1.   (1) $Na(s) + \frac{1}{2}Cl_2(g) \rightarrow Na(g) + Cl(g)$; $\Delta H = +230$ kJ
        (2) $Na(g) + Cl(g) \rightarrow Na^+(g) + Cl^-(g)$; $\Delta H = +147$ kJ
        (3) $Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s)$; $\Delta H = -411$ kJ
    need $Na^+(g)$ and $Cl^-(g)$ on the LHS; reverse (2) $\rightarrow$ (4)
    need $NaCl(s)$ on the RHS; use (3) as is
    need $Na(g)$ and $Cl(g)$ on the LHS to cancel $Na(g)$ and $Cl(g)$ on the RHS in (4); reverse (1) $\rightarrow$ (5)
        (4) $Na^+(g) + Cl^-(g) \rightarrow Na(g) + Cl(g)$; $\Delta H = -147$ kJ
        (3) $Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s)$; $\Delta H = -411$ kJ
        (5) $Na(g) + Cl(g) \rightarrow Na(s) + \frac{1}{2}Cl_2(g)$; $\Delta H = -230$ kJ
    adding (4)+(3)+(5) gives $Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$; $\Delta H = -788$ kJ

2. ΔH_f^0 of $VCl_3(s)$ is ΔH for $V(s) + 1.5Cl_2(g) \rightarrow VCl_3(s)$
        (1) $V(s) + 2Cl_2(g) \rightarrow VCl_4(l)$; $\Delta H = -569.4$ kJ
        (2) $VCl_3(s) \rightarrow VCl_2(s) + \frac{1}{2}Cl_2(g)$; $\Delta H = +128.9$ kJ
        (3) $2VCl_3(s) \rightarrow VCl_2(s) + VCl_4(l)$; $\Delta H = +140.2$ kJ
    need $V(s)$ on the LHS; use (1) as is
    need $VCl_4(l)$ on the LHS to cancel $VCl_4(l)$ on the RHS in (1); reverse (3) $\rightarrow$ (4)
    need $VCl_2(s)$ on the RHS to cancel $VCl_2(s)$ on the LHS in (4); use (2) as is
        (1) $V(s) + 2Cl_2(g) \rightarrow VCl_4(l)$; $\Delta H = -569.4$ kJ
        (4) $VCl_2(s) + VCl_4(l) \rightarrow$ $2VCl_3(s)$; $\Delta H = -140.2$ kJ
        (2) $VCl_3(s) \rightarrow VCl_2(s) + \frac{1}{2}Cl_2(g)$; $\Delta H = +128.9$ kJ
    adding (1)+(4)+(2) gives $V(s) + 1.5Cl_2(g) \rightarrow VCl_3(s)$; $\Delta H = -580.7$ kJ

3.   (1) $2Al_2O_3(s) \rightarrow 4Al(s) + 3O_2(g)$; $\Delta H = +3340$ kJ
    (a) ΔH_f^0 for $Al_2O_3(s)$ is ΔH for
        (2) $2Al(s) + 1.5O_2(g) \rightarrow Al_2O_3(s)$
        but (2) is (1) divided by 2 and reversed
        hence ΔH_f^0 for $Al_2O_3(s)$ is -3340 kJ/2 = -1670 kJ

    (b) $10.0 \text{ g } Al_2O_3 \times \dfrac{1 \text{ mol } Al_2O_3}{102 \text{ g } Al_2O_3} \times \dfrac{3340 \text{ kJ}}{2 \text{ mol } Al_2O_3} = 164 \text{ kJ}$

4. $\Delta T = (25.00 - 21.18)°C = 3.82°C$
    Q = heat lost by 50.0 g water + heat lost by calorimeter
        $(50.0 \text{ g} \times 4.18 \text{ J/g·°C} \times 3.82°C) + (75.0 \text{ J/°C} \times 3.92°C) = 1085 \text{ J}$
    since the temperature decreased, ΔH for the process
    $KClO_3(s) \rightarrow K^+(aq) + ClO_3^-(aq)$ is   + Q/mol $KClO_3$

    $\text{Q/mol } KClO_3 = \dfrac{1085 \text{ J}}{3.17 \text{ g}} \times \dfrac{122.6 \text{ g}}{1 \text{ mol}} = 42.0 \text{ kJ/mol}$

    $\Delta H = +42.0 \text{ kJ}$

5. heat given off $= 2.48$ g Cu x $\dfrac{128.4 \text{ kJ}}{63.55 \text{ g Cu}} = 5.01$ kJ

$Q = 5.01$ x 10^3 J $= (125$ g x 4.18 J/g·°C x $\Delta T) + (75.0$ J/°C x $\Delta T)$
   5.01 x 10^3 J $= (522.5 + 75.0)$ $\Delta T = 597.5$ ΔT

