③

also called U

Internal energy (E) →

products

Reactants

Since the products have more energy than reactants,

(a) the diagram represents an increase in internal energy.

(b) therefore the sign of ΔE is positive.

   (or you could say ΔU is positive)

(c) If no work is done, is the process exo or endo?

$$\Delta E = q + w = q + \emptyset$$

Since ΔE is positive, q must be positive, so the process is endothermic.

④

(i)

q

w

(ii)

w

(iii)

q

w

(i) $q = \ominus$
     $w = \oplus$
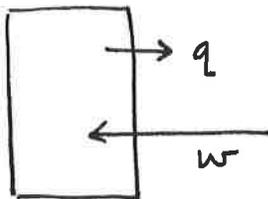
(ii) $q = 0$
     $w = \oplus$

(iii) $q = \oplus$
      $w = \ominus$

$\Delta E = q + w$
will come out positive since $|w| > |q|$ and w is positive.

$\Delta E = q + w$
will come out positive
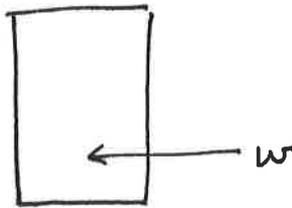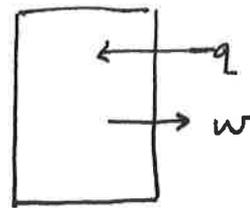
$\Delta E = q + w$
will come out positive since $|q| > |w|$ and q is positive.

(a) The third one is endothermic since q is positive.

(b) ΔE is not negative for any of these. it is $\oplus$ for all 3.

(c) ΔE is positive for all 3 systems, so all 3 systems experience a net gain in internal energy.

(17) BTU = British Thermal Unit (a unit of energy)

1 BTU = heat energy req'd to raise temp of 1 lb H_2O by 1 °Fahrenheit.

Convert to Joules!

1 lb = 453.59 g

$°F = \left(\frac{9}{5}\right)(°C) + 32$ ← so when something has a temp change of 1 °F, it has a temp change of $\left(\frac{5}{9}\right)$ degrees Celsius

$$q = mc\Delta T = (453.59\,g)\left(\frac{4.184\,J}{g\,°C}\right)\left(\frac{5}{9}\,°C\right) = \boxed{1054\ Joules}$$

$$\boxed{1\ BTU = 1054\ J}$$

(18) Power = energy / time

1 Watt = 1 Joule per second $\left(1\,\frac{J}{s}\right)$

(a) how many J in 1 kW-hr?

$$(1\,kW\text{-}hr)\left(\frac{1000\,W}{1\,kW}\right)\left(\frac{1\,J/s}{1\,W}\right)\left(\frac{3600\,s}{hr}\right) = 3.6 \times 10^6\ Joules\ \text{in one kW-hr}$$

(b) An adult radiates heat to surroundings at same rate as 100 W incandescent bulb! How much energy (in kcal!) radiated per day?

100 W = 100 J/s

$$(24\,hr)\left(\frac{3600\,s}{hr}\right)\left(\frac{100\,J}{s}\right)\left(\frac{1\,cal}{4.184\,J}\right)\left(\frac{1\,kcal}{1000\,cal}\right) = \boxed{2065\ kcal}$$

In other words, ≈ 2000 Dietary Calories are needed to maintain body temperature since we radiate heat. !!!

(27) Calculate ΔE (aka ΔU)
and classify as endo or exothermic

(a) $q = (0.763 \text{ kJ})\left(\frac{1000 \text{ J}}{\text{ kJ}}\right) = 763 \text{ J}$
$w = -840 \text{ J}$

$q = .763 \text{ kJ}$

$w = -840 \text{ J}$
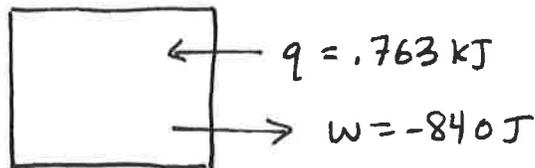
$\Delta E = q + w = 763 \text{ J} + (-840 \text{ J})$

$\Delta E = -77$ Joules   (or $-.077$ kJ)
endothermic since q was positive

(b) a system releases 66.1 kJ to surroundings
and surroundings do 44.0 kJ work on system.
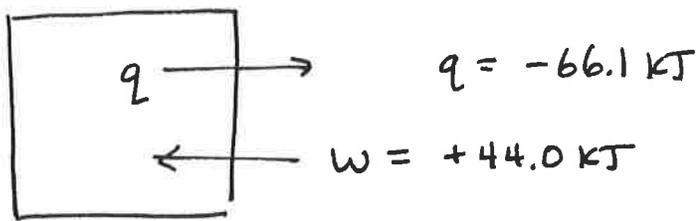
$q = -66.1 \text{ kJ}$

$w = +44.0 \text{ kJ}$

$\Delta E = q + w = -66.1 \text{ kJ} + 44.0 \text{ kJ} = $

-22.1kJ
exothermic since
q is negative; heat
energy released
by system

(c) System absorbs 7.25 kJ heat from
surroundings while volume remains constant.

if $\Delta V = 0$,   no p-v work was done, so $w = 0$

$\Delta E = q + w = $ $7.25 \text{ kJ} + 0 = \boxed{7.25 \text{ kJ}}$

endothermic since
q was positive

(42) Which has higher enthalpy?

