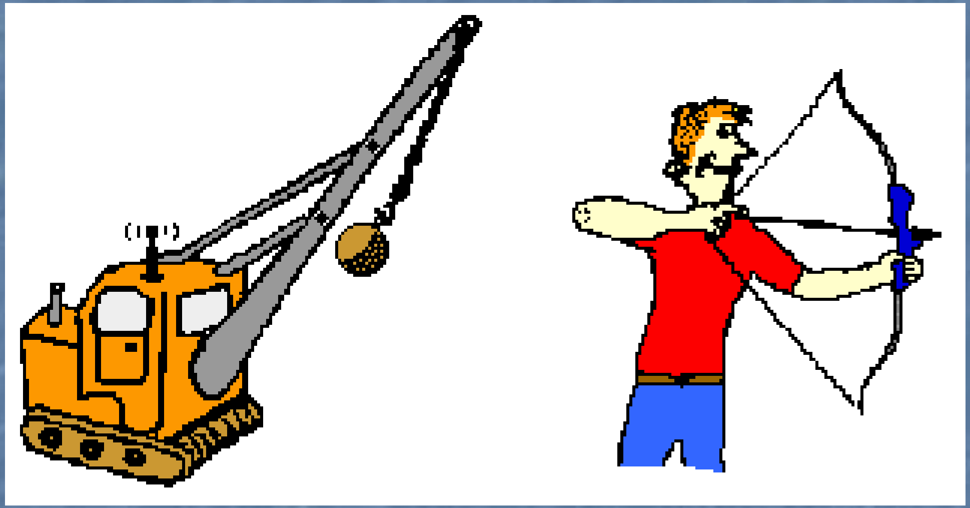
Temperature, Heat, Expansion

Objectives

- Distinguish aspects of heat flow (endothermic, exothermic, potential energy, kinetic energy, heat and temperature).
- Identify heat flow, potential and kinetic energy, phase changes, and heating or cooling for phase diagrams of a substance.
- Define temperature, heat flow direction, thermal expansion, and specific heat.
- Calculate / measure heat changes in a system using a calorimeter.
- Understand thermal expansion and its importance regarding water.

I. Energy

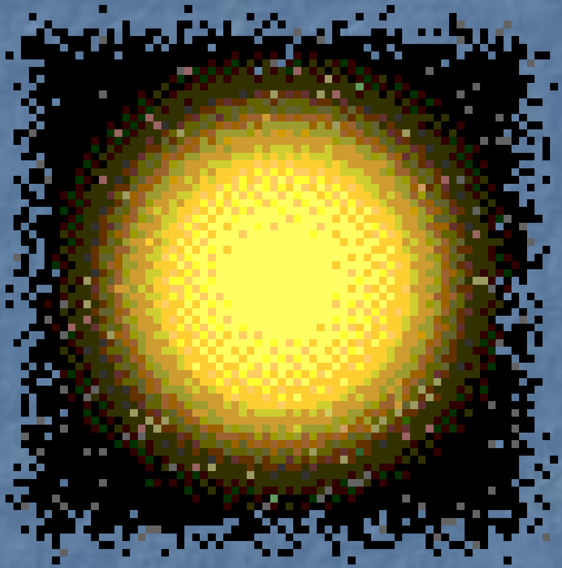
The capacity to do work ("joules" → j)



- A. Potential Energy (PE) → stored energy
- B. Kinetic Energy (KE) → energy of motion

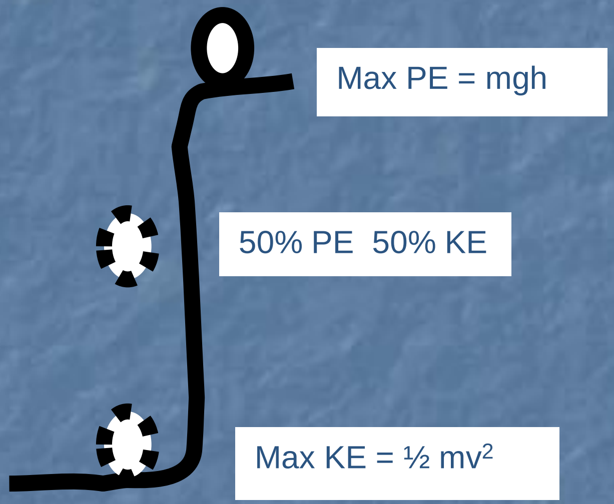
Potential Energy (PE) → stored energy

- 1. An objects energy just prior to falling (*position*)
- 2. Food, Heat
- 3. Fuel
- 4. Ammunition



Kinetic Energy (KE) → energy of motion

- 1. A falling object (MOTION)
- 2. **Temperature**
  - The average motion of molecules



Temperature

The measure of heat intensity: describes the average kinetic energy (KE) of the molecules in a system

- Units: measured in degrees
  - Fahrenheit (°F)
  - Celsius (°C); also called centigrade
  - **Kelvin (K)**
    - "Absolute temperature"
    - Begins with 0 K (*theoretical temperature*)
    - includes volume and motion

Heat

mass

A form of energy that takes into account the quantity of matter

- Units: calories (Kcal) or joules (KJ)
- Calorimeter → instrument used to measure heat
- Physical and chemical changes involve changes in energy:

  (1) Exothermic - releases energy or heat

  (2) Endothermic - gains/absorbs energy or heat

Temperature versus Heat

Compare the temperature and heat of a match and a bonfire.

Temperature versus Heat

Temperature involves the motion of molecules only

Heat incorporates **temperature** and mass (amount)

Example: *burning a match verses a bonfire*

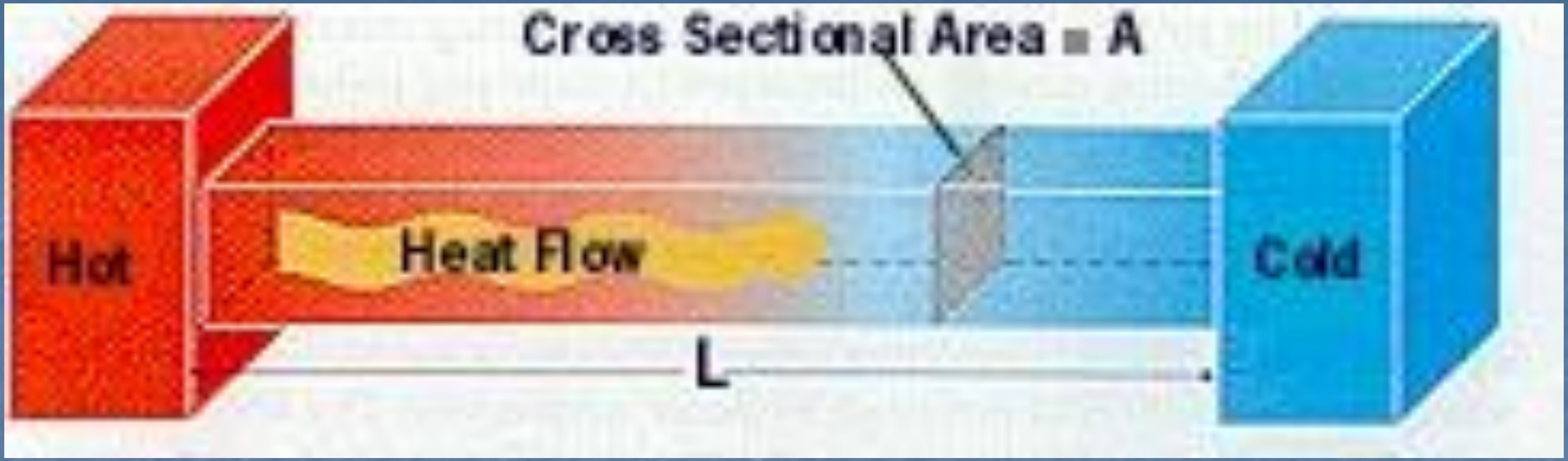
- *Temperature for both is the same (~400-600 °C)*
- *The bonfire contains MUCH more heat*

Temperature versus Heat

Temperature involves the average KE of molecules ONLY

Heat incorporates temperature and mass

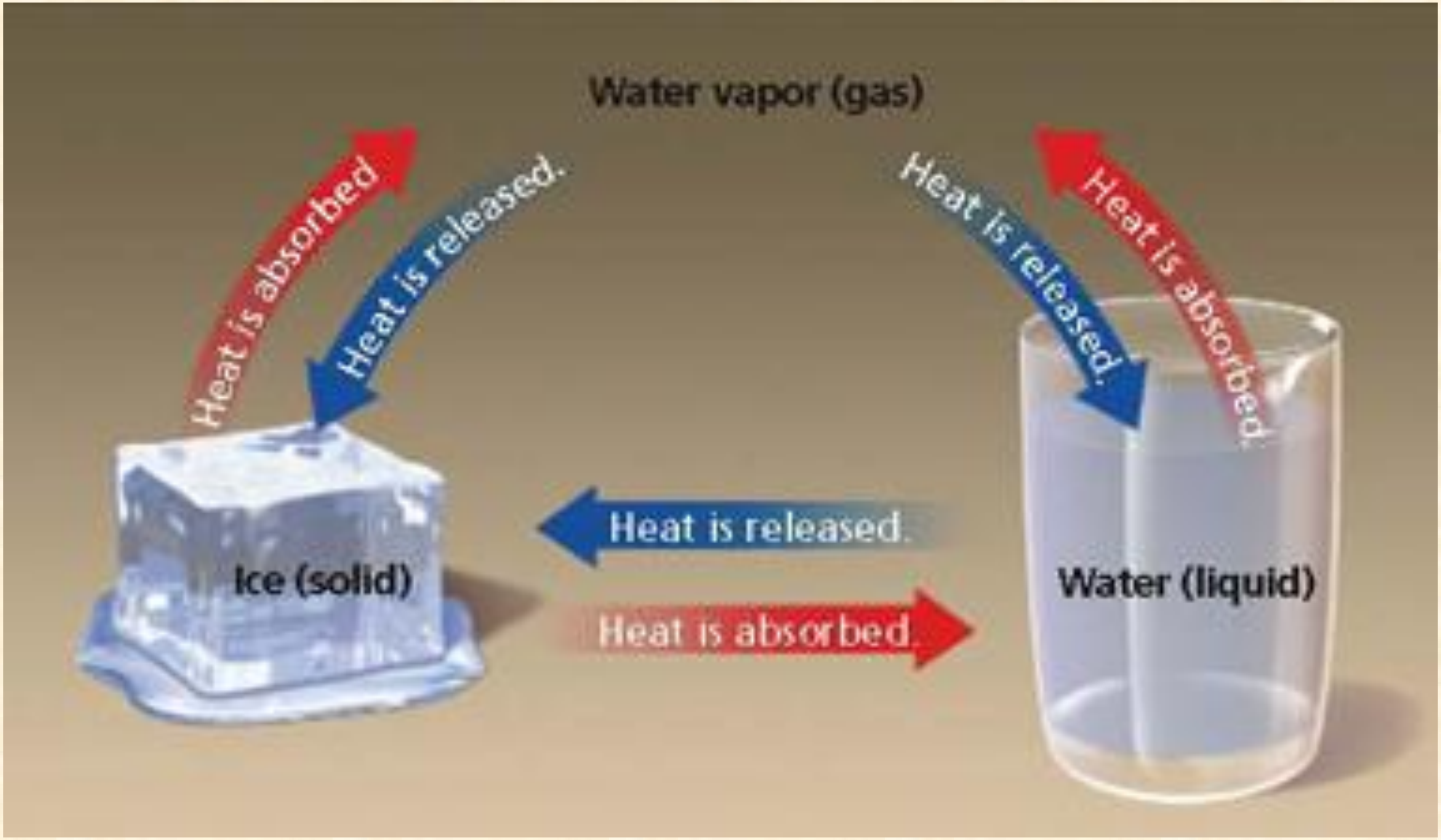
- example one: *burning one match verses five matches*
  - Temperature for both is the same (~400-600 °C)
  - Heat of one match = ~500 cal
  - Heat of five matches = ~2500 cal
- example two: *burning a match verses a bonfire or house*
  - The temperature of a bonfire or burning house is almost the same as a burning match (~400-600 °C)
  - The heat of a match is ~500 cal while
  - the heat of a bonfire or house is thousands of times greater



