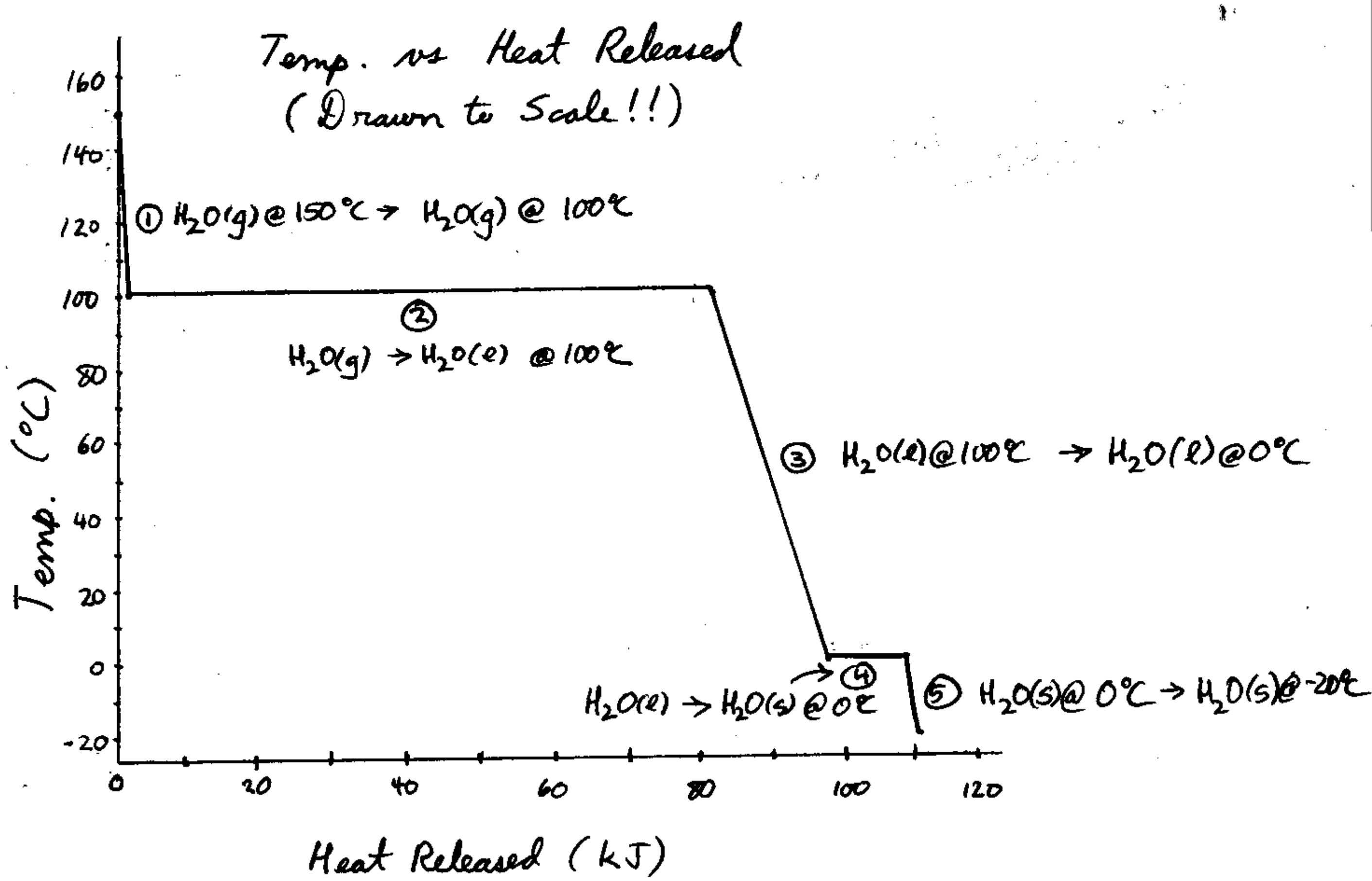
1.