(a)   1 mole $CO_{2(s)}$ *or*   1 mole $CO_{2(g)}$ (at same Temp)

↑
this one!
sublimation (solid → gas)
is endothermic.

(b)   2 moles hydrogen atoms   *or*   1 mole H_2

↑
this one.

$$H\bullet \quad + \quad H\bullet \quad \rightarrow \quad H-H \quad + \text{ energy}$$
(energy is released when bonds form)
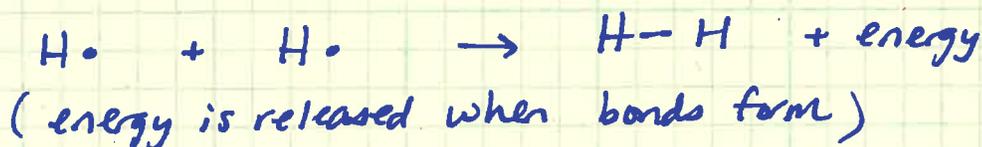
(c) (1 mole $H_{2(g)}$ and 0.5 mole $O_{2(g)}$) at 25°C
    or   (1 mole H_2O) at 25°C

↗
this has more enthalpy

$$H_2 \; + \; 0.5\, O_{2} \rightarrow H_2O_{(g)} \; + \text{ energy}$$
     (g)   (g)
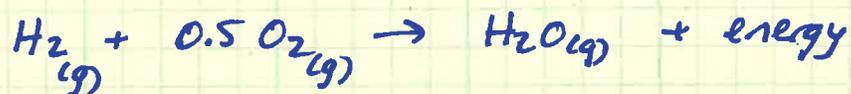
combustion of hydrogen releases energy since
$H_2O_{(g)}$ has lower potential energy than $H_{2(g)}$ & $O_{2(g)}$

(d) 1 mole $N_{2(g)}$ at 100°C   *or*   1 mole $N_{2(g)}$ at 300°C

↑
this one.
higher temperature → higher
kinetic energy
of molecules

(44.) $2\ CH_3OH_{(g)} \longrightarrow 2\ CH_{4(g)} + O_{2(g)}$ $\Delta H = +252.8\ kJ$
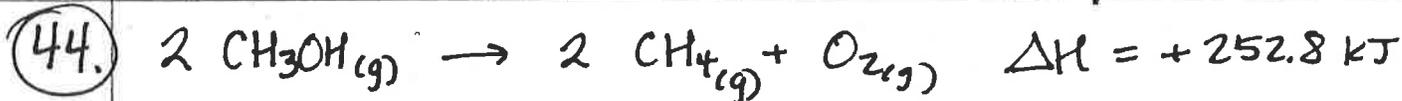
(a) The rxn is [endothermic] since ΔH is positive.

(b) calculate heat transferred if $24.0\ g$ CH_3OH decomposes @ const. P

$$(24.0\ g\ CH_3OH)\left(\frac{1\ mole}{32.042\ g}\right)\left(\frac{252.8\ kJ}{2\ mole\ CH_3OH}\right) = 94.676\ kJ \rightarrow \boxed{\begin{array}{c} 94.7\ kJ \\ absorbed \end{array}}$$

(c) if the enthalpy change is $82.1\ kJ$, what mass of methane can be produced from methanol?

$$(82.1\ kJ)\left(\frac{2\ mole\ CH_4}{252.8\ kJ}\right)\left(\frac{16.0428\ g}{mole}\right) = \boxed{10.4\ grams\ CH_4}$$

(d) If $38.5\ g$ CH_4 react completely w/ O_2, to form CH_3OH, How much energy will be released?

(If we reverse the rxn, $2\ CH_4 + O_2 \rightarrow 2\ CH_3OH$, $\Delta H = -252.8\ kJ$) this rxn is exo.

$$(38.5\ g\ CH_4)\left(\frac{1\ mole}{16.0428\ g}\right)\left(\frac{252.8\ kJ}{2\ mole\ CH_4}\right) = \boxed{303\ kJ}$$

(45.) $Ag^{+1}_{(aq)} + Cl^{-1}_{(aq)} \longrightarrow AgCl_{(s)}$ $\Delta H = -65.5\ kJ$

(a) Calculate ΔH for production of 0.450 mole $AgCl$

$$(.450\ mole\ AgCl)\left(\frac{-65.5\ kJ}{1\ mole\ AgCl}\right) = -29.475 \rightarrow \boxed{-29.5\ kJ}$$

or you could also say that $29.5\ kJ$ are released.. but they asked for "ΔH"

