CHM 123 - Lecture (Friday 10:30am)

Enthalpy and PV Work

H - state function — so how can $H = q$?

q - not a state function

How do internal energy and enthalpy differ?

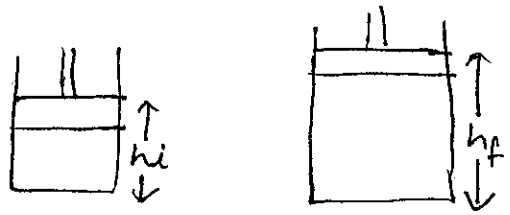
$\Delta E = q + w$

$H = q$

Ans: work.

Pressure is defined as Force per unit Area

$P = F/A$



$|Work| = Force \times distance$

$= F \times \Delta h$

$= (P \times A) \times \Delta h$

$= P \times \Delta V$

$P = F/A \Rightarrow F = P \times A$

$\Delta V = A \times \Delta h$

W = "−" as system is expanding

$W = -P\Delta V = -RT\Delta n$ (for gas laws)

if only (PV) work .....

$\Delta E = q + (w) = q - (P\Delta V)$ → to maintain same sign convention.

For a constant vol. rxn or process, $\Delta V = 0$ hence $\Delta E = q$

II) At constant pressure

$\Delta E = q_p + w = q_p - P\Delta V$

$q_p = \Delta E + P\Delta V$

q_p ~ enthalpy

Enthalpy defined $\Rightarrow H = E + PV$ (q_p transferred at constant pressure)

$\Delta H = \Delta E + P\Delta V$

Thermochemical Equations

$\Delta H = H_{final} - H_{initial} = H_{products} - H_{reactants}$

ΔH_{rxn} – enthalpy or heat of reaction.

$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)} \quad \Delta H = -571.66\ kJ$

~~Significance of~~ Coefficients of equation represent # of moles of reactants and products producing the energy change.

ΔH°
- standard enthalpy change
- enthalpy change at 1 bar pressure and 25°C.

RULES OF ENTHALPY

- Enthalpy is an extensive property – magnitude of ΔH depends on amounts of reactants consumed.

$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)} \quad \Delta H = -890\ kJ$

$2CH_{4(g)} + 4O_{2(g)} \rightarrow 2CO_{2(g)} + 4H_2O_{(l)} \quad \Delta H = -1780\ kJ$

- Enthalpy change of a reaction is equal in magnitude but opposite in sign to ΔH for the reverse reaction

$$CO_{2(g)} + 2H_2O_{(l)} \rightarrow CH_{4(g)} + 2O_{2(g)} \qquad \Delta H = +890\ kJ$$

- Enthalpy change for a reaction depends on the states of the reactants and products.

$$H_2O_{(l)} \rightarrow H_2O_{(g)} \qquad \Delta H = +44\ kJ$$

$$H_2O_{(s)} \rightarrow H_2O_{(g)} \qquad \Delta H = +50\ kJ$$

Bond Enthalpies

- during chemical reaction bonds are broken and made.
- breaking bonds requires energy input (endothermic) ΔH = +ve.
- formation of bonds releases energy (exothermic ΔH = -ve)
- weaker bonds broken and stronger bonds formed.

Hess's Law

- we can calculate ΔH for a reaction using ΔHs for other known reactions.
- ΔH is a state function - result is the same no matter how we arrive at the final state.

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ $\quad \Delta H = -393.5\,kJ$

$CO_{(g)} + \frac{1}{2} O_{2(g)} \rightarrow CO_{2(g)}$ $\quad \Delta H = -283.0\,kJ$

what is ΔH for $C_{(s)} + \frac{1}{2} O_{2(g)} \rightarrow CO_{(g)}$???

$C_{(s)} + O_{2(g)} \rightarrow \cancel{CO_{2(g)}}$ $\quad \Delta H = -393.5\,kJ$

$+ \quad \cancel{CO_{2(g)}} \rightarrow CO_{(g)} + \frac{1}{2} O_{2(g)}$ $\quad \Delta H = +283.0\,kJ$

$C_{(s)} + \frac{1}{2} O_{2(g)} \rightarrow CO_{(g)}$ $\quad \Delta H = -110.5\,kJ$