Heat flows from warm to cold

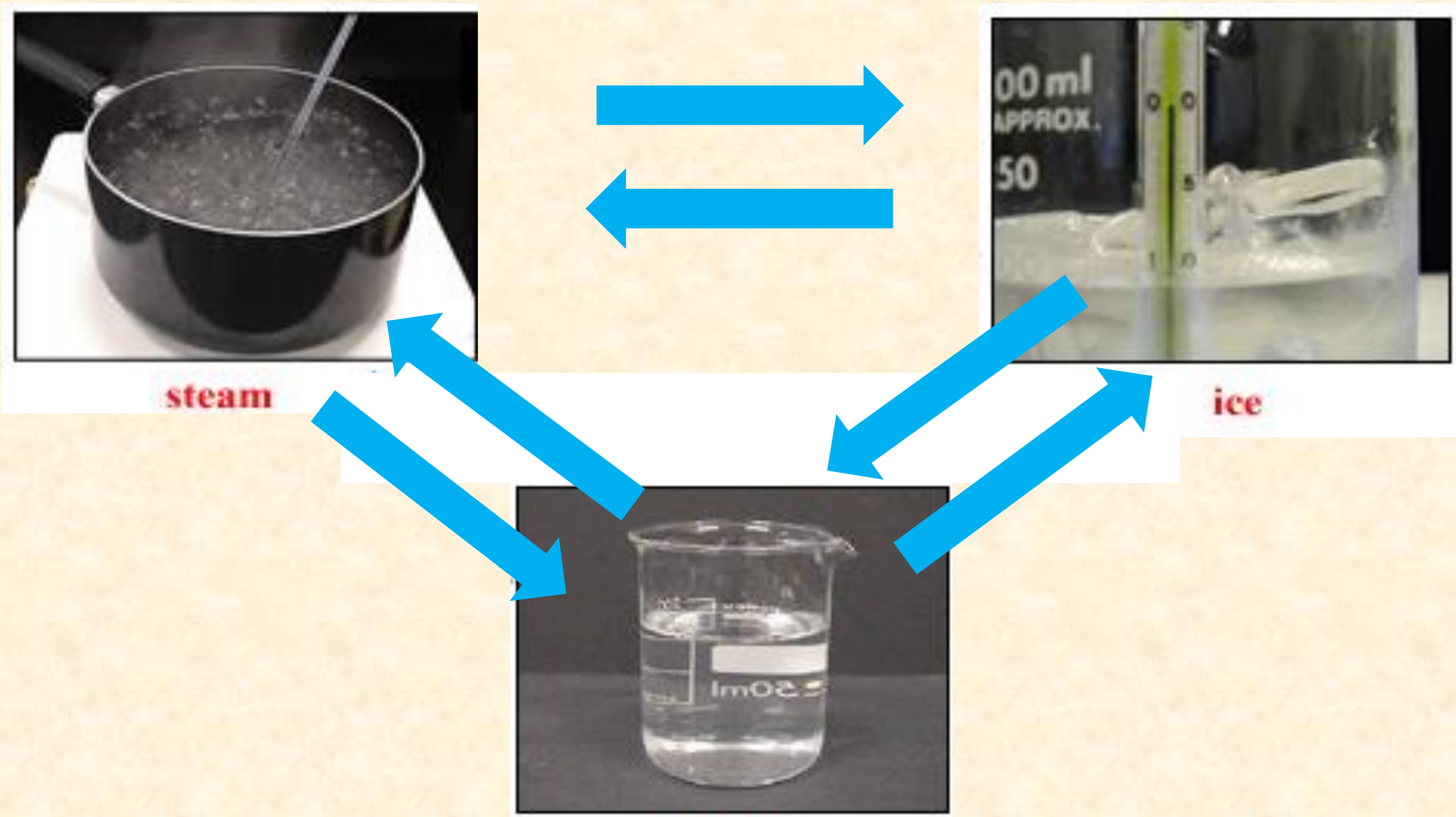
This represents KE

Energy Interactions



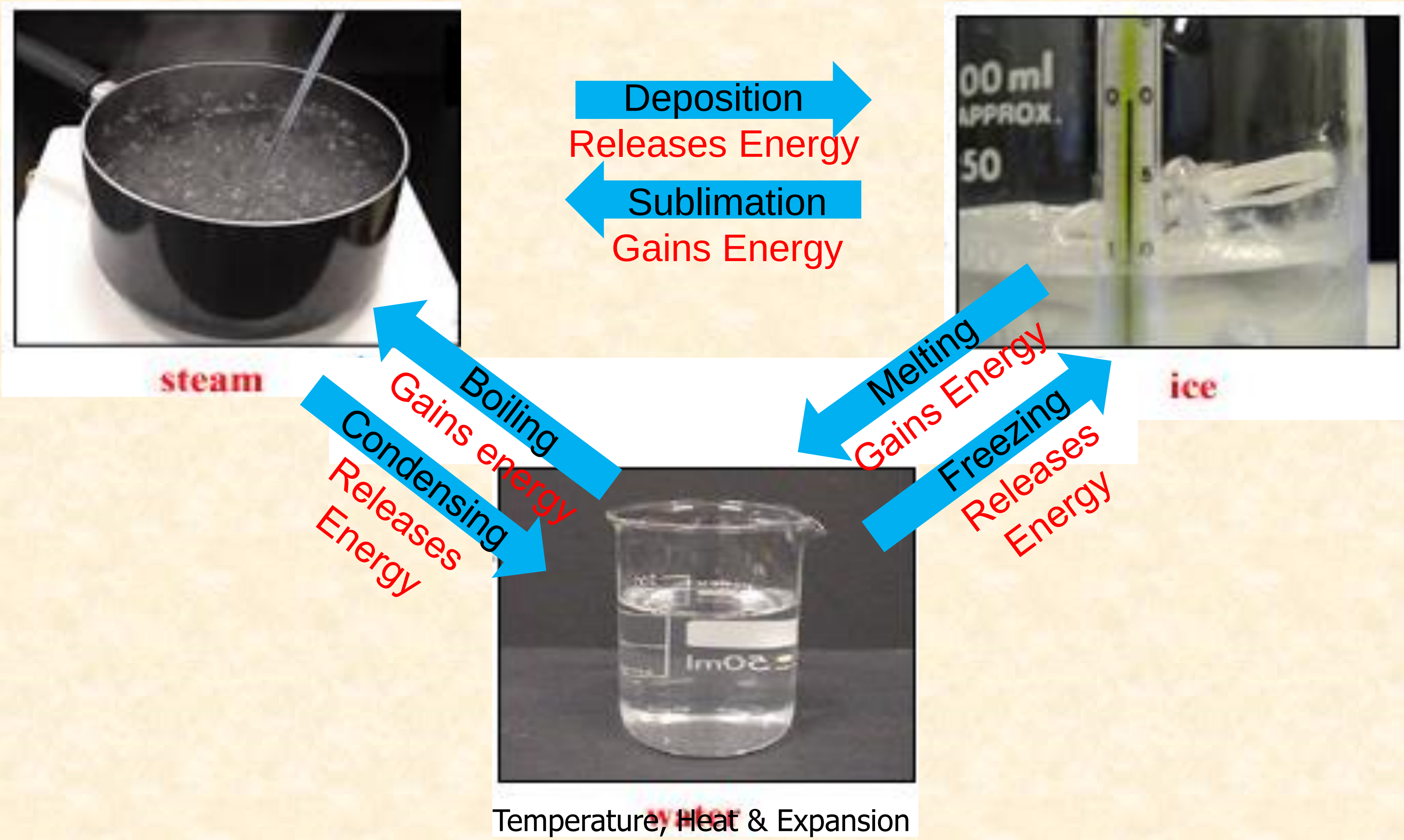
Energy Through the Water Cycle

State if energy released or gained for each phase change.



Energy Through the Water Cycle

Names of processes are Enrichment



Heating and Cooling Curves

http://somup.com/cFX6DGni0X (1:12)

Inquiry Questions:

- Why doesn't the temperature change as ice melts? As water boils?
- What kind of energy relationships are occurring as ice melts and then water boils?
- What do we call the stages when ice melts and water boils?

Heating and Cooling Curves

Inquiry Questions:

- Why doesn't the temperature change as ice melts? As water boils? (*PE vs. KE*)
- What kind of energy relationships are occurring as ice melts and then boils? (*Endothermic*)
- What do we call the stages when ice melts (*melting point*) and water boils (*boiling point*)?

