Heat

Physical Science 20 – Ms. Hayduk

Heat Terminology

System Definitions

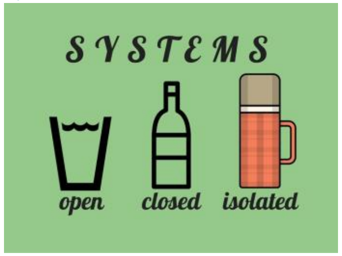
- **System**: the part of the universe being studied (big – Earth, or small – one atom)
- **Surroundings**: the part of the universe outside the system

We look at transfer of energy/matter between system and surroundings.

System Definitions

- **Open system:** free exchange of matter and energy with surroundings
- **Closed system**: free exchange of energy only
- **Isolated system**: no interaction with surroundings

System Definitions



Heat as Energy

- **Kinetic energy**: associated with motion
- **Potential energy**: stored energy associated with forces of attraction or repulsion
- **Thermal energy:** from random molecular motion
  - Proportional to temperature, but not the same thing
  - Measured in Joules (1 kg·m^2/s^2)

Conservation of Energy

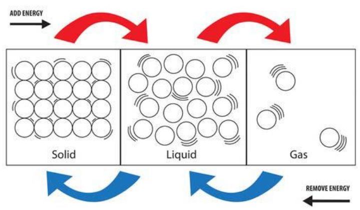
- In interactions between system and surroundings, energy cannot be created or destroyed, only transferred
- Total energy (q) will remain constant

$$q_{system} + q_{surroundings} = 0$$

- Basically: all energy that leaves the system goes into the surroundings, and vice versa

Heat and Temperature

Particle Theory



Temperature

- Measure of **average molecular kinetic energy** of a substance
- Measured in degrees Celsius (°C), degrees Fahrenheit (°F) or Kelvin (K)
- Kelvin and degrees Celsius are the same increment (interchangeable units)

Heat

- Energy transferred between system and its surroundings due to **difference in temperature**
- Always moves from **high** temperature to **low** temperature
- Heat "flows" until average molecular kinetic energy is the same between system and surroundings

Heat

- Variable: q Units: J (Joules)
- A system cannot "contain" heat – it has internal energy that can be transferred as heat

Heat versus Temperature

Heat (q)	Temperature (T)
Measure of energy flowing into or out of a system	Measure of random motion of particles in a substance
Energy that is transferred during a temperature change	Indication of direction heat will flow

Calorimetry

Calorimetry

- The science of determining changes in energy of a system by measuring heat exchanged with surroundings
- **Calorimeter**: well-insulated vessel (as close to an isolated system as possible)

Calorimetry

- Inside, a heated or cooled object (or chemical reaction) – the system – is submerged in water –the surroundings. The temperature change of the water can be used to determine the heat released by the system.
- The calorimeter needs to prevent heat loss to the surroundings so the temperature change within it can be accurately recorded.

Calorimetry

- Coffee cup calorimeter →→→
- Styrofoam is a great insulator
- Scientists that are doing more sophisticated experiments may use fancier equipment, but this device is very accurate for experiments at the high school level!

Heat Capacity

Quantity of Heat

- Amount of heat needed to change temperature of a substance depends on:
  - How much temperature will change
  - Quantity of substance
  - Nature of substance

Heat Capacity

- Physical property of a substance
- **Molar heat capacity**: describes amount of heat needed to increase temperature of <u>one mole</u> of a substance by one degree (J/mol·°C or J/mol·K)
- **Specific heat capacity**: same, but for <u>one gram</u> instead of <u>one mole</u> (J/g·°C or J/g·K)

Specific Heat Capacity

- Unique for each substance
- Higher for compounds than elements, due to more complex structures
- Describes to how easy it is to heat up a substance, and how long it takes to cool
- Think: why do Toronto and Vancouver have warmer winters than Regina?

