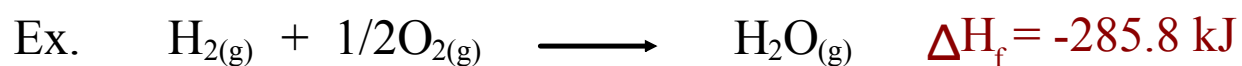


# Reference Energy State

Reference energy state - elements are defined as the reference point at which the potential energy is shown to be zero.

Therefore:  $E_p$  of  $H_{2(g)} = 0$  kJ

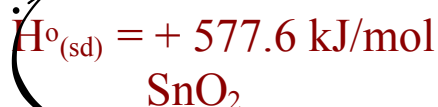
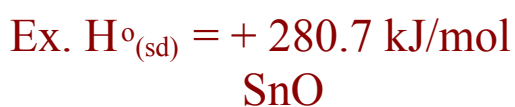


\*allows us to describe the enthalpy change for a formation reaction from zero to a final value

# Thermal Stability

Thermal Stability - the tendency of a compound to resist decomposition when heated.

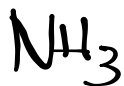
- the more endothermic the simple decomposition (sd), the more stable the compound.



Therefore  $\text{SnO}_2$  is more stable.

\*Normally not given the  $H_{sd}$ , but given the  $H_f$

Which is more stable, ammonia or butane?



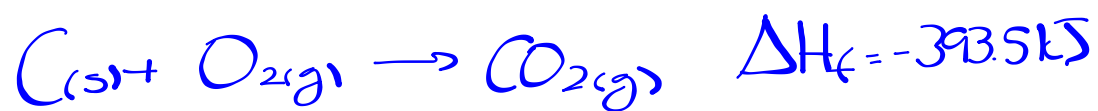
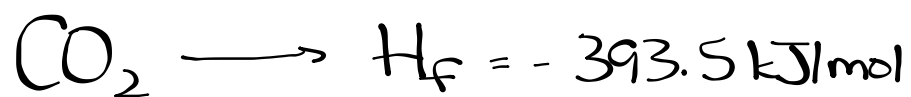
$H_f = -45.9 \text{ kJ/mol}$

$H_{sd} = 45.9 \text{ kJ/mol}$



$H_f = -125.6 \text{ kJ/mol}$

$H_{sd} = 125.6 \text{ kJ/mol}$

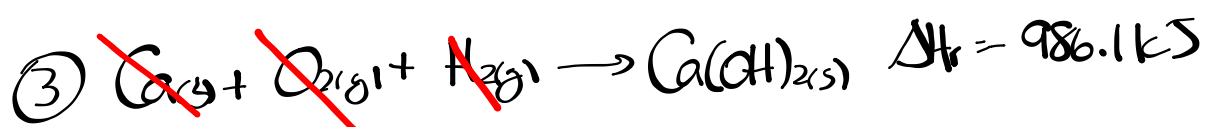
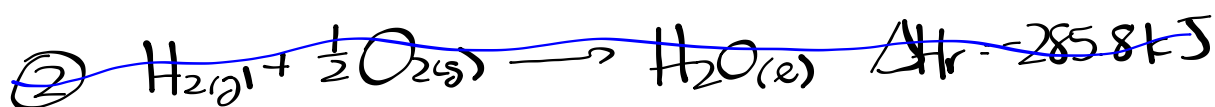


## Predicting $\Delta H_r$ Using Formation Reactions

The Standard Enthalpy Change ( $\Delta H_r^\circ$ ) for a reaction can be found by writing the formation equation and corresponding standard enthalpy change for each compound in the given equation and then applying Hess' Law.



Step 1: Write formation equations (with standard enthalpy change) each compound in the given equation.

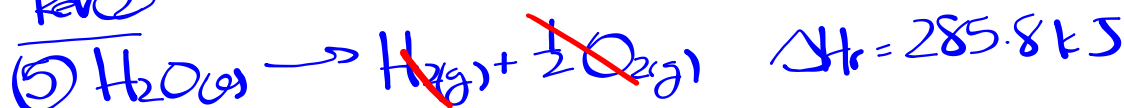


Step 2: Apply Hess' Law

Rev ①



Rev ②



③ + ④ + ⑤



## Enthalpies of Formation to Predict $\Delta H_r$

$$\Delta H_r = \Delta H_{f, \text{Ca(OH)}_2} + (-\Delta H_{f, \text{CaO}}) + (-\Delta H_{f, \text{H}_2\text{O}})$$

$$\Delta H_r = \Delta H_{f, \text{Ca(OH)}_2} - (\Delta H_{f, \text{CaO}} + \Delta H_{f, \text{H}_2\text{O}})$$

$$\Delta H_r = \Delta H_{\text{fp}} - \Delta H_{\text{fr}}$$

products                      reactants

$$\Delta H_r = \sum n H_{\text{fp}} - \sum n H_{\text{fr}}$$

knowing that  $\Delta H = nH$

Ex. What is the standard molar enthalpy of combustion of methane fuel?



$$\Delta H_r = \sum n H_{f,p} - \sum n H_{f,r}$$

$$\Delta H_r = \left[ \overset{\text{CO}_2}{(1 \text{ mol})} \left( -393.5 \frac{\text{kJ}}{\text{mol}} \right) + (2 \text{ mol}) \left( -241.8 \frac{\text{kJ}}{\text{mol}} \right) \right] -$$

$$\left[ (1 \text{ mol}) \left( -74.4 \frac{\text{kJ}}{\text{mol}} \right) + (2 \text{ mol}) \left( 0 \frac{\text{kJ}}{\text{mol}} \right) \right]$$