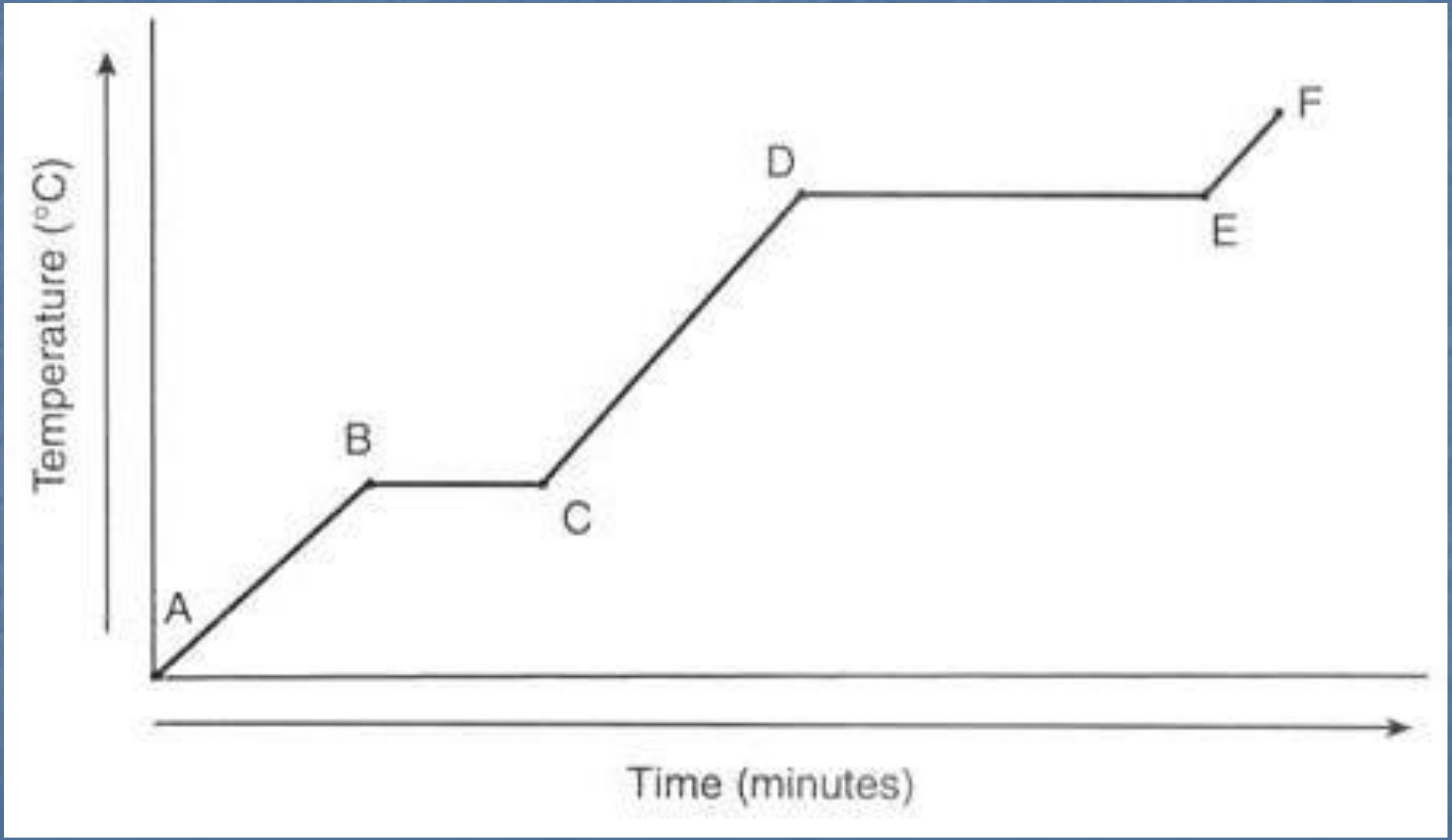
Phase Change Diagrams

- Endothermic $\Delta H = +$
  - Heating Curve
  - Heat flows **into** the system from the surroundings

- Exothermic $\Delta H = -$
  - Cooling Curve
  - Heat flows **from** the system into the surroundings

Phase Change Diagrams

Endothermic "Heating Curve"



Phase Change Diagrams

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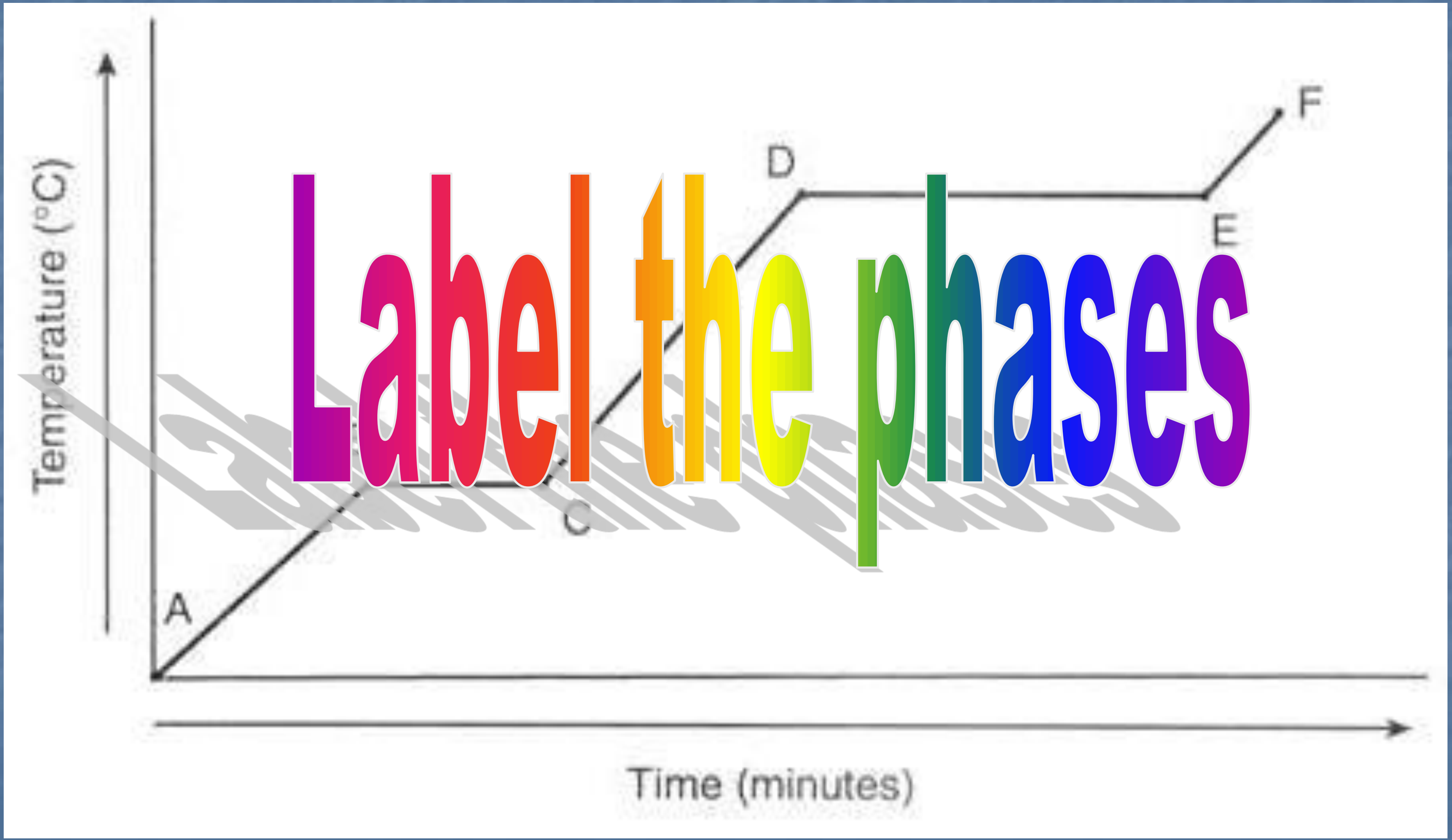
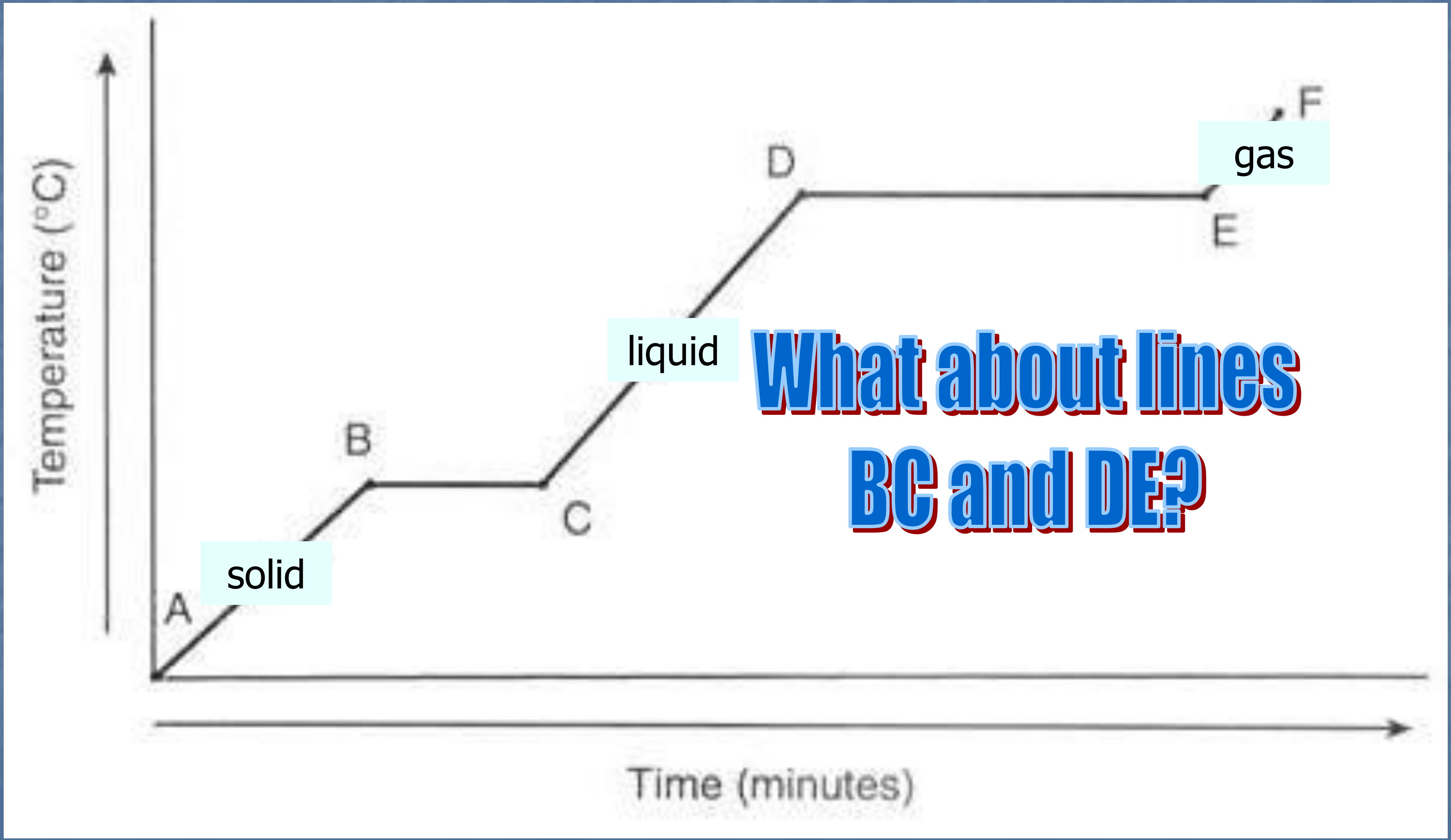