(b) calculate ΔH for formation of $9.00\ g$ $AgCl$

$$(9.00\ g\ AgCl)\left(\frac{1\ mole}{143.321\ g}\right)\left(\frac{-65.5\ kJ}{1\ mole\ AgCl}\right) = \boxed{-4.11\ kJ}$$

(c) calculate ΔH if 9.25×10^{-4} mole $AgCl$ dissolves into H_2O if $AgCl_{(s)}$ is dissolving into H_2O, the reverse "rxn" would occur:

$$AgCl_{(s)} \longrightarrow Ag^{+1}_{(aq)} + Cl^{-1}_{(aq)}\qquad \Delta H = +65.5\ kJ$$
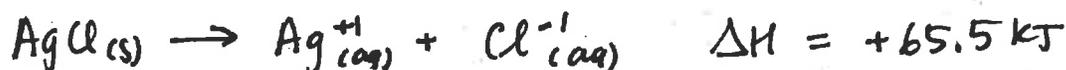
this is $\Delta H_{solution}$ of $AgCl$

$$(9.25 \times 10^{-4}\ mole)\left(\frac{65.5\ kJ}{mole\ AgCl}\right) = 0.06059\ kJ \rightarrow \boxed{\begin{array}{c}0.0606\ kJ\\ or\ 60.6\ J\end{array}}$$

← this amt of. energy is absorbed.

(50) Two solids, "Object A" and "Object B" left in boiling water til they come to same temp as boiling H_2O. Then, both removed, placed in separate beakers containing $\wedge$ H_2O at 10.0°C.

      1000 g

Object A increases the cold H_2O temp by 3.50°C (10.0°C → 13.5°C)
Object B increases the cold H_2O temp by 2.60°C (10.0°C → 12.6°C)

(a) Which Object had the higher heat capacity?

> Object A had the higher heat capacity. Though both "A" and "B" started at 100°C (temp of boiling H_2O), A transferred more energy to the cold water than B did
> (A caused a larger ΔT; +3.50°C vs +2.60°C)
> Therefore A contained more thermal energy than B at a given temp,

> $\boxed{\text{So A had the higher heat capacity.}}$

> (We could solve for it :
>
> $$q = mc\,\Delta T \qquad\qquad q = C\,\Delta T \qquad\qquad \text{or } C = \frac{q}{\Delta T}$$
>   ↑ specific heat   ↑ heat capacity
>
> q (from A) $= mc\,\Delta T = (1000\,g)(1\,\frac{cal}{g\cdot c})(3.5°c) = 3500\ cal$
> q (from B) $= mc\,\Delta T = (1000\,g)(1\,cal/g\,c)(2.6°c) = 2600\ cal$
>
> $$C_A = \frac{|q_A|}{|\Delta T_A|} = \frac{3500\ cal}{100°C - 13.5°C} = \frac{3500\ cal}{86.5°C} = 40.5\ cal/°C$$
>
> $$C_B = \frac{|q_B|}{|\Delta T_B|} = \frac{2600\ cal}{100 - 12.6°C} = \frac{2600\ cal}{87.4°C} = 29.7\ cal/°C$$
>
> So $C_A > C_B$!)

(b) What can we say about the specific heats?

> Not much ... we don't know the mass of either solid (or whether they are equal), so we can't compare the specific heats.
>
> A could have the higher heat capacity due to having a higher specific heat (c), or a higher mass, or both!
>
> $$C = mc \qquad\qquad \text{A has a higher "mc" than B.}$$
>   ↑ heat capacity   ↑ specific heat

(51) (a) specific heat of liquid H_2O: $\boxed{1.000 \, cal/g\,^\circ C}$ or $\boxed{4.184 \, J/g\,^\circ C}$

(b) <u>molar</u> heat capacity of $H_2O_{(\ell)}$

$$\left(\frac{1.000 \, cal}{g\,^\circ C}\right)\left(\frac{18.0152 \, g}{mole}\right) = \boxed{\frac{18.02 \, cal}{mole \, ^\circ C}} \qquad (or \; 75.38 \, J/mole \, ^\circ C)$$

(c) heat capacity of $185 \, g \; H_2O$

$$q = mc\Delta T \qquad / \qquad q = C\Delta T \qquad\qquad C = \text{heat capacity} = mc$$

$\uparrow$ specific heat $\qquad\qquad$ $\uparrow$ heat capacity

$$C = mc = (185g)\left(1.000 \, \tfrac{cal}{g\,^\circ C}\right) = 185 \, \tfrac{cal}{^\circ C} \qquad \left(or \; 774 \, \tfrac{J}{^\circ C}\right)$$

(d) how many kJ needed to raise temp of $10.00 \, kg \; H_2O$ from $24.6\,^\circ C$ to $46.2\,^\circ C$?

$$q = mc\Delta T = (10.00 \, kg)\left(\frac{1000 \, g}{1 \, kg}\right)\left(1 \, \tfrac{cal}{g\,^\circ C}\right)(46.2\,^\circ C - 24.6\,^\circ C)$$

$$q = 216000 \, cal$$

$$\left(216000 \, cal\right)\left(\frac{4.184 \, J}{cal}\right)\left(\frac{1 \, kJ}{1000 \, J}\right) = 903.7 \rightarrow \boxed{904 \, kJ}$$

(52.) (a) Which Substance in Table 5.2 (page 176) requires the smallest amount of energy to increase temp of 50.0 g substance by 10 Kelvin?

