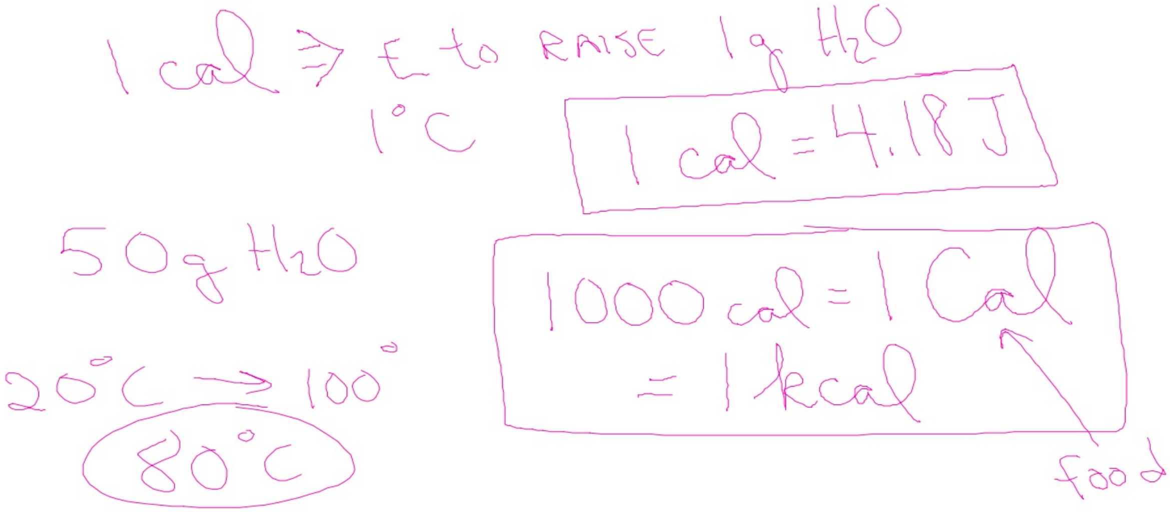


The purpose of this slide is to relate back to the introductory activity on heat where students were asked to find the amount of heat necessary to get a sample of water to the boiling point. The definitions of heat units are also discussed.

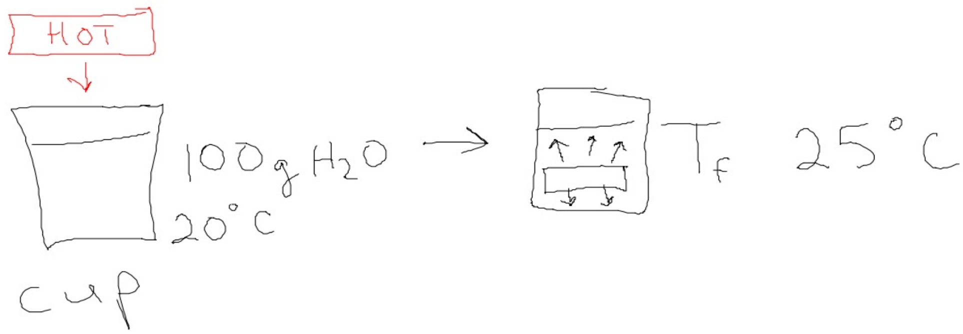
HEAT → Q

"ENERGY that flows between 2 objects at different temp."

Definition of heat and how it always flows between objects of different temepratures until the two objects are the same temperature.

Water is used as a gauge to see how much heat flows out or into other stuff

CALORIMETRY

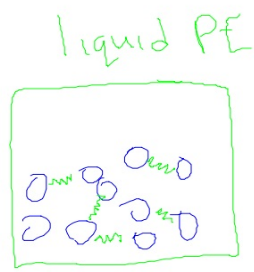


Introduction of the calorimetry concept. Since water's properties are so well known (1 cal is the amount of heat necessary to raise 1 g of water 1 degree Celcius) we use water to measure how much heat comes out of or goes into other objects (and chemical reaction).

Internal Energy, E

$$E = KE + PE$$

$$\Delta E = \Delta KE + \Delta PE$$



The internal energy of an object is comprised of two parts--how fast the molecules are moving or jiggling (KE) and how strongly the molecules in a solid or liquid are attracted to one another (PE). Intermolecular forces (PE) are like little springs that hold solids and liquids together.

$$\Delta KE \propto \Delta T$$

$$Q = m C \Delta T$$

table 16-2

$$\Delta PE \Rightarrow \text{Phase } \Delta$$

$s \rightleftharpoons l$ melting (fusion)

$l \rightleftharpoons g$ boiling (vaporization)

$$Q = n \Delta H_{vap}$$

$$Q = n \Delta H_{fus}$$

kJ/mol

table 16-6

The heat necessary to change how fast the molecules are moving (KE) is given in the green equation. The heat necessary to overcome PE energy in a solid or liquid is given in red. Ideal gases don't have any PE--only KE.

"C" is the specific heat capacity of a substance and is found in a table or measured in the lab. Heats of vaporization or fusion are associated with phase changes.

79. How much heat is absorbed by a 2000 kg granite boulder as energy from the sun causes its temperature to change from 10°C to 29°C?

$Q = mC\Delta T$

$Q = (2000 \times 10^3 g)(0.803 \frac{J}{g°C})(29°C - 10°C)$

$Q = 3.05 \times 10^7 J$

Table 16-2

Specific Heats of Common Substances at 298 K (25°C)	
Substance	Specific heat J/(g·°C)
Water(l) (liquid)	4.184
Water(s) (ice)	2.03
Water(g) (steam)	2.01
Ethanol(l) (grain alcohol)	2.44
Aluminum(s)	0.897
Granite(s)	0.803
Iron(s)	0.449
Lead(s)	0.129
Silver(s)	0.235
Gold(s)	0.129

Note that the mass was given in "kg" and had to be converted to "g" to get the units to cancel. Also note how the units "g" and "C" cancel, leaving units of heat, "J."

343 g CH_3CH_2OH

$$\frac{343\text{ g}}{} \Big| \frac{1\text{ mol}}{46\text{ g}} = 7.5\text{ mol}$$

$Q = n\Delta H_{vap}$

$Q = (7.5\text{ mol})(38.6\ \tfrac{\text{kJ}}{\text{mol}})$

$Q = 290\text{ kJ}$

Table 16-6

Standard Enthalpies of Vaporization and Fusion			
Substance	Formula	ΔH°_{vap} (kJ/mol)	ΔH°_{fus} (kJ/mol)
Water	H_2O	40.7	6.01
Ethanol	C_2H_5OH	38.6	4.94
Methanol	CH_3OH	35.2	3.22
Ammonia	NH_3	23.3	5.66

82. How much heat is required to vaporize 343 g of liquid ethanol at its boiling point? $\Delta H_{vap} = 38.6$ kJ/mol

83. How much heat is evolved when 1255 g of water condenses to a liquid at 100°C? $\Delta H_{cond} = -40.7$ kJ/mol

convert to moles

$Q = n\Delta H_{vap}$

Heats of fusion and vaporization are given in our text in kJ/mol, so you have to convert mass into moles to use this equation. Note that "n" means "number of moles" and it's different than "m" which means mass.