Hess's Law

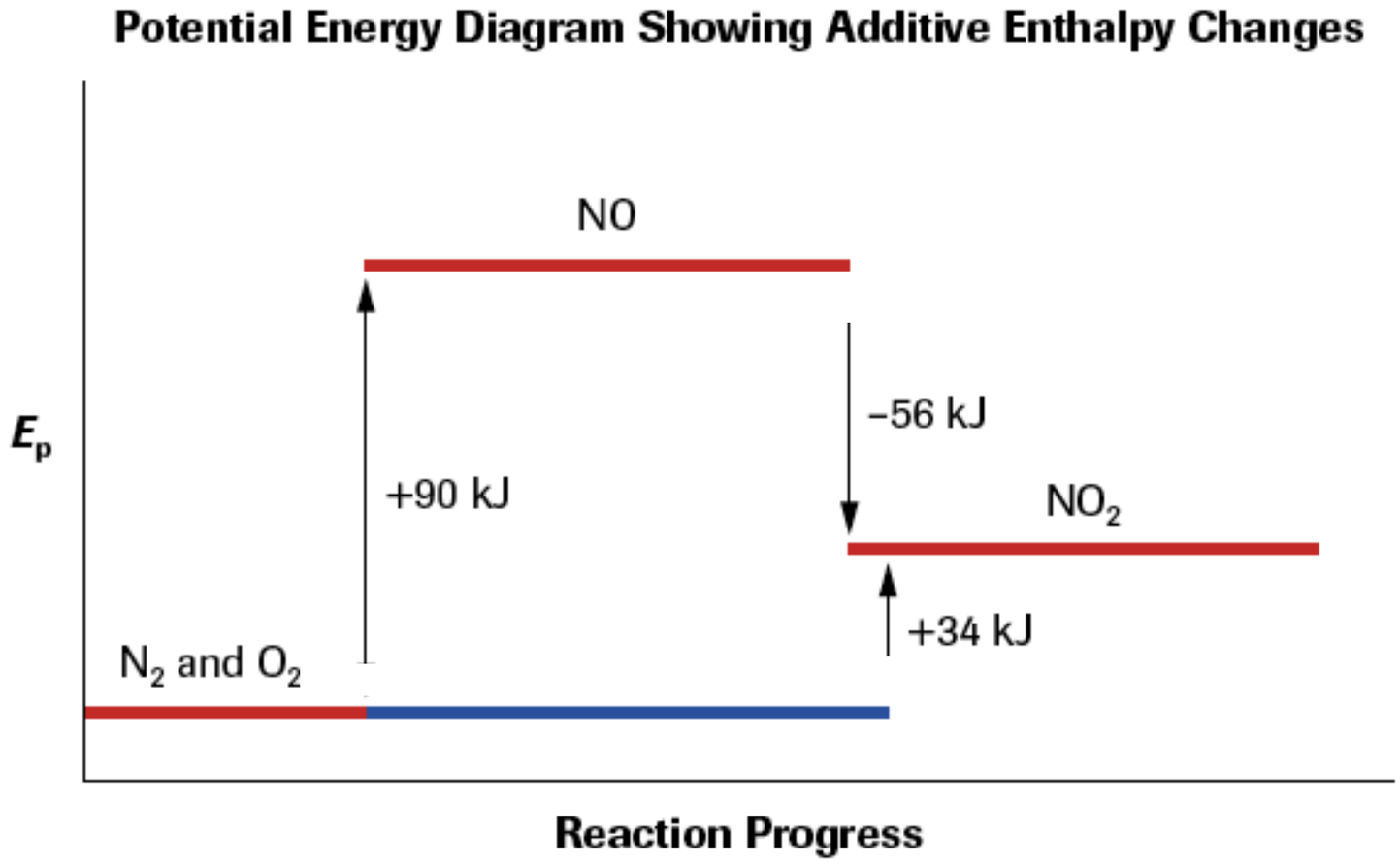
- Calorimetry is an accurate technique for determining enthalpy changes. But what if a reaction is:
  - too dangerous to perform in a lab
  - too slow to perform in a lab
  - impossible to perform in a calorimeter

Other Techniques

- Aside from calorimetry (q=mcT), scientists can use the following to determine enthalpy changes.
  - Hess's law
  - Enthalpy of Formation data

Hess's Law of Summation

For any reaction that can be written in steps, the $\Delta H°$ is the same as the sum of the values of the $\Delta H°$ for each individual step.



Rule #1

N_2O_4 (g) → 2 NO_2 (g) $\Delta H° = +57.93$ kJ

2 NO_2 (g) → N_2O_4 (g) 

1. Reversing a chemical reaction causes a **sign** change in front of the $\Delta H°$ value.

Rule #2

2. If the coefficients of a chemical equation are altered **by multiplying or dividing** then the ΔH is altered the same way

$C_2H_2 + 5/2\ O_2 \rightarrow CO_2 + H_2O \qquad \Delta H = -1299\ kJ$

$2C_2H_2 + 5\ O_2 \rightarrow 2CO_2 + 2H_2O \qquad \Delta H = -2598\ kJ$

Rule #3

3. When cancelling compounds for Hess's Law, the <u>state</u> of the compounds is important.

$H_{2(g)} + ½\ O_{2(g)} \rightarrow H_2O_{(g)}$

$\Delta H° = +57.93\ kJ$

$K_2SO_{4(s)} + 2H_2O_{(l)} \rightarrow H_2SO_{4(aq)} + 2KOH_{(s)}$

$\Delta H° = +342.4kJ$

These two CANNOT cancel each other out

ΔH° Rules

1. If all the coefficients of an eq^{n} are multiplied or divided by a common factor, the ΔH° must be changed likewise.

2. When a reaction is reversed, the sign of ΔH° must also be reversed.

3. When cancelling compounds for Hess's Law, the <u>state</u> of the compounds is important.

Enthalpy

$\Delta H°$ values of unknown reactions can be solved when other known reactions are given.

Example #1

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} \quad \Delta H^\circ = ?$$

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

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

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

Example #2

$$2\ Fe_{(s)} + 3/2\ O_{2(g)} \rightarrow Fe_2O_{3(s)}\ \Delta H° = ?$$

$$Fe_2O_{3(s)} + 3\ CO_{(g)} \rightarrow 2\ Fe_{(s)} + 3\ CO_{2(g)}\ \Delta H° = -26.7\ kJ$$

Flip! $2\ Fe_{(s)} + 3\ CO_{2(g)} \rightarrow Fe_2O_{3(s)} + 3\ CO_{(g)}\ \Delta H° = +26.7\ kJ$

$$CO_{(g)} + ½\ O_{2(g)} \rightarrow CO_{2(g)}\ \Delta H° = -283.0\ kJ$$

x3! $3CO_{(g)} + 3/2\ O_{2(g)} \rightarrow 3CO_{2(g)}\ \Delta H° = -849.0\ kJ$

$$2\ Fe_{(s)} + 3/2\ O_{2(g)} \rightarrow Fe_2O_{3(s)}\ \Delta H° = -822.3\ kJ$$