Specific Heat Capacity

$$q = mc\Delta T$$

q = heat energy transferred (-) /absorbed (+)

m = mass of substance (g)

c = specific heat capacity of substance (J/g·°C or J/g ·K)

ΔT = temperature change of substance, $T_{final} - T_{initial}$ (°C or K)

Example 1: Specific Heat

It takes 132 J of energy to heat a sample of iron from 12.3°C to 33.9°C. What is the **mass** of the iron sample?

Example 2: Specific Heat

152.4 mL of ethanol is cooled from 60.0°C to 43.7°C. The density of ethanol is 0.789 g/mL. How much **energy** is transferred?

Example 3: Specific Heat

A 3.50 g sample of pure gold releases 33.1 J of heat. Its initial temperature was 78.2°C. What is the final temperature of the sample?

Example 4: Specific Heat

A 170.8 g sample of iron is heated to 102.1°C. The sample of iron is then submerged in a calorimeter full of water at 10.6°C. The final temperature of both substances is 24.2°C. What is the volume of water used?

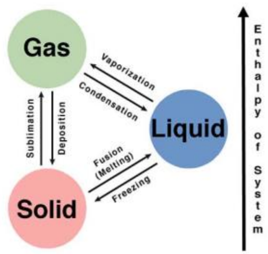
Heat Transfer in Systems

Endothermic/Exothermic

- **Endothermic**: heat is absorbed by the system
- **Exothermic**: heat is released by the system
- Can occur during physical changes <u>and</u> chemical changes

Phase Changes

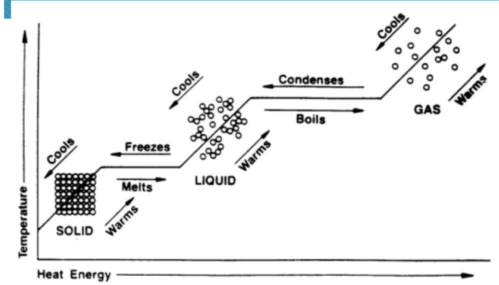
Phase Changes



Energy in Phase Changes

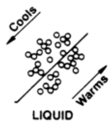
- Physical change – no new substance produced
- Heat is absorbed or released during the process, depending on the energy change in the system
- Endothermic processes: sublimation, fusion, evaporation (to higher energy phase)
- Exothermic processes: freezing, deposition, condensation (to lower energy phase)

Heating and Cooling Curves



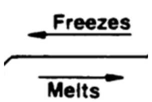
Phase Change Process

- Between phase changes:
  - All energy entering/leaving system is changing molecular motion
  - One state (solid, liquid, gas)
  - Temperature is changing

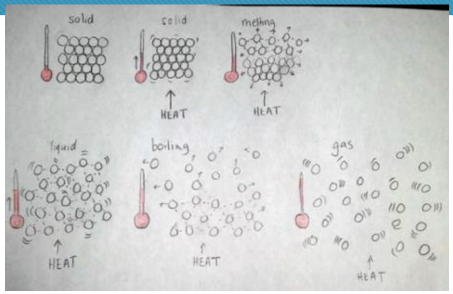


Phase Change Process

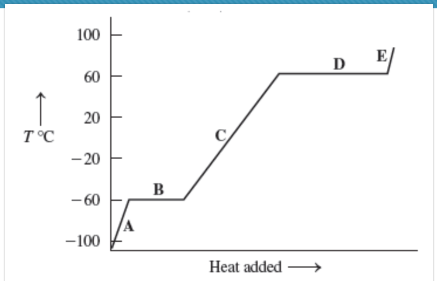
- During phase changes (plateaus):
  - All energy entering/leaving system is breaking/forming bonds between molecules (**latent heat**)
  - Two states at the same time
  - Temperature is constant until phase change is complete



Phase Changes



Example: Heating Curve



Example: Heating Curve Con'd

1. When is kinetic energy changing? When is potential energy changing?
2. What state is the substance at C?
3. What is happening at D?
4. What is the melting point of the substance?
5. What is the boiling point of the substance?
6. Does it take more energy to melt or boil the substance?

