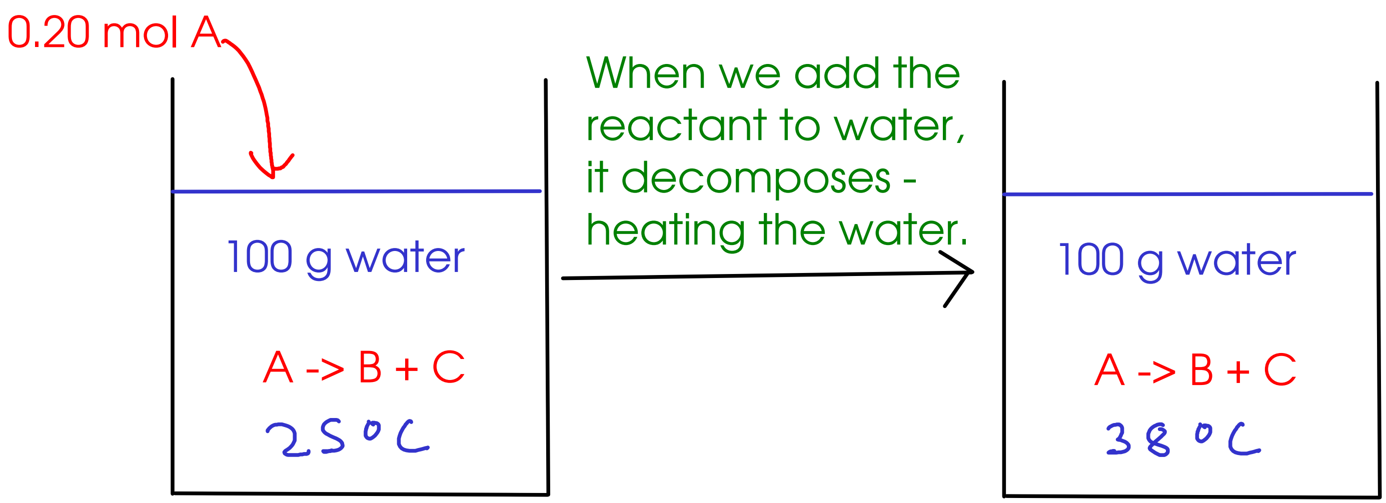
CALORIMETRY

- the measurement of heat. But how do we measure heat?



... what is Q for this reaction?

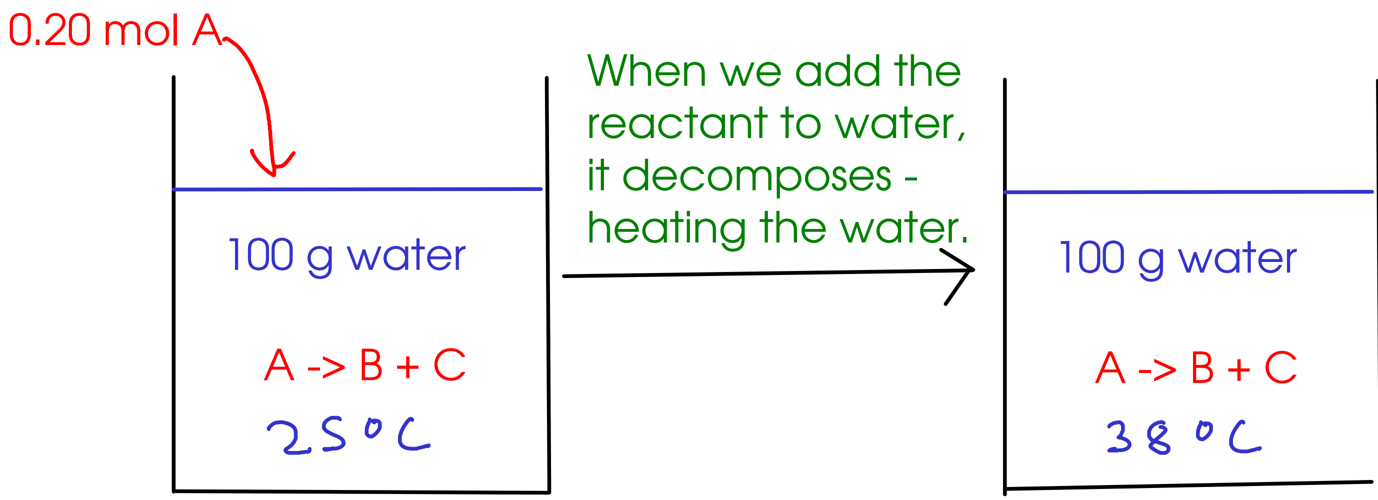
Assuming that no heat is lost from the water to the surrounding air,

$$\underbrace{Q_r}_{\text{reaction}} + \underbrace{Q_w}_{\text{water}} = 0$$

Conservation of energy. The terms add to zero because they have opposite signs.

... if we knew something about the WATER, we could use that to find the heat of the REACTION!

We can look up the water's SPECIFIC HEAT and use it to relate the temperature change of the water to Q.



When we add the reactant to water, it decomposes - heating the water.

Specific heat of water:

$4.184 \; J/g\,°C$

$$Q_r + Q_w = 0$$

$$Q_w = m_w \times s_w \times \Delta T_w$$
$$= (100\,g)(4.184\; J/g\,°C)(38°C - 25°C)$$
$$= +5439.2\; J$$

$$Q_r + Q_w = 0;\quad Q_r + 5439.2\; J = 0;\quad Q_r = -5439.2\; J$$

Usually, reaction heats are reported on a per mole basis:

$$Q = \frac{Q_r}{mol\ A} = \frac{-5439.2\; J}{0.20\ mol\ A} = -27196\ \frac{J}{mol\ A} = \boxed{-27\ \frac{kJ}{mol\ A}}$$

This kind of number is usually called the HEAT OF REACTION!

One problem ...

PATH. The amount of energy required for a process depends on how the process is carried out.

Example: Driving from Florence to Columbia. How much energy is required? (gas)

RAV4 gasoline only (33 mpg) vs hybrid (38 mpg) The gasoline-only model will use more fuel than the hybrid even though they start and end from exactly the same place. So the fuel usage is what we call a PATH FUNCTION, while the location of the car is a STATE FUNCTION.

- Heat of reaction (Q) depends on how the reaction is done.

  - we need (for reporting) some kind of standard condition. At constant pressure, we can define a state function called ENTHALPY (H)

$$H = U + PV$$

$$\Delta H = Q_{\text{constant pressure}}$$

... we record the "enthalpy change of reaction" in our data books.

$$\Delta H_r$$

SINCE the enthalpy change does NOT depend on path, this means that we can use standard values for enthalpy to predict the heat change in reactions that we have not tested in a calorimeter.

THERMOCHEMICAL EQUATIONS

- is like a regular chemical equation, except that phase labels are REQUIRED and the enthalpy for the reaction is given along with the equation.

$$CH_3COCH_3(l) + 4O_2(g) \longrightarrow 3CO_2(g) + 3H_2O(l) \; ; \Delta H = -1800 \text{ kJ}$$

- Why are phase labels required? Because phase changes either absorb or release energy.

$\Delta H = -1800$ kJ ... what does this mean?

$$\boxed{\begin{aligned} 1 \text{ mol } CH_3COCH_3 &= -1800 \text{ kJ} \\ 4 \text{ mol } O_2 &= -1800 \text{ kJ} \\ 3 \text{ mol } CO_2 &= -1800 \text{ kJ} \\ 3 \text{ mol } H_2O &= -1800 \text{ kJ} \end{aligned}}$$

We treat the enthalpy change as if it's another product of the reaction!

USING A THERMOCHEMICAL EQUATION

$$CH_3COCH_3(l) + 4O_2(g) \longrightarrow 3CO_2(g) + 3H_2O(l) \; ; \Delta H = -1800 \text{ kJ}$$

What would be the enthapy change when 25 g of water are produced by the reaction?

1) Convert 25 grams of water to moles. Use FORMULA WEIGHT.
2) Convert moles water to enthalpy change. Use THERMOCHEMICAL EQUATION.

① H_2O: H - 2 x 1.008
O - 1 x 16.00

18.016 g H_2O = mol H_2O

② 3 mol H_2O = -1800 kJ

$$25 \text{ g } H_2O \times \frac{\text{mol } H_2O}{18.016 \text{ g } H_2O} \times \frac{-1800 \text{ kJ}}{3 \text{ mol } H_2O} = \boxed{-830 \text{ kJ}}$$

① ②

Notes:
1) This is an exothermic reaction. Energy is lost to the surroundings. Not surprising since this is a combustion (burning) reaction!
2) -830 kJ would be the heat observed in a calorimetry experiment with this reaction provided pressure remained constant.