① $Q = mc\Delta T$

$Q = 35g \times 2.01\ J/g°C \times -50°C$

$Q = -3517.5\ J$

$Q = -3.5\ kJ$

② $Q = -L_v m$

$Q = -40.8\ kJ/mol \times \frac{1 mol}{18.02 g} \times 35g$

$Q = -79.2\ kJ$

③ $Q = mc\Delta T$

$Q = 35g \times 4.184\ J/g°C \times -100°C$

$Q = -14644\ J$

$Q = -14.7\ kJ$

④ $Q = -L_F m$

$Q = -6.03\ kJ/mol \times \frac{1 mol}{18.02 g} \times 35g$

$Q = -11.7\ kJ$

⑤ $Q = mc\Delta T$

$Q = 35g \times 2.01\ J/g°C \times -20°C$

$Q = -1407\ J$

$Q = -1.4\ kJ$

⑥ $Q_T = Q_1 + Q_2 + Q_3 + Q_4 + Q_5$

$Q_T = -3.5 kJ - 79.2 kJ - 14.7 kJ - 11.7 kJ - 1.4 kJ$

$Q_T = -110.526\ kJ$

2. gold: $Q_G = ?$ (− value ← released)
$m_G = 25g$
$c_G = 0.129\ J/g°C$
$\Delta T_G = T_F - 800°C$

water: $Q_W = ?$ (+ value ← absorbed)
$m_W = 200g$
$c_W = 4.184\ J/g°C$
$\Delta T_W = T_F - 10°C$

Due to heat transfer $Q_W = -Q_G$

$$m_W c_W \Delta T_W = -m_G c_G \Delta T_G$$

$$200g \times 4.184\ J/g°C \times (T_F - 10°C) = -25g \times 0.129\ J/g°C \times (T_F - 800°C)$$

$$836.8 \times (T_F - 10°C) = -3.225\ (T_F - 800°C)$$

$$836.8\ T_F - 8368°C = -3.225\ T_F + 2580°C$$

$$840.025\ T_F = 10948°C$$

$$T_F = 13.0329°C$$

3. Units will be a problem - suggest converting to kJ

$$120 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} = 502.08 \text{ kJ}$$

Q_T = total heat = 502.08 kJ

Q_w = heat used to warm water

Q_s = heat used to create steam

Find Q_w:

$Q_w = ?$

$m = 0.750 \text{ L} \rightarrow 750 \text{ mL} \rightarrow 750 \text{ g}$

$c = 4.184 \text{ J/g°C}$

$\Delta T = 100 - 15 = 85°C$

$Q_w = mc\Delta T$

$Q_w = 750 \text{ g} \times 4.184 \text{ J/g°C} \times 85°C$

$Q_w = 266\,730 \text{ J}$

$Q_w = 266.73 \text{ kJ}$

Find Q_s available for change of state

$Q_T = Q_w + Q_s$

$Q_s = Q_T - Q_w$

$= 502.08 \text{ kJ} - 266.73 \text{ kJ}$

$= 235.35 \text{ kJ}$

Find mass of steam created

$Q_s = 235.35 \text{ kJ}$

$L = 40.8 \text{ kJ/mol}$ *

$m = ?$

$Q = L_v m$

$$m = \frac{Q}{L_v}$$

$$m = \frac{235.35 \text{ kJ}}{\left(\frac{40.8 \text{ kJ}}{1 \text{ mol}}\right)\left(\frac{1 \text{ mol}}{18.02 \text{ g}}\right)}$$ *

$m = 103.95 \text{ g steam}$

* options for converting!