Latent Heat

Latent Heat

- **Latent heat**: energy is put into or removed from a system but no temperature change occurs
- All energy used to break or form bonds
- Different value for fusion (melting) and vaporization

Latent Heat

$$q = mL$$

q is heat in J

L_v (L_f) is heat of vaporization (fusion) in J/g

m is mass in g

Example 1: Latent Heat

How much energy does it take to melt 100 g of ice into water at 0.0°C? (L_f = 333.7 J/g)

Example 2: Latent Heat

How much energy does it take to melt 10.0 g of snow at 0.0°C and heat it to boiling? (L_f = 333.7 J/g)

Example 3: Latent Heat

You want to thaw a 1.0 kg chicken that is -2.8°C to room temperature (20°C).

Melting point = -2.8°C L_f = 247 J/g

c = 3.32 J/g·°C

a. How much energy does this take?
b. In a 700W microwave, how long would this take on full power? (P = E/t)

Enthalpy

Enthalpy

- Total energy content of a system
- Unmeasurable; but more relevant are <u>changes</u> in enthalpy, which is measurable
- Change in heat of a system is enthalpy of reaction, ΔH
  - Δ is Greek letter Delta, which means "change"

Enthalpy of Reaction

- ΔH: difference between energy of reactants and of products
- Shows if energy has transferred into or out of the system
- Energy is absorbed to break bonds/released when new bonds form
  - Energy absorbed > energy released = endothermic
  - Energy released > energy absorbed = exothermic

Thermochemical Equations

- **Thermochemical equations**: chemical equation that include:
  - All reactants and products
  - States of all reactants and products
  - Energy change of the reaction (ΔH)

Thermochemical Equations

$4\,Fe\,(s) + 3\,O_2\,(g) \rightarrow 2\,Fe_2O_3\,(s)$ ΔH = -1625 kJ

a. What type of reaction is this?
b. Write the mole ratios.
c. Is this endothermic or exothermic?

Thermochemical Equations

$2\ H_2S\ (g) + 3\ O_2\ (g) \rightarrow 2\ SO_2\ (g) + 2\ H_2O\ (g) \quad \Delta H = -1040\ kJ$

- Enthalpy is based on coefficients in the reaction
- Above, based on 2 mol H_2S, 3 mol O_2, etc…
- Questions:
  - Is this endothermic or exothermic?
  - How much energy is released/absorbed if 6 mol of O_2 reacts?

Example 1: TC Equations

Solid sodium bicarbonate reacts with hydrochloric acid (HCl) to make a sodium chloride solution, water and carbon dioxide gas. The reaction absorbed 11.8 kJ of heat for each mole of sodium bicarbonate.

Write the thermochemical equation for the reaction. (Endothermic or exothermic?)

Example 2: TC Equations

Gasoline contains ethanol (C_2H_5OH) which completely combusts to produce 1235 kJ of heat per mole of ethanol.

a. Write the thermochemical equation. (Endothermic or exothermic?)
b. How much heat is released when 50.00 g of ethanol completely combusts?

Enthalpy of Solution

Enthalpy of Solution

- Also called **heat of solution**
- Solid dissolving in water absorbs or releases energy
- Experimentally determined by finding q (using $q = mc\Delta T$) and dividing by number of moles of reactant:

$$\Delta H = \frac{q}{n}$$

Example 1: Heat of Solution

Magnesium sulfate can be dissolved in water to make a hot pack.

$MgSO_4\ (s) \rightarrow Mg^{2+}\ (aq) + SO_4^{2-}\ (aq) \quad \Delta H = -91.4\ kJ$

If 5.00 g of $MgSO_4$ is dissolved in 100.0 g of water at 25.0°C:

a. How much energy is released?
b. What temperature is the final solution? (Assume c = 4.184 J/g·°C)

Example 2: Heat of Solution

You dissolve 5.0 g of sugar ($C_6H_{12}O_6$) in a 250.0 mL of coffee at 92.00°C and the temperature drops to 91.72°C.

a. What assumptions can you make?
b. How much energy is transferred between the system and surroundings?
c. What is the experimental heat of solution in kJ/mol of sugar?
d. Is this endothermic or exothermic?