Phase Change Diagrams

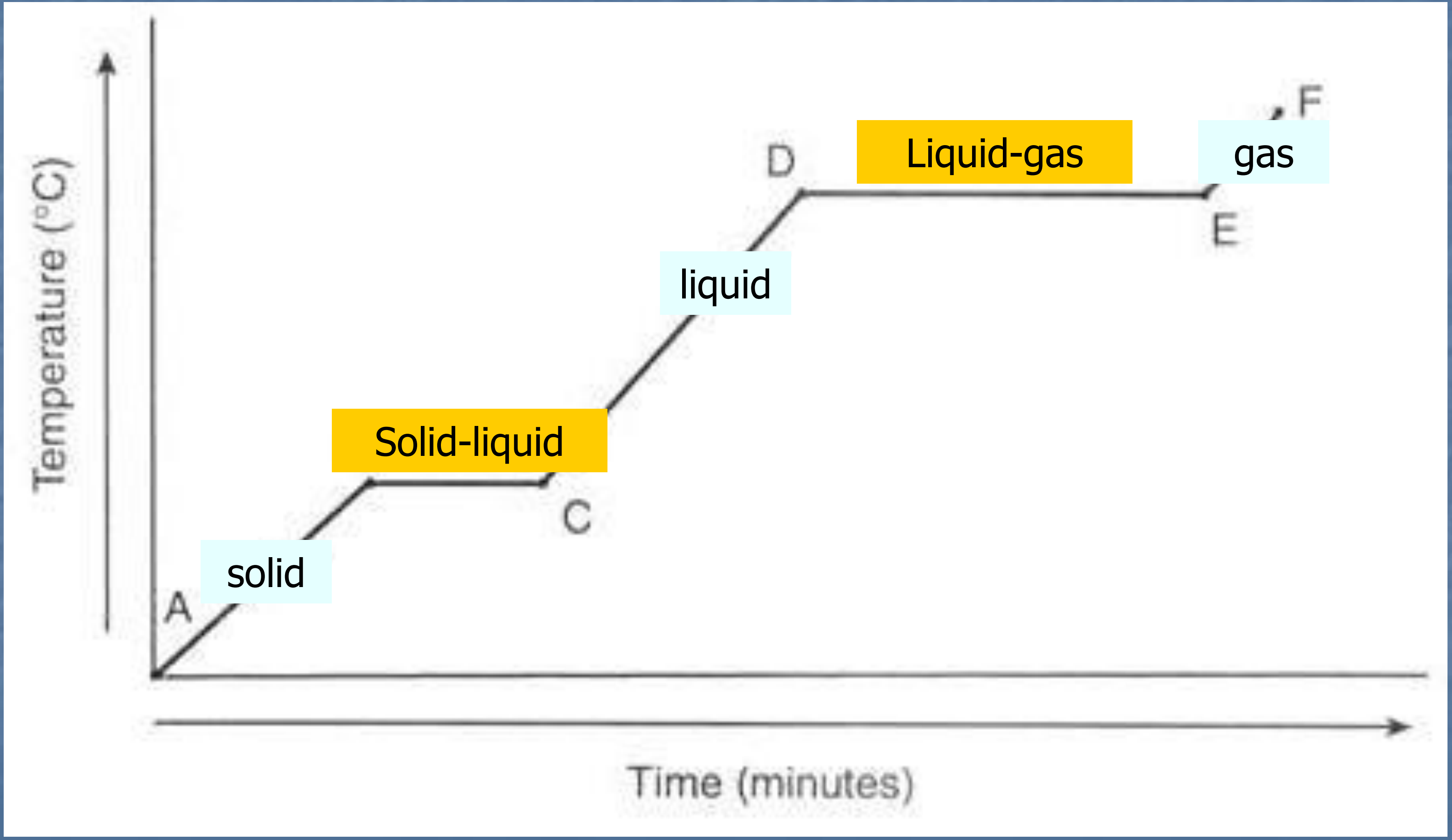
Endothermic "Heating Curve"



What about lines BC and DE?

Phase Change Diagrams

Endothermic "Heating Curve"



Phase Change Diagrams

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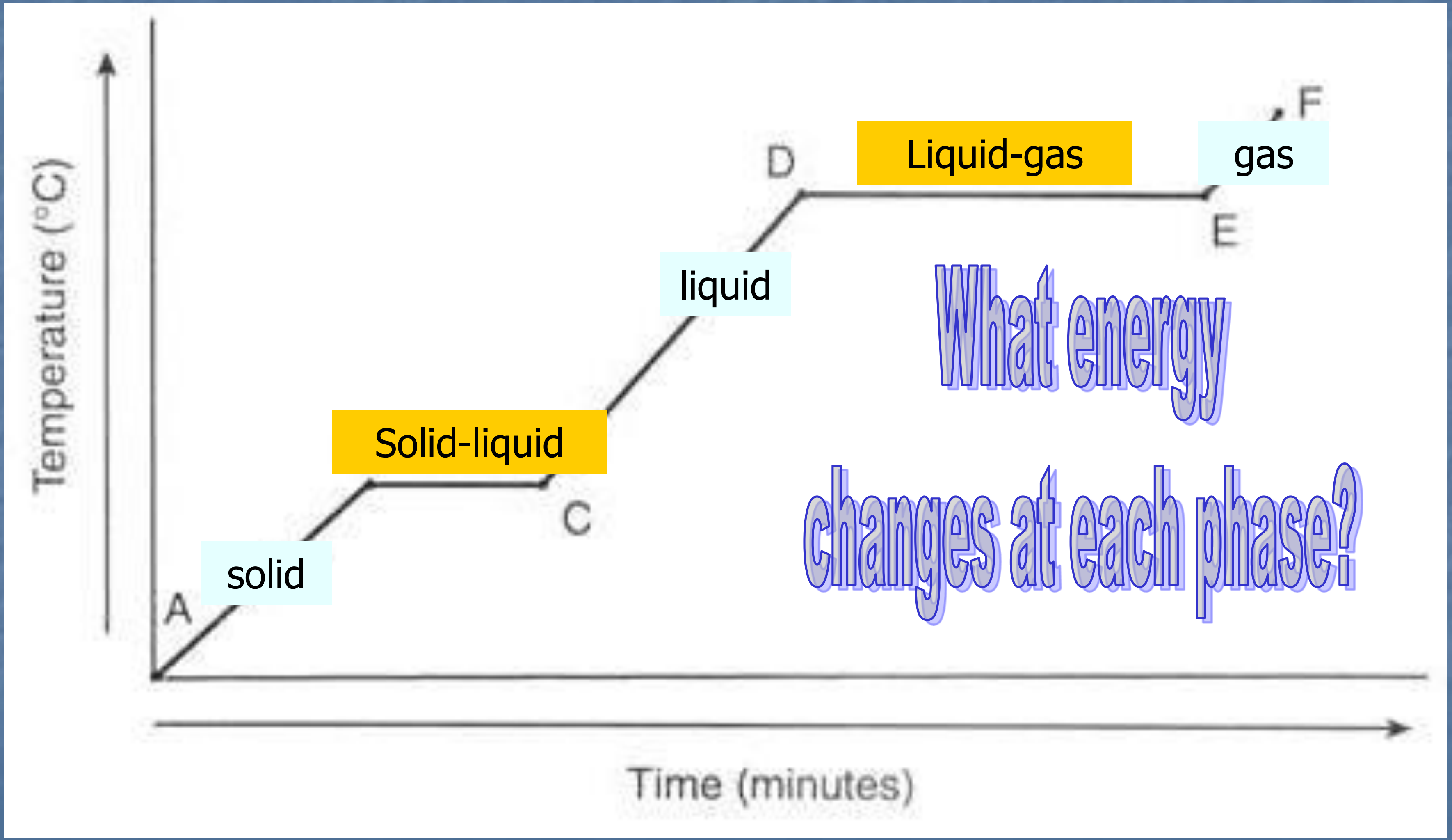
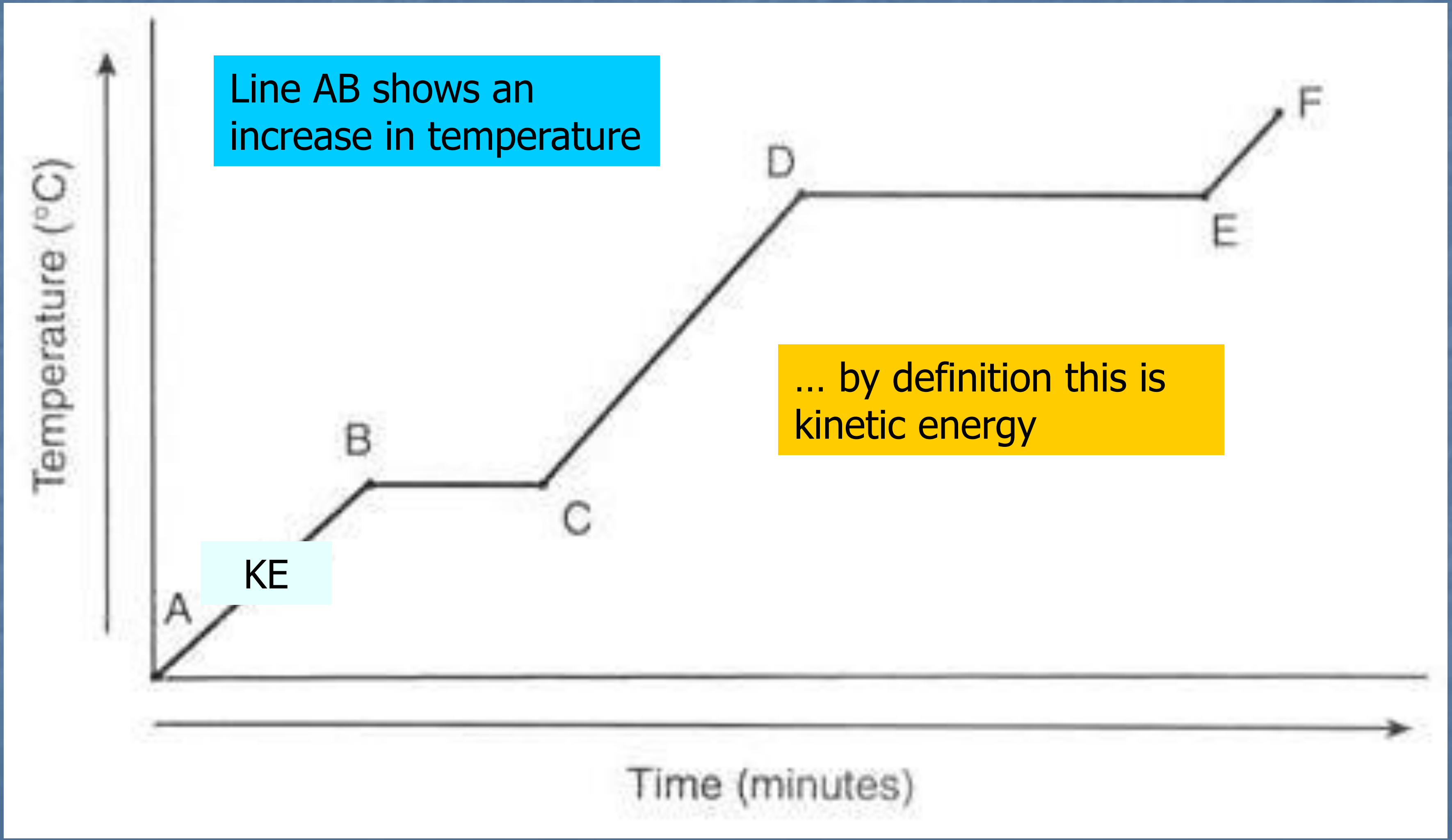


