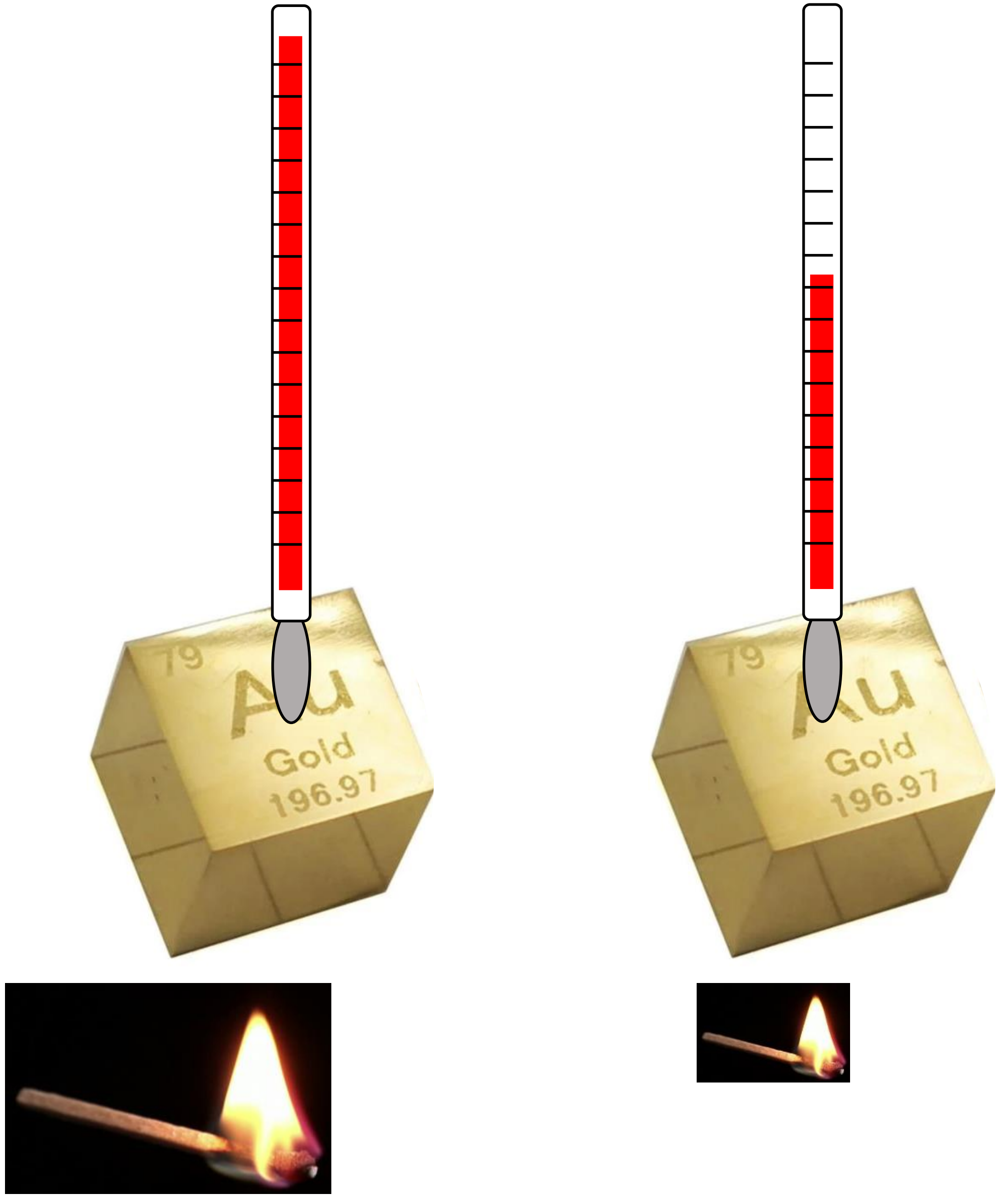
Announcements for Wednesday, 13NOV2024

- Week 10 Homework Assignments available on eLearning
  - Graded and Timed Quiz 10 – "Reactions in aqueous solution" due **tonight at 6:00 PM (EST)**
- Exam 2 is now available for reviewing through ***Gradescope***
- Requests for Exam Question Regrades Now Open
  - Wednesday, 13NOV2024, 12:01 AM (EST) – Friday, 15NOV2024, 11:59 PM (EST)
  - MUST be submitted through ***Gradescope*** (do not email instructors)
  - see Canvas announcement from Nov 12 for regrading policies and procedure
  - after the deadline, Exam 2 grades will not be changed

