# Thermochemistry Practice Problems
## Answers

Fossum
Chem 1A

1. Brass $8.40 \ g/cm^3$, SHC $0.385 \ J/g°C$

$$(14.5 \ cm^3)\left(\frac{8.40g}{1 \ cm^3}\right) = 121.8 \ g \ brass$$

$$q_{brass} = S_b M_b \Delta T_b \qquad\qquad q_{water} = S_w M_w \Delta T_w$$

$$q_{brass} = -q_{water}$$

$$\Delta T_b = (T_f - 152°C) \qquad \Delta T_w = T_f - 23.7°C$$

final temp is the same for both.

SO:

$$S_b m_b (T_f - T_i)_b = -S_w m_w (T_f - T_i)_w$$

distribute...

$$S_b m_b T_f - S_b m_b T_{ib} = -S_w m_w T_f + S_w m_w T_{iw}$$
$$+ S_w m_w T_f \quad + S_b m_b T_{ib} \qquad + S_w m_w T_f \quad + S_b m_b T_{ib}$$

get $T_f$ on one side by itself.

$$S_b m_b T_f + S_w m_w T_f = S_w m_w T_{iw} + S_b m_b T_{ib}$$

factor out $T_f$

$$\frac{T_f (S_b m_b + S_w m_w)}{(S_b m_b + S_w m_w)} = \frac{S_w m_w T_{iw} + S_b m_b T_{ib}}{(S_b m_b + S_w m_w)}$$

divide by that

$$T_f = \frac{S_w m_w T_{iw} + S_b m_b T_{ib}}{(S_b m_b + S_w m_w)}$$

now, plug in values.

$$T_f = \frac{(4.184 \ J/g°C)(138g)(23.7°C) + (.385 \ J/g°C)(121.8g)(152°C)}{[(.385 \ J/g°C)(121.8g) + (4.184 \ J/g°C)(138g)}$$

$$T_f = \frac{13684 \ J + 7128 \ J}{46.89 \ J/°C + 577.4 \ J/°C} = \frac{20812 \ J}{624.3 \ J/°C} = \boxed{33.3°C}$$

1. alternate method: put in numbers right away.

$q_{brass} = S_b M_b \Delta T_b = (.385 \, J/_{g \cdot c})(121.8g)(T_f - 152°C)$

$q_{brass} = 46.\overline{8}9 \, T_f \, J \, - \, 71\overline{2}7.7 \, J$

$q_w = S_w M_w \Delta t_w = (4.184 \, J/_{g \cdot c})(138g)(T_f - \$23.7°C)$

$q_w = 577.4 \, T_f \, J \, - \, 13\overline{6}84 \, J$

then $q_b = -q_w$ so $(46.89 T_f - 7127.7) = -(577.4 \, T_f - 13684)$

Subtracting terms from both sides

$46.89 \, T_f + 577.4 \, T_f = +13684 + 7127.7$

$T_f (46.89 + 577.4) = T_f (624.3) = 20812$

$T_f = \dfrac{20812}{624.3} = 33.3°C$

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2.    $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$

a. $\left(1.00 \times 10^6 \, kJ\right)\left(\dfrac{1 \, mol \, CH_4}{-890.3 \, kJ}\right)\left(\dfrac{16.04 \, g \, CH_4}{1 \, mol \, CH_4}\right)\left(\dfrac{1 \, kg}{1000g}\right) =$

$= \boxed{18.0 \, kg \, methane}$

b. find # moles!   $PV = nRT$   $n = \dfrac{PV}{RT}$

$n = \dfrac{[(748/760) \, atm](1.03 \times 10^3 \, L)}{\left(0.08206 \, \frac{L \cdot atm}{K mol}\right)(294.\overline{9}5 \, K)} = 41.\overline{8}8 \, mol \, CH_4$

$\left(41.\overline{8}8 \, mol \, CH_4\right)\left(\dfrac{-890.3 \, kJ}{1 \, mol \, CH_4}\right) = \boxed{-3.73 \times 10^4 \, kJ}$

2c. $q = Sm\Delta T$     $q_{rxn} = -q_{water}$

$$m = \frac{q}{S\Delta T} = \frac{(+3.73 \times 10^4 \, KJ)(\frac{1000J}{1KJ})}{(4.184 \, J/g°C)(60.8 - 227.7°C)} = 2.3\overline{3}9 \times 10^5 \, g \, H_2O$$

$$\left(2.3\overline{3}9 \times 10^5 \, g \, H_2O\right)\left(\frac{1 \, mL}{1.00 \, g}\right)\left(\frac{1 \, L}{1000 \, mL}\right) = \boxed{234 \, L} \, H_2O$$