This would be the substance with the lowest "c" value (specific heat) which is <u>mercury</u>! $\quad c = 0.14 \, J/g \cdot K$

$$\boxed{Hg_{(\ell)}}$$

(b) Calculate the energy req'd !

$$q = mc\Delta T = (50.0 \, g)\left(0.14 \, \tfrac{J}{g \cdot K}\right)(10 \, K) = \boxed{70 \; Joules}$$

(55) $NaOH_{(s)} \longrightarrow Na^{+1}_{(aq)} + OH^{-1}_{(aq)}$

6.50 g NaOH into 100.0 g H_2O

$T_i = 21.6°C$

$T_f = 37.8°C$

$> \Delta T = 37.8°C - 21.6°C = 16.2°C$

Find ΔH_{soln}

$q_{H_2O} = m \, c \, \Delta T_{H_2O} = 100.0g \left(1 \frac{cal}{g°C}\right)(16.2°C) = 1620 \, cal$

$\left(\begin{array}{l} or, if you include the 6.50 g NaOH, \\ q = (106.5g)\left(1 \frac{cal}{g°C}\right)(16.2°C) = 1725.3 \, cal \end{array}\right)$

$(1620 \, cal)\left(\frac{4.184 J}{cal}\right)\left(\frac{1 kJ}{1000 J}\right) = 6.77808 \, kJ$ absorbed by H_2O

$(6.50 g \, NaOH)\left(\frac{1 mole}{39.9971 g}\right) = 0.16251$ moles NaOH

$q_{H_2O} = -q_{NaOH}$

The H_2O absorbed 6.778 kJ, So the NaOH released 6.778 kJ

$q_{NaOH} = -q_{H_2O} = -6.77808 \, kJ$

$\Delta H = \frac{-6.77808 \, kJ}{.16251 \, moles} = -41.7 \, kJ/mole$

$\left(\begin{array}{l} or \; -44.4 \frac{kJ}{mole} \\ if include the \\ NaOH \, mass \end{array}\right)$

(56) $NH_4NO_{3(s)} \longrightarrow NH_4^{+1} + NO_3^{-1}_{(aq)}$
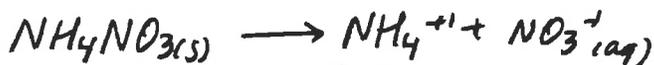
4.25 g in 60.0 g H_2O

$T_i = 22.0°C$, $T_f = 16.9°C$

(a) $\Delta T = T_f - T_i = 16.9°C - 22.0°C = -5.1°C$

$q_{H_2O} = mc\Delta T = (60.0g)\left(4.184 \frac{J}{g°C}\right)(-5.1°C) = -1280.3 \, Joules$

$(-1280.3 J)\left(\frac{1 kJ}{1000 J}\right) = -1.280304 \, kJ$

$\left(\begin{array}{l} or, if you include the \\ 4.25g in the mass, \\ you get \end{array}\right)$

(b) (The "rxn" was endo thermic. the H_2O released 1.28 kJ, and the NH_4NO_3 absorbed the 1.28 kJ)

$(4.25 g \, NH_4NO_3)\left(\frac{1 mole}{80.0432 g}\right) = 0.053096$ moles NH_4NO_3

$q_{NH_4NO_3} = -q_{H_2O} = +1.280304 \, kJ$

$\frac{1.280304 \, kJ}{0.053096 \, mole} = \boxed{24 \, kJ/mole}$

$\left(\begin{array}{l} or \; 26 \, kJ/mole \\ if you include the mass \end{array}\right)$

(58.) 1.800g phenol (C_6H_5OH) burn in a calorimeter.

$C_{calorimeter}$ = 11.66 kJ/°C T_i = 21.36°C, T_f = 26.37°C

(a)

$$C_6H_5OH + 7\,O_2 \longrightarrow 6\,CO_2 + 3\,H_2O$$
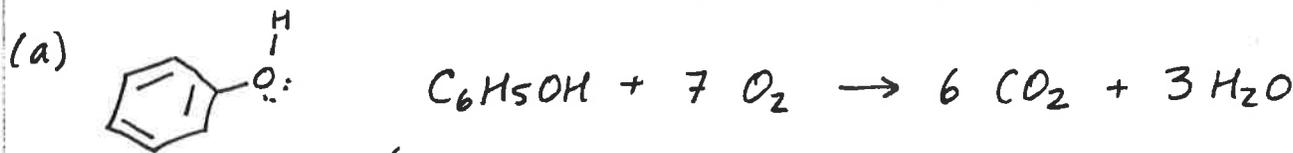
(b) $q = C\Delta T = \left(11.66\,\frac{kJ}{°C}\right)(26.37°C - 21.36°C) = 58.4166\ kJ$

$\underbrace{\qquad}_{5.01°C}$

q calorimeter = 58.4166 kJ   ← the calorimeter absorbs

q rxn = - q calorimeter = -58.4166 kJ   the heat released by rxn.