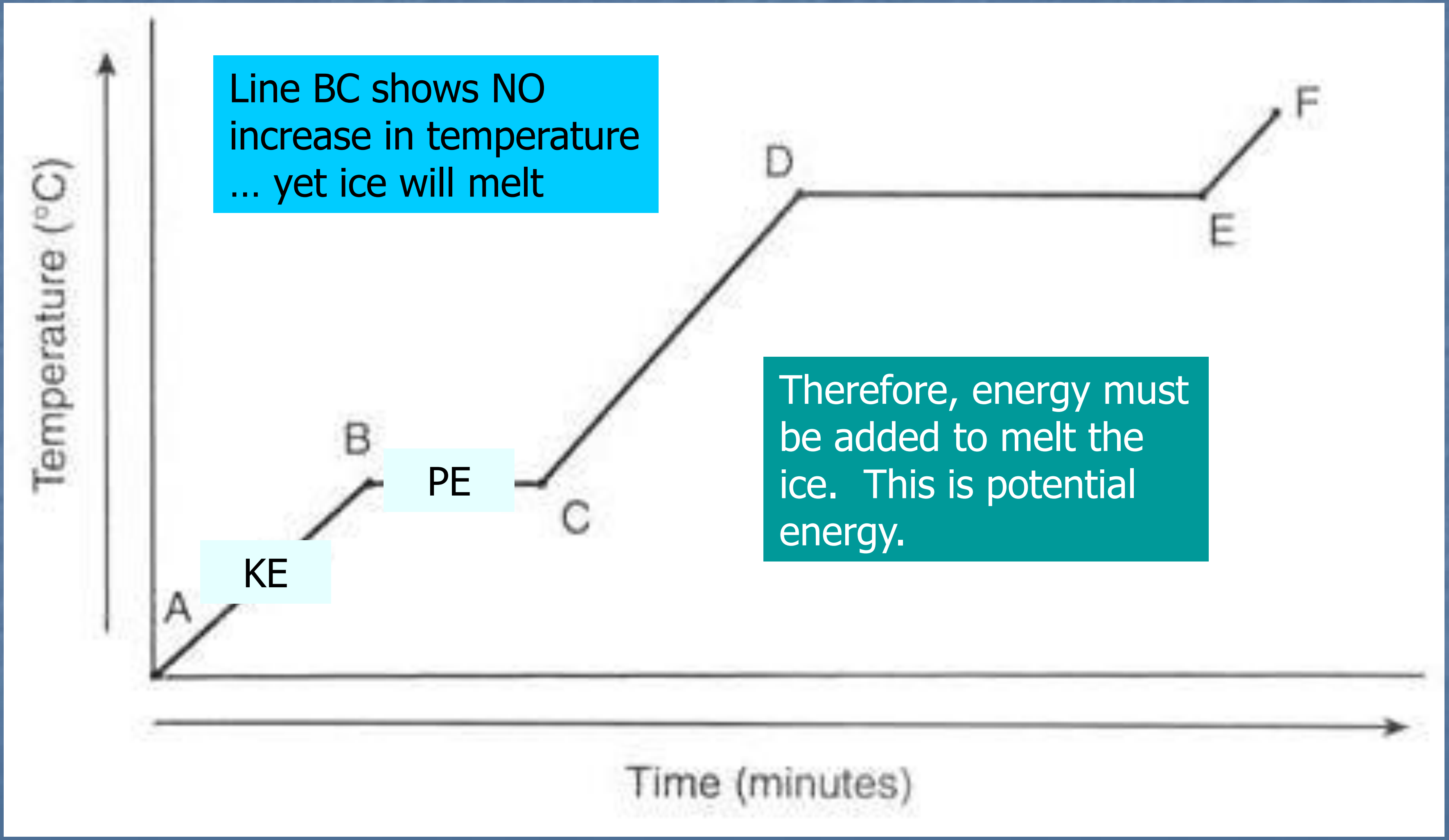
Phase Change Diagrams

Endothermic "Heating Curve"



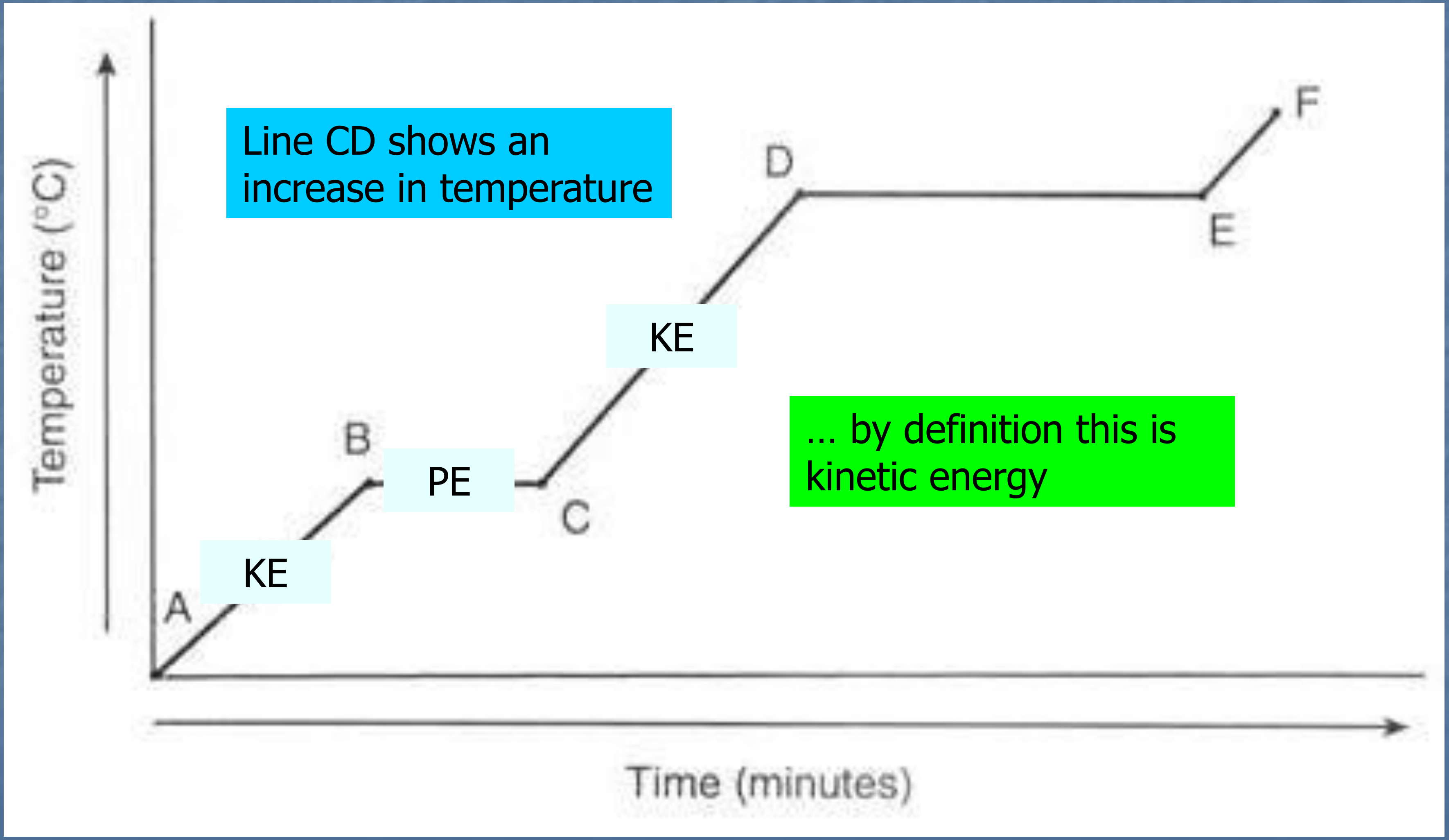
Phase Change Diagrams

Endothermic "Heating Curve"



