Chem 111

Lecture 15

Announcements

- Breanne recitation session → HASA 126, Wed (10/13) from 5 - 6pm

Homework

- Continue Reading Chapter 5
- OWL online homework

Recap

- Beer’s Law
- Thermodynamics – Kinetics v Potential
- Units
- System V Surroundings
- Temperature
- Thermal Equilibrium
- Endo and Exothermic Process
- Heat Cap city

Let’s Practice

How much heat is needed to warm 250k g of water (about a cup of water) from 22 °C (about room temperature) to near its boiling point, 98 °C?

Specific Heat H_2O = 4.18 J/gK

$$q = mC\Delta T$$

heat

$$= (\text{mass})(\text{SH})(\text{change in temp})$$

$$= (250\,\cancel{g})\left(4.18\,\frac{J}{\cancel{g}\cancel{K}}\right)(98-22\,\cancel{K})$$

$$\Delta K = \Delta °C$$

$$= 7.9 \times 10^4\ J$$

Page 217

- Assume energy is transferred as only heat
- Energy is only transferred within the system
- q_{water} is positive value ✓
- q_{metal} is negative value ✓
- $q_{water} = -q_{metal}$

$$q_{water} + q_{metal} = 0$$

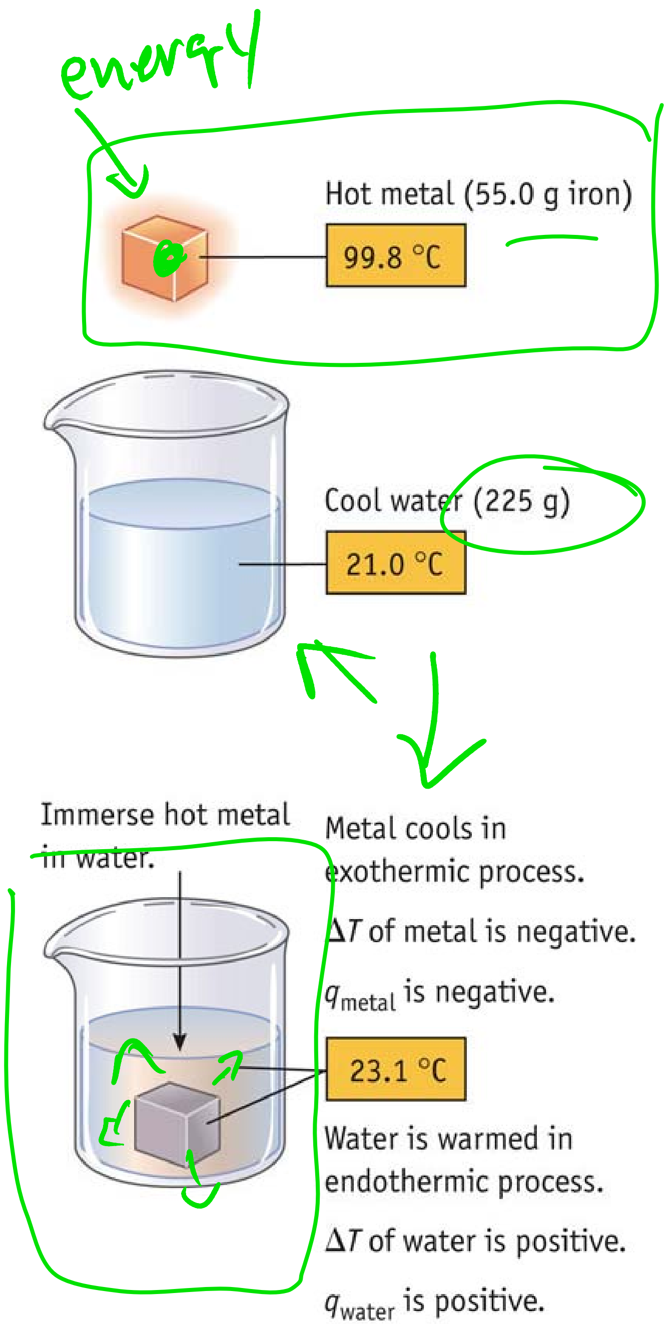
- Water and metal end up at the same temperature

