Chapter 5

Energy

First Law of Thermodynamics

Energy is neither created nor destroyed, it can change forms.

$\Delta E_{system} + \Delta E_{surroundings} = 0$

System= what you are studying: reactions, ice melting, coffee cooling

Surroundings= everything else

System can be open, closed or isolated

Open

Closed

Isolated

$\Delta E = q - P\Delta V$ (heat transferred + work)

$w = -P\Delta V$

$\Delta E = q + w$

When ΔV is positive (expansion) the system is doing work on the surroundings. = −w

When the ΔV is negative (compression), the surroundings did work on the system. = +w

When heat is transferred out of the system to the surroundings, −q

When heat is added to the system from the surroundings, = +q

1. When coffee (system) cools, q for coffee is

A. Positive B. Negative C. Zero

2. When Ice melts(system) in a drink (surroundings), the q for the ice is

A. Positive B. Negative C. Zero

$= q - P\Delta V$ (heat transferred + work)

$-P\Delta V$

$= q + w$

n ΔV is positive (expansion) the system is doing work on the oundings. = −w

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n heat is transferred out of the system to the surroundings, −q

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1. When coffee (system) cools, q for coffee is

A. Positive B. Negative C. Zero

2. When Ice melts(system) in a drink (surroundings), the q for the ice is

A. Positive B. Negative C. Zero

What is the change in internal energy (ΔE) when a system is heated with 35 J of energy while it does 20 J of work?

$\Delta E = q - P\Delta V$ (heat transferred + work)

$w = -P\Delta V$

$\Delta E = q + w$

$\Delta E = 35 + (-20)$

$\Delta E = 15J$

What is the change in internal energy (ΔE) when a system is cooled by removing with 35 J of energy while it does 20 J of work?

a. -55J b. +55K c. –20J d. +15J e. –15J

$\Delta E = q - P\Delta V$ (heat transferred + work)

$w = -P\Delta V$

$\Delta E = q + w$

$\Delta E = -35 + (-20)$

$\Delta E = -55J$

$\Delta E = q + w$

According to the first law of thermodynamics, which of the changes, A–D, will **always** increase the internal energy (ΔE will be positive) of a system? (q = energy transferred, and w = work done)

a. $q<0, w<0$

b. $q<0, w>0$

c. $q>0, w>0$

d. $q>0, w<0$

e. None of these

Phase changes

- Which of these processes are endothermic and which of these are exothermic?

  - Fusion: Liquid —> Solid
  - Sublimation: Solid —> Gas
  - Melting: Solid —> Liquid
  - Vaporization: Liquid —> Gas
  - Deposition: Gas—> Solid
  - Condensation: Gas —> Liquid

Phase changes

- Endothermic processes (require heat energy)
    - Melting: Solid —> Liquid +q
    - Vaporization: Liquid —> Gas +q
    - Sublimation: Solid —> Gas +q
- Exothermic processes (release heat)
    - Deposition: Gas—> Solid -q
    - Condensation: Gas —> Liquid -q
    - Fusion: Liquid —> Solid -q

System = Boiling water

Boiling water is:

A. Endothermic

B. Exothermic

C. Neither

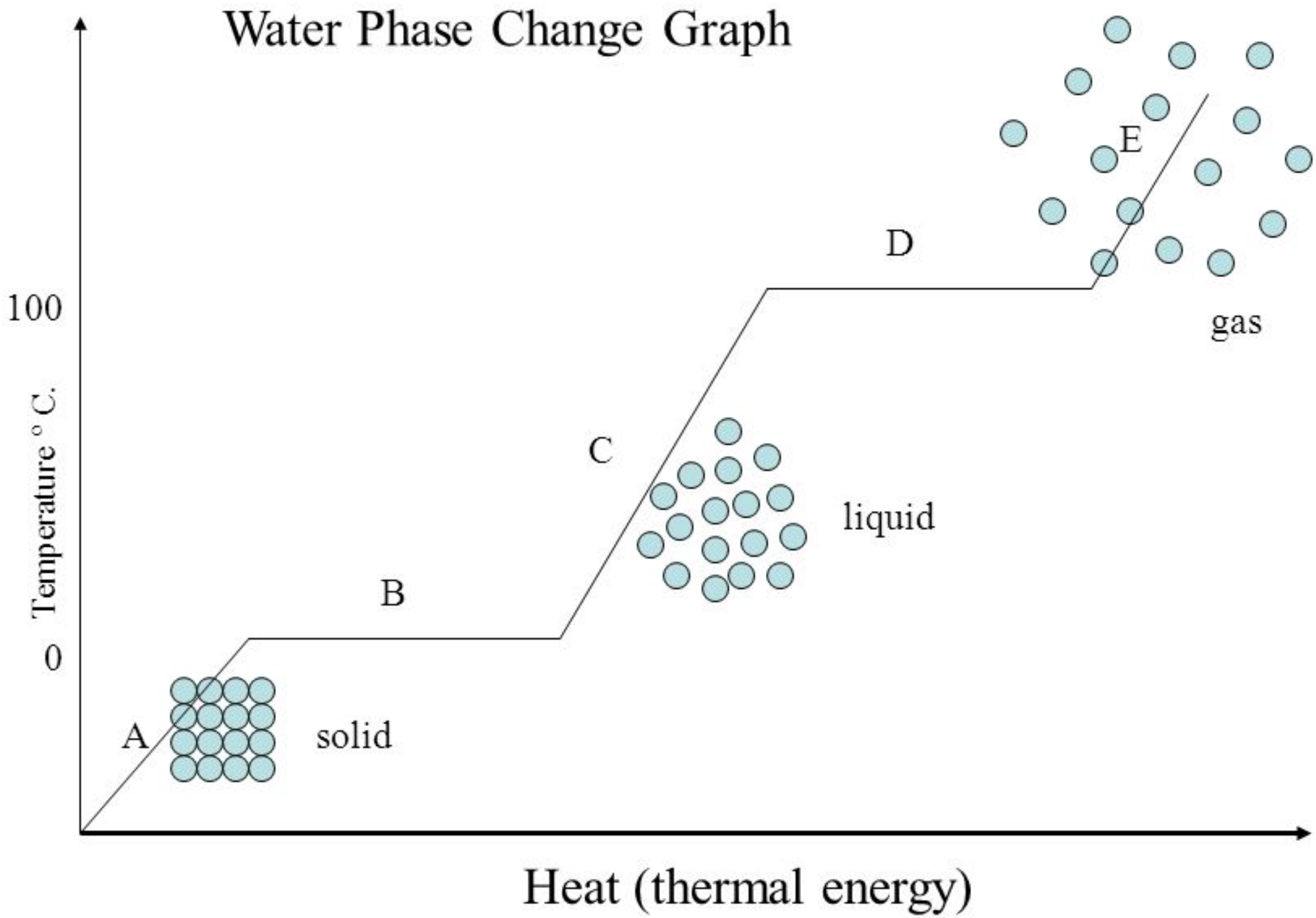
D. $\Delta E_{system} = 0$

E. Can not be determined

The system = pot of water

In order for the pot of water to boil, q for the system must be

A. Positive
B. Negative
C. Zero
D. It is impossible to determine


Water Phase Change Graph
Temperature ° C.
100
0
A
solid
B
C
liquid
D
E
gas
Heat (thermal energy)

Calculate separate q values, then add them up.

From ice (-20°C) to vapor (120°C):

Ice at -20°C to Ice at 0°C: $q = \text{mass} \times Cp_{ice} \times \Delta T$

Ice to liquid: $q = \text{mass} \times \Delta H_{fusion}$

Liquid at 0°C to liquid at 100°C: $q = \text{mass} \times Cp_{liquid} \times \Delta T$

Liquid to Gas: $q = \text{mass} \times \Delta H_{vaporization}$

Gas at 100°C to gas at 120°C: $q = \text{mass} \times Cp_{gas} \times \Delta T$

Boiling point	373 K
Melting point	273 K
Enthalpy of vaporization	2,260 J/g
Enthalpy of fusion	334 J/g
Specific heat capacity (solid)	2.11 J/(g · K)
Specific heat capacity (liquid)	4.18 J/(g · K)
Specific heat capacity (gas)	2.08 J/(g · K)

How much energy is required to change 100 grams of ice at -20°C to steam at 120°C? (Note ΔT in K is the same as the ΔT in °C.) $\Delta T = T_{final} - T_{initial}$

<u>From ice at -20°C to vapor at 120°C</u>

Ice at -20°C to Ice at 0°C: q = 100g x 2.11J/gK x 20K

Ice to liquid: q = 100 g x 334J/g

Liquid at 0°C to liquid at 100°C: q= 100g x 4.18 J/gKx 100K

Liquid to Gas: q = 100 g x 2,260J/g

Gas at 100°C to gas at 120°C: q = 100 x 2.08J/gK x 20K

Using the following data for water, determine how much energy is needed to change 100 g of ice at –10°C to steam at 225°C.

Boiling point	373 K
Melting point	273 K
Enthalpy of vaporization	2,260 J/g
Enthalpy of fusion	334 J/g
Specific heat capacity (solid)	2.11 J/(g · K)
Specific heat capacity (liquid)	4.18 J/(g · K)
Specific heat capacity (gas)	2.08 J/(g · K)

a. 3.29×10^2 kJ b. 98.2 kJ c. 48.8 kJ d. 203.3 kJ e. none of these

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