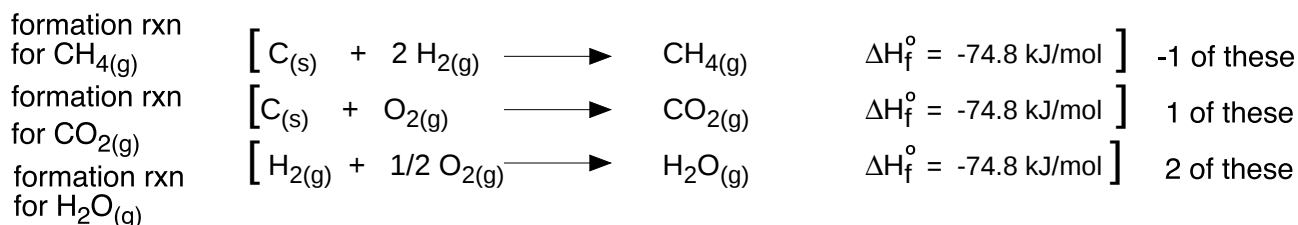


Let's look at the thermochemistry of chemical reactions. A lot of this stuff is just knowing the mechanics of how to use tabulated data to answer questions.

- **Exothermic reactions** give off energy, $\Delta H < 0$.
- **Endothermic reactions** take in energy, $\Delta H > 0$.

Hess' Law says that if you add up a bunch of chemical reactions to get a final reaction, you can add up the heats of reaction (ΔH_{rxn}) the same way and get the ΔH_{rxn} for the final reaction. Chemists used this idea to make a couple of data tables (heats of formation) from which most ΔH_{rxn} 's can be calculated. The idea is to use formation reactions to generate the reaction you want. Formation reactions do not have to have any basis in reality; you just make the molecule you want out of elements. Elemental species have no formation reactions, so JUST IGNORE THEM!

For example: Calculate the heat of reaction for: $\text{CH}_{4(g)} + 2 \text{O}_{2(g)} \longrightarrow \text{CO}_{2(g)} + 2 \text{H}_2\text{O}_{(g)}$



$$\Delta H_{\text{rxn}}^\circ = (-1)(-74.8 \text{ kJ}) + (1)(-393.5 \text{ kJ}) + (2)(-241.8) = -802.3 \text{ kJ/mol}$$

$-802.3 < 0$ so it's an exothermic reaction.

Okay, so notice that the reactants formation reactions always get reversed. So they get a sign change. Also note that whatever stoichiometric coefficient a molecule has shows up as how many formation reactions you need. In math shorthand this is just:

$$\Delta H_{\text{rxn}}^\circ = \left(\sum v_i \Delta H_f^\circ \right)_{\text{products}} - \left(\sum v_i \Delta H_f^\circ \right)_{\text{reactants}}$$

It's just the sum of the heats of formation of the product times their stoichiometric coefficients minus the sum of the heats of formation of the reactants times their coefficients.

- Watch out for water: $\text{H}_2\text{O}_{(l)}$ and $\text{H}_2\text{O}_{(g)}$ have different values.

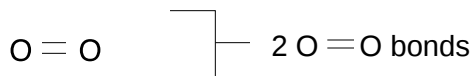
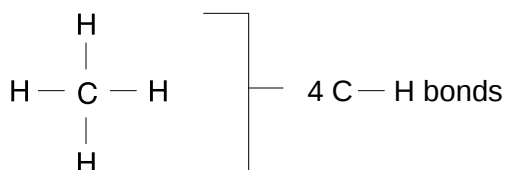
Heat of reaction from bond energies

You can also calculate heats of reaction by keeping track of the bonds broken and formed during a chemical reaction. If we assume that we know the energy for types of bonds like C—H (and we do know this because they are in tables), then just add up the energy required to break the ones that break and the energy given off by those that form. That's kind of an important point--it takes energy to break bonds (energy into the system is positive) and energy is given off when they form (energy out of the system is negative).

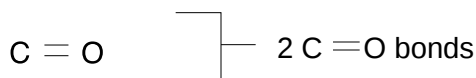
For example: Let's calculate the heat of combustion for methane using bond energies.



look at the bonds that are broken:



look at the bonds that are formed:



Using a table of your choice you have to add up all of the bond energies.