3. $$\left(\frac{-5.48 \times 10^3 \, KJ}{1 \, mol \, octane}\right)\left(\frac{1 \, mol \, C_8H_{18}}{114.22 \, g \, C_8H_{18}}\right)\left(\frac{.703 \, g \, C_8H_{18}}{1 \, mL \, C_8H_{18}}\right) \times$$

$$\times \left(\frac{1000 \, mL}{1 \, L}\right)\left(\frac{3.785 \, L}{1 \, gal}\right) = \boxed{-1.28 \times 10^5 \quad \frac{KJ}{gal \, octane}}$$

4. $\Delta H_{sol} \, NaOH: \dfrac{-42 \, KJ}{1 \, mol \, NaOH}$

$$\left(.500 \, L \, NaOH\right)\left(\frac{7.0 \, mol \, NaOH}{1 \, L}\right) = 3.5 \, mol \, NaOH$$

$$\left(3.5 \, mol \, NaOH\right)\left(\frac{-42 \, KJ}{1 \, mol \, NaOH}\right) = -147 \, KJ \text{ released.}$$

transfer this to water —
↳ +147 KJ (the solution, actually)

$q = Sm\Delta T$     $\Delta T = \dfrac{q}{Sm} = \dfrac{147 \times 10^3 \, J}{(4.184 \, J/g°C)(500 \, g)} =$

$\Delta T = 70.3 \, °C$ (this is only approximate —
$d_{solution} \neq d_{H_2O}$, and $S_{solution} \neq S_{water}$)

because 500 mL ≈ 500 g for the water solutions

So   $21° + 70.3 \, °C \Rightarrow \boxed{\sim 91 \, °C}$

5. $\left(1.078g\ C_8H_6O_4\right)\left(\dfrac{1\ mol}{166.13g}\right)\left(\dfrac{-3.224\times10^3\ kJ}{mol\ C_8H_6O_4}\right)$

$= -20.9\overline{2}04\ kJ$ given off using this mass of
                                                    o-phthalic acid.
calorimeter absorbs $+20.9\overline{2}04\ kJ$

$\Delta T_{cal} : T_f - T_i \overset{\oplus}{=} 5.80\ °C$

heat capacity: $\dfrac{\#kJ}{\Delta T} = \dfrac{20.9\overline{2}04\ kJ}{5.80\ °C} = \boxed{3.61\ kJ/°C}$

6.  $\Delta H_f°\ O_{2\ (g)}:\quad 0.0\ kJ/mol$
    $\Delta H_f°\ CO_{2\ (g)}:\quad -393.5\ kJ/mol$
    $\Delta H_f°\ H_2O_{(\ell)}:\quad -285.84\ kJ/mol$

$\Delta H_{rxn} = prod - react\quad = -8326\ kJ$

$-8326\ kJ = \left[12(-393.5) + 14(-285.84)\right] - \left[2(\Delta H_f°\ C_6H_{14}) + 19(0)\right]$

$-8326 = (-4722 - 4001.8) - (2(\Delta H_f\ C_6H_{14}))$

$-8326 = -8723.8 - 2(\Delta H_f\ C_6H_{14})$

$+8723.8\qquad\quad +8723.8$
_____

$\dfrac{39\overline{7}.8}{-2} = \dfrac{-2\ \Delta H_f\ C_6H_{14}}{-2}\qquad \boxed{\Delta H_f°(C_6H_{14}) = -199\ kJ/mol}$

Get a balanced equation:

7. $2CH_3CHO_{(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 4H_2O_{(\ell)}$

find $\Delta H_{rxn}$: $\sum \Delta H_f^\circ$ prod $- \sum \Delta H_f^\circ$ react.

$\Delta H_{rxn}^\circ = [4(-393.5) + 4(-285.84)] - [2(-166) + 5(0)]$ kJ

$\Delta H_{rxn}^\circ = (-2717.36) - (-332) = -2385.36$ kJ/mol

find moles of each: this is a limiting reactant problem.