$q = m C \Delta T$

$m_w C_w \Delta T_w \ominus m_m C_m \Delta T_m$

$(23.1 - 21)$ $\quad$ 4 $\quad$ $(23.1 - 99.8)$

energy



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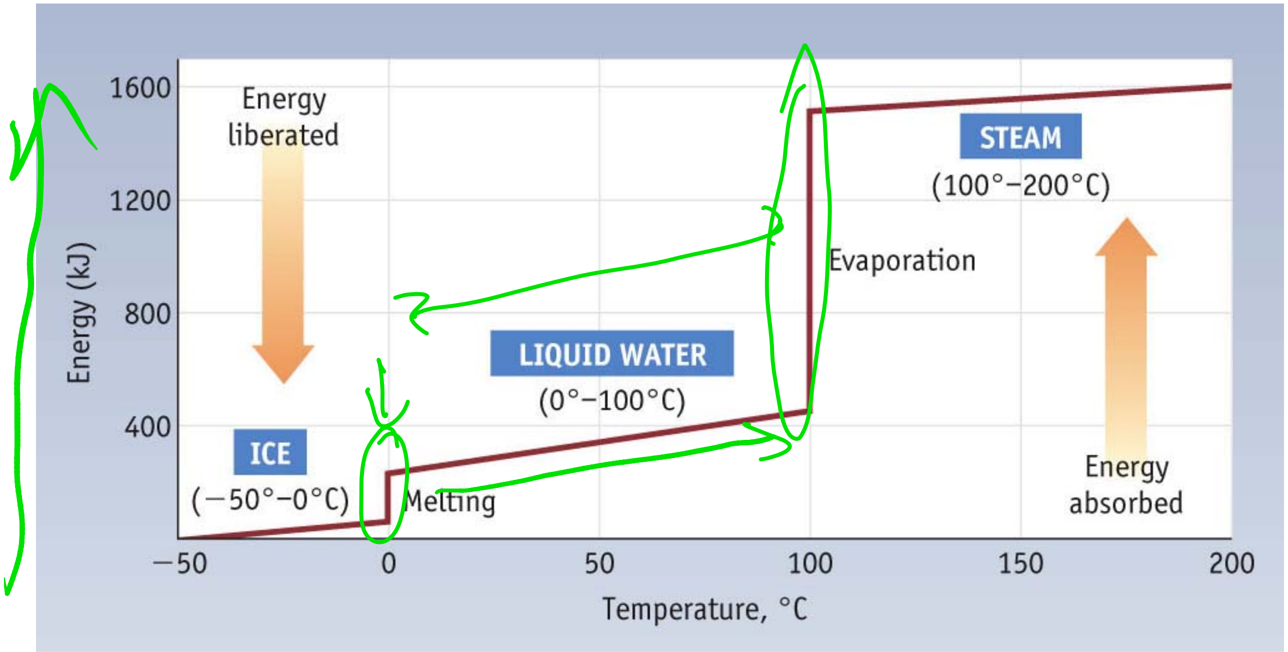
Phase Change

Heat of Fusion: Heat required to convert a pure substance from a solid to a liquid.

Heat of Vaporization: Heat required to convert pure substance from a liquid to a gas.