$(1.800\ g\ C_6H_5OH)\left(\dfrac{1\,mole}{94.11289}\right) = 0.019126$ moles C_6H_5OH

Heat of combustion in kJ/g $= \dfrac{-58.4166\,kJ}{1.800\,g} = \boxed{-32.5\ kJ/g}$

59m next page Heat of combustion in kJ/mole $= \dfrac{-58.4166\,kJ}{.019126\,mole} = \boxed{-3050\ kJ/mole}$

these could be reported as positive values instead.

(59) **(60)** $\Delta H_c\ (C_6H_5COOH) = 26.38\ kJ/g$

2.760 g benzoic acid burn in bomb cal

T_i = 21.60°C, T_f = 29.93°C $\Delta T = 8.33°C$

$|q_{combustion}| = \left(26.38\,\dfrac{kJ}{g}\right)(2.760\,g) = 72.8088$ kJ released by rxn (combustion)

q calorimeter $= C_{calorimeter}\ \Delta T_{calorimeter}$

72.8088 kJ $= C(8.33°C)$ $C_{calorimeter} = 8.74055\ kJ/°C$

(a) $\boxed{8.74\ kJ/°C}$

(b) Now 1.440 g of a new organic substance are combusted in Same calorimeter. T_i = 22.14°C, T_f = 27.09°C

$q_{cal} = C_{cal}\ \Delta T = (8.74055\ kJ/°C)(4.95°C) = 43.2657$ kJ

$\dfrac{43.2657\ kJ}{1.440\ g} = 30.0456 \longrightarrow \boxed{30.0\ kJ/g}$ (or negative 30.0 kJ/g;)

(c) If some of the calorimeter's water is lost, the heat capacity of the calorimeter will <u>decrease</u>.

$q = mc\,\Delta T$ with a lower m, less heat will be required to increase the temperature/ achieve a given ΔT.

Crap! (a) and (b) weren't assigned....

(59) glucose: $C_6H_{12}O_6$

$\Delta H_{combustion} = 15.57$ kJ/g

3.500 g glucose burn, $T_i = 20.94°C$, $T_f = 24.72°C$

Energy released by glucose $= (3.500g)\left(15.57\frac{kJ}{g}\right) = 54.495$ kJ

$\Delta T_{calorimeter} = 24.72°C - 20.94°C = 3.78°C$

$q_{calorimeter} = C_{calorimeter} \Delta T$

↑ heat capacity of calorimeter

$C_{calorimeter} = \dfrac{q}{\Delta T} = \dfrac{54.495 \, kJ}{3.78°C} = 14.417 \longrightarrow \boxed{14.4 \dfrac{kJ}{°C}}$

(b) if the glucose sample were twice as large (massive?) then the glucose would have released twice as much energy, so the ΔT would have been twice as large.

$\Delta T_{cal} = \dfrac{q}{C_{cal}}$

(60c) see previous page

(64) Given: ① $2C + O_2 \rightarrow 2CO \qquad \Delta H = -221.0 \, kJ$

② $2C + O_2 + 4H_2 \rightarrow 2CH_3OH \qquad \Delta H = -402.4 \, kJ$
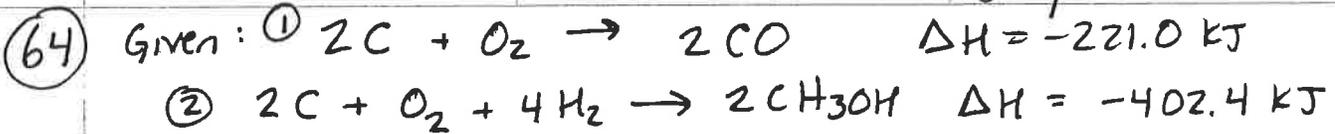
Find ΔH for this rxn: $CO + 2H_2 \rightarrow CH_3OH$

First reverse rxn ① and multiply by $\frac{1}{2}$ to get 1 CO on left:

$1 CO \rightarrow 1C + \frac{1}{2}O_2 \qquad \Delta H = -\frac{1}{2}(-221.0 \, kJ) = +110.5 \, kJ$

Next multiply rxn ② by $\frac{1}{2}$ to get 1 CH_3OH on right

$1 C + \frac{1}{2}O_2 + 2H_2 \rightarrow 1 CH_3OH \qquad \Delta H = \frac{1}{2}(-402.4 \, kJ) = -201.2 \, kJ$

add these 2 rxns

$1CO + 1\cancel{C} + \frac{1}{2}\cancel{O_2} + 2H_2 \rightarrow \cancel{1C} + \cancel{\frac{1}{2}O_2} + 1CH_3OH$

$\boxed{CO + 2H_2 \rightarrow CH_3OH} \qquad \Delta H = 110.5 + -201.2 = \boxed{-90.7 \, kJ}$

answer.