$(5.93 \text{ g } CH_3CHO)\left(\dfrac{1 \text{ mol}}{44.05g}\right) = 0.1346 \text{ mol } CH_3CHO$

$n_{O_2} = \dfrac{PV}{RT} = \dfrac{(1.00atm)(9.29L)}{\left(0.08206 \frac{L \cdot atm}{mol \cdot K}\right)(298K)} = 0.3799 \text{ mol } O_2$

\* See below

$(0.3799 \text{ mol } O_2)\left(\dfrac{2 \text{ mol } CH_3CHO}{5 \text{ mol } O_2}\right) = 0.1520 \text{ mol } CH_3CHO$ needed to react with all of the $O_2$ – we

only have $0.1346$ mol $CH_3CHO$, so its <u>limiting</u>.

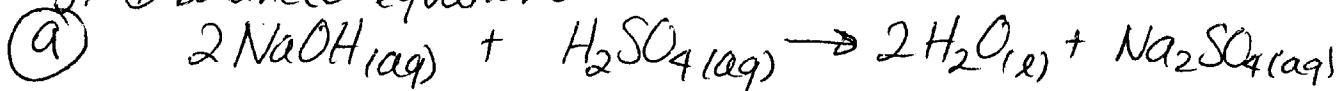$(0.1346 \text{ mol } CH_3CHO)\left(\dfrac{-2385.36 \text{ kJ}}{2 \text{ mol } CH_3CHO}\right) = \boxed{-161 \text{ kJ produced}}$

↑ from balanced equation.

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\* finding limiting reactant: alternate method

need: $\dfrac{5 \text{ mol } O_2}{2 \text{ mol } CH_3CHO} = \dfrac{2.5 \text{ mol } O_2}{1 \text{ mol } CH_3CHO}$   have: $\dfrac{0.3799 \text{ mol } O_2}{0.1346 \text{ mol } CH_3CHO} = \dfrac{2.82 \text{ } O_2}{1 \text{ } CH_3CHO}$

have too much $O_2$ – so $CH_3CHO$ is limiting.

8. (1) balanced equation
(a)     $2 NaOH_{(aq)} + H_2SO_{4(aq)} \rightarrow 2 H_2O_{(\ell)} + Na_2SO_{4(aq)}$

(2) find $\Delta H^0_{rxn}$. $= \left( \Delta H^0_f \text{ prod} - \Delta H^0_f \text{ react.} \right)$

$\Delta H^0_f$ $NaOH_{(aq)} = -470.114$ kJ         $H_2O_{(\ell)} = -285.83$ kJ
$H_2SO_{4(aq)} = -909.27$ kJ         $Na_2SO_4 {(aq)} = -1390$ kJ

$\Delta H^0_{rxn} = \left[ 2(-285.83) + (-1390) \right] - \left[ 2(-470.114) + (-909.27) \right]$ kJ

$\Delta H^0_{rxn} = (-1961.66) - (-1849.498)$ kJ $= \boxed{-112 \text{ kJ}}$

(3) Given specific amounts of 2 reactants:
    Limiting reactant problem! Find moles of each.

$(.300 L) \left( \dfrac{2.00 \text{ mol } NaOH}{1 L} \right) = 0.600$ mol $NaOH$

$(0.400 L) \left( \dfrac{1.00 \text{ mol } H_2SO_4}{1 L} \right) = 0.400$ mol $H_2SO_4$

look @ mole ratios...
need: $\dfrac{2 \text{ mol } NaOH}{1 \text{ mol } H_2SO_4}$        have: $\dfrac{0.600 \text{ mol } NaOH}{0.400 \text{ mol } H_2SO_4} = \dfrac{1.5 \text{ mol } NaOH}{1 \text{ mol } H_2SO_4}$

not enough $NaOH$ --- so $NaOH$ is limiting.
Use # moles of the limiting reactant we have
to finish problem.

$(0.600 \text{ mol } NaOH) \left( \dfrac{-112 \text{ kJ}}{2 \text{ mol } NaOH} \right) = \boxed{-33.6 \text{ kJ}}$
                                    ↗                                              (given off)
            from balanced equation

8b) Resulting solution: 300. mL NaOH + 400. mL $H_2SO_4$
(700. mL total) $\left(\dfrac{1.05g}{1 mL}\right)$ = 735 g solution

$q = Cm\Delta T$  $\qquad$ $q = +33.6 kJ$ or $+33600 J$
$\qquad\qquad\qquad\qquad$ (heat is <u>absorbed</u> by the water)

$\Delta T = \dfrac{q}{Cm} = \dfrac{33600 J}{(4.2 \frac{J}{g°C})(735g)} = +10.88 °C$

$\Delta T = T_f - T_i$  $\qquad$ So $T_f = \Delta T + T_i$
$\qquad\qquad\qquad\qquad$ $T_f = 11°C + 25°C$ ⟹ $\boxed{36°C}$