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Temperature is constant throughout a change of state

$q = \boxed{m\,C}\,\Delta T$

Internal energy

Internal energy (U): is the total energy of the system

First Law of Thermodynamics:

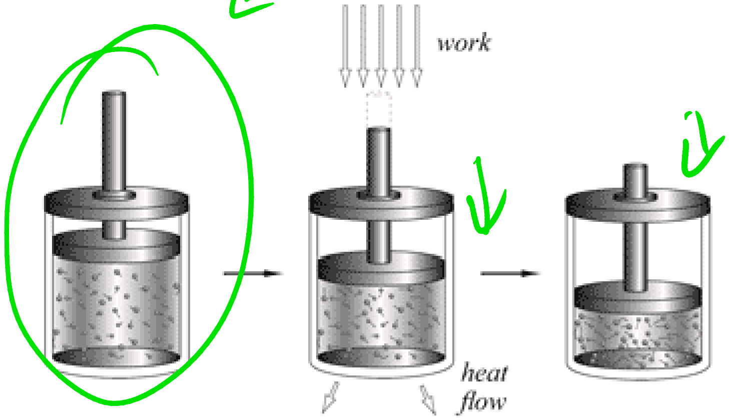
$$\Delta U = q + w$$

up ↑

heat work

↓

← temperatur =g

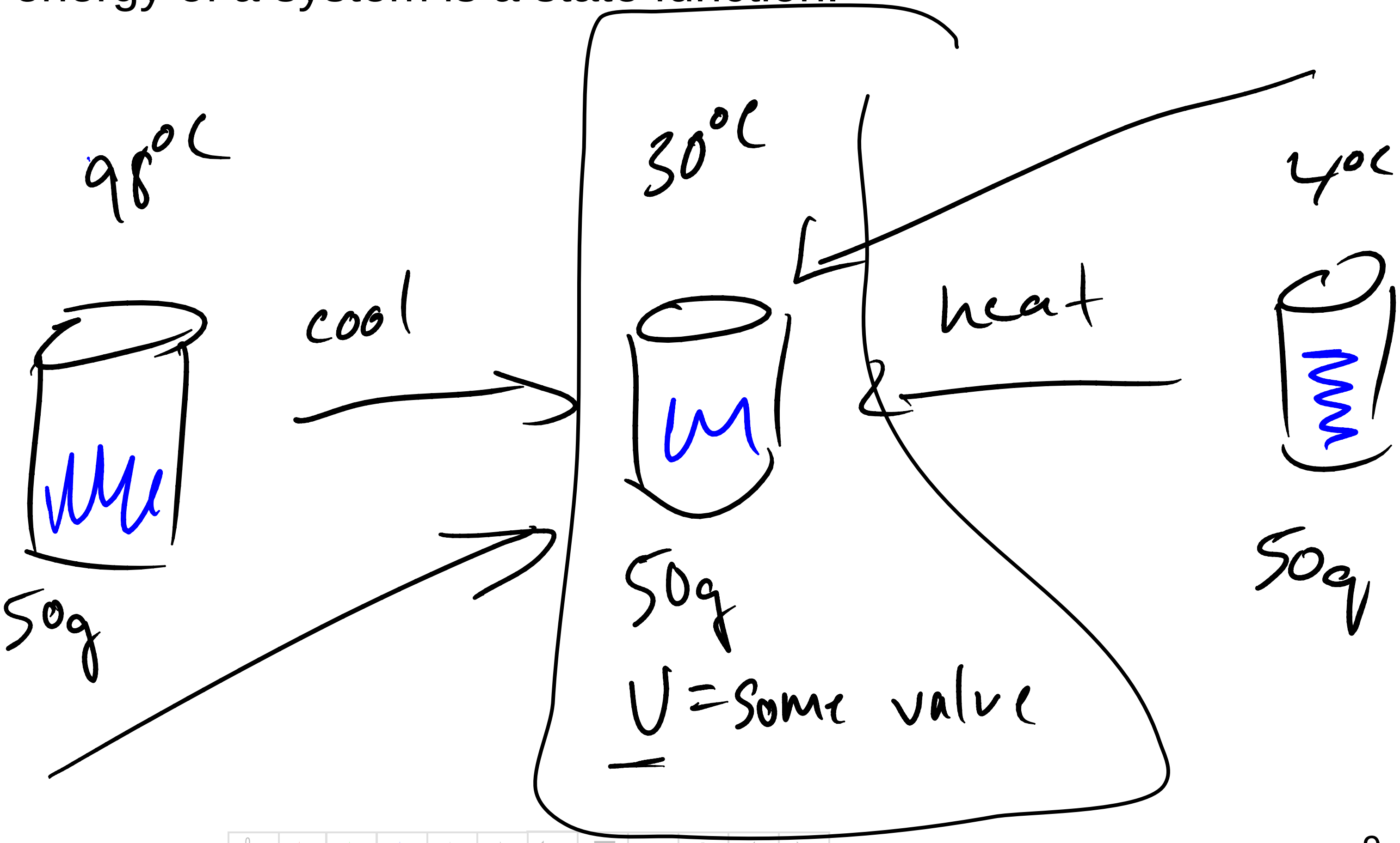


Change	Sign Convention	Effect on U_{system}
Heat: System ← Surrounding	$q>0$ (+)	U increases
Heat: System → Surrounding	$q<0$ (-)	U decreases
Work: System ← Surrounding	$w>0$ (+)	U increases
