

GENERAL CHEMISTRY II - Chem 206

Be curious...

- How do they determine how many calories are in food?
- Why do we put salt on the roads in winter?
- What does Beano[®] do?
- Why do stalactites and stalagmites form in caves?
- Why does red cabbage make salad dressing turn pink?
- Why do we clean kettles & coffeemakers with vinegar?

...And learn to address your own questions!

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GENERAL CHEMISTRY II - Chem 206

WHY STUDY CHEMISTRY ?

- learn how substances tend to behave, & why
- learn to figure out how everyday stuff works

BUILD SKILLS:

- learn to think on multiple levels
- learn to apply knowledge
- learn to attack problems

Section 01: Tues 18:00-20:15

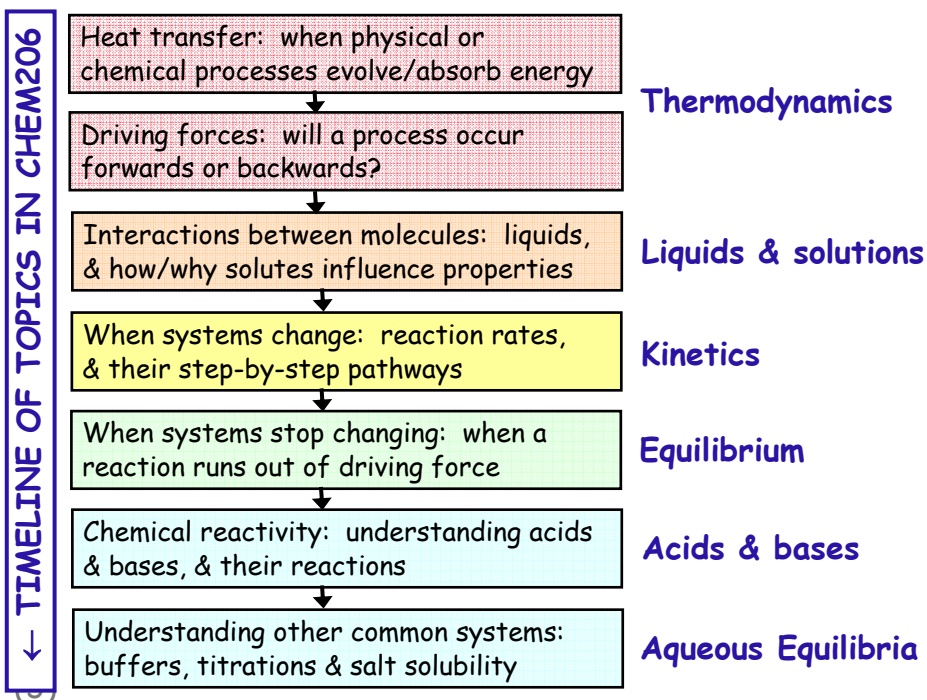
Lecture 01: Tues. Sept. 04/07

Professor: Dr. Carrie Rogers, SP-201.17, x5838

<http://artsandscience.concordia.ca/facstaff/p-r/rogers>

Office hours: Mon.-Fri. 13:15-14:15

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FIRST TOPIC:

WHAT DRIVES CHEMICAL REACTIONS?

Thermodynamics:

ENERGETICS OF CHEMICAL PROCESSES

CHAPTER 6: THERMOCHEMISTRY

CHAPTER 19: SPONTANEITY

Review on your own before next class:

- Balanced chemical equations (Ch.4)
- Product-favoured vs. reactant favoured rxns (p.197)
- Kelvin & Celsius temperature scales (Section 1.6)
- States of matter & changes of state (Section 1.5)

CHAPTER 6 Energy & Chemical Reactions & Physical Processes

Chapter Outline:

- 6.1 Energy: basic principles
- 6.2 Specific heat capacity & heat transfer
- 6.3 Energy & changes of state
- 6.4 The 1st law of thermodynamics
- 6.5 Enthalpy changes for chemical reactions
- 6.6 Calorimetry
- 6.7 Hess's law
- 6.8 Standard enthalpies of formation
- 6.9 Product- or reactant-favoured reactions & thermochemistry

Chapter Goals:

- Assess heat transfer associated with changes in temperature and changes of state.
- Apply the first law of thermodynamics.
- Define and understand the state functions enthalpy and internal energy
- Calculate energy changes occurring in chemical reactions and learn how these changes are measured

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The Law of Conservation of Energy = "the 1st law"

Energy can be neither created nor destroyed,
BUT can be *transferred & transformed*

↘

i.e., converted from
one form into another

The energy of the universe is constant.