ANY GENERAL QUESTIONS? Feel free to see me after class!

q vs. ΔT

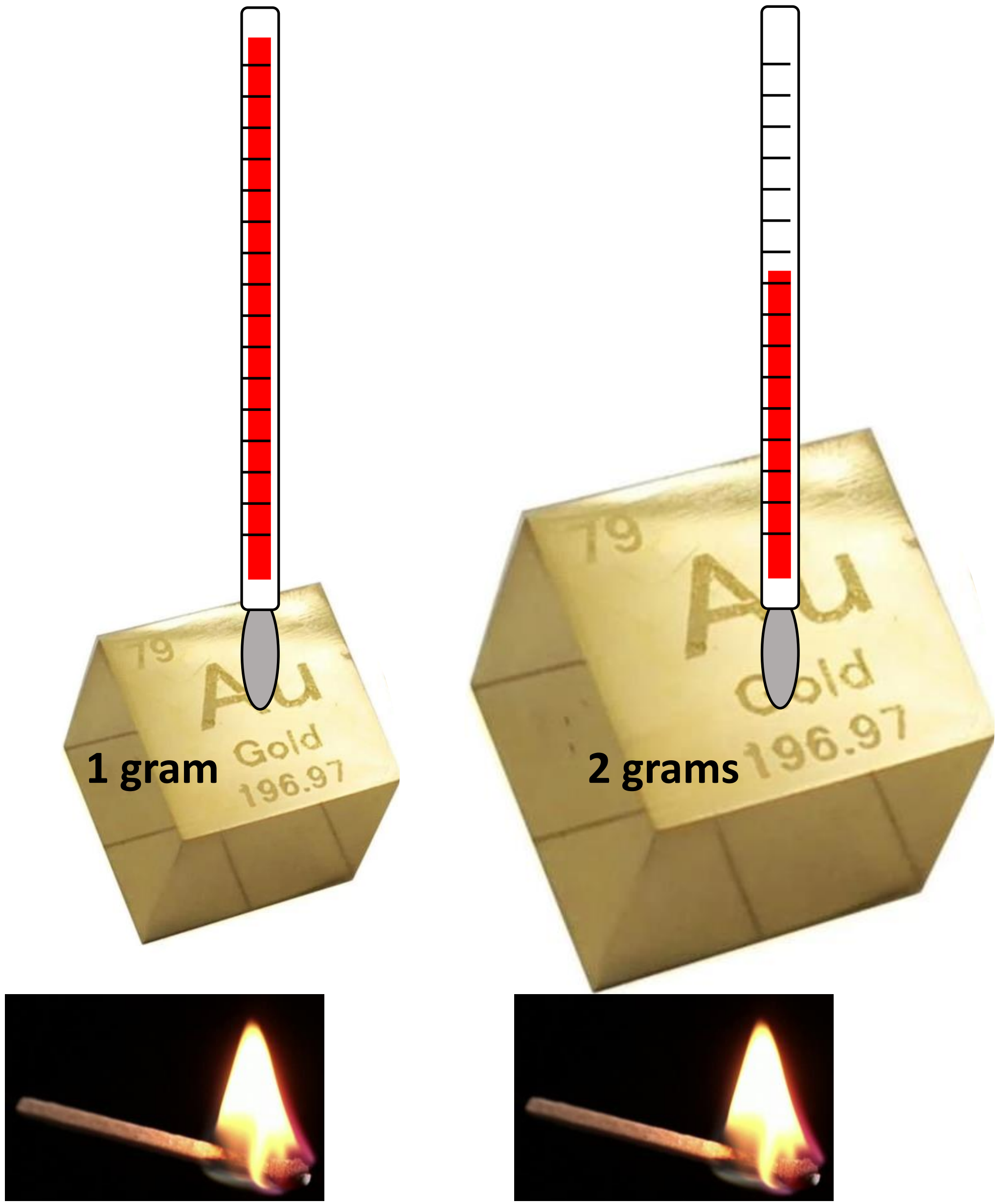


same masses of matter

different amounts of heat added

different ΔT

mass vs. ΔT



different masses of matter

same amount of heat added

different ΔT

composition vs. ΔT

different matter

same masses

same amounts of heat added

different ΔT

Heat Capacity (C)