$\Delta T = \dfrac{5.01 \text{ x } 10^3}{597.5} = 8.38°C$

since the reaction is exothermic, $T_{final} = (22.50 + 8.38) = 30.88°C$

6. $\Delta T = (25.54 - 24.00)°C = 1.54°C$. Total volume $= (155 + 260)$ mL $= 415$ mL
   Assuming density $= 1.00$ g/mL, total mass $= 415$ g
   $Q = (415$ g x 4.184 J/g·°C x $1.54°C) + (40.4$ J/°C x $1.54°C)$

   $= 2736$ J $= 2.736$ kJ
   mol $H_3PO_4 = 0.155$ L x 0.1250 M $= 0.01938$
   mol $Sr(OH)_2 = 0.260$ L x 0.1000 M $= 0.02600$

   0.01938 mol H_3PO_4 needs 0.01938 mol H_3PO_4 x $\dfrac{3 \text{ mol } Sr(OH)_2}{2 \text{ mol } H_3PO_4}$

   $= 0.02907$ mol $Sr(OH)_2$
   Therefore, $Sr(OH)_2$ is the limiting reactant

   $\dfrac{Q}{\text{mol } Sr(OH)_2} = \dfrac{2.736 \text{ kJ}}{0.0260} = 105.2$ kJ

   $\Delta H = -Q$ for 3 moles of $Sr(OH)_2$; $\Delta H = -(105.2$ x $3) = -316$ kJ

7. $\Delta T = (25.64 - 23.40)°C = 2.24°C$
   heat gained on 6.00 kg H_2O increasing in temperature by $2.24°C$
   $= 6.00$ x 10^3 g x 4.184 J/g·°C x $2.24°C = 5.623$ x 10^4 J $= 56.23$ kJ
   heat gained by calorimeter increasing in temperature by $2.24°C$
   $= 3.180$ kJ/°C x $2.24°C = 7.12$ kJ
   total heat gained $= 63.35$ kJ $=$ heat given off in the reaction

   Q/mol $C_{12}H_{22}O_{11} = \dfrac{63.35 \text{ kJ}}{3.85 \text{ g}}$ x $\dfrac{342.2 \text{ g}}{1 \text{ mol}} = 5.631$ x 10^3 kJ

   since the reaction gave off heat, $\Delta H = -5.631$ x 10^3 kJ and the reaction is
       $C_{12}H_{22}O_{11} + 12O_2 \rightarrow 12CO_2 + 11H_2O$; $\Delta H = -5631$ kJ
   $\Delta H = \Sigma \Delta H_f^0{}_{products} - \Sigma \Delta H_f^0{}_{reactants}$
   $-5631 = [12(-393.5) + 11(-285.8)] - \Delta H_f^0{}_{sucrose}$
   $\Delta H_f^0{}_{sucrose} = -4722 + (-3144) + 5631 = -2235$ kJ

8.   let ΔT = change in temperature
heat gained by 5.00 kg water increasing in temperature by ΔT
$= 5.00 \times 10^3$ g x 4.18 J/g·°C x $\Delta T = 20.9\ \Delta T \times 10^3$ J $= 20.9\ \Delta T$ kJ
heat gained by calorimeter increasing in temperature by ΔT
$= 3.612$ kJ/°C x $\Delta T = 3.612\ \Delta T$ kJ
since heat given off $= 69.3$ kJ, $(20.9 + 3.612)\Delta T = 69.3$
$24.51\ \Delta T = 69.3$ $\Delta T = 69.3/24.51 = 2.83$°C
Final temperature $= (23.00 + 2.83)$°C $= 25.83$°C

9. ΔH_f HI(g) $= 25.9$ kJ/mole
        (1) $\frac{1}{2}H_2(g) + \frac{1}{2}I_2(s) \rightarrow$ HI(g); $\Delta H = 25.9$ kJ
        ΔH of sublimation of $I_2 = 62.4$ kJ/mole
        (2) $I_2(s) \rightarrow I_2(g)$; $\Delta H = 62.4$ kJ
        (3) $I_2(g) \rightarrow I_2(s)$; $\Delta H = -62.4$ kJ
        (4) $\frac{1}{2}I_2(g) \rightarrow \frac{1}{2}I_2(s)$; $\Delta H = -31.2$ kJ
        (1) + (4) gives:
        $\frac{1}{2}H_2(g) + \frac{1}{2}I_2(g) \rightarrow$ HI(g); $\Delta H = (25.9 - 31.2)$ kJ $= -5.3$ kJ

