Totally Epic AP Chem Review: Enthalpy Calculations!

You Must Know: 6 Ways to Calculate Enthalpy (ΔH)

1. Calorimetry
2. Heat of formation, ΔH_f°
3. Heat of reaction, ΔH_{rxn}°
4. Stoichiometry
5. Bond energies
6. Hess's Law

1. **Calorimetry:** technique used to experimentally determine the change in energy of a chemical reaction or phase change by putting it in contact surroundings of known heat capacity. (e.g. H_2O)

➔ The ~~energy~~ heat change in the water is equal and opposite to the ~~heat~~ energy change by the system!

- The system can be an object, a phase change, or a chemical reaction.

$$+q_{H_2O} = -q_{system}$$

$$+[\mathrm{mC\Delta T}]_{H_2O} = -[\mathrm{mC\Delta T}]_{system}$$

→ If water bath increases in temperature, it gained energy → chemical reaction or phase change lost energy. (exo, $-\Delta H$)

→ If water bath decreases in temperature, it lost energy → chemical reaction or phase change gained energy. (endo, $+\Delta H$)

To calculate the enthalpy of reaction (or phase change), you will need to divide q_{system} by the moles of reaction.

For example,

For the dissolution of a salt (your solute):

$$\Delta H_{soln} = \frac{q_{soln}}{mol_{solute}}$$

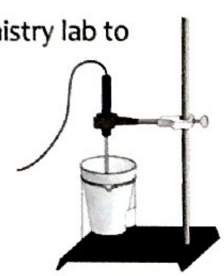
For the enthalpy of neutralization for the following reaction: $3\ NaOH(aq) + H_3PO_4(aq) \rightarrow 3\ H_2O(l) + Na_3PO_4(aq)$

$$\Delta H_{neut} = \frac{q_{rxn}}{mol_{acid}} = \frac{q_{rxn}}{\frac{1}{3}mol_{base}}$$ (assuming neither reactant is in excess)

Warning: Experimental Error with Calorimetry!

- We assume in the above equation that **ALL** energy lost by the system is gained *only* by the water, but that's not true! The calorimeter can also absorb heat, or heat can be lost to the surrounding air.
- Both of these errors would lead to a calculated heat (q) that was SMALLER than the actual heat exchange, and thus the calculated ΔH would be ↓ than the actual ΔH.
  in magnitude

Coffee cup calorimeter: Styrofoam cups are commonly used as insulators in the high school chemistry lab to measure temperature changes without a loss of energy to the surroundings.



2. **Standard enthalpy (heat) of formation** (ΔH_f°): change in enthalpy that accompanies the formation of 1 mole of the compound in its standard state from its component elements their standard states.

Example

$S_{(s)} + \frac{3}{2} O_{2(g)} \rightarrow SO_3(g)$ $\Delta H_f^\circ = -396$ kJ/mol$_{rxn}$

Note: you will see fractional coefficients to ensure only 1 mole of compound is formed.

The ΔH_f° for elements (in their standard state) **is always** Ø **kJ/mol$_{rxn}$!**

3. **Heat of reaction**: ΔH°_{rxn} (aka Big Momma's Equation)

You can use the heat of formation of the reactants and products to find the total enthalpy change in a reaction, according to the following equation:

$$\Delta H^\circ_{rxn} = \sum\left[\Delta H_f^\circ(\text{products})\right] - \sum\left[\Delta H_f^\circ(\text{reactants})\right]$$

4. **Stoichiometry:** heat is an extensive property. Amount matters!

5. **Bond energies:** Breaking chemical bonds requires an input of energy to overcome the attractive forces. When new bonds are formed, energy is released. The difference between input and output determines whether a process is endothermic or exothermic. This can be quantified in the following reaction:

$\Delta H = \Sigma$ Energy to break bonds $-$ Σ Energy released forming new bonds
$\Delta H = \Sigma$ E (bonds in reactants) $-$ Σ E (bonds in products)

There are three steps to calculating enthalpy change through use of bond energies.

1. Draw the Lewis dot structures for the reactants and products.
2. Identify the type and number of bonds being broken and bonds being formed.
3. Subtract the sum of bond energies formed from the sum of bond energies broken.

Endothermic: $+\Delta H^\circ$ (system gains energy)	Exothermic: $-\Delta H^\circ$ (system loses energy)
➢ Breaking a chemical bond is always endothermic ➢ Hint: "**end**-ing a bond is **end**othermic"	➢ Forming a chemical bond is always exothermic (more stable)

6. **Hess's Law:** the overall enthalpy change in a reaction is the sum of all the reactions (and is independent of the route taken)

Rule 1: If you reverse the reaction, then change the sign of ΔH.

Rule 2: If you multiply the reaction by a coefficient, then multiply ΔH by same coefficient.

Strategy: Find things in your goal equation that appear in only one of the available reactions and make them match by flipping equations or multiplying/dividing coefficients. Then arrange equations to cancel out things that don't appear in the "goal." **Whatever you do to the equation, you must do to ΔH!**

Let's Practice!

1. The specific heat (in J/(g °C)) of solid aluminum is 0.89, of solid iron is 0.45, of liquid mercury is 0.14, and of carbon graphite is 0.71. When the same amount of heat is applied to one gram of these substances, which one will reach the highest temperature? Explain.

Hg (l) will reach the highest temperature, b/c it has the lowest specific heat capacity, thus the same amount of heat energy added will raise the temperature the most.