(66) given: ① $N_2 + O_2 \rightarrow 2NO \qquad \Delta H_1 = 180.7 \, kJ$

② $2NO + O_2 \rightarrow 2NO_2 \qquad \Delta H_2 = -113.1 \, kJ$

③ $2N_2O \rightarrow 2N_2 + O_2 \qquad \Delta H_3 = -163.2 \, kJ$

Find ΔH for: $N_2O + NO_2 \rightarrow 3NO$

use rxn ③ to get N_2O on left: multiply by $\frac{1}{2}$

$\rightarrow \quad N_2O \rightarrow N_2 + \frac{1}{2}O_2 \qquad \Delta H = (-163.2)(\frac{1}{2}) = -81.6 \, kJ$

multiply rxn ② by $(-\frac{1}{2})$ to get 1 NO_2 on left

$\rightarrow \quad NO_2 \rightarrow NO + \frac{1}{2}O_2 \qquad \Delta H = -113.1(-\frac{1}{2}) = +56.55 \, kJ$

use rxn ① as is to cancel N_2 and O_2 and get 2 more NO

$N_2 + O_2 \rightarrow 2NO \qquad \Delta H = (180.7)(1) = 180.7 \, kJ$

total

$N_2O + NO_2 \rightarrow 3NO \qquad \Delta H = -81.6 + 56.55 + 180.7$

$= 155.65 \, kJ$

$\boxed{155.7 \, kJ}$

(71) Thermite! $2 Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + 2 Fe_{(s)}$

$\Delta H_{formation}^{\circ}$: 0 -822.16 $-1669.8 \ kJ$ 0

from App. C
p. 1059-1061 $Al_{(s)}$ and $Fe_{(s)}$ are elements in std state, so zero ΔH_f°

$\Delta H_{rxn} = \Sigma \Delta H_{prod} - \Sigma \Delta H_{f \ react}$

$= \Delta H_f (Al_2O_3) + 2 \Delta H_f (Fe) - 2 \Delta H_f (Al) - 1 \Delta H_f (Fe_2O_3)$

$= -1669.8 \ kJ + 2(0) - 2(0) - (-822.16 \ kJ)$

$= -847.64 \ kJ \rightarrow \boxed{-847.6 \ kJ}$

(74) Use Appendix C to Calculate ΔH_{rxn}

(a) $CaO_{(s)} + 2HCl_{(g)} \rightarrow CaCl_{2(s)} + H_2O_{(g)}$
     -635.5 -92.30 -795.8 -241.82 $\leftarrow \Delta H_f^{\circ}$ in kJ/mole
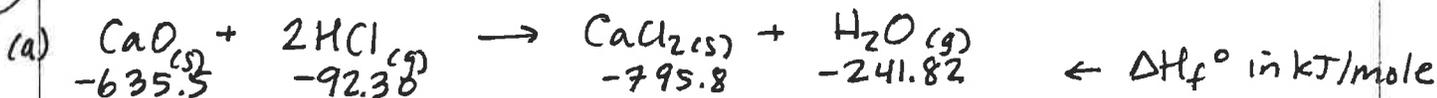
$\Delta H = (-795.8) + (-241.82) - (-635.5) - 2(-92.30) = -217.52$

$\boxed{-217.5 \ kJ}$
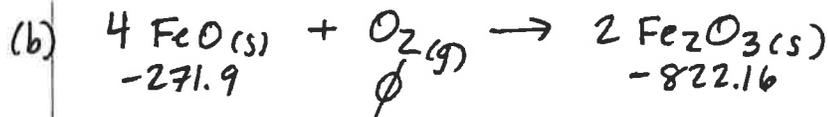
(b) $4 FeO_{(s)} + O_{2(g)} \rightarrow 2 Fe_2O_{3(s)}$ units could
     -271.9 0 -822.16 be kJ/mole.

$\Delta H = 2(-822.16) - 4(-271.9) - 1(0) = -556.72 \rightarrow \boxed{-556.7 \ kJ}$

(c) $2 CuO_{(s)} + NO_{(g)} \rightarrow Cu_2O_{(s)} + NO_{2(g)}$
     -156.1 90.37 -170.7 33.84
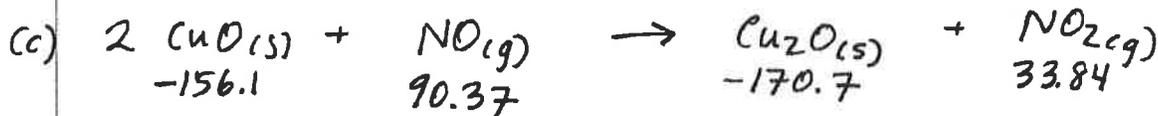
$\Delta H = 1(-170.7) + 1(33.84) - 2(-156.1) - 1(90.37) = 84.97 \rightarrow \boxed{85.0 \ kJ}$

