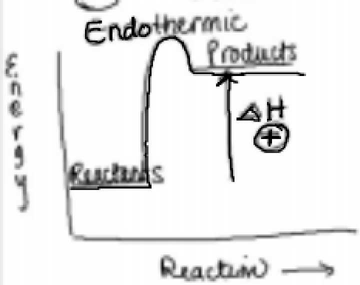
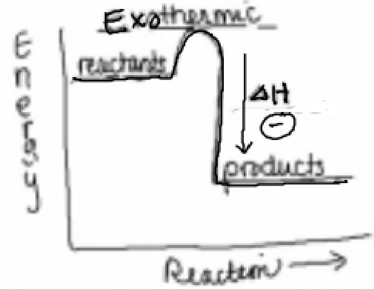
We evaluate the energy (q) by calculating ΔH. (Heat)

Sign of ΔH

(+) endothermic — heat is absorbed - energy required for rxn

(−) exothermic — heat is released - energy is produced by rxn





$$\Delta H_{reaction} = \Delta H_{Products} - \Delta H_{reactants}$$

$\Delta H_{formation} = \emptyset$ if pure element or diatomic ($H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2$)

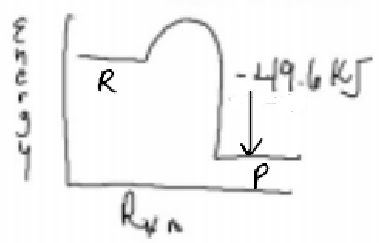
coefficient from balanced reaction = moles.

$$2\,NaBr(s) + Cl_2(g) \longrightarrow 2\,NaCl(s) + Br_2(g)$$

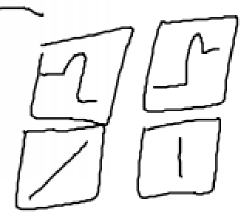
$$\Delta H_{rxn} = \left(2\,mol \cdot -385.9\ KJ/mol + 1\,mol \cdot \emptyset\right) - \text{(subtract)}$$

$$\left(2\,mol \cdot -361.1\ KJ/mol + \emptyset\right)$$

$$\Delta H_{rxn} = -771.8\ KJ - {}^{-}722.2\ KJ = \boxed{-49.6\ KJ}$$



Exothermic



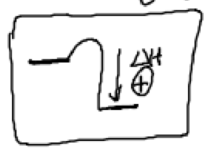
$\underline{2}\ C(s) + \underline{1}\ O_2(g) \rightarrow \underline{2}\ CO(g)$ What is the ΔH_{rxn}?

Compound	ΔH°f (kJ/mol)
$CH_4(g)$	– 74.8
$CO_2(g)$	– 393.5
$CO(g)$	– 110.5

$$(2\ \text{mol} \cdot -110.5\ \text{kJ/mol}) - [(2\ \text{mol} \cdot 0) + (1\ \text{mol} \cdot 0)]$$

$$-221.0 \quad - \quad 0 = \boxed{-221.0\ \text{kJ}}$$

Exo



Compound	ΔH°f (kJ/mol)	Compound	ΔH°f (kJ/mol)
$CH_4(g)$	– 74.8	$NaHCO_3(s)$	– 947.7
$CO_2(g)$	– 393.5	$NaOH(s)$	– 426.7
$CO(g)$	– 110.5	$NH_3(g)$	– 46.2
$HCl(g)$	– 92.3	$NH_4Cl(s)$	– 315.4
$H_2O(g)$	– 241.8	$NO(g)$	+ 90.4
$H_2O(l)$	– 285.8	$NO_2(g)$	+ 33.9
$H_2S(g)$	– 20.1	$SO_2(g)$	– 296.1
$H_2SO_4(l)$	– 811.3	$SO_3(g)$	– 395.2
$MgSO_4(s)$	– 1278.2	$SnCl_4(l)$	– 545.2
$MnO(s)$	– 384.9	$SnO(s)$	– 286.2
$MnO_2(s)$	– 519.7	$SnO_2(s)$	– 580.7
$NaCl(s)$	– 411.0	$ZnO(s)$	– 348.0
$NaF(s)$	– 569.0	$ZnS(s)$	– 202.9

$\underline{1}\ H_2S(g) + \underline{1}\ ZnO(s) \rightarrow \underline{1}\ H_2O(g) + \underline{1}\ ZnS(s)$

$$[(1\ \text{mol} \cdot -241.8\ \text{kJ/mol}) + (1\ \text{mol} \cdot -202.9\ \text{kJ/mol})] - [(1\ \text{mol} \cdot -20.1\ \text{kJ/mol}) + (1\ \text{mol} \cdot -348.0\ \text{kJ/mol})]$$

$$\Delta H_{rxn} = \boxed{-76.6\ \text{kJ}}$$

Calorimetry $q = E = \Delta H$

$KCl(aq) \rightarrow K^{1+}(aq) + Cl^{1-}(aq)$

Example:

12.8g KCl dissolves in 75.0g H_2O, the temp drops from 31.0°C → 21.6°C.

Calculate ΔH for this process

System: $KCl + H_2O$

$q = (\text{mass})(\Delta T)(C_p)$ H_2O

$q = (75.0g)(-10.4°C)(4.184\ J/g°C) =$ ~~$-2949.72J$~~ $-2950J$ or $-2.95KJ$

absorbed by H_2O

$12.8g\ KCl \times \frac{1 mol\ KCl}{74.55g\ KCl} \times \frac{2.95KJ}{1 mol\ KCl} = 17.2KJ$

Practice:

$NaI(s) \rightarrow Na^{1+}(aq) + I^{1-}(aq)$

25.7g NaI 80.0g H_2O, 20.5°C → 24.4°C,

$Energy_{H_2O} = (\text{mass})(T_f - T_i)(C_{p_{H_2O}})$

$q = (80.0g)(24.4 - 20.5°C)(4.184\ J/g°C) = 1.31KJ$ H_2O

$25.7g\ NaI \times \frac{1 mol\ NaI}{149.90g\ NaI} \times \frac{-1.31KJ}{1 mol\ NaI} = -0.225KJ$

Homework - problems 1-19, 24-26