Announcements

Wednesday, October 28, 2009

Exam 2 is next Monday (Nov 2) covering chapters 4, 5, and 6 (thru today's lecture)

Ch 6 MasteringChemistry due NEXT Friday Nov 6. BUT you should still do appropriate problems for exam studying.

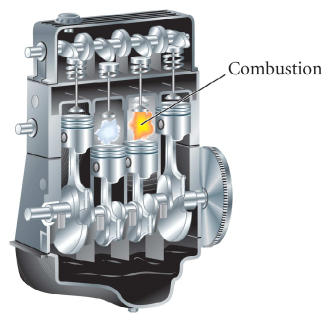
Thermochemistry lab report due a week from Monday (Nov 9).

Reaction work

What if the system is a **chemical reaction** that gives off or absorbs heat?

internal energy — $\Delta E_{rxn} = q + w$ (q: heat transferred; w: work)

The most common work a chemical reaction does is expanding gases against an exterior pressure (pressure-volume work).



$w = -P\Delta V$

But usually we are mostly interested in the heat exchanged with the surroundings

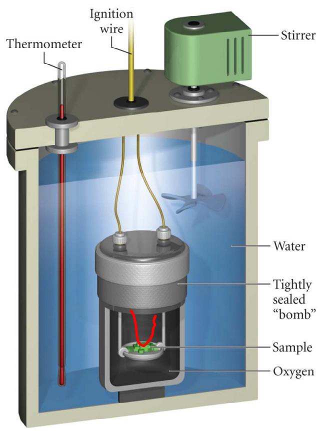
If $\Delta V = 0$, $w = 0$ and $\Delta E_{rxn} = q_V$ (constant volume) — no work

If P is constant, we can use a different state function called enthalpy to only consider heat transferred.

Constant-volume calorimetry *measurement of heat*

A **bomb calorimeter** is used to measure ΔE_{rxn}

The Bomb Calorimeter



high pressure (30 atm)

$\Delta E_{rxn} = q + w$

$w = -P\Delta V$

If $\Delta V = 0$, $w = 0$ and $\Delta E_{rxn} = q_V$ (constant volume)

- q is a measurable energy transfer for a certain experimental process
- State functions like ΔE are more general - they're given for a balanced chemical equation or per mole of a substance

$q_{cal} = C_{cal}\,\Delta t$ (C_{cal} is given) *heat capacity of calorimeter*

$q_{rxn} = -q_{cal}$ *sign change*

$\Delta E_{rxn} = q_{rxn} / n$ *moles*

Constant-volume calorimetery

$q_{cal} = C_{cal}\,\Delta t$ (C_{cal} is given)

$q_{rxn} = -q_{cal}$

$\Delta E_{rxn} = (q_{rxn}) / n$

2.14 g C_5H_{12} is combusted in a bomb calorimeter and the temperature rises from 25.50 °C to 41.23 °C. The heat capacity of the calorimeter is 6.23 kJ/°C. What is ΔE for the combustion of C_5H_{12} in kJ/mol?

$\Delta t = 15.73°C$

$q_{cal} = (6.23\ kJ/°C)(15.73°C) = 97.9979\ kJ$ ~~98.0 kJ~~

$q_{rxn} = -97.9979\ kJ$

$2.14\,g\ C_5H_{12} \times \frac{mol}{72.146\,g} = .02966\ mol\ C_5H_{12}$

$\Delta E = \frac{-97.9979}{0.02966\ mol} = -3.31 \times 10^3\ kJ/mol$

ΔH, Change in enthalpy

It's more convenient to do a reaction open to the atmosphere, where pressure is constant.

The heat exchanged in a specific experiment is q_P and its related state function for a general process is **ΔH** : change in enthalpy.