(d) $4 NH_{3(g)} + O_{2(g)} \rightarrow 2 N_2H_{4(g)} + 2 H_2O_{(l)}$
     -46.19 0 95.40 -285.83
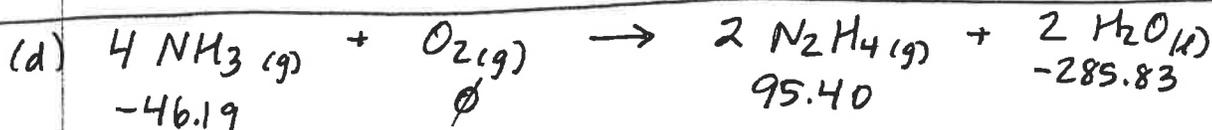
$\Delta H = 2(95.40) + 2(-285.83) - 4(-46.19) - (0) = \boxed{-196.10 \ kJ}$

(75) $C_3H_6O_{(l)} + 4 O_{2 (g)} \longrightarrow 3 CO_{2 (g)} + 3 H_2O_{(l)}$ $\Delta H = -1790 \, kJ$

ϕ $-393.5 \, kJ$ $-285.83 \, kJ$

acetone

("propanone")

:O:

Heats of formation! looked up CO_2 and H_2O in appendix C

Q: calculate $\Delta H_{formation}$ of C_3H_6O

$\Delta H_{rxn} = 3 \Delta H_f (CO_2) + 3 \Delta H_f (H_2O) - 4 \Delta H_f (O_2) - 1 \Delta H_f (C_3H_6O)$

$-1790 \, kJ = 3(-393.5 \, kJ) + 3(-285.83 kJ) - 4(0) - X$

$X = -248.05 \, kJ \longrightarrow \boxed{-248 \, kJ = \Delta H_{form} \; C_3H_6O}$

"or kJ/mole"

(77) 2,3,4 trimethyl pentane (C_8H_{18})

C-C-C-C-C
  | | |
  C   C   C

Combustion of 1 mole yields 5064.9 kJ if $H_2O_{(g)}$ and $CO_{2 (g)}$ produced

(a) Balanced eqn for comb. of 1 mole C_8H_{18}

$C_8H_{18 (l)} + \frac{25}{2} O_{2 (g)} \longrightarrow 8 CO_{2 (g)} + 9 H_2O_{(g)}$

↖ have to do a fraction since they say to write rxn for 1 mole C_8H_{18}.

(b) Balanced eqn for formation of 1 mole C_8H_{18} from its elements:

$8 C_{(s)} + 9 H_{2 (g)} \longrightarrow C_8H_{18 (l)}$ $\Delta H = \Delta H_f (C_8H_{18})$

(c) Use Data from Table 5.3 (page 184, same as Appendix C...) to calculate ΔH_f of C_8H_{18} (trimethyl pentane)

$\Delta H_{rxn} = -5064.9 \, kJ = 8 \Delta H_f (CO_2) + 9 \Delta H_f (H_2O) - \Delta H_f (C_8H_{18}) - \frac{25}{2} \Delta H_f (O_2)$

$-5064.9 \, kJ = 8(-393.5) + 9(-241.8) - X - \frac{25}{2}(0)$

$\boxed{X = -259.3 \quad kJ/mole}$

heat of formation of 2,3,4 trimethylpentane

(on p.184 it is -241.8
  on p.1060 it is -241.82
  if you use -241.82 you
  get -259.5 kJ)

(83) (a) Chicken noodle soup:

2.5 g fat
14 g carbohydrate
7 g protein

Acc to page 188,

1 g fat contains 9 Cal energy
1 g carbohydrate has 4 Cal
1 g protein has 4 cal
(when metabolized)

$(2.5 g \; fat)\left(\dfrac{9 \; Cal}{g}\right) + (14 g \; carb.)\left(\dfrac{4 \; cal}{g}\right) + (7 g \; protein)\left(\dfrac{4 \; cal}{g}\right)$

$= 106.5 \approx \boxed{107 \; Calories}$ (w/ extra SF)

(89) Earth population: 6.8 billion (end of 2009)

What mass glucose $(C_6H_{12}O_6)$ needed to provide
$\underline{\dfrac{1500 \; cal \, / person}{Day}}$ for a year?

$C_6H_{12}O_{6 \, (s)} + 6 \, O_{2 (g)} \longrightarrow 6 \, CO_{2 (g)} + 6 \, H_2O_{(\ell)} + 2803 \, kJ$

