PROBLEM SET #5

5.1 Potential vs Kinetic — Energy of motion

Energy of position

They can be equal in terms of magnitudes

5.3 Ⓐ $E_K = \frac{1}{2} mv^2$ $E_K = \frac{\left[45g \frac{1kg}{1000g}\left(\frac{61m}{s}\right)^2\right]}{2} = 83.72\ J = \boxed{84\ J}$

Ⓑ $83.72\ J \times \frac{1\ cal}{4.184\ J} = \boxed{20\ cal}$ Ⓒ $E_K \rightarrow 0$ transferred to sand, heat ball deformation

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5.8 A 100 watt light bulb uses current from wiring, Energy travels through the filament (W) producing heat and light (both energy)

- A human gets energy from Food. Food stuffs have chemical energy stored in them to do work (via muscles) and produce heat. The heat in both the bulb and human is wasted, as in not doing work.

5.11 Ⓐ System, well-defined (the chemicals), not the Surroundings (i.e. container, Air, liquid medium, etc). as in ↑

Ⓑ A closed system can exchange heat, but not mass w/ the Surroundings

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5.17 Ⓐ In any chemical or physical Δ, energy cannot be created or destroyed (it can Δ form)

Ⓑ Internal E (E) of a closed system is the total of all E_K's, E_p

Ⓒ Internal E increases when work is done on the system by the surroundings, or when heat is transferred to the system by the surroundings.

5.20 Ⓐ Calculate the Δ in internal E. use $\Delta E = q + w$

For exothermic q is negative, endothermic q is positive.

For Ⓐ q is positive, w is negative $\Delta E = {}^{+}900\,J - 422\,J$

$= +478\,J.$

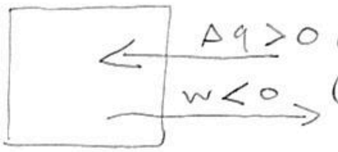
endothermic

Ⓑ q is negative (loses heat to surroundings - hence cooling the system)

So, work is ~ 0. $\Delta E = -3140\,J$ exothermic

Ⓒ q is negative, $w \simeq 0$ so, $-8.65\,kJ$ exothermic

5.22 Ⓐ



$\Delta q > 0$ (not exothermic - heat added to system)

$w < 0$ (< 0 = w on surroundings)

Ⓑ $\overset{endo}{\Delta E}$ is positive if ($\Delta E = q + w$) if magnitude of w is > than q.

$\Delta E = +q - w$

Ⓒ $\overset{exo}{\Delta E}$ is negative if → ($\Delta E = q + w$) if magnitude of q is greater than w

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5.27 Ⓐ At constant pressure there is little to no work $\Delta E = q + \underset{\overset{\approx}{0}}{w}$

So heat transfer is obvious.

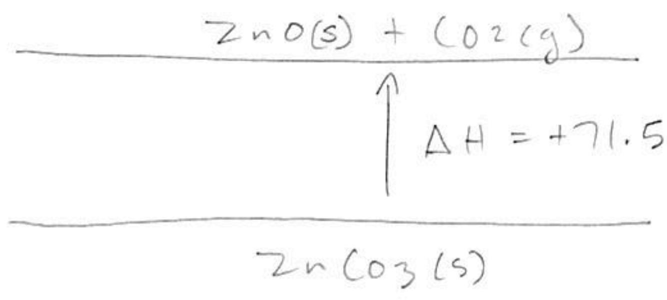
Ⓑ H is state during constant pressure. State is based on present state (constant pressure)

Therefore, ΔH = heat loss or gain (not work)

Ⓒ ΔH is negative (enthalpy decreases (total $E_p + E_K$)) so heat is lost to the surroundings - exothermic!

5.30 Ⓐ $ZnCO_3(s) \longrightarrow ZnO(s) + CO_2(g)$ $\Delta H^\circ = 71.5\,kJ$

Ⓑ



5.38 At constant volume ($\Delta V = 0$), $\Delta E = q_v$. According to the definition of enthalpy, $H = E + PV$, so $\Delta H = \Delta E + \Delta(PV)$. For an ideal gas at constant temperature and volume $\Delta PV = V\Delta P = RT\Delta n$. For this reaction, there are 2 mol of gaseous product and 3 mol of gaseous reactants, so

$\Delta n = -1$. Thus $V\Delta P$ or $\Delta(PV)$ is negative. Since $\Delta H = \Delta E + \Delta(PV)$, the negative $\Delta(PV)$ term means that ΔH will be smaller or more negative than ΔE.