- in general, the amount of heat needed to change the temperature of a substance by a specified amount

$$C = \frac{q\ (J)}{\Delta T\ (°C)}$$

units: $\mathbf{\frac{J}{°C}}$

or

$$\mathbf{q = C \times \Delta T}$$

  - substances with **high heat capacities generally require more heat** to bring about a temperature change than substances with low heat capacities
    - consider heating an empty aluminum pot on the stove vs. the pot filled with water

Specific Heat Capacity (C_s)

- **specific heat capacity (C_s) =** the amount of heat required to change the temperature **of 1 gram of a substance by 1 °C**

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

heat (J)
+ (absorbed)
– (released)

mass (g)
always +

specific heat ($\frac{J}{g{\cdot}°C}$)
always +

temp. change (°C)
+ (heat absorbed)
– (heat lost)

- **molar heat capacity =** the amount of heat required to change the temperature **of 1 mole of a substance by 1 °C**
- different substances have different specific heat capacities
- How much heat must 55.0 g gold lose to lower its temperature by 80.0 °C?

$$q = m \times C_s \times \Delta T = (55.0\ g)(0.128\ J/g{\cdot}°C)(-80.0\ °C)$$

$$q = -563\ J \text{ or "563 J of heat lost"}$$

TABLE 9.2 Specific Heat Capacities of Some Common Substances

Substance	Specific Heat Capacity, C_s (J/g · °C)*
Elements	
Lead	0.128
Gold	0.128
Silver	0.235
Copper	0.385
Iron	0.449
Aluminum	0.903
Compounds	
Ethanol	2.42
Water	4.18

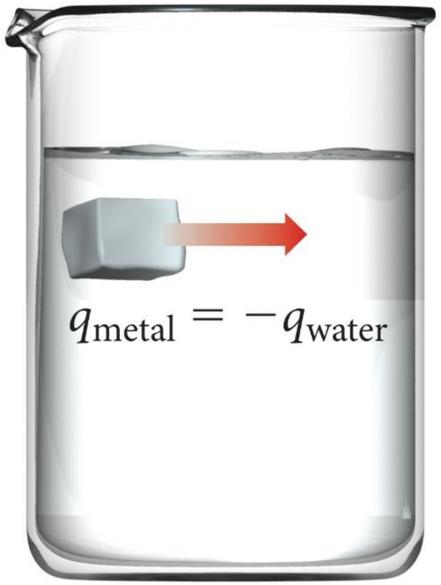
Quantifying Heat Transfers between System and Surroundings

Consider the following scenario:

25 g aluminum metal at 95.0 °C is immersed in 200. g water at 25.0 °C. The specific heats of Al and H_2O are 0.903 J/g·°C and 4.18 J/g·°C, respectively.

Upon immersion

- What happens to the temperature of the water? It increases.
- Why? The water **absorbed heat** from the metal.
- What happens to the temperature of the aluminum? It decreases.
- Why? The metal **released heat** into the water.
- When will the aluminum stop releasing heat and when will the water stop absorbing heat? When the temperatures become equal…the Al reaches **thermal equilibrium** with H_2O.
- How does the amount of heat gained by the water relate to the amount of heat lost by the Al? They have the same magnitude.
- What is the mathematical relationship between the heat gained by the water and the heat lost by the metal? $q_{metal} = -q_{water}$
- How can we determine the final temperature of both the Al and H_2O at thermal equilibrium?

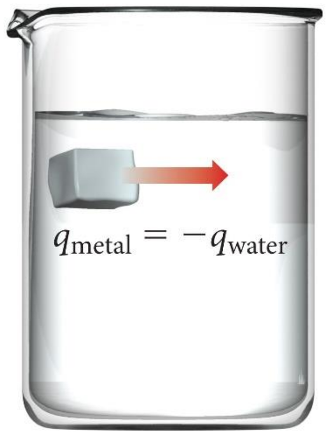


Quantifying Heat Transfers between System and Surroundings

How can we determine the final temperature of both the Al and H_2O at thermal equilibrium?