Work: System → Surrounding	$w<0$ (-)	U decreases

State Function

The value of a state function does not depend on the particular history of the sample, only its present condition.

The internal energy of a system is a state function.



Enthalpy

Many changes occur under constant pressure.

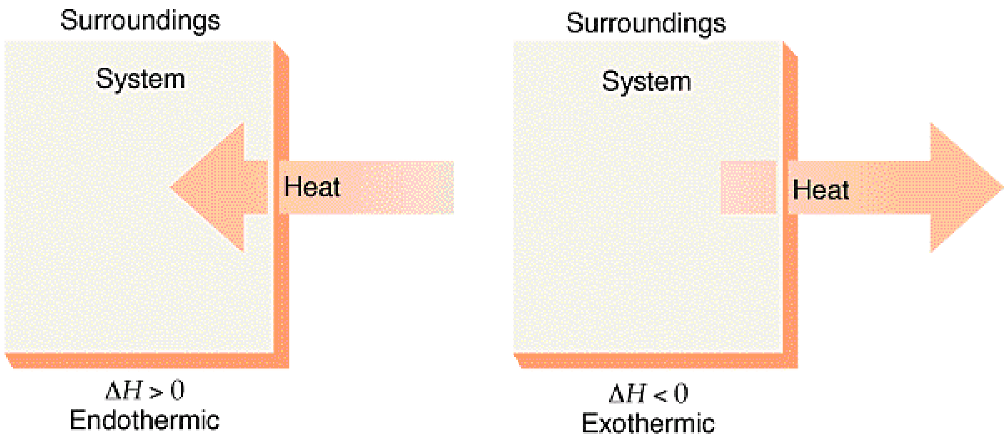
So most energy transfers that take place are in the form of heat.

Enthalpy is the heat absorbed or released under constant pressure.

It is a state function.

The change in enthalpy, ΔH, equals the heat, q_p, gained or lost by the system when the process occurs under constant pressure.