$$\Delta E_{\text{universe}} = 0$$

" the first law of thermodynamics "

IMPORTANT:

- the **system** you are interested in CAN gain or lose E
- through energy exchange with its **surroundings**

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6.1 ENERGY = capacity to do work or produce heat

1.) Potential energy: energy due to position

- gravitational energy
- electrostatic energy: due to separation of charges
- chemical potential energy:
 - due to attraction of e⁻s and nuclei in molecules...
 - changed by rearranging bonds

2.) Kinetic energy: energy due to motion of the object

- mechanical energy: of a moving macroscopic object
- electric energy: e⁻s moving through a conductor
- sound: compression & expansion of spaces b/w molecules
- thermal energy: kinetic energy available to molecules because of sample's temperature

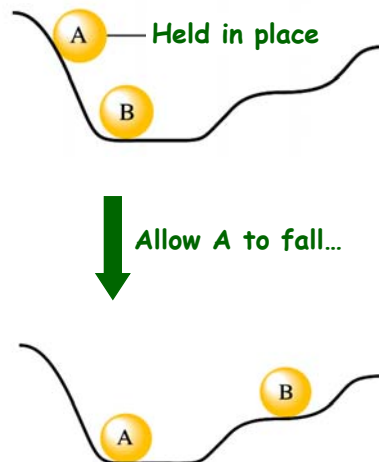
Zumdahl's Figure 6.1 (cf Kotz Fig. 6.4)

POTENTIAL ENERGY

Where did A's lost energy go?

Potential energy lost by ball A is converted to:

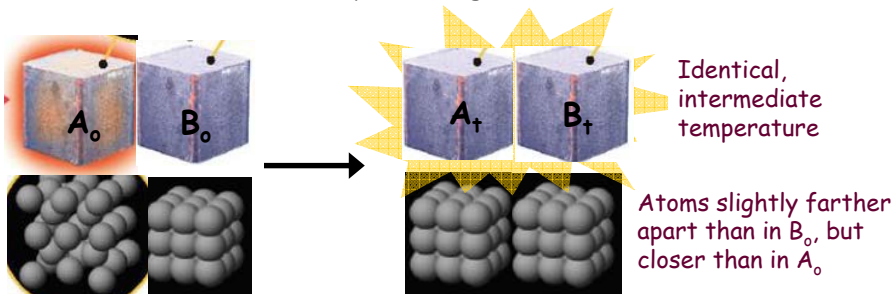
- ↑ in potential energy of ball B
- random motions of components of hill → frictional heating increases hill's temperature!



Heat is transferred because of collisions

When two objects at different temperatures touch

- Hotter substance's molecules have higher kinetic E
- Collisions at interface between hot & cold material:
 - higher E molecules transfer E to lower E molecules
 - ⇒ hotter substance has lost some thermal E, cooler has gained some
- As this continues, the temperature gradient lessens...



Thermal equilibrium: when objects reach same temperature

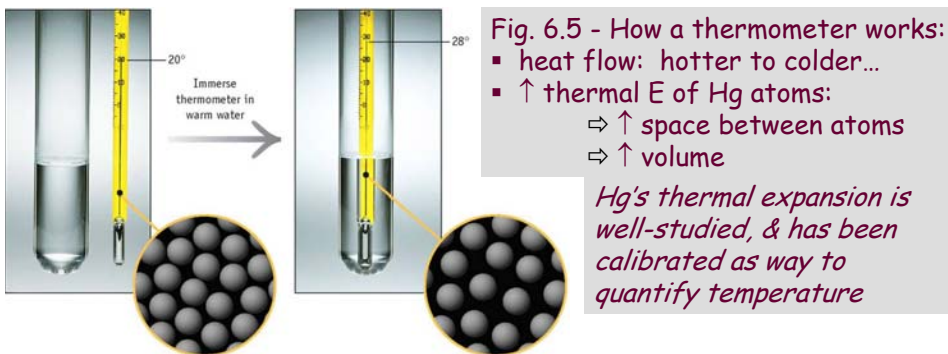
When thermal E throughout entire system is uniform, collisions no longer result in a net transfer of E...

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TEMPERATURE \neq HEAT: not the same!

TEMPERATURE describes average thermal energy...
...HEAT is ENERGY, in transit !

- Temperature reflects random motions of particles
i.e., higher T ⇒ higher thermal E ⇒ more motion
- Heat involves a transfer of energy between two objects because of a temperature difference



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Q: WHICH FRYING PAN WILL GET THE HOTTEST?