$\left(\dfrac{1500 \; Cal}{person \cdot Day}\right)\left(365 \; Days\right)\left(6.8 \times 10^9 \; people\right)\left(\dfrac{1 \; Kcal}{1 \; Cal}\right)\left(\dfrac{4.184 \, kJ}{1 \; kcal}\right)\left(\dfrac{1 \; mole \; C_6H_{12}O_6}{2803 \; kJ}\right)$

food calorie

$\left(\dfrac{180.1572 g}{mole}\right)$

$= 1.0 \times 10^{15}$ grams $C_6H_{12}O_6$

or $\boxed{1.0 \times 10^{12} \; Kg \; C_6H_{12}O_6}$

(97) $Ca^{+2}_{(aq)} + 2 \, HCO_3^{-1}_{(aq)} \longrightarrow CaCO_{3 (s)} + CO_{2 (g)} + H_2O_{(\ell)}$
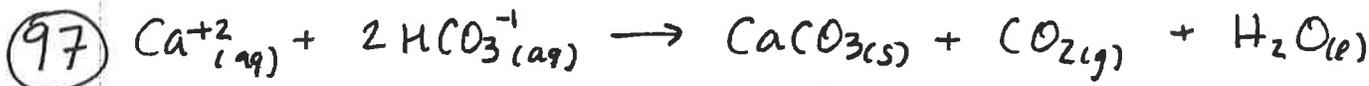
limestone!
stalactite and stalgmite formation

when
1 mole $CaCO_3$ forms,
Rxn performs 2.47 kJ of P-V work as the $CO_{2 (g)}$ pushes back the atmosphere.

And 38.95 kJ heat are absorbed from environment.

Work was done by the CO_2: it did 2.47 kJ work on surroundin
So work done on "rxn" is negative: $W = -2.47 \, kJ$
Rxn absorbs 38.95 kJ so $q = \boxed{+38.95 \, kJ} = \Delta H_{rxn}$

$\Delta U \; (\Delta E) = q + W = 38.95 \, kJ + (-2.47 \, kJ) = \boxed{36.48 \, kJ}$

$\Delta E \; or \; \Delta U$

(#99) Passive solar with bricks!

given. $\begin{cases} \text{1 brick} = 1.8 \text{ kg} \\ \text{specific heat of brick} = 0.85 \dfrac{J}{g°C} \end{cases}$ $\left(\text{or } \dfrac{0.85 \ J}{g \cdot K} \right)$

Q: How many bricks are needed, to have the same heat capacity as 1.7×10^3 gallons H_2O?

$$C = m \cdot c$$

↗ heat capacity $\dfrac{cal}{°C}$ or $\dfrac{J}{°C}$

↑ mass (g)

↖ specific heat $\dfrac{cal}{g°C}$ or $\dfrac{J}{g°C}$

Heat capacity of 1 brick = $m \cdot c = (1800 g)\left(.85 \dfrac{J}{g°C}\right) = \underline{\underline{1530 \dfrac{J}{°C}}}$

Heat capacity of the 1700 gallons $H_2O = mc$

$$m = (1700 \text{ gallons})\left(\dfrac{3.7854 L}{gal}\right)\left(\dfrac{1000 mL}{L}\right)\left(\dfrac{1 g}{mL}\right) = 6\ 435\ 180 \ g$$

↑ density of H_2O

$$mc = (6435180 g)\left(\dfrac{4.184 J}{g°C}\right) = 2.6925 \times 10^7 \ J/{°c}$$

$$2.6925 \times 10^7 \dfrac{J}{°C} = x \ (1530 \ J/{°c})$$

$$x = 17597.9 \rightarrow \boxed{18000 \ \text{bricks!}}$$

(111) Estimate:

5.5×10^{16} g CO_2 /yr   "fixed" by photosynthesis.
(assume all this CO_2 is converted to glucose).

(a) Calculate <u>energy</u> stored by photosynthesis (per yr)

$$6 CO_2 + 6 H_2O \longrightarrow C_6H_{12}O_6 + 6O_2 \quad \Delta H = +2803 kJ$$

opposite of ΔH from problem 89, since photosynthesis is the opposite of cellular respiration of glucose!
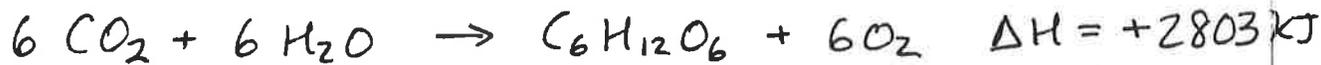
$$\left(\frac{5.5 \times 10^{16} g\, CO_2}{yr}\right)\left(\frac{1 \,mole\, CO_2}{44.0098 g}\right)\left(\frac{2803\, kJ}{6\, CO_2\, moles}\right)$$

$$= 5.838 \times 10^{17} \ kJ/yr \longrightarrow \boxed{5.8 \times 10^{17}\, kJ/yr}$$

energy stored by photosynthesis.

(b) Energy stored in MW (1 Watt = 1 Joule per second)

$$\left(\frac{5.838 \times 10^{17}\, kJ}{yr}\right)\left(\frac{1\, yr}{365\, days}\right)\left(\frac{1\, day}{86400\, seconds}\right)\left(\frac{1\, MJ}{1000\, kJ}\right)$$

$$= 1.851 \times 10^7 \longrightarrow 1.9 \times 10^7\, MJ/s$$

or $\boxed{1.9 \times 10^7\, MW}$

A large nuclear power plant produces $\approx 10^3$ MW

$$\frac{1.9 \times 10^7\, MW}{10^3\, MW/plant} = 1.9 \times 10^4 \quad or \quad \boxed{19000 \text{ nuclear power plants.}}$$

Wow!