Problem 1.53. Look up the enthalpy of formation of atomic hydrogen in the back of this book. This is the enthalpy change when a mole of atomic hydrogen is formed by dissociating 1/2 mole of molecular hydrogen (the more stable state of the element). From this number, determine the energy needed to dissociate a single H_2 molecule, in electron-volts.

Eg. Problem 1.53

For $H(g)$ $\quad \Delta_f H = 217.97$ kJ/mole

(This is taking $\frac{1}{2}$ mole of H_2 and creating mole of H).

$$H_i = U_i + p_i V_i \qquad p_i = 1\,\text{atm}$$

$$H_f = U_f + p_f V_f \qquad p_f = p_i$$

$$\Delta H = \Delta U + p\,\Delta V$$

$$\Delta U = \Delta H - p\,\Delta V$$

what is ΔV?

Ideal gas: $\quad V_i = \dfrac{n_i R T_i}{p_i}$, $\quad n_i = \frac{1}{2}$ mole

$$V_f = \frac{n_f R T_f}{p_f}, \quad n_f = 1\text{ mole}$$

assume $T_i = T_f$ = room temperature

$$\therefore \Delta V = (n_f - n_i)\frac{R T_i}{p_i} = \left(\frac{1}{2}\right)\frac{R T_i}{p_i} = V_i$$

$\therefore \ \Delta U = \Delta H - p\Delta V$

$$= \Delta H - p_i \cdot \frac{1}{2}\frac{RT_i}{p_i}$$

$$\Delta U = \Delta H - \frac{1}{2}RT$$

$$\Delta U = \frac{217.97\times10^3 J}{mole} \times 1\,mole$$

$$-\left(\tfrac{1}{2}mole\right)(8.314\ J/mol\cdot K)(300K)$$

$\Delta U = 216.7\times10^3 J$ for 1/2 mole

$$\frac{\Delta U}{molecule} = \frac{216.7\times10^3 J}{6.02\times10^{23}\,molecules \times \frac{1}{2}} \times \frac{1\,eV}{1.602\times10^{-19} J}$$

$$= 4.5\,eV$$

Problem 1.51 (HW, Glucose)

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$$

The ΔH_f in the back of the book is for creating 1 mole of substance from it's stable elemental constituents.

What is the net effect?

LHS: Imagine breaking sugar down into C, H_2 O_2. What ΔH is required?

RHS: forming CO_2 and H_2O, get ΔH back. What's the net?