$$\Delta H = H_{final} - H_{intitial} = q_p$$



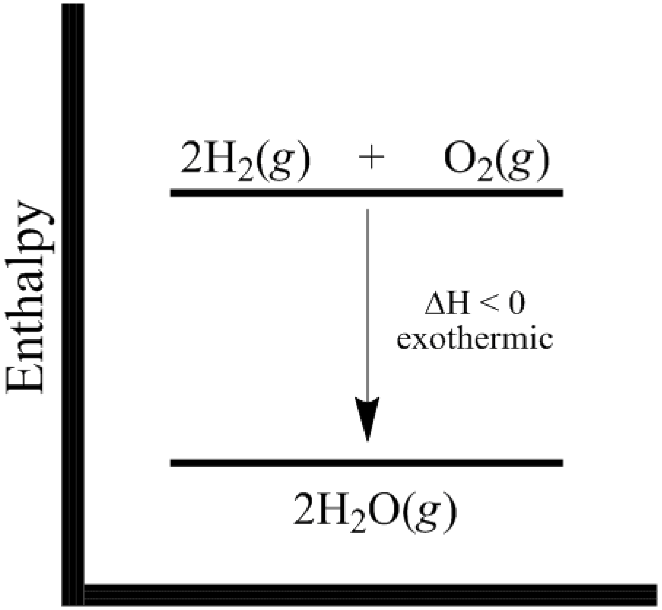
Enthalpies of Reaction

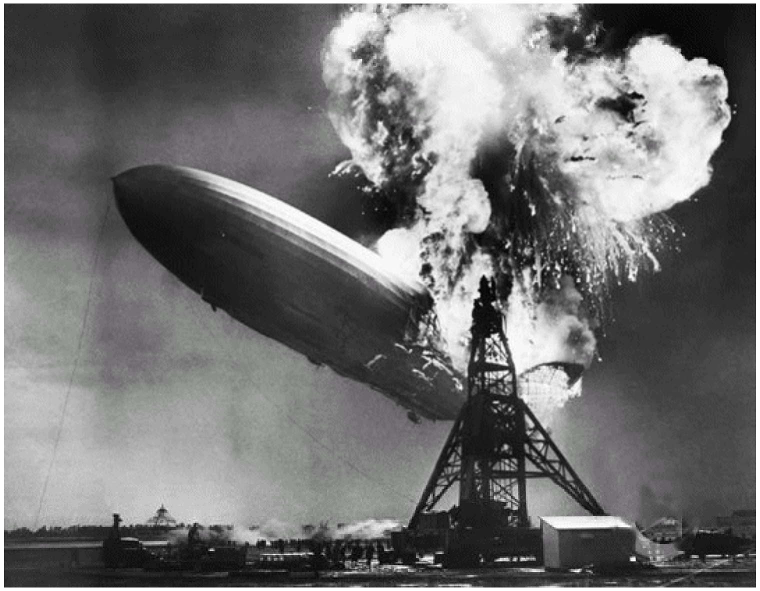
$\Delta H = H_{final} - H_{initial}$

So for a reaction

$\Delta H_{rxn} = H(products) - H(reactants)$

$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$ $\Delta H^o_{rxn} = -483.6\ kJ$





Enthalpies of Reaction

Enthalpy is an extensive property. The magnitude of ΔH is directly proportional to the amount of reactant consumed in the process.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g) \qquad \Delta H^o = -802\ \text{kJ}$$

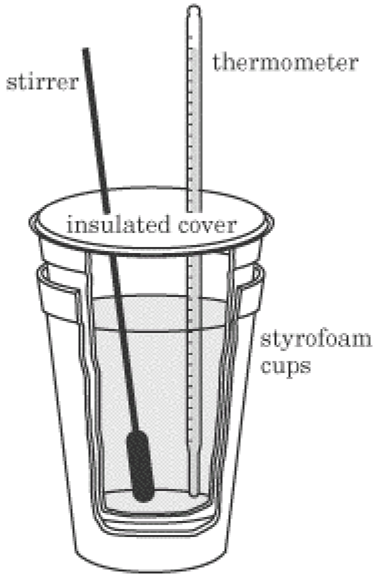
The enthalpy change for a reaction is equal in magnitude but opposite in sign to ΔH for the reverse reaction.

$$CO_2(g) + 2\,H_2O(g) \rightarrow CH_4(g) + 2O_2(g) \qquad \Delta H^o = 802\ \text{kJ}$$

The enthalpy change for a reaction depends on the state of reactants and products.

$$2H_2O(g) \rightarrow 2H_2O(l) \qquad \Delta H^o = -88\ \text{kJ}$$

Calorimetry



The measurement of heat flow.
Apparatus used to measure is called a calorimeter.

Not sealed, so constant pressure.