2. Use the data regarding the standard enthalpies of formation to calculate ΔH°_{rxn} for the following reaction:

$$2\ C_3H_6(g) + 9\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(l)$$

Substance	ΔH°_f(kJ/mol)
$C_3H_6(g)$	20.9 kJ/mol_{rxn}
$CO_2(g)$	-393.5 kJ/mol_{rxn}
$H_2O(l)$	-286 kJ/mol_{rxn}

$$\Sigma \Delta H^\circ_{rxn} = \Sigma \Delta H^\circ_{pr} - \Sigma \Delta H^\circ_{re}$$

$O_{2(g)}$ in std state

$$= [6(-393.5) + 6(-286)] - [-2(20.9) + 9(\emptyset)]$$

$$= -4077 - 41.8 = \boxed{-4119\ kJ/mol_{rxn}}$$

s.f.

3. When 1.095 g of NaOH is dissolved in 150.0 g of water initially at 23.50°C in a coffee-cup calorimeter, the final temperature is found to be 25.32°C. Assume the specific heat of the solution is the same as that of water (4.184 J/g°C) and no heat is absorbed by the calorimeter.
   a. What is the enthalpy of dissolution, ΔH_{soln}?

$$q_{rxn} = -q_{cal} = -mC\Delta T$$

$$= -(150.0 + 1.095)(4.184\ \tfrac{J}{g°C})(25.32 - 23.50)$$

$$= -1150\ J = -1.15\ kJ$$

$$1.095\ g\ NaOH \times \frac{1\ mol}{39.998\ g} = 0.02738\ mol\ NaOH$$

$$\Delta H_{rxn} = \frac{q}{mol_{NaOH}}$$

$(1\ mol_{rxn} : 1\ mol_{NaOH})$

$$= \frac{-1.15\ kJ}{0.02738\ mol}$$

$$= \boxed{-42.0\ kJ/mol_{rxn}}$$

   b. If heat was absorbed by the calorimeter, what effect would it have on the calculated heat of reaction? Justify your answer.

Some of the heat of the reaction being absorbed by the calorimeter would decrease the temperature change of the water, decreasing the magnitude of calculated q but not affecting the number of mol_{rxn}, so calculated ΔH_{rxn} would be smaller in magnitude (less negative) than actual ΔH_{rxn}.

4. The heat of formation of copper(I) chloride is –137 kJ/mol$_{rxn}$.
   a. Write the balanced chemical equation. [Hint: use the definition of a heat of formation!]

$$Cu_{(s)} + \tfrac{1}{2}Cl_{2(g)} \rightarrow CuCl_{(s)}$$

   b. How many joules are released when 4.46 grams of copper react with excess chlorine to produce copper(I) chloride?

$$4.46\text{ g Cu} \times \frac{1\text{ mol Cu}}{63.55\text{ g Cu}} \times \frac{1\text{ mol}_{rxn}}{1\text{ mol Cu}} \times \frac{-137\text{ kJ}}{1\text{ mol}_{rxn}} \times \frac{1000\text{ J}}{1\text{ kJ}} = -9.61\times10^{3}\text{ J}$$

$$\Rightarrow \boxed{9.61\times10^{3}\text{ J released}}$$

5. Given the following information, find the heat of formation for methane: $C(s) + 2H_2(g) \rightarrow CH_4(g)$

flip $\left[CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)\right]$ $\Delta H^\circ_{comb} = -891 \text{ kJ mol}^{-1} \times -1 = 891$

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H^\circ_f = -394 \text{ kJ mol}^{-1}$ ✓

$2 \times \left(H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l)\right)$ $\Delta H^\circ_f = -286 \text{ kJ mol}^{-1} \times 2 = -572$

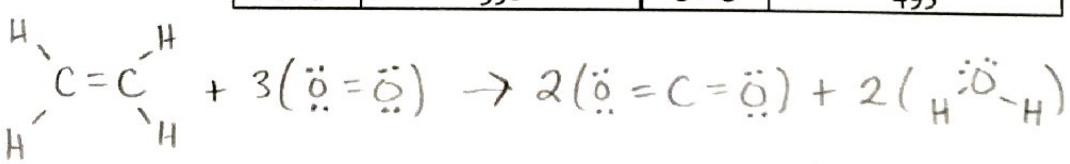
$$\cancel{CO_{2(g)}} + \cancel{2H_2O_{(l)}} \rightarrow CH_{4(g)} + \cancel{2O_{2(g)}}$$
$$C_{(s)} + \cancel{O_{2(g)}} \rightarrow \cancel{CO_{2(g)}}$$
$$2H_{2(g)} + \cancel{O_{2(g)}} \rightarrow \cancel{2H_2O_{(l)}} +$$
$$C_{(s)} + 2H_{2(g)} \rightarrow CH_{4(g)}$$

$$\Delta H^\circ_{rxn} = 891 - 394 - 572 = \boxed{-75 \text{ kJ/mol}_{rxn}}$$

6. The flammable gas ethene, C_2H_4, combusts as such: $C_2H_4(g) + 3O_2 \rightarrow 2CO_2(g) + 2H_2O(l)$

Given the table of bond energies below, what is the enthalpy change for this reaction?

Bond	Average Bond Energy (kJ/mol)	Bond	Average Bond Energy (kJ/mol)
C - H	413	C = O	799
C - C	347	H - O	467
C = C	614	H - H	432
C - O	358	O = O	495



$$\Delta H_{rxn} = \Sigma BE(\text{react}) - \Sigma BE(\text{prod})$$
$$= [4(C\text{-}H) + 1(C=C) + 3(O=O)] - [4(C=O) + 4(O\text{-}H)]$$
$$= [4(413) + 614 + 3(495)] - [4(799) + 4(467)]$$
$$= 3751 - 5064 = \boxed{-1313 \text{ kJ/mol}_{rxn}}$$