8c) To give off 100. KJ of energy:

$(-100. KJ)\left(\dfrac{2 \text{ mol NaOH}}{-112 KJ}\right) = \boxed{1.79 \text{ mol NaOH needed}}$

$(-100. KJ)\left(\dfrac{1 \text{ mol } H_2SO_4}{-112 KJ}\right) = \boxed{0.893 \text{ mol } H_2SO_4 \text{ needed}}$

9.  Convert 250. g ice @ -25°C to steam @ 250.°C.
ice -25°C → ice 0°C → water 0°C ⟶ water 100°C → steam 100°C
5 processes.  $\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad$ ↓
$\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad$ Steam 250°C

① ice warming  $q = Cm\Delta T$
$\quad q = \left(\dfrac{2.1 J}{g°C}\right)(250.g)(0-(-25°C)) = 13125 J$

② ice melting
$\quad (250.g)\left(\dfrac{333 J}{g}\right) = 83250 J$

③ water warming  $q = Sm\Delta T$
$\quad q = (4.2 \frac{J}{g°C})(250.g)(100°C) = 105000 J$
$\qquad\qquad\qquad\qquad\qquad\qquad$ ↑
$\qquad\qquad\qquad\qquad\qquad (100-0°C)$

9
continued  ④ water vaporizing

$$(250. g)\left(\frac{2260 J}{g}\right) = 56\bar{5}000 J$$

⑤ steam warming   $q = Cm\Delta T$

$$q = (2.0 \, J/g°C)(250. g)(250°C - 100°C) = 7\bar{5}000 J$$

Now — add!

| | |
|---|---|
| $13\underline{1}25 J$ | ± 1000 |
| $83\underline{2}50$ | ± 100 |
| $1\bar{0}5000$ | ± 10000 ← largest uncert. |
| $56\bar{5}000$ | ± 1000 |
| $75000$ | ± 1000 |
| $84\underline{1}375 J$ | ± 10000 |

$$8.4 \times 10^5 \, J$$
$$\text{or } 8.4 \times 10^2 \, KJ$$

10.   136 KJ  →what mass of water can be heated from 25°C to 100°C and vaporized @ 100°C ?

Let mass of water = $x$ g
136 KJ is total energy.    $q_1 + q_2 = 136 \, kJ$

①E to heat      $q = Cm\Delta T = q_1$

$$q_1 = \left(\frac{4.2 \, J}{g°C}\right)(x \, g)(75°C) = 3\bar{1}5 \, x \, J = q_1$$

②E to vaporize $= q_2$

$$q_2 = (x \, g)\left(\frac{2260 J}{g}\right) = 22\bar{6}0 \, x \, J = q_2$$

$$q_1 + q_2 = 136000 J = 3\bar{1}5x \, J + 22\bar{6}0x \, J$$
$$\frac{13600 \, J}{2575} = \frac{2575 \, x \quad J}{2575}$$

$$x = 52.8 \, g \quad \text{water}$$

11. $q_{rxn} = -q_{cal} = -(q_w + q_{bomb})$     $\Delta T = T_f - T_i$
     $\Delta T = +9.19°C$

$q_w = C_w m_w \Delta T_w = (4.184 \, J/_{g°C})(775 \, g)(9.19°C) = 29799.5 \, J$

$q_b = C\Delta T = (893 \, J/_{°C})(9.19°C) = 8206.67 \, J$

$q_{cal} = 38\overline{0}06.17 \, J = 38.\overline{0}06 \, kJ$

$q_{rxn} = -38.\overline{0}06 \, kJ$     per 1.500 g $C_7H_6O_2$

$\left(1.500g \; C_7H_6O_2\right)\left(\dfrac{1 \, mol}{122.118g}\right) = 0.0122\overline{8}32 \, mol \; C_7H_6O_2$

$\dfrac{-38.\overline{0}06 \, kJ}{0.0122\overline{8}32 \, mol} = -3094 \, kJ/mol = \boxed{-3.09 \times 10^3 \, kJ/_{mol \; C_7H_6O_2}}$

12.     $2 C_7H_6O_{2\,(s)} + 15 O_{2\,(g)} \rightarrow 14 CO_{2\,(g)} + 6 H_2O_{(\ell)}$

$-3.09 \times 10^3 \, kJ/_{mol \; C_7H_6O_2}$     but above equation shows
                                                $2 C_7H_6O_2$.