The heat gained by the solution, q_{soln}, is readily calculated from the mass of the solution, it specific heat, and the temperature change.

$$q_{soln} = C_{soln} m_{soln} \Delta T_{soln}$$

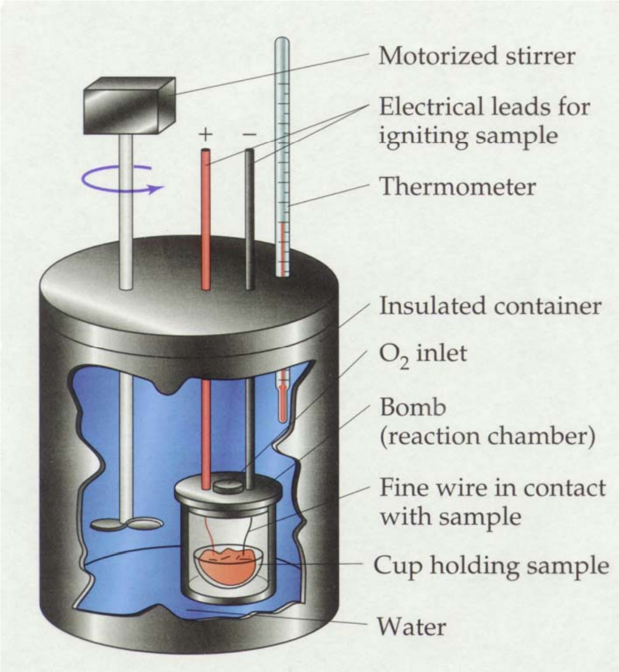
Calorimeter is isolating the system from surroundings, therefore the heat of reaction, q_{rxn}, must be equal in magnitude and opposite in sign from $q_{soln.}$

$$q_{rnx} = - q_{soln} = - \{C_{soln} m_{soln} \Delta T_{soln}\}$$

Bomb Calorimetry

Usually used with organic compounds.

$q_{rxn} = -C_{calorimeter}\Delta T$



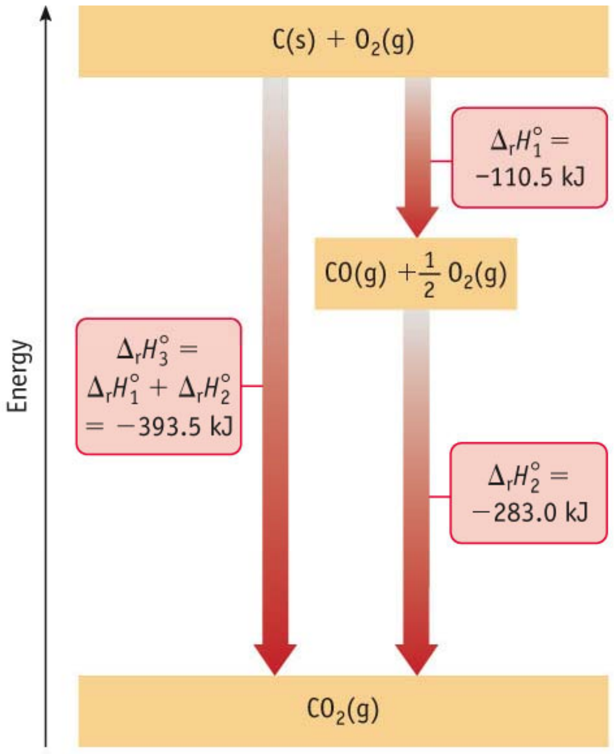
Hess' Law

Remember enthalpy is a state function.

$\text{C} + \tfrac{1}{2}\text{O}_2 \rightarrow \text{CO} \qquad \Delta H^\circ = -110.5\ \text{kJ}$

$\text{CO} + \tfrac{1}{2}\text{O}_2 \rightarrow \text{CO}_2 \qquad \Delta H^\circ = -283.0\ \text{kJ}$

$\text{C} + \text{O}_2 \rightarrow \text{CO}_2 \qquad \Delta H^\circ = -393.5\ \text{kJ}$



Hess's Law states that if a reaction is carried out in a series of steps, ΔH for the reaction will be equal to the sum of the enthalpy changes for the individual steps.