Since ΔH only involves the heat given off or absorbed, it has the same sign conventions as q:

Endothermic reaction: heat is transferred from surroundings to system (inwards) feels cold
ΔH is: positive
in system

Exothermic reaction: heat is transferred from system to surroundings (outwards) feels hot
ΔH is: negative

Thermochemical equations

ΔH_{rxn} is always given per a balanced chemical equation.
extensive property : depends on amount of substance

Thermochemical equation:

- balanced chemical equation
- thermochemical state function (ΔH_{rxn})
- molar interpretation (coefficients = # moles)

for 2 mol H_2O

$2 H_2(g) + O_2(g) \longrightarrow 2 H_2O(g)$; $\Delta H_{rxn} = -483.7$ kJ for 2 mol H_2O or 2 mol H_2

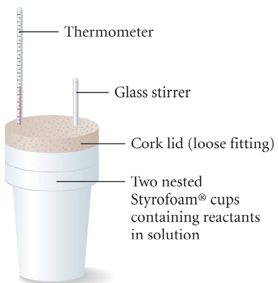
How much heat (in kJ) is given off when 20.0 g H_2 react with excess oxygen?

$$20.0\text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.016\text{ g } H_2} \times \frac{-483.7\text{ kJ}}{2\text{ mol } H_2} = \boxed{-2.40\times10^3\text{ kJ}}$$

Constant-pressure calorimetry

ΔH_{rxn} can be measured using a constant-pressure calorimeter

The Coffee-Cup Calorimeter



We assume all of the heat lost by the reaction is gained by the solution (or vice-versa), and since there is a varying amount of solution, we use its specific heat capacity (C_s or s)

(g) (J/g·°C) (°C)

$q_{soln} = (m_{soln})(C_{s,soln})(\Delta t)$

$q_{rxn} = -m\,C_s\,\Delta t$

will be given on exam →

$q_{surr} = m\,C_s\,\Delta t$

$q_{sys} = -q_{surr}$

If the reaction is made only of dilute solutions, you can make these assumptions:

- $C_{s,soln} = C_{s,water}$ = 4.184 J/g°C
- $m_{solutions} = m_{water}$
- $d_{solutions}$ = 1 g/mL

15 mL HCl(aq)
20 mL NaOH(aq)

$m_{soln} \approx$ 35 g assumed.

Constant-pressure calorimetry problem

50.0 mL of 0.500 *M* $H_2SO_4(aq)$ were added to 20.0 mL of 0.500 *M* NaOH in a coffee-cup calorimeter. The temperature rose from 25.00 °C to 27.20 °C. Assume C_s = 4.184 J/g°C. What is q_{rxn} in kJ?

$q_{soln} = (70.0\,g)(4.184\,J/g°C)(2.20°C)\left(\frac{1\,kJ}{1000\,J}\right) = 0.644336$ kJ

assumed d = 1 g/mL

$q_{rxn} = -0.644336$ kJ

What is ΔH_{rxn} for the following balanced chemical equation?

$H_2SO_4(aq) + 2\,NaOH(aq) \rightarrow Na_2SO_4(aq) + 2\,H_2O(l)$; ΔH_{rxn} = ?

Find the limiting reactant (LR):

50.0 mL H_2SO_4 × $\frac{L}{1000\,mL}$ × $\frac{0.500\,mol\,H_2SO_4}{1\,L}$ × $\frac{1\,mol\,Na_2SO_4}{1\,mol\,H_2SO_4}$ = ~~.025 mol Na_2SO_4~~

20.0 mL NaOH × $\frac{1\,L}{1000\,mL}$ × $\frac{0.500\,mol\,NaOH}{1\,L}$ × $\frac{1\,mol\,Na_2SO_4}{2\,mol\,NaOH}$ = .005 mol Na_2SO_4

limiting reactant

What is n_{LR} for the reaction that was performed? 0.01 mol NaOH

ΔH_{rxn} = coefficient$_{LR}$ x $\frac{q_{rxn}}{n_{LR}}$ ← eqn will be given on exam

$(2\,mol)\left(\frac{-0.644336\,kJ}{0.01\,mol}\right) = -129$ kJ