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)

2000 Jeep Cherokee vs 2008 Toyota Prius. The Jeep will use much more fuel than the Prius 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 is a STATE FUNCTION.

- so the heat of reaction 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.

$$\Delta H_r$$

157 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.

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

- 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!

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

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

1 - Convert 25 g water to moles. Use FORMULA WEIGHT.
2 - Convert moles water to enthalpy change. Use THERMOCHEMICAL EQUATION.

1 H_2O - H: 2×1.008
O: 1×16.00

18.016 g H_2O = mol H_2O

2 3 mol H_2O = -1800 kJ

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

This is an EXOTHERMIC reaction ... this energy is released to the surroundings.

As long as we're at constant pressure, -830 kJ is also equal to the HEAT, Q.

A few more terms related to enthalpy:

- Enthalpy of vaporization / heat of vaporization: The enthalpy change on vaporizing one mole of a substance. (from liquid to vapor)

- Enthalpy of fusion / heat of fusion: The enthalpy change when a mole of liquid changes to the solid state.

Phase changes require energy, too!

 Example problems:

FORMULA WEIGHTS in g/mol in purple

$$\underset{2.016}{2H_2(g)} + \underset{32.00}{O_2(g)} \longrightarrow \underset{18.02}{2H_2O(g)}\ ;\ \Delta H = -484\ kJ$$

Calculate the enthalpy change for the combustion of 1000. g of hydrogen gas.

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1 - Convert 1000. g hydrogen gas to moles. Use FORMULA WEIGHT.
2 - Convert moles hydrogen gas to enthalpy change. Use THERMOCHEMICAL EQUATION.

1 $2.016\ g\ H_2 = mol\ H_2$ 2 $2\ mol\ H_2 = -484\ kJ$

$$1000.\ g\ H_2 \times \underbrace{\frac{mol\ H_2}{2.016\ g\ H_2}}_{1} \times \underbrace{\frac{-484\ kJ}{2\ mol\ H_2}}_{2} = \boxed{-12\bar{0}000\ kJ} \text{ per } 1000\ g\ H_2$$