How much energy per 2 mol $C_7H_6O_2$?
(Twice as much)
$\left(2 \, mol \; C_7H_6O_2\right)\left(\dfrac{-3.09 \times 10^3 \, kJ}{1 \, mol \; C_7H_6O_2}\right) = \boxed{\dfrac{-6.18 \times 10^3 \, kJ}{2 \, mol \; C_7H_6O_2}}$

Since the rxn was done in a bomb calorimeter,
the heat given off $= \Delta E$     (w=0 because volume
is constant).
So $\Delta E$ for the rxn as written $= -6.18 \times 10^3 \, kJ = \Delta E$

$$\Delta E = \Delta H + w \qquad w = -P\Delta V = -\Delta n_{gas} RT$$

$$\Delta n_{gas} = n_{prod} - n_{react} = 14 - 15 = -1$$

(count only gas molecules) $\qquad T = 31.69 + 273.15 K$

$$-\Delta n RT = -(-1 \text{ mole})(8.314 \tfrac{J}{K \cdot mol})(304.84 K) = +2534 J$$

$$w = +2.534 \text{ kJ} \qquad \text{work is } \oplus \text{ - done on the system}$$
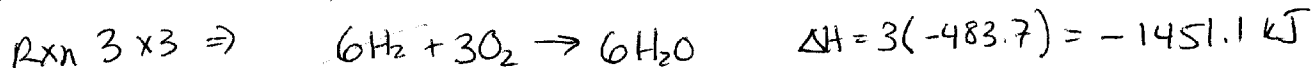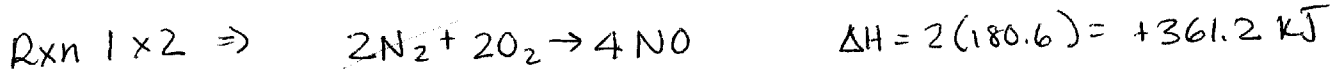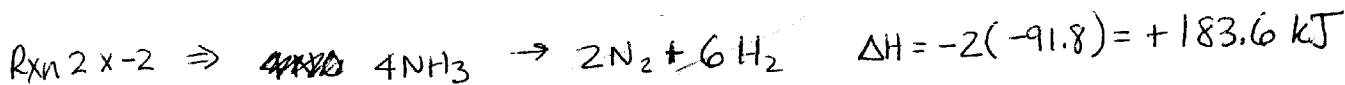
$$\Delta H = \Delta E - w = -6.18 \times 10^3 \text{ kJ} - 2.534 \text{ kJ} = -6.18 \times 10^3 \text{ kJ}$$

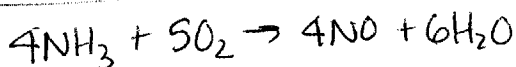work is negligible compared with the amount of heat generated!
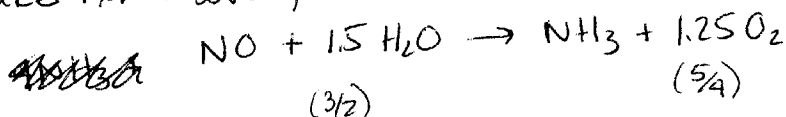
13. a.  Rxn 2 : reverse and x2
Rxn 1 :   x2
Rxn 3 :   x3

Rxn 2 x-2 $\Rightarrow$ ~~4NNO~~ $4NH_3 \rightarrow 2N_2 + 6H_2$ $\qquad \Delta H = -2(-91.8) = +183.6 \text{ kJ}$

Rxn 1 x2 $\Rightarrow$ $\qquad 2N_2 + 2O_2 \rightarrow 4NO$ $\qquad \Delta H = 2(180.6) = +361.2 \text{ kJ}$

Rxn 3 x3 $\Rightarrow$ $\qquad 6H_2 + 3O_2 \rightarrow 6H_2O$ $\qquad \Delta H = 3(-483.7) = -1451.1 \text{ kJ}$

add _____

$\qquad 4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O \qquad \boxed{\Delta H = -906.3 \text{ kJ}}$

b.  Take rxn above, reverse and $\div 4$

~~4NO+6H~~ $\qquad NO + 1.5 H_2O \rightarrow NH_3 + 1.25 O_2 \qquad \Delta H = \dfrac{-(-906.3 \text{ kJ})}{4}$

$\qquad\qquad (3/2) \qquad\qquad\qquad (5/4)$

$$\boxed{\Delta H = {}^{+}226.6 \text{ kJ}}$$