T_{final} of Al = T_{final} of water

$$\boldsymbol{q_{Al} = -\,q_{water}}$$



$q_{Al} = m_{Al} \times Cs_{Al} \times \Delta T_{Al}$

$q_{water} = m_{water} \times Cs_{water} \times \Delta T_{water}$

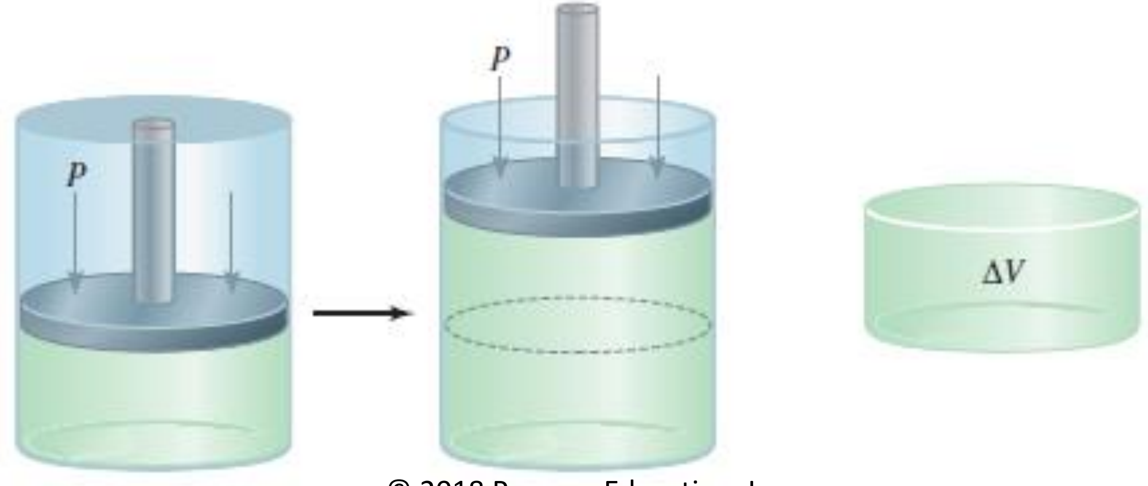
$m_{Al} \times Cs_{Al} \times (T_{final} - T_{initial}) = -\,[m_{water} \times Cs_{water} \times (T_{final} - T_{initial})]$

$25\text{g} \times 0.903\ \text{J/g}\cdot°\text{C} \times (T_{final} - 95.0°\text{C}) = -\,[200.\text{g} \times 4.18\ \text{J/g}\cdot°\text{C} \times (T_{final} - 25.0\ °\text{C})]$

$T_f = 26.8\ °\text{C}$

Quantifying Work (w)

P-V work = work due to changes in the volume of a system against a constant external pressure from the surroundings



$$\textbf{work}\ (\boldsymbol{w}) = -\boldsymbol{P_{ext}\Delta V}$$

P = external pressure in atm

ΔV = change in volume in Liters

1 L·atm = 101.3 J

the volume of a system can expand, contract, or be constant

- when the system expands, V_{final} > V_{inital}
  - ΔV (+), *w* (–), the **system** does work ***ON the surroundings***
- when the system is compressed, V_{final} < V_{inital}
  - ΔV (–), *w* (+), the **system** has work done on it ***BY the surroundings***
- when volume is constant, V_{final} = V_{inital}, ΔV = 0 and *w* = 0

Try This On Your Own

- A 1.40-L gaseous system absorbs 75 J of heat and expands its volume to 2.00 L against an external pressure of 1.02 atm. What is the change in internal energy for this process? 1 L·atm = 101.3 J

A Summary of the Important Sign Conventions

$$\Delta E = q + w = q - P\Delta V$$

ΔE	–	internal energy of the system *decreases*
	+	internal energy of the system **increases**
q	–	the system *releases heat* into the surroundings (lowers E)
	+	the system **absorbs heat** from the surroundings (raises E)
w	–	the *system does work* on its surroundings (lowers E)
	+	the **surroundings does work** on the system (raises E)
ΔV	–	the system is *compressed* by the surroundings (*w* is (+))
	+	the system **expands** against the surroundings (*w* is (–))