10. Let final temperature $= T$

Heat gained on melting $= 125$ g x $\dfrac{1\ \text{mol}}{60.0\ \text{g}}$ x $\dfrac{11.7\ \text{kJ}}{1\ \text{mol}}$ $= 24.38$ kJ

Heat gained on warming from 16.7°C to T:

$= 125$ g x $\dfrac{1\ \text{mol}}{60.0\ \text{g}}$ x $\dfrac{0.125\ \text{kJ}}{\text{°C·mol}}$ x (T - 16.7)°C

$= (0.2604\ \text{T} - 4.349)$ kJ

Heat lost on cooling from 33.8°C to T:

$= 755$ g x $\dfrac{1\ \text{mol}}{60.0\ \text{g}}$ x $\dfrac{0.125\ \text{kJ}}{\text{°C·mol}}$ x (33.8 - T)°C

$= (53.16 - 1.573\ \text{T})$ kJ

Heat gained = heat lost
       $0.2604\ \text{T} - 4.349 + 24.38 = 53.16 - 1.573\ \text{T}$
            $(0.2604\ 1.573)\text{T} = 53.16 + 4.349 - 24.38$
                   $1.833\ \text{T} = 33.129$
                        $\text{T} = 18.1$°C

CHECK:

$24.38 + [\ \dfrac{125}{60}$ x 0.125 x (18.1-16.7)] $= \dfrac{775}{60}$ x 0.125 x (33.8-18.1)

$24.7 = 24.7$

11. (a) From Appendix C: ΔH_f^0 Al_2O_3 = -1676 kJ; ΔH_f^0 Fe_2O_3 = -825.5 kJ
for the reaction, $\Delta H = \Delta H_f^0$ Al_2O_3 - ΔH_f^0 Fe_2O_3 = [-1676 - (-825.5)] kJ
= - 850.5 kJ

(b) For the reaction as written, 1 mol Al_2O_3 (102 g) and 2 mol Fe (111.7 g) are formed when
850.5 kJ of heat are evolved. Since these are the only materials present, Q (the heat evolved,
850.5 kJ) must be equal to the sum of (mass x specific heat x ΔT) for the Al_2O_3 and Fe.
850.5 x 10^3 J = 102 g x 0.79 J/g·°C x ΔT + 111.7 g x 0.45 J/g·°C x ΔT
850.5 x 10^3 = (80.58 + 50.27) ΔT = 130.9 ΔT
ΔT = 850.5 x 10^3/130.9 = 6.5 x 10^{3}°C
Final temperature = (25 + 6.5 x 10^3)°C = 6.5 x 10^{3}°C
(only approximate; see below)

(c) The final temperature is well above the melting point of iron and hence we would expect the
iron to melt. However, we should calculate if, after reaching 1535°C (the mp of Fe), there is
enough heat (270 J/g) to melt the Fe.
let T_{final} = 1535°C   then ΔT = (1535-25)°C = 1510°C
Q = (102 x 0.79 x 1510) + (111.7 x 0.45 x 1510) = 200 x 10^3 J
hence there are (850.5 - 200) x 10^3 J = 650.5 x 10^3 J still available to melt the Fe.
heat to melt 111.7 g Fe = 111.7 g x 270 J/g = 30 x 10^3 J
Therefore there is enough heat to melt the iron and (650.5 - 30) x 10^3 J (620.5 x 10^3) J are
left.

We can also calculate if the Al_2O_3 would be melted.
mp of Al_2O_3 = 2072°C; heat of fusion of Al_2O_3 = 1071 J/g
Assuming that molten Fe has the same specific heat as solid Fe, we can calculate the
amount of additional heat needed to reach the melting point of Al_2O_3.
ΔT = 2072 - 1535 = 537°C
Q = [(102 x 0.79 x 537) + (111.7 x 0.45 x 537)] J
  = 70 x 10^3 J
Hence, after reaching 2072°C, (620 - 70) x 10^3 = 550 x10^3 J are left.
Heat to melt the 102 g of Al_2O_3 = 102 g x 1071 J/g = 109 x 10^3 J
Heat left = (550 - 109) x 10^3 J = 441 x 10^3 J
Again, assuming molten Al_2O_3 has the same specific heat as the liquid, we can calculate the
approximate final temperature.
441 x 10^3 J = (102 x 0.79 x ΔT) + (111.7 x 0.45 x ΔT)
ΔT = 441000/130.9 = 3370°C
Final temperature = 3370 + 2072 = 5442°C (5.4 x 10^{3}°C)

**When this reaction used to be used for welding rails, it would have been done in the
open and much of the heat would have been lost and maybe the Al_2O_3 would not have
melted.**