Magnesium carbonate decomposes according to the following reaction:

$MgCO_3\,(s) \rightarrow MgO\,(s) + CO_2\,(g) \quad \Delta H_{rxn} = +117.3\ kJ/mol$

a). Is the reaction exo- or endothermic?

endothermic b/c ΔH is positive

b). What is ΔH_{rxn} for the reverse reaction?

$-117.3\ kJ/mol$

c). What is ΔH when 5.35 mol of CO_2 reach with an excess of CO_2?

$$\Delta H = n \cdot \Delta H_{rxn} = 5.35\,mol \cdot (-117.3\ kJ/mol)$$

$$= -628\ kJ$$

d). What is ΔH when 35.5 g of CO_2 reach with an excess of MgO?

$$\#\,moles\,(CO_2) = 35.5\,g \cdot \frac{1\,mol}{44.01\,g} = 0.807\,mol$$

$$\Delta H = 0.807\,mol \cdot -117.3\ kJ/mol = -94.7\ kJ$$

When 1 mol of KBr decomposes to its elements, 394 kJ of heat is absorbed

a). Write a balanced thermochemical equation

$KBr(s) \rightarrow K(s) + \frac{1}{2}Br_2(g)$

b). What is ΔH_{rxn}

394 kJ/mol

c). How much heat is released when 10.0 kg of KBr forms from its elements

$\frac{1}{2}Br(g) + K(s) \rightarrow KBr(s) \quad \Delta H_{rxn} = -394\ kJ/mol$

$\#\ moles\ (KBr) = 10.0\ kg \cdot \frac{1000\ g}{1\ kg} \cdot \frac{1\ mol}{119.0\ g} = 84.03\ mol$

$\Delta H = 84.03\ mol \cdot (-394\ kJ/mol) = 33109\ kJ$

d). How much heat is released when 1 mol of Br_2 is produced?

$\Delta H = \frac{1\ mol\ Br_2}{1/2\ mol\ Br_2} \cdot (394\ kJ/mol)$

What mass of CH_4 must burn to emit 267 kJ of heat

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

$\Delta H_{rxn} = -802.3$ kJ/mol

Plan: $q \rightarrow$ moles $\rightarrow$ mass

$$\Delta H = n \cdot \Delta H_{rxn} \Rightarrow n = \frac{\Delta H}{\Delta H_{rxn}} = \frac{-267\ \text{kJ}}{-802.3\ \text{kJ/mol}}$$

$$n = 0.333\ \text{mol}$$

$$\text{mass}(CH_4) = 0.333\ \text{mol} \cdot \frac{16.04\ \text{g}}{1\ \text{mol}} = 5.34\ \text{g}$$

2g (CH_4) reacts with 5g O_2. How much heat released?

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

\# moles (CH_4) $= \rightarrow 0.125$ mol

\# moles (O_2) $= \rightarrow 0.156$ mol

$$\Delta H = n \cdot \Delta H_{rxn} = 0.156 \text{ mol} \cdot \frac{1 \text{ mol reaction}}{2 \text{ mol } O_2} \cdot (-802.3 \text{ kJ/mol})$$

$$= -62.6 \text{ kJ}$$

Heat Capacity

when a system absorbs heat, its temperature changes by ΔT

$$q \propto \Delta T$$

$$q = C \cdot \Delta T$$

$\hookrightarrow$ heat capacity [J/°C]

Specific heat capacity C_s

Amount of heat required to raise the temp. of 1g of substance by 1°C

[J/g°C]

$$q = m \cdot C_s \cdot \Delta T$$

Molar heat capacity C_m

Raise the temp. of 1mol by 1°C

[J/mol °C]

$$q = n \cdot C_m \cdot \Delta T$$

You heat the same amount of Ag and Fe on the same hot plate

$C_s(Ag) = 0.235\ J/g°C$

$C_s(Fe) = 0.449\ J/g°C$

Ag has the higher temp. b/c it has the lower heat capacity

How much heat is required to warm 1.50L of water from 25°C to 100°C ($d = 1\ g/cm^3$
$= 1\ g/mL$)

$$\text{mass}(H_2O) = 1.50L \cdot \frac{1000\ mL}{1L} \cdot 1\ g/mL = 1500\ g$$

$$q = m \cdot C_s \cdot \Delta T$$

$$q = 1500\ g \cdot 4.18\ J/g°C \cdot (100°C - 25°C)$$

$$q = 95\ kJ$$