Three separate 75.0 g samples of copper, iron and aluminum are available, each initially at 24.0°C.

If 2.00 kJ of heat is applied to each sample, which will end up at the highest temperature?

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6.2 Specific heat capacity & heat transfer

Quantity of heat transferred when T changes depends on:

- identity AND quantity of material
- size of T change

SUBSTANCES HAVE CHARACTERISTIC RESPONSE TO HEAT:

- **Specific heat capacity (C_s)**
= energy required to raise the temperature of 1 g of material by 1°C

NOTE: 1 calorie (cal) = 4.18 J
& 1 dietary Calorie = 1 kcal

TABLE 6.1 (Zumdahl)
The Specific Heat Capacities of Some Common Substances See K Table 6.1

Substance	Specific Heat Capacity (J/°C · g)
H ₂ O(l)	4.18
H ₂ O(s)	2.03
Al(s)	0.89
Fe(s)	0.45
Hg(l)	0.14
C(s)	0.71

Kotz calls specific heat capacity "C", but see common conventions next...

How much E does it take to warm up a material?

- **Specific heat capacity (C_s)**
= energy required to raise temp. of **1 g** of material by 1°C
- **Molar heat capacity (C_m)**
= energy required to raise temp. of **1 mole** of material by 1°C
- **Heat capacity (C)**
= energy required to raise temp. of **an object** by 1°C

TOTAL ENERGY transferred as heat = q

$q = \text{heat capacity} \times \Delta T$ where ΔT is $T_{\text{final}} - T_{\text{initial}}$

**** use dimensional analysis to determine your approach!**

e.g., $q = \underbrace{C_s \times m}_{C} \times \Delta T$ or $q = \underbrace{C_m \times n}_{C} \times \Delta T$ etc...

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Q: WHICH FRYING PAN WILL GET THE HOTTEST?

Three separate 75.0 g samples of copper, iron and aluminum are available, each initially at 24.0°C .

If 2.00 kJ of heat is applied to each sample, which will end up at the highest temperature?

Specific heat capacities

Cu $0.39 \text{ J}/(\text{g}^\circ\text{C})$

Fe $0.45 \text{ J}/(\text{g}^\circ\text{C})$

Al $0.89 \text{ J}/(\text{g}^\circ\text{C})$

TWO APPROACHES

(one long...& one SHORT):

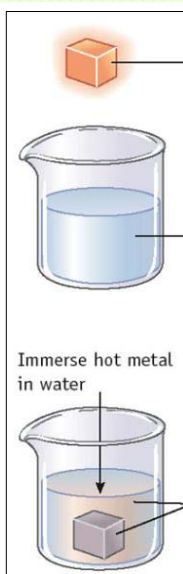
- 1.) Use $q = C_s \times m \times \Delta T$
→ calculate final temperature for each sample

or...

- 2.) Use logic without calculations! Just explain it...

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A typical heat transfer question: a "hot rock problem"



Imagine you are stranded in the wilderness, and you want some warm water to bathe with. You put a rock into your campfire for a while, and then drop the rock into a bucket of cool lake water to warm it up.

Some information you'll need:

Your rock

$$m = 1250 \text{ g}$$

$$C_s = 0.425 \text{ J}\cdot\text{g}^{-1}\text{K}^{-1}$$

Your cool water

$$V = 3.25 \text{ L}$$

$$C_s = 4.184 \text{ J}\cdot\text{g}^{-1}\text{K}^{-1}$$

$$d = 0.9997 \text{ g}\cdot\text{mL}^{-1}$$

If the water was initially at 12°C , & its final temperature is 36°C :

a) What is the rock's final temperature?

b) What was its original temperature?

Fig. 6.10

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Simplifying assumption:

- bucket (immediate surroundings) is perfectly insulated
- thus: can ignore rest of universe

Keeping track of heat flow: "q"

A matter of perspective:

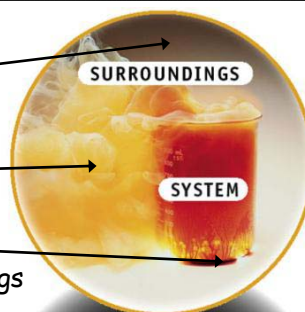
SYSTEM	vs.	SURROUNDINGS
The matter which we are studying		Everything else
• can be contained in a flask, etc... or just have imaginary boundary		• all matter that can exchange energy with system

System + Surroundings = the whole UNIVERSE

Pay most attention to the immediate surroundings...