5.50 Following the logic in Solution 5.49, the dissolving process is endothermic, ΔH is positive. The total mass of the solution is (60.0 g H_2O + 3.88 g NH_4NO_3) = 63.88 = 63.9 g. The temperature change of the solution is 23.0 - 18.4 = 4.6°C. The heat lost by the water is

$$63.88 \text{ g } H_2O \times \frac{4.184 \text{ J}}{1 \text{ g}\bullet{}^\circ\text{C}} \times 4.6^\circ\text{C} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 1.229 = 1.2 \text{ kJ}$$

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Thus, 1.2 kJ is absorbed when 3.88 g $NH_4NO_3(s)$ dissolves.

$$\frac{+1.229 \text{ kJ}}{3.88\ NH_4NO_3} \times \frac{80.05 \text{ g } NH_4NO_3}{1 \text{ mol } NH_4NO_3} = +25.36 = +25 \text{ kJ/mol } NH_4NO_3$$

5.51 *Analyze/Plan*. Follow the logic in Sample Exercise 5.8. *Solve*:

$q_{bomb} = -q_{rxn}$; $\Delta T = 30.57^\circ C - 23.44^\circ C = 7.13^\circ C$

$$q_{bomb} = \frac{7.854 \text{ kJ}}{1^\circ\text{C}} \times 7.13^\circ\text{C} = 56.00 = 56.0 \text{ kJ}$$

At constant volume, $q_V = \Delta E$. ΔE and ΔH are very similar.

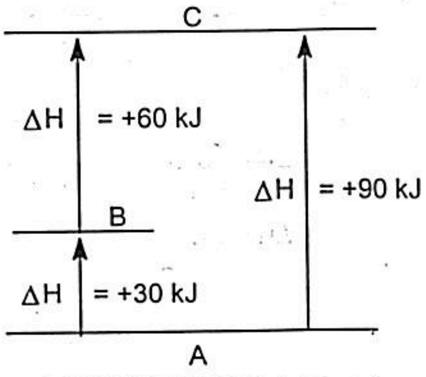
$$\Delta H_{rxn} \approx \Delta E_{rxn} = q_{rxn} = -q_{bomb} = \frac{-56.0 \text{ kJ}}{2.20 \text{ g } C_6H_4O_2} = -25.454 = -25.5 \text{ kJ/g } C_6H_4O_2$$

$$\Delta H_{rxn} = \frac{-25.454 \text{ kJ}}{1 \text{ g } C_6H_4O_2} \times \frac{108.1 \text{ g } C_6H_4O_2}{1 \text{ mol } C_6H_4O_2} = -2.75 \times 10^3 \text{ kJ/mol } C_6H_4O_2$$

5.57 (a) *Analyze/Plan*. Arrange the reactions so that in the overall sum, B appears in both reactants and products and can be canceled. This is a general technique for using Hess's Law. *Solve*:

A → B	ΔH = +30 kJ
B → C	ΔH = +60 kJ
A → C	ΔH = +90 kJ

(b)



Check. The process of A forming C can be described as A forming B and B forming C.

5.60 $3H_2(g) + 3/2\ O_2(g) \rightarrow 3H_2O(g)$ $\Delta H = 3/2(-483.6\ \text{kJ})$

$O_3(g) \rightarrow 3/2\ O_2(g)$ $\Delta H = 1/2(-284.6\ \text{kJ})$

$3H_2(g) + O_3(g) \rightarrow 3H_2O(g)$ $\Delta H = -867.7\ \text{kJ}$

5.67

(a) $1/2\ N_2(g) + 3/2\ H_2(g) \rightarrow NH_3(g)$ $\Delta H_f^\circ = -80.29\ \text{kJ}$

(b) $1/8\ S_8(s) + O_2(g) \rightarrow SO_2(g)$ $\Delta H_f^\circ = -269.9\ \text{kJ}$

(c) $Rb(s) + 1/2\ Cl_2(g) + 3/2\ O_2(g) \rightarrow RbClO_3(s)$ $\Delta H_f^\circ = -392.4\ \text{kJ}$

(d) $N_2(g) + 2H_2(g) + 3/2\ O_2(g) \rightarrow NH_4NO_3(s)$ $\Delta H_f^\circ = -365.6\ \text{kJ}$

5.68

(a) $1/2\ H_2(g) + 1/2\ Br_2(l) \rightarrow HBr(g)$ $\Delta H_f^\circ = -36.23\ \text{kJ}$

(b) $Ag(s) + 1/2\ N_2(g) + 3/2\ O_2(g) \rightarrow AgNO_3(s)$ $\Delta H_f^\circ = -124.4\ \text{kJ}$

(c) $2Hg(l) + Cl_2(g) \rightarrow Hg_2Cl_2(s)$ $\Delta H_f^\circ = -264.9\ \text{kJ}$

(d) $2C(s,\ gr) + 1/2\ O_2(g) + 3H_2(g) \rightarrow C_2H_5OH(l)$ $\Delta H_f^\circ = -277.7\ \text{kJ}$

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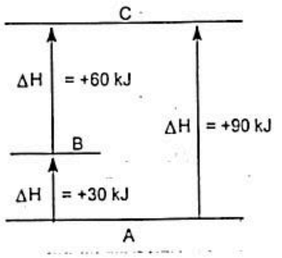

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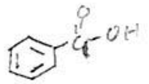
Additional Calorimetry PROBLEMS

$\Delta T = T_f - T_i$ $55 - 4 = 51$

1) $q = ms\Delta t$ $q = (25.0\,g)(0.902\,J/g°C)(51°C)$

$q = \boxed{1150\,J}$

2) Benzoic Acid  has a molar mass of $\sim 122\,g/mol$.

