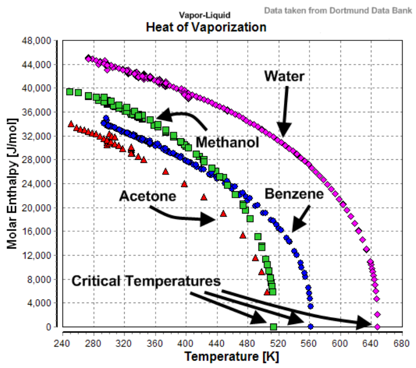
6.4 - Enthalpies of Formation



Enthalpies of Formation

Standard enthalpy of formation, ΔH_f^o:the enthalpy change for the reaction in which a compound is made from its constituent elements in their elemental forms.

The degree symbol indicates that the process has been done under standard conditions - because thermodynamic processes depend on pressures and concentrations of reactants, a precise reference frame must be used to compare properties.

Conventional Definitions of Standard States

For a compound:

For a gas, pressure is exactly 1 atm.

For a solution, concentration is exactly 1 *M*.

Pure substance in a condensed state, its standard state is a pure liquid or solid. Example: FeO is solid.

For an Element

The form in which it exists at 1 atm and 25°C. Example, nitrogen: $N_2(g)$, potassium: $K(s)$.

Note: Pure elements in their standard states have an ΔH_f^o of zero.

Calculation of ΔH

We can use Hess's law in this way:

$$\boxed{\Delta H = \Sigma\, n\, \Delta H_f^\circ{}_{\text{products}} - \Sigma\, m\, \Delta H_f^\circ{}_{\text{reactants}}}$$

where *n* and *m* are the stoichiometric coefficients.

Σ is the mathematical symbol: sum (add together).

1. Enthalpy of Formation Example

Calculate ΔH° for the following reaction:

$2\,Na(s) + 2\,H_2O(l) \rightarrow 2\,NaOH(aq) + H_2(g)$

Find the following information in Appendix 4 (Pages A19 - A 22):

	ΔH_f° (kJ/mol)
$Na(s)$	0
$H_2O(l)$	–286
$NaOH(aq)$	–470
$H_2(g)$	0

1. Enthalpy of Formation Example

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$$\Delta H^\circ = \Sigma\, n\, \Delta H_f^\circ{}_{\text{products}} - \Sigma\, m\, \Delta H_f^\circ{}_{\text{reactants}}$$

$2\,NaOH(aq)$ - $2\,H_2O(l)$

2(-470kJ/mol)-(2(-286 kJ/mol)) = $\boxed{-368 \text{ kJ/mol}}$

2. Methanol Example

Methanol (CH_3OH) is added to gasoline to improve combustion. What is the standard enthalpy of combustion per gram of methanol?

Balanced reaction:

$2\ CH_3OH_{(l)} + 3\ O_{2(g)} \rightarrow 2\ CO_{2(g)} + 4\ H_2O_{(l)}$

Find the following information in Appendix 4 (Pages A19 - A 22):

	ΔH_f° (kJ/mol)
$CH_3OH_{(l)}$	–239
$O_{2(g)}$	0
$CO_{2(g)}$	–394
$H_2O_{(l)}$	–286

2. Methanol Example

$2\ CH_3OH_{(l)} + 3\ O_{2(g)} \rightarrow 2\ CO_{2(g)} + 4\ H_2O_{(l)}$

Find the following information in Appendix 4 (Pages A19 - A 22):

$\Delta H^\circ = \Sigma\ n\ \Delta H_f^\circ{}_{\text{products}} - \Sigma\ m\ \Delta H_f^\circ{}_{\text{reactants}}$

$2\ CO_{2(g)} + 4\ H_2O_{(l)} - 2\ CH_3OH_{(l)}$

2(-394 kJ/mol) + 4(-286 kJ/mol) - (2(-239 kJ/mol))

= -1450 kJ/mol

Since 2 moles of methanol were involved, dividing the energy by twice the molar mass yields:

$$\frac{-1450\, kJ/mol}{64.08\, g} = \boxed{22.6\, kJ/g}$$

Homework

Read 6.5 - 6.6

6.4 Problems in your Booklet
Due: Next Class.

Unit Review???