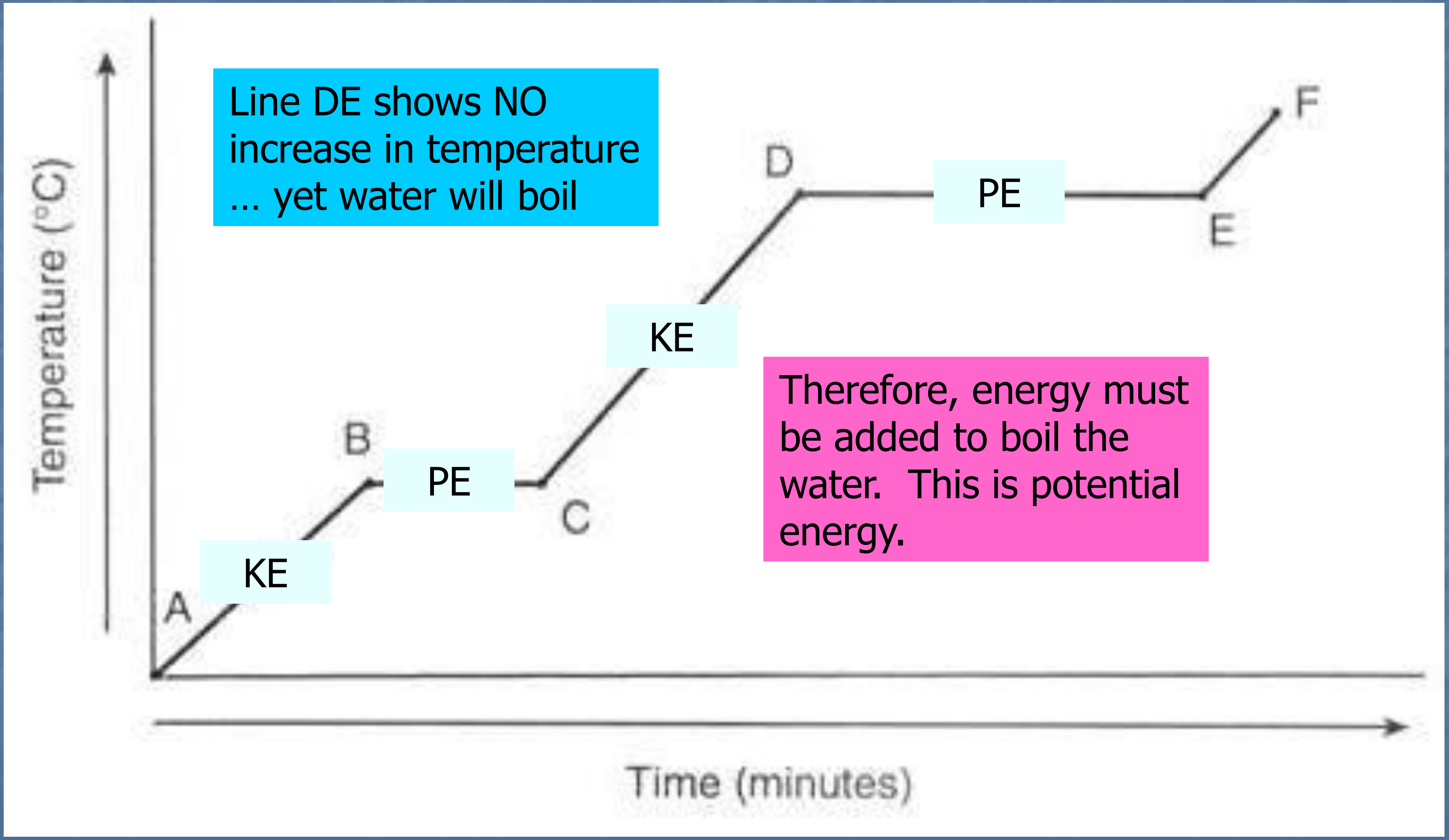
Phase Change Diagrams

Endothermic "Heating Curve"



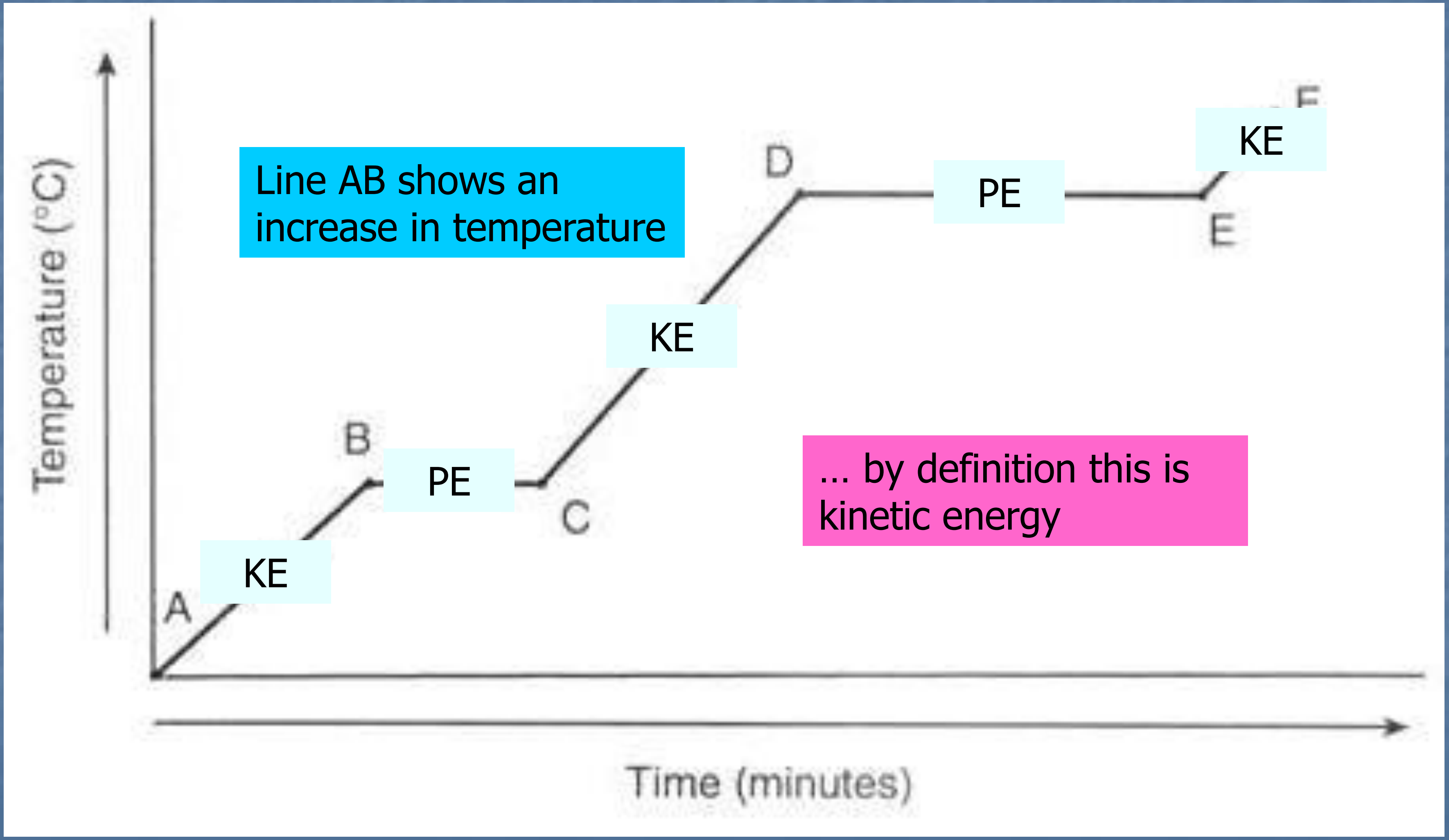
Phase Change Diagrams

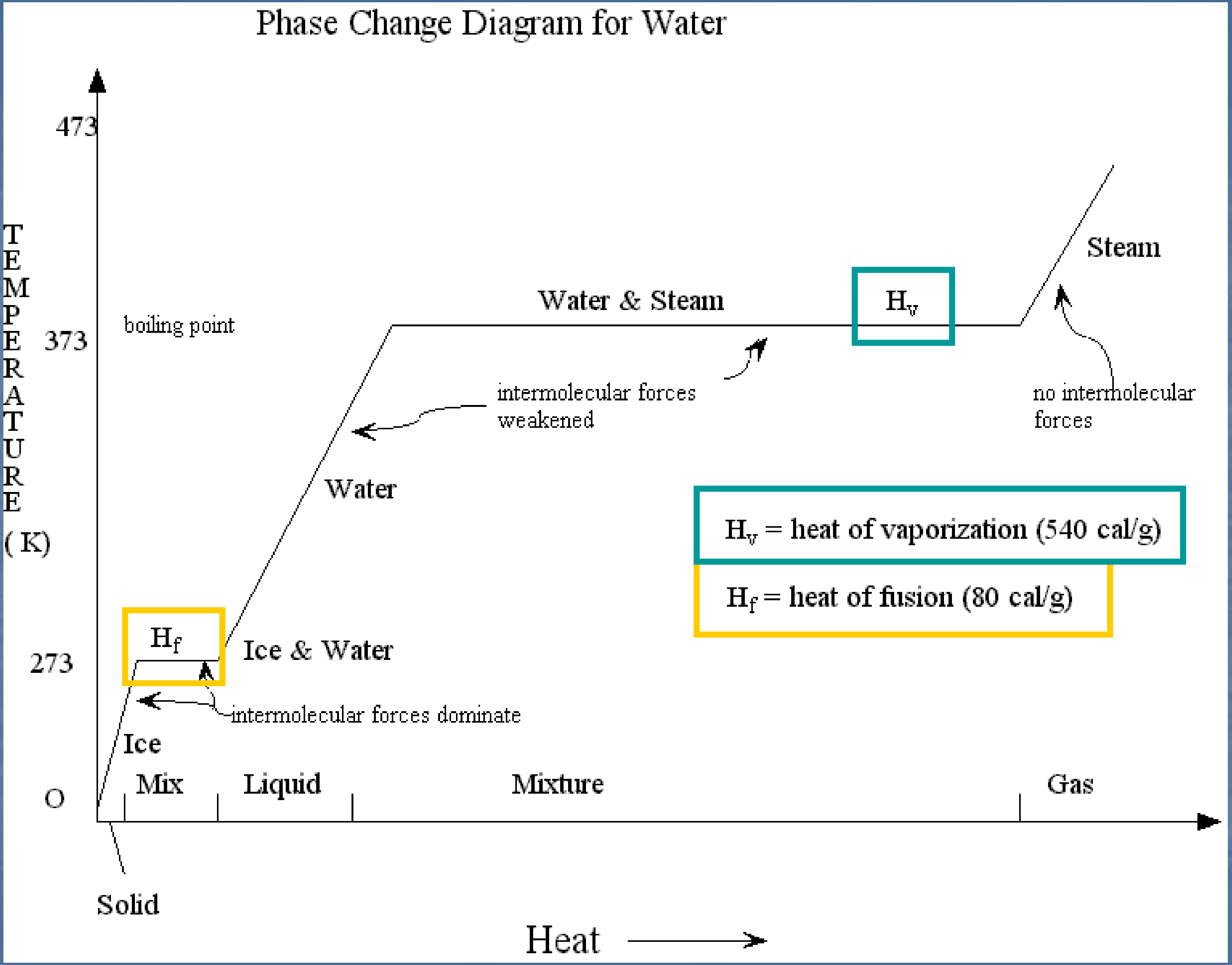
Endothermic "Heating Curve"