3. Heat energy of 120 kcal will be used to
alternate

- first warm water from 15°C → 100°C
- then to convert water to steam

$$120\,\text{kcal} \times \frac{4.184\,\text{kJ}}{1\,\text{kcal}} = 502.08\,\text{kJ} \leftarrow \text{heat available}$$

Heat required to warm water:

$Q = ?$
$m = 0.750\,\text{L} \rightarrow 750\,\text{mL} \rightarrow 750\,\text{g}$
$c = 4.184\,\text{J/g°C}$
$\Delta T = (100 - 15) = 85°\text{C}$

$Q = mc\Delta T$

$Q = 750\,\text{g} \times 4.184\,\text{J/g°C} \times 85°\text{C}$

$Q = 266\,730\,\text{J}$

$Q = 266.730\,\text{kJ}$

Heat remaining to evaporate water:

$502.080\,\text{kJ} - 266.730\,\text{kJ} = 235.350\,\text{kJ}$

Mass of water evaporated

$Q = 235.350\,\text{kJ}$

$$L_v = \frac{40.8\,\text{kJ}}{1\,\text{mol}} \times \frac{1\,\text{mol}}{18.02\,\text{g}} = 2.264\,\text{kJ/g}$$

$m = ?$

$Q = L_v \times m$

$$m = \frac{Q}{L_v}$$

$$m = \frac{235.350\,\text{kJ}}{2.264\,\text{kJ/g}}$$

$m = 103.946\,\text{g}$

4. This is a thermal equilization problem

metal
$Q_m = ?$
$m_m = 32 g$
$c_m = ?$
$\Delta T_m = (20.468 - 100)°C$
$= -79.532°C$

water
$Q_w = ?$
$m_w = 500 mL \rightarrow 500 g$
$c_w = 4.184 J/g°C$
$\Delta T_w = (20.468 - 20)°C$
$= 0.468°C$

- since metal cools (releases heat) Q_m is a − value
- since water warms (absorbs heat) Q_w is a + value
- since heat lost by metal must equal heat gained by water

$$-Q_m = Q_w$$

$$\therefore -m_m c_m \Delta T_m = m_w c_w \Delta T_w$$

$$\therefore c_m = \frac{m_w c_w \Delta T_w}{-m_m \Delta T_m}$$

$$c_m = \frac{500 g \times 4.184 J/g°C \times 0.468°C}{-32 g \times -79.532°C}$$

$$c_m = 0.385 J/g°C$$

$\therefore$ the metal is **copper** ~~zinc~~
($c = 0.388 J/g°C$)

5. $H_2O(s)$ @ $-25°C \rightarrow H_2O(s)$ @ $0°C$

a) $Q_1 = mc\Delta T$

$Q_1 = 355g \times 2.033 J/g°C \times 25°C$

$Q_1 = 18043 J$

$Q_1 = 18.043 kJ$

b) $H_2O(s)$ @ $0°C \rightarrow H_2O(l)$ @ $0°C$

$Q_2 = L_F m$

$Q_2 = \frac{6.02 kJ}{1 mol} \times \frac{1 mol}{18.02g} \times 355g$

$Q_2 = 118.596 kJ$

c) $H_2O(l)$ @ $0°C \rightarrow H_2O(l)$ @ $100°C$

$Q_3 = mc\Delta T$

$Q_3 = 365g \times 4.184 J/g°C \times 100°C$

$Q_3 = 148532 J$

$Q_3 = 148.532 kJ$

d) $H_2O(l)$ @ $100°C \rightarrow H_2O(g)$ @ $100°C$

$Q_4 = L_v m$

$Q_4 = \frac{40.6 kJ}{1 mol} \times \frac{1 mol}{18.02g} \times 355g$

$Q_4 = 799.834 kJ$

e) $H_2O(g)$ @ $100°C \rightarrow H_2O(g)$ @ $250°C$

$Q_5 = mc\Delta T$

$Q_5 = 355g \times 2.010 J/g°C \times 150°C$

$Q_5 = 107033 J$

$Q_5 = 107.033 kJ$

$$Q_T = Q_A + Q_B + Q_C + Q_D + Q_E$$

$$Q_T = 18.043kJ + 118.596kJ + 148.532kJ + 799.834kJ + 10.703kJ$$

$$Q_T = 1192.037kJ$$

Heat Energy vs Temp for H_2O