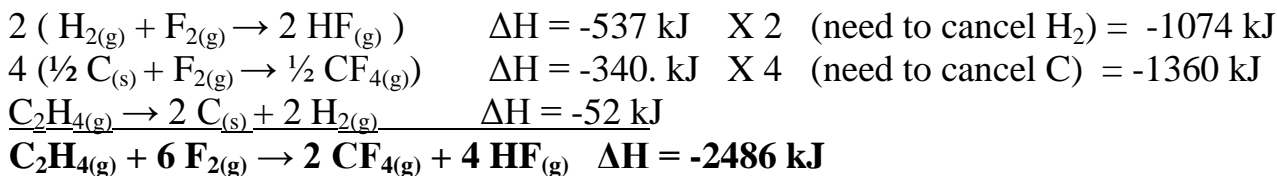
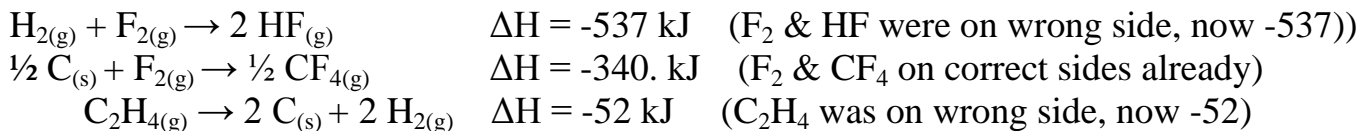
**If not, reverse the entire reaction, and change the sign of ΔH .

2) Check to see, if all of the unwanted compounds will cancel completely.

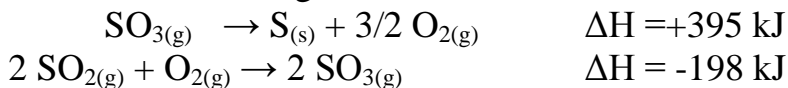
**If not, multiply an entire reaction by a number so that they do cancel completely and multiply ΔH by that same number.



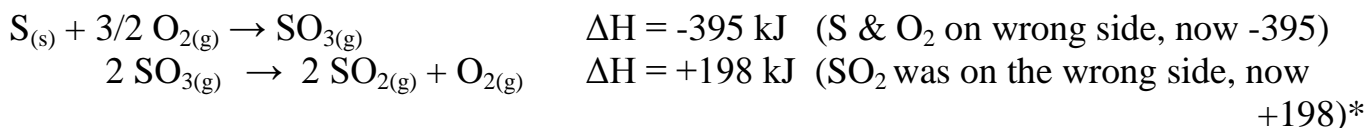
Find ΔH_r° for: $\text{C}_2\text{H}_{4(g)} + 6 \text{F}_{2(g)} \rightarrow 2 \text{CF}_{4(g)} + 4 \text{HF}_{(g)}$



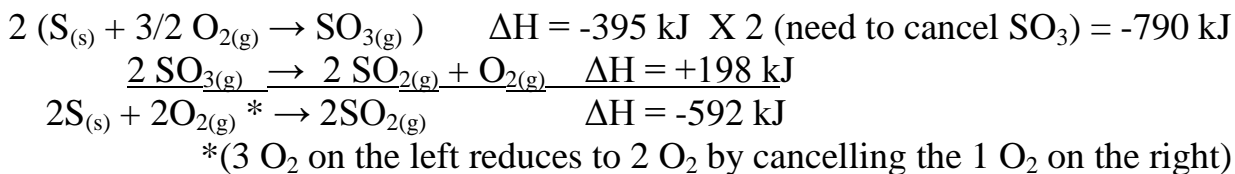
Ex. 2) Given the following data:



Calculate ΔH for the reaction: $\text{S}_{(s)} + \text{O}_{2(g)} \rightarrow \text{SO}_{2(g)}$



*Now for the second reaction O_2 is on the wrong side, but there is other O_2 in the first reaction!



End of Notes (Assignment #51 is a Review Assignment. There are no notes for this assignment.)