Fumes (product!) are part of the system

Beaker is part of immediate surroundings



Due to first law: $q_{\text{sys}} + q_{\text{surr}} = 0$
Thus: $q_{\text{sys}} = -q_{\text{surr}}$

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& Now back to our "hot rock problem":

a) What is the final temperature of the rock?

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When heat flows between two objects, it flows only until the two objects reach the same temperature.

b) What was the original temperature of the hot rock?

To figure this out: break problem down further, and ask:

How much energy was transferred as heat from rock to water?

*1°C & 1 K = same size (just offset)
⇒ ΔT is same in EITHER unit*

*(2 SFs....but to do the calculation on the next page
we'll use an extra digit to minimize rounding error)*

This quantity of heat came from the rock,

Thus, $q_{\text{rock}} = -q_{\text{water}} = -330 \text{ kJ}$ ← heat flow according to rock

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....What was the original temperature of the hot rock?

Given $q_{\text{rock}} = -326000 \text{ J}$ (using 3 SFs to minimize rounding error...
& Joules to match units of C_s)

So, $-326000 \text{ J} = C_s \times m \times \Delta T$ now using data for the rock...

$$-326000 \text{ J} = (0.425 \text{ J}\cdot\text{g}^{-1}\cdot\text{K}^{-1}) \times (1250 \text{ g}) \times \Delta T$$

$$\Delta T = \frac{-326000 \text{ J}}{(0.425 \text{ J}\cdot\text{g}^{-1}\cdot\text{K}^{-1}) \times (1250 \text{ g})}$$

$$= -613.6 \text{ K}$$

$$\Delta T = -610 \text{ K} \text{ (2SF based on rules using original data's SF)}$$

i.e., rock's temperature dropped by 610 K, and ended up at 36°C ,
therefore, must have started at 646°C . Mathematically:

$$\Delta T = T_f - T_i \Rightarrow -610^\circ\text{C} = 36^\circ\text{C} - T_i$$

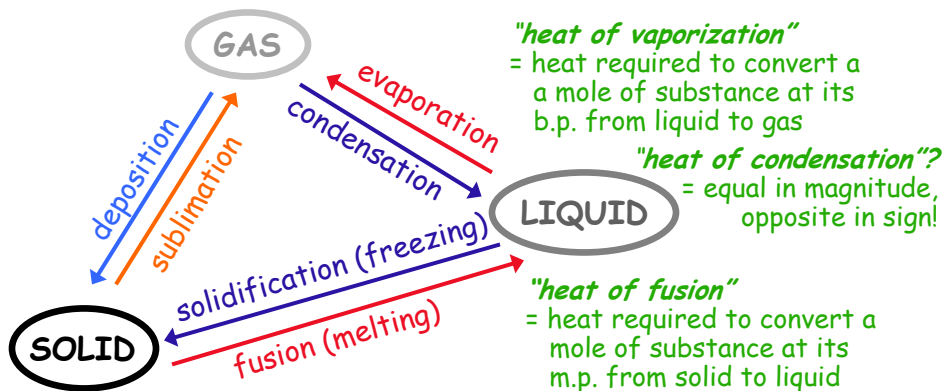
$$T_i = 646^\circ\text{C} \text{ (3SF via addition rules)}$$

Final Answer: The rock was at 646°C when it was dropped into
the water. By transferring 330 kJ of heat to the water, it
warmed up the water to a temperature of 36°C .

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6.3 Energy and changes of state (physical processes)

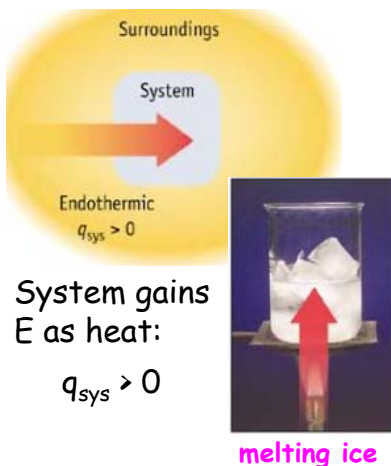
- If molecules brought closer together: heat is released
(can interact more \Rightarrow energetically favourable)
- If molecules forced farther apart: heat must be absorbed
(disrupt attractions between molecules \Rightarrow energetically unfavourable)



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Heat flow is described from system's perspective

Endothermic process:
Heat flows **into** system
(from surroundings).



Kotz Fig.6.8

Exothermic process:
Heat flows **out** of system
(into surroundings).