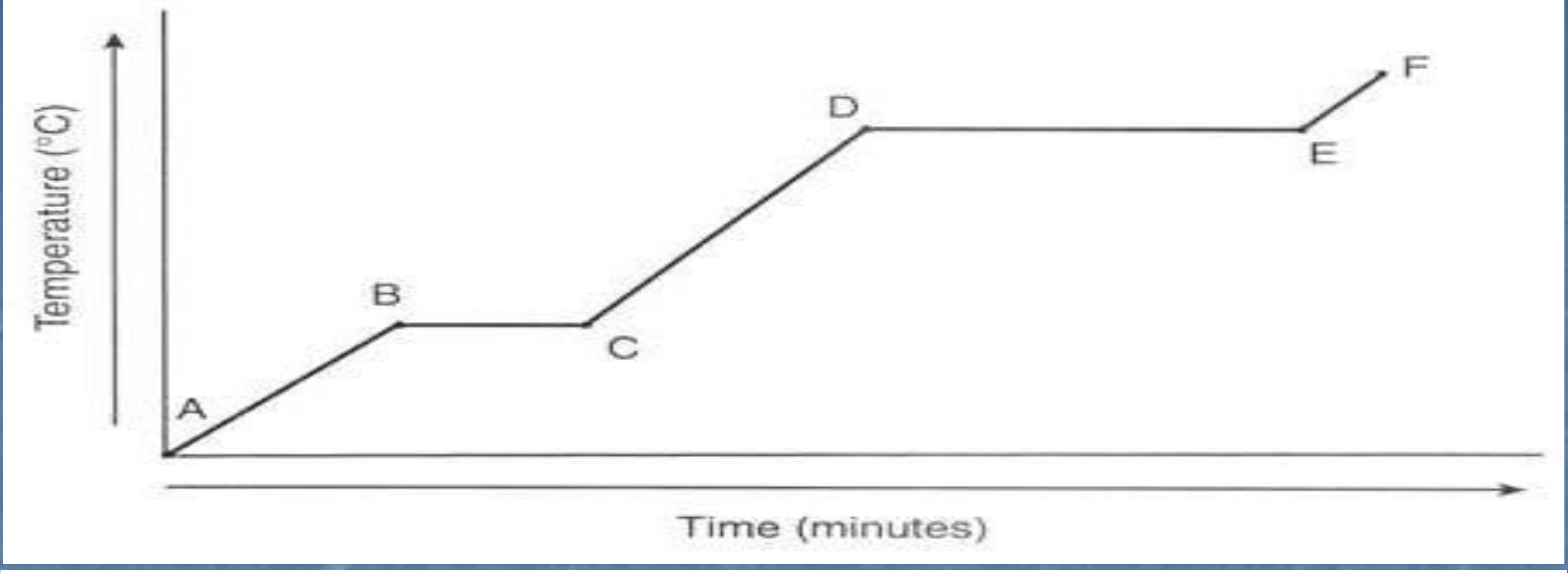
Phase Change Diagrams

Endothermic "Heating Curve"

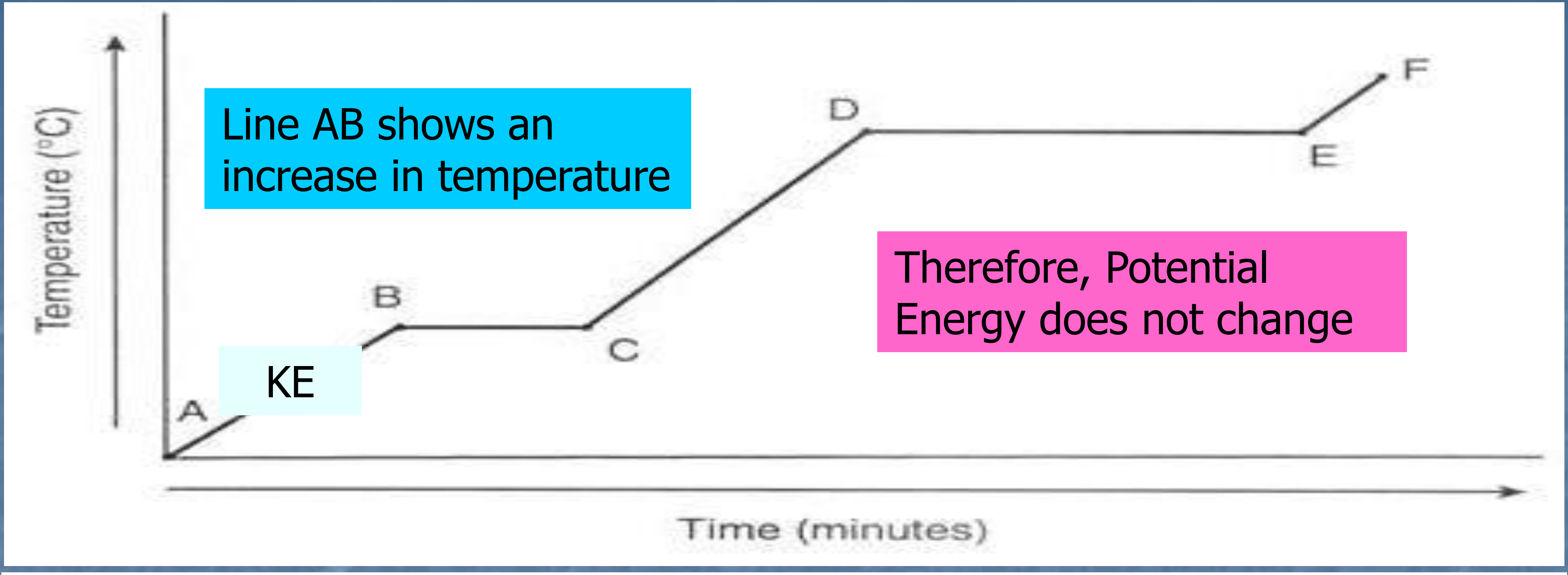




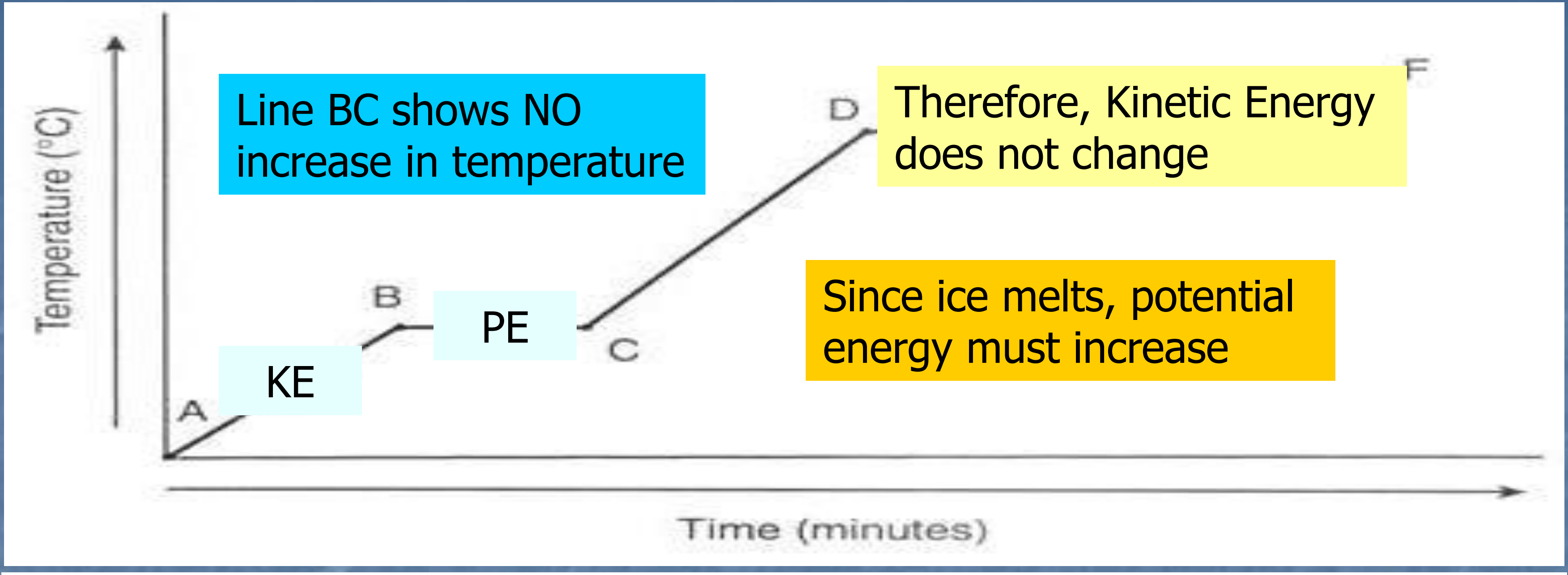
Phase Change Diagram for Water
TEMPERATURE (K)
473
373
273
O
boiling point
Water & Steam
H_v
Steam
intermolecular forces weakened
no intermolecular forces
Water
H_v = heat of vaporization (540 cal/g)
H_f = heat of fusion (80 cal/g)
H_f
Ice & Water
intermolecular forces dominate
Ice
Mix
Liquid
Mixture
Gas
Solid
Heat