$q_{Rxn} = 1.0236\,g_{\text{benzoic A}} \times \frac{1\,mol}{122\,g} \times \frac{-3226.7\,kJ}{1\,mol} = \underline{-27.073\,kJ}$

$q(Rxn) = -[q(H_2O) + q(bomb)]$ $-27{,}073\,J = -(ms\Delta T + C_{bomb}\Delta T)$

$-27{,}073\,J = -[(1000\,g \cdot 4.184\,J/g°C \cdot 3.81°C) + (C_{bomb} \cdot 3.81°C)]$

$-27{,}073\,J = -15941\,J + -3.81°C \cdot C_{bomb}$

$-11{,}132\,J = -3.81°C\,C_{bomb}$ / $C_{bomb} = \frac{-11{,}132\,J}{-3.81°C} =$

$\boxed{2922\,J/°C \text{ or } 2.92\,kJ/°C}$

3) $q_{H_2O} = -(q_{bomb})$ $ms\Delta T = -(C\Delta T)$ $\Delta T_{H_2O} = 32 - 50 = -18°C$

$(150\,g)(4.184\,J/g°C)(-18°C) = -(C(12°C))$ $\Delta T_{bomb} = 32 - 20 = 12°C$

$\boxed{941\,J/°C = C}$

4) $q(Rxn) = -(q(H_2O) + q(cal))$ / $q(Rxn) = -[(ms\Delta T) + (C\Delta T)]$

$q_{Rxn} = -[(1200\,g)(4.184\,J/g°C)(1.83°C) + (1726\,J/°C)(1.83°C)]$

$q_{Rxn} = -(9188.1\,J + 3158.6\,J)$

$q_{Rxn} = \frac{-12346\,J}{0.1111\,g} = \boxed{-111{,}125\,J/g \text{ or } -111.125\,kJ/g}$

Calorimetry PRACTICE (w FOOD?)

① $q(p.corn) = -[q(H_2O)]$ (no Ccal info available.)

$q(p.corn) = -[50.0\,g \cdot 4.184\,J/g°C \cdot 24°C] = -5020.8\,J$

$= \boxed{\frac{5.02\,kJ}{gram}}$ or $\frac{1200\,cal}{gram}$ or $\frac{1.2\,Cal\,(kcal)}{gram}$

② $110\,Cal = 110{,}000\,cal \times \frac{4.184\,J}{cal} = 460240\,J$ or

$\frac{460.24\,kJ}{30\,g} = \boxed{\frac{15.34\,kJ}{gram}}$

much heat was lost to the surroundings, no insulation (combustion occurred outside of cal.

③ $q(snack)_{3\,grams} = -[q(H_2O)]$ no Ccal data avail.

$\Delta T = 32.4°C - 25.0°C = 7.4°C$

$q(snack) = -(2000\,g \cdot 4.184\,J/g°C \cdot 7.4°C) = -61923.2\,J$

$61{,}923.2\,J \; \frac{1\,cal}{4.184\,J} \; \frac{1\,Cal}{1000\,cal} = \frac{14.8\,Cal}{3\,grams} = 4.93\,Cal/gram$

④ $q(Rxn)_{1.0\,g} = -[1000\,g \; 4.184\,J/g°C \; 8.05°C] = \frac{-33681.2\,J}{gram}$

$\frac{-33681.2\,J}{gram} \; \frac{65.8\,grams}{mol} = \frac{-2{,}216{,}223\,J}{mol}$ or $\frac{-2{,}216\,kJ}{mol}$

Additional Calorimetry Problems

$\Delta T = T_f - T_i$

$55 - 4 = 51$

① $q = ms\Delta t$ $\quad q = (25.0\,g)(0.902\,J/g°C)(51°C)$

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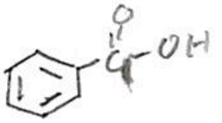
2) Benzoic Acid  has a molar mass of ~122 g/mol

$q_{Rxn} = 1.0236\,g\ \text{benzoic A} \times \frac{1\,mol}{122\,g} \times \frac{-3226.7\,kJ}{1\,mol} = \underline{-27.073\,kJ}$

$q(Rxn) = -[q(H_2O) + q(bomb)]$ $\quad -27{,}073\,J = -(ms\Delta T + C_{bomb}\Delta T)$

$-27{,}073\,J = -[(1000\,g \cdot 4.184\,J/g°C \cdot 3.81°C) + (C_{bomb} \cdot 3.81°C)]$

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$-11{,}132\,J = -3.81°C\ C_{bomb}$ / $C_{bomb} = \frac{-11{,}132\,J}{-3.81°C} =$

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