While a change of state is occurring: T is constant

Below the melting point: adding heat causes crystal to warm up

⇒ more rapid vibration of molecules (thermal motions)

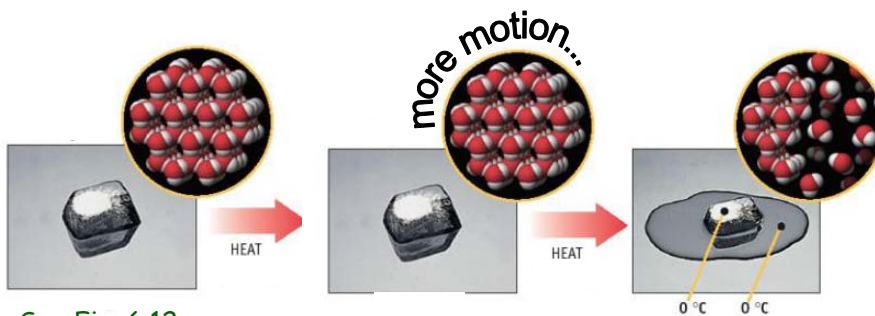
⇒ crystal expands, but ~no molecules change position or leave crystal...

At the melting point:

⇒ any further increase in thermal E overcomes $\text{H}_2\text{O}-\text{H}_2\text{O}$ attractions

⇒ all heat input is used to break intermolecular forces

⇒ sample T stays same until ALL of sample has changed phase



See Fig. 6.12

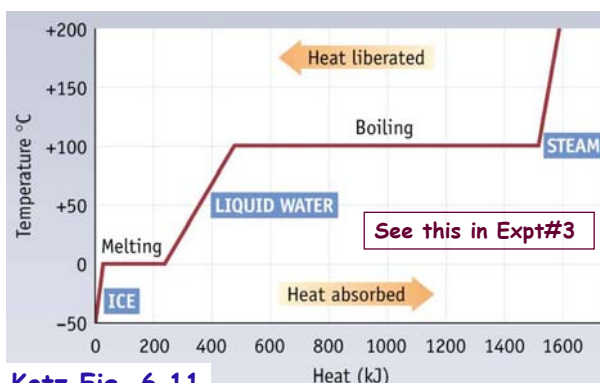
While a change of state is occurring: T is constant

Imagine heating water from -50°C to $+200^{\circ}\text{C}$:

- add heat \Rightarrow T increases

EXCEPT:

- while melting (& again while boiling):
 - \Rightarrow all heat used to break intermolecular forces
 - \Rightarrow sample T stays same until ALL sample has changed phase



Let water cool down again: heat flows into surroundings \Rightarrow T \downarrow

EXCEPT: while condensing (& again while freezing)

- heat released by phase change balances with heat loss \Rightarrow T stays same!

Check out Kotz worked Example 6.3

(23) calculate E required to heat 500. g ice chunk from -50°C to $+200^{\circ}\text{C}$

EX: Will all the ice in your cola melt?

Kotz 6th Ch.6 # 78: You take a cola from the refrigerator, and pour 240 mL of it into a glass. Its temperature is initially 10.5°C , then you add one 45 g ice cube. Which of the following describes the system when thermal equilibrium is reached?

- The temperature is 0°C and some ice remains.
- The temperature is 0°C and no ice remains.
- The temperature is higher than 0°C and no ice remains.

Determine the final temperature and the amount of ice remaining, if any.

Assumptions:

- The cola has the same density & specific heat capacity as water.
 1.00 g/mL $4.184\text{ J}/(\text{g}^{\circ}\text{C})$
- The glass can be ignored, and it is perfectly insulated.
- Can look up enthalpy of fusion of water (ice!): $\Delta H_{\text{fus}} = +333\text{ J/g}$

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EX: Will all the ice in your cola melt?

Approach: start with logic...

- Determine energy loss required to cool cola to 0°C.
- Determine energy input required to melt one ice cube.

So, which process involves more energy? Will all the ice melt?

- Assume: $q_{\text{cola}} = -q_{\text{ice}}$ for the mass of ice melted...
 $m_{\text{cola}} C \Delta T_{\text{cola}} = m_{\text{ice-melted}} C \Delta T_{\text{ice}}$

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ASSIGNED READINGS:

BEFORE NEXT CLASS: Review and...

Read 206 material: Ch.6 sections 6.1-6.3

Work on Ch.6 problems: #1-26, 62-63, 70-78

BEFORE FIRST LAB:

*Review from 205: Ch.12 concepts, & read 12.6-12.9

Remember:

- Labs start week of Jan.7th: Tues aft., Wed aft., Thurs eve.
arrive prepared: lab coat & glasses
completed Expt.#1 prelab
completed tutorial homework
- Chem101 seminars next week

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