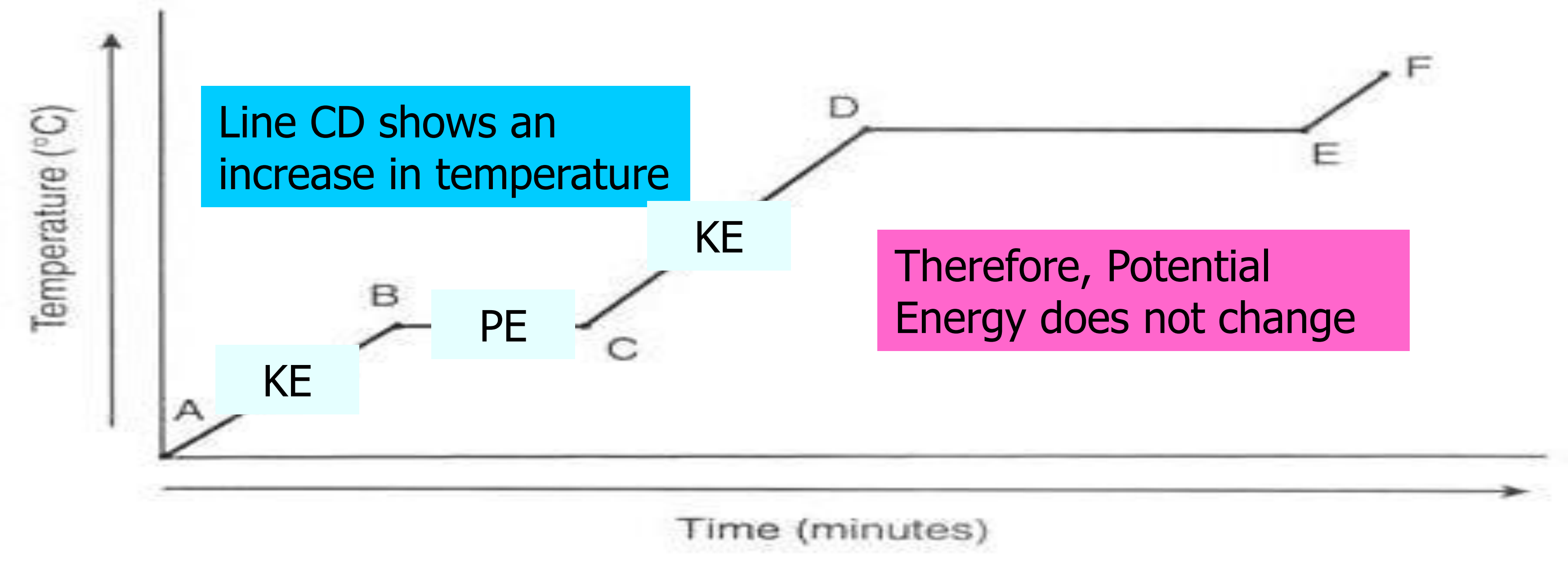
	Graph	Temp	Phase	KE	PE
	A→B	-20 → 0	s		
	B→C	0	s → l		
	C→D	0 → 100	l		
	D→E	100	l → g		
	E→F	100 → 110	g		



	Graph	Temp	Phase	KE	PE
	A→B	-20 → 0	s	**increase**	**constant**
	B→C	0	s → l		
	C→D	0 → 100	l		
	D→E	100	l → g		
	E→F	100 → 110	g		



	Graph	Temp	Phase	KE	PE
	A → B	-20 → 0	s	**increase**	**constant**
Melting pt	B → C	0	s → l	**constant**	**increase**
	C → D	0 → 100	l		
	D → E	100	l → g		
	E → F	100 → 110	g		



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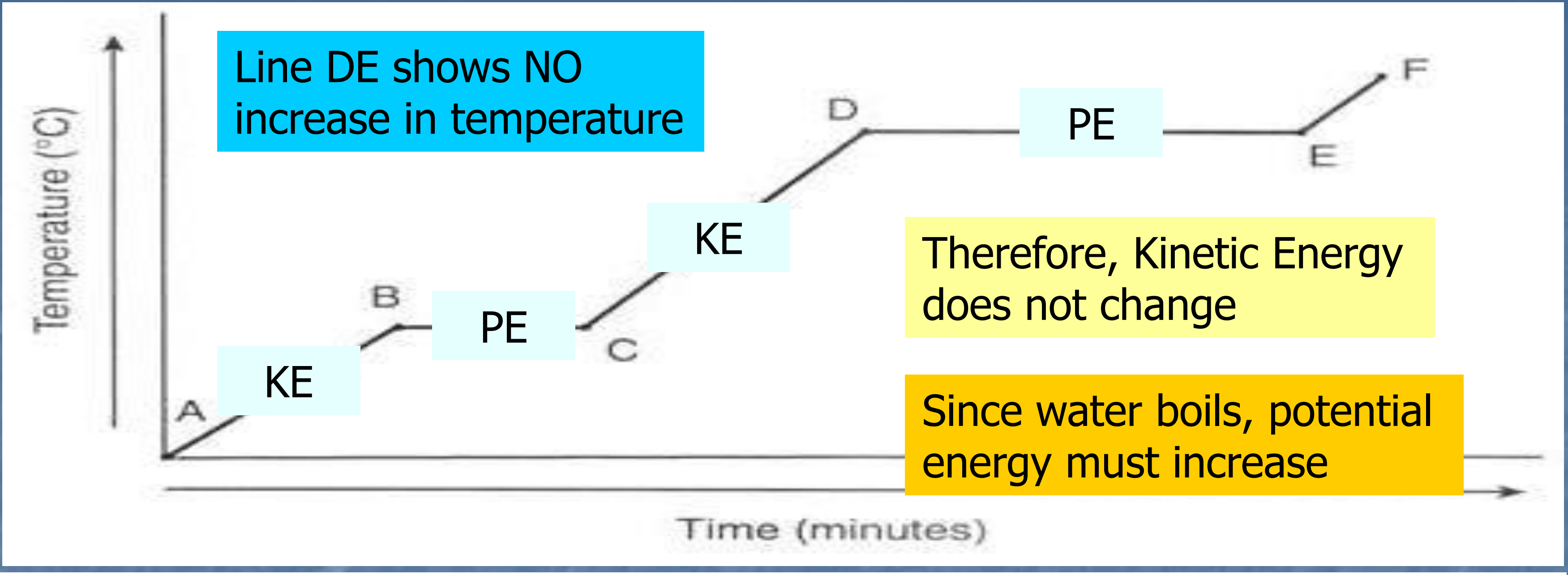

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Melting pt	B→C	0	s → l	**constant**	**increase**
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boiling pt	D→E	100	l → g	**constant**	**increase**
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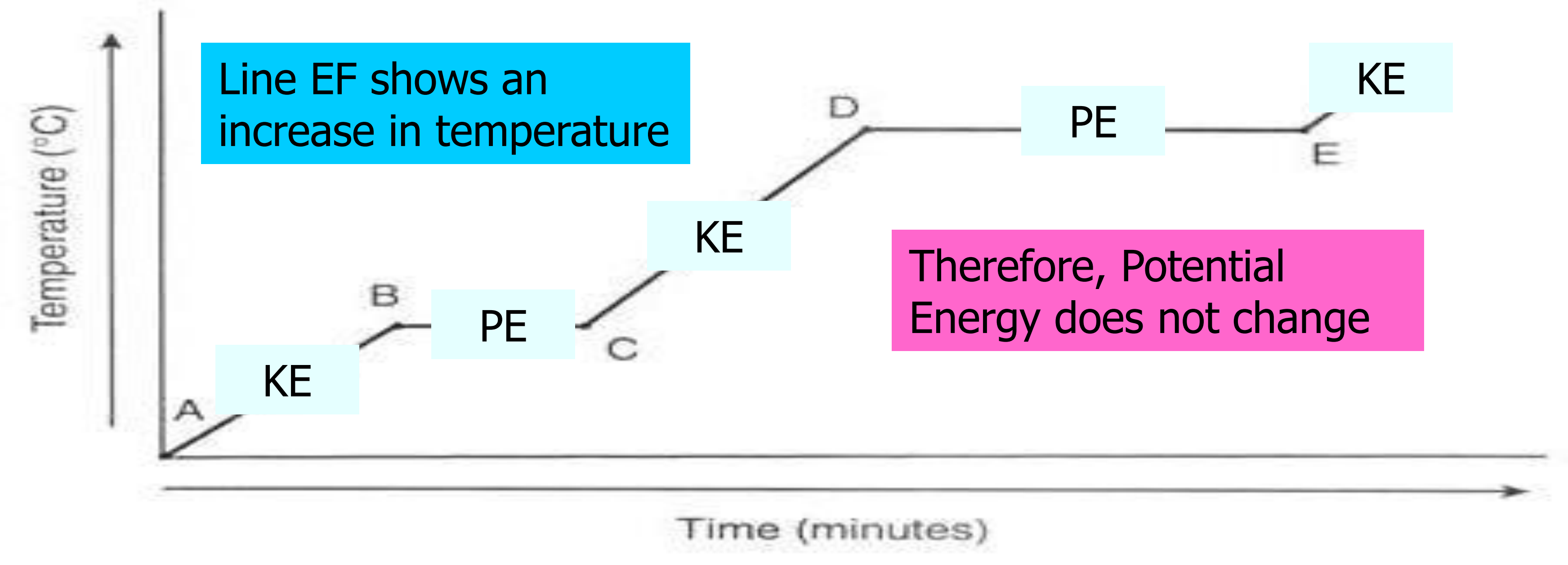

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