<u>Bond</u>	<u>Bond dissociation enthalpy kJ/mol</u>	<u># broken</u>	<u>#formed</u>	<u>Total energy exchanged</u>
C — H	415	4		1660
O = O	498	2		996
C = O	745		-2	-1490
O — H	464		-4	-1856

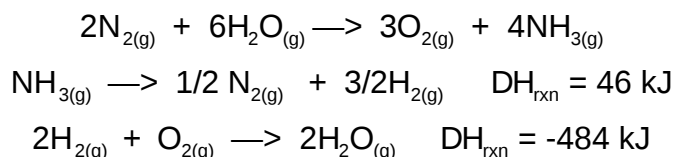
Add these guys up and you get a $\Delta H_{\text{rxn}} = -690 \text{ kJ/mol}$.

Student Packet: 7.3: Heats of Reaction: AssessMe questions

- 1) A mixture of powdered aluminum and rust (Fe_2O_3) will show no indication of a reaction upon sitting for years. Does this mean that the mixture is at equilibrium?
- A) Yes, since at equilibrium no reaction is possible.
 B) Yes, a lack of reactivity defines the equilibrium point.
 C) No, this simply indicates that the reaction rate is slow.
 D) No, the mixture might be at equilibrium, however, it is equally likely that the lack of reactivity is associated with slow reaction kinetics.
 E) I'm clueless
- 2) Consider the reaction: $\text{N}_{2(g)} + 3\text{H}_{2(g)} \rightarrow 2\text{NH}_{3(g)}$
 The equilibrium constant expression for this reaction is:
- A) $K_{\text{eq}} = [\text{NH}_3]/([\text{N}_2][\text{H}_2])$ B) $K_{\text{eq}} = [\text{NH}_3]^2/([\text{N}_2][\text{H}_2]^3)$
 C) $K_{\text{eq}} = [\text{N}_2][\text{H}_2]^3/[\text{NH}_3]^2$ D) $K_{\text{eq}} = [\text{N}_2][\text{H}_2]^{3/2}/[\text{NH}_3]$ E) I'm clueless
- 3) The alkenes are compounds of carbon and hydrogen with the general formula C_nH_{2n} . If 0.561 grams of any alkene is burned in excess oxygen, what number of moles of H_2O is formed?
- A) 0.0400 mole B) 0.0600 mole
 C) 0.0800 mole D) 0.400 mole E) I'm clueless
- 4) A sample of 9.00 g of aluminum metal is added to an excess of HCl. The volume of hydrogen gas produced at standard temperature and pressure is:
- A) 22.4 liters B) 11.2 liters C) 7.46 liters D) 3.37 liters E) I'm clueless
- 5) What ratio of concentrations of 0.500 M HNO_2 ($K_a = 4.0 \times 10^{-4}$) and 0.500 M NaNO_2 must be mixed to prepare 1.0 L of a solution buffered at pH 3.55?
- A) .015
 B) .03
 C) 1
 D) 6.67
 E) I'm clueless
- 6) Which of the following would not make a good buffering system?
- A) SO_4^{2-} and H_2SO_4 B) HCO_3^- and H_2CO_3
 C) NH_3 and NH_4^+ D) CH_3COO^- and CH_3COOH E) I'm clueless

Student Packet: 7.3: Heats of Reaction: ShowMe questions

- 1) Calculate the change in enthalpy for the production of ammonia from nitrogen gas and water vapor, based on the provided data:



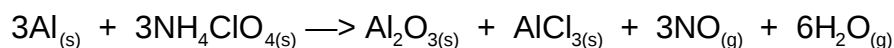
- 2) Calculate the enthalpy of combustion of ethane (C_2H_6) based on the following enthalpies of formation:

$$\text{DH}_f \text{C}_2\text{H}_{6(g)} = -84.7 \text{ kJ/mol}$$

$$\text{DH}_f \text{H}_2\text{O}_{(g)} = -242 \text{ kJ/mol}$$

$$\text{DH}_f \text{CO}_{2(g)} = -393.5 \text{ kJ/mol}$$

- 3) A mixture of aluminum and ammonium perchlorate is sometimes used as rocket fuel. One of the reactions between these two species is the following:



Using the following data, calculate DH° for this reaction.

$$\begin{aligned}
 \text{DH}_f^\circ \text{NH}_4\text{ClO}_{4(s)} &= -295 \text{ kJ/mol} \\
 \text{DH}_f^\circ \text{Al}_2\text{O}_{3(s)} &= -1675 \text{ kJ/mol} \\
 \text{DH}_f^\circ \text{AlCl}_{3(s)} &= -704 \text{ kJ/mol} \\
 \text{DH}_f^\circ \text{NO}_{(g)} &= 90 \text{ kJ/mol} \\
 \text{DH}_f^\circ \text{H}_2\text{O}_{(g)} &= -242 \text{ kJ/mol}
 \end{aligned}$$

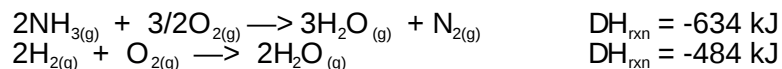
- 4) Calculate the enthalpy of formation of carbon monoxide gas, based on the following data:

$$\text{DH}_{\text{comb}} (\text{C}_{(s)}) = -393.7 \text{ kJ/mol}$$

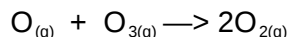
$$\text{DH}_{\text{comb}} (\text{CO}_{(g)}) = -283.3 \text{ kJ/mol}$$

Student Packet: 7.3: Heats of Reaction: TestMe questions

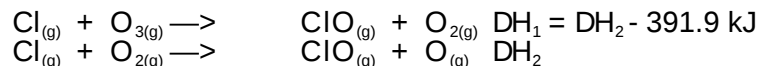
- 1) Given the following data:

What is the enthalpy of formation of $\text{NH}_{3(g)}$?

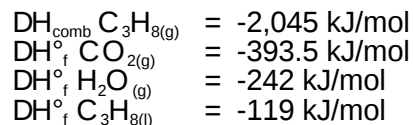
- A) 634 kJ/mol B) 317 kJ/mol C) -46 kJ/mol D) -92 kJ/mol E) -184 kJ/mol
- 2) The enthalpy of combustion of propane gas ($\text{C}_3\text{H}_{8(g)}$) is -2,045 kJ/mol. What is the enthalpy of formation of $\text{C}_3\text{H}_{8(g)}$? $\text{DH}^\circ_f \text{CO}_{2(g)} = -393.5 \text{ kJ/mol}$, $\text{DH}^\circ_f \text{H}_2\text{O}_{(g)} = -242 \text{ kJ/mol}$
- A) -103.5 kJ/mol B) 1410 kJ/mol C) 2560 kJ/mol D) 948 kJ/mol E) -624 kJ/mol
- 3) The decomposition of ozone gas ($\text{O}_{3(g)}$) occurs by reaction with atomic oxygen ($\text{O}_{(g)}$):



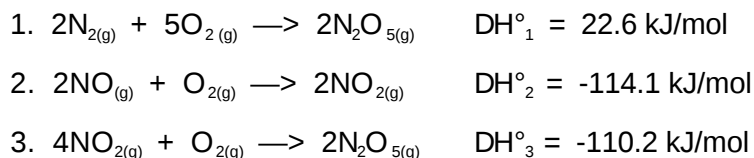
In the upper atmosphere, this process is initiated by the reaction of atomic chlorine ($\text{Cl}_{(g)}$) with ozone. Based on the following reactions, calculate the change in enthalpy for the decomposition of ozone by reaction with atomic oxygen.



- A) -391.9 kJ B) -783.8 kJ C) 391.9 kJ D) 783.8 kJ E) 196.0 kJ
- 4) Based on the following data, calculate the enthalpy of vaporization of liquid propane (C_3H_8).



- A) 2380 kJ/mol B) 1950 kJ/mol C) 224 kJ/mol D) 14 kJ/mol E) -88 kJ/mol
- 5) Calculate the standard enthalpy of formation of $\text{NO}_{(g)}$ based on the following reactions:



- A) -202 kJ/mol
B) 90.2 kJ/mol
C) -247 kJ/mol
D) -81.6 kJ/mol
E) -47.5 kJ/mol