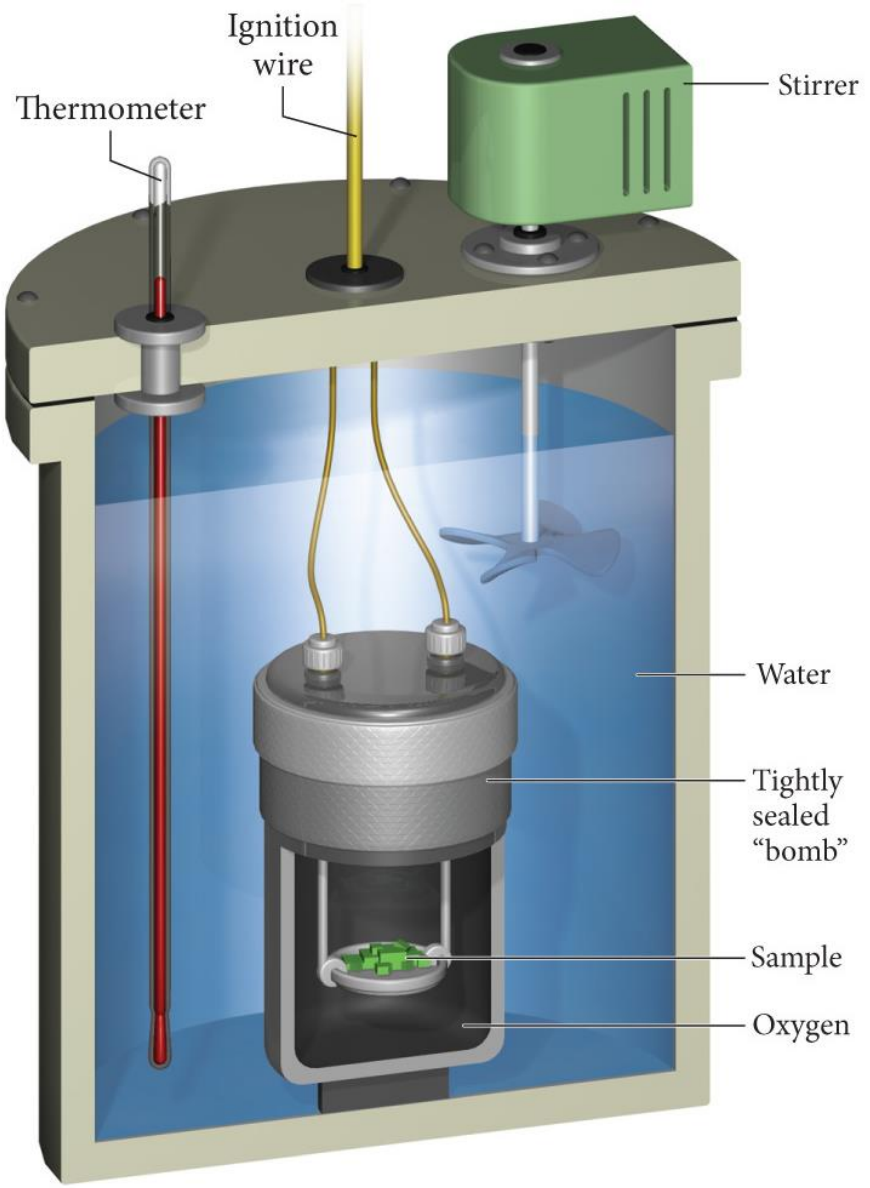
Constant-Volume Calorimetry

- an experimental technique that allows the direct measurement of ΔE for a chemical reaction (ΔE_{rxn}) by forcing all of ΔE to manifest as heat rather than work
- ΔE_{rxn} is measured by measuring the temperature change ***of the surroundings***
  - the reaction takes place in a container of ***constant volume***
  - since ΔV is 0, $\Delta E = q + w \rightarrow \Delta E_{rxn} = q_V$ (heat at constant volume)

Constant-Volume Calorimetry (continued)

use of a bomb calorimeter

- heat transfer between the reaction in the bomb (rxn) and the surrounding calorimeter is measured by the change in temperature (ΔT) of the calorimeter
  - system = the reaction of interest
  - surroundings = the entire calorimeter



Constant-Volume Calorimetry (continued)

