Energy and Chemical Reactions

Thermochemistry:
heat changes with chemical reactions

Units of energy: Joule (SI) & KJ

calorie (imperial) 1 cal = 4.18 J
nutritional Calorie = 1000 calories

Calorimetry

The measurement of heat changes

Definitions:

specific heat & heat capacity

Specific Heat (s)

Heat required to raise temperature of 1 g of a substance by 1^{o}C

A physical property, constant units: J/g.^{o}C

For water: s = 4.18 J/g.^{o}C

Heat Capacity(C)

Heat required to raise temperature of X g of a substance by 1^{o}C

includes mass term

C = specific heat × mass of substance
(units: J/^{o}C)

To Calculate C

need mass and specific heat

C = m x s

J/EC **gram** **J/gEC**

Calculating total heat (q)

For any substance $q = m \text{ x } s \text{ x })T$

$$)T = T_{final} - T_{initial}$$

How can this equation be used?

1. **calculate heat needed to warm matter**
2. **calculate heat lost when matter cools**

Bath Time

How much heat energy is required to heat 100,000 grams bath water from 25.0 to 55.0 EC ?

q = c x m x)T

q = 4.18 x 100,000 x 30

= 1.25 x 10^7 J

Remember !

density of water = 1.0 g/mL

for water only: mass = volume

100 g water same amount as 100 mL water

Another Example

A 500.0 g block of aluminum cools from 100.0 to 50.0EC.
How much heat is lost ?

q = c x m x)T

$c_{(Al)}$ = 0.900 J/g.EC

q = 0.900 x 500.0 x 50.0
= 2.25 x 10^4 J

Heats of Reaction

Heat is released or absorbed during a chemical change (reaction)

heat released: exothermic reaction

heat absorbed: endothermic reaction

Enthalpy (H)

Heat content of a substance

high heat content: oil, gasoline, wood

low heat content: water

Heat released during chemical reaction =
heat of reaction = $)H = H_{products} - H_{reactants}$

Thermochemical Equations

Show heat evolved during a reaction

$$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

$$2\ H_2 + O_2 \rightarrow 2\ H_2O + 570\ kJ$$

$$2\ \text{moles}\ H_2 + 1\ \text{mole}\ O_2 \rightarrow 570\ kJ$$

$$2\ H_2 + O_2 \rightarrow 2\ H_2O \qquad \Delta H = 570\ kJ$$

Thermochemical Equations

Show heat evolved during a reaction

....combustion of octane

$2\ C_8H_{18} + 25\ O_2$ ¸ $16\ CO_2 + 9\ H_2O \quad + 1.1 \times 10^4$ kJ

2 moles octane or 228 g

Thermochemical Equations

$2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 9\ H_2O + 1.1 \times 10^4$ kJ

228 g octane produce **1.1×10^4 kJ** of heat

How much heat would 1 g of octane produce?

Thermochemical Equations

How much heat would 1 g of octane produce?

set up ratio $\frac{228 \text{ g octane}}{1.1 \times 10^4 \text{ kJ}} = \frac{1 \text{ g octane}}{? \text{ kJ}}$

$$? \text{ kJ} = 228 \text{ g} \times 1.1 \times 10^4 \text{ kJ} = 2.5 \times 10^6 \text{ kJ}$$

Other Energy Rich Compounds

methane CH_4

acetylene C_2H_4

sugars $C_6H_{12}O_6$ $C_{11}H_{22}O_{12}$

energy stored in chemical bonds

Hess's Law

Chemical equations can be added
)H values also added
Use to find)H for a reaction

H_2O (s) ¸ H_2O (l))H = + 6 kJ
H_2O (l) ¸ H_2O (g))H = + 44 kJ

H_2O (s) ¸ H_2O (g))H = + 50 kJ

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H_2O (s) ¸ H_2O (l))H = + 6 kJ
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Hess's Law

Chemical equations can be added
)H values also added
Use to find)H for a reaction

H_2O (s) ,   )H = + 6 kJ
                , H_2O (g))H = + 44 kJ

H_2O (s) , H_2O (g)   )H = + 50 kJ

Hess's Law

<u>Rule 1.</u> When an equation is reversed)H sign is also reversed

H_2O (s) ‚ H_2O (l))H = + 6 kJ

H_2O (l) ‚ H_2O (s))H = - 6 kJ

Hess's Law

<u>Rule 2</u>. When equation multiplied by integer, multiply ΔH by same amount

$$H_2O\,(s) \rightarrow H_2O\,(l) \qquad \Delta H = +6\ kJ$$

$$2\,H_2O\,(s) \rightarrow 2H_2O\,(l) \qquad \Delta H = +12\ kJ$$

Hess's Law

Use Hess's Law and rules to find)H for a new equation, using others

How would you find)H for:

$2\ S\ +\ 3\ O_2\ \ ,\ \ 2\ SO_3$

Hess's Law

Need related equations and ΔH values

A. $S + O_2 \rightarrow SO_2 \quad \Delta H = -297\ kJ$
B. $2\ SO_2 + O_2 \rightarrow 2\ SO_3 \quad \Delta H = -198\ kJ$

Somehow rearrange above to form

$2\ S + 3\ O_2 \rightarrow 2\ SO_3 \quad \Delta H = ?$

Hess's Law

First, try adding equations A and B

A. $S + O_2 \rightarrow SO_2 \quad \Delta H = -297\ kJ$

B. $2\ SO_2 + O_2 \rightarrow 2\ SO_3 \quad \Delta H = -198\ kJ$

$$S + O_2 + 2\ SO_2 + O_2 \rightarrow SO_2 + 2\ SO_3$$

$$= \quad S + 2\ O_2 + SO_2 \rightarrow 2\ SO_3$$

This is NOT:

$$2\ S + 3\ O_2 \rightarrow 2\ SO_3$$

Hess's Law

Need 2 S and 3 O_2 on left side

Multiply equation A by 2 then add B

A. 2 X (S + O_2 ¸ SO_2)) H = -297 kJ

B. 2 SO_2 + O_2 ¸ 2 SO_3) H = -198 kJ

2 S + 2 O_2 + 2 SO_2 + O_2 ¸ 2 SO_2 + 2 SO_3

= 2 S + 2 O_2 + 2 SO_2 + O_2 ¸ 2 SO_2 + 2 SO_3

= 2 S + 3 O_2 + 2 SO_2 ¸ 2 SO_2 + 2 SO_3

= 2 S + 3 O_2 ¸ 2 SO_3

Hess's Law

Multiply)H of equation A by 2 then add to)H of equation B

A. 2 X)H = -297 kJ

B.)H = -198 kJ

Y **2 S + 3 O_2 ¸ 2 SO_3)H = -792 kJ**

Hess's Law

Can also use heats of formation to find)H for a specific reaction

Heat associated with the formation of 1 mole of compound from its elements at standard conditions

Values in Table 6.4

Hess's Law

$\frac{1}{2}\ N_2 + 3/2\ H_2 \rightarrow NH_3 \quad \Delta H = -46\ kJ/mol$

Also written: $\Delta H_f^{o}(NH_3) = -\ 46\ kJ$

Hess's Law

$$) H_f^o(NH_3) = - 46 \text{ kJ}$$

standard conditions

write as an equation &
find) H as before