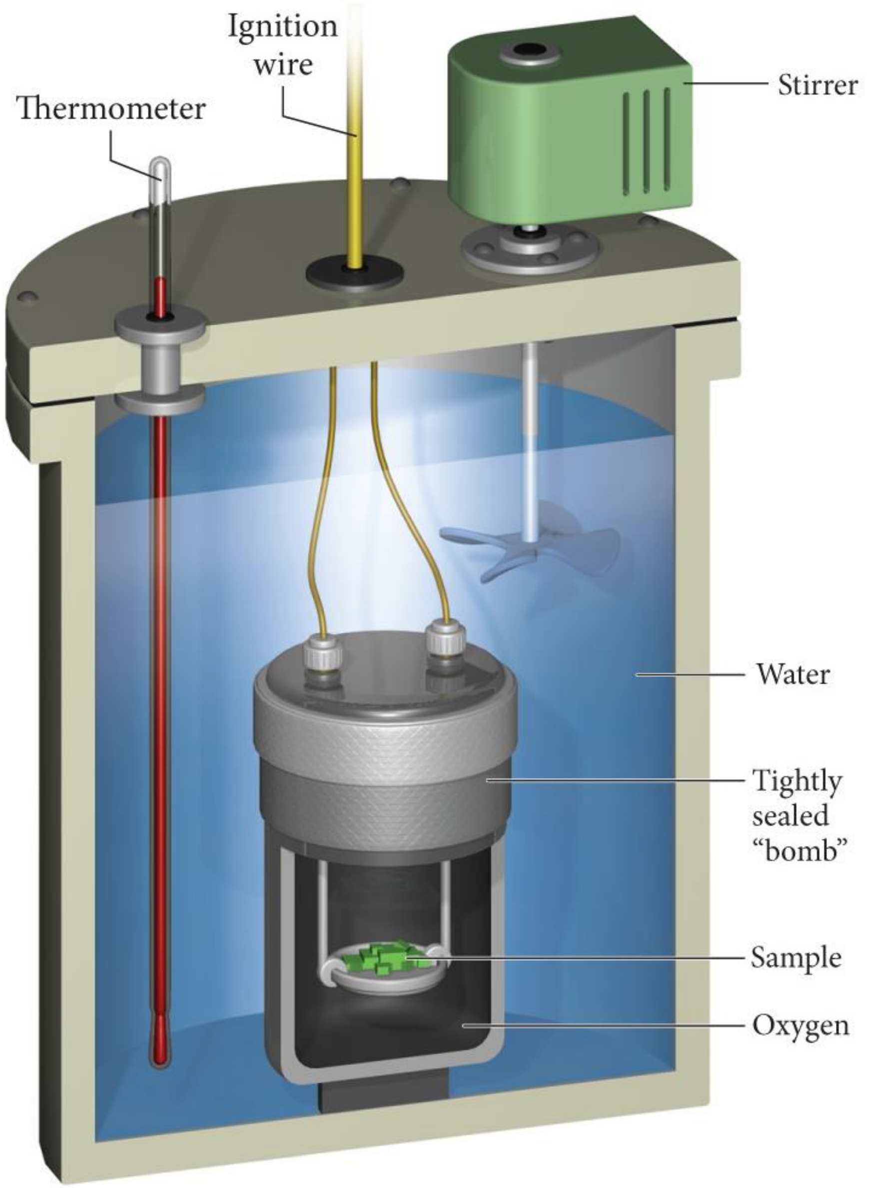
system = the reaction of interest
surroundings = the entire calorimeter

$$\Delta E_{rxn} = -q_{calorimeter}$$

- if ΔE_{rxn} (–), q_{cal} (+) and temperature of calorimeter **increases**
- if ΔE_{rxn} (+), q_{cal} (–) and temperature of calorimeter **decreases**

$$q_{cal} = C_{cal} \times \Delta T$$

- C_{cal} is the heat capacity of the ***entire calorimeter assembly*** (i.e., the water, the walls of the calorimeter, etc.)
- specific amounts of reactants are consumed and the resulting ΔE is for those specific amounts
  - to get ΔE per mole of reactant, ΔE must be divided by the amount of reactant actually reacted



Try This On Your Own

- When **1.550 g** of liquid hexane (C_6H_{14}) undergoes combustion in a bomb calorimeter, the temperature of the calorimeter rises from 25.87 °C to 38.13 °C. Find ΔE_{rxn} for the combustion of ***1 mole*** of hexane in kJ. The heat capacity of the bomb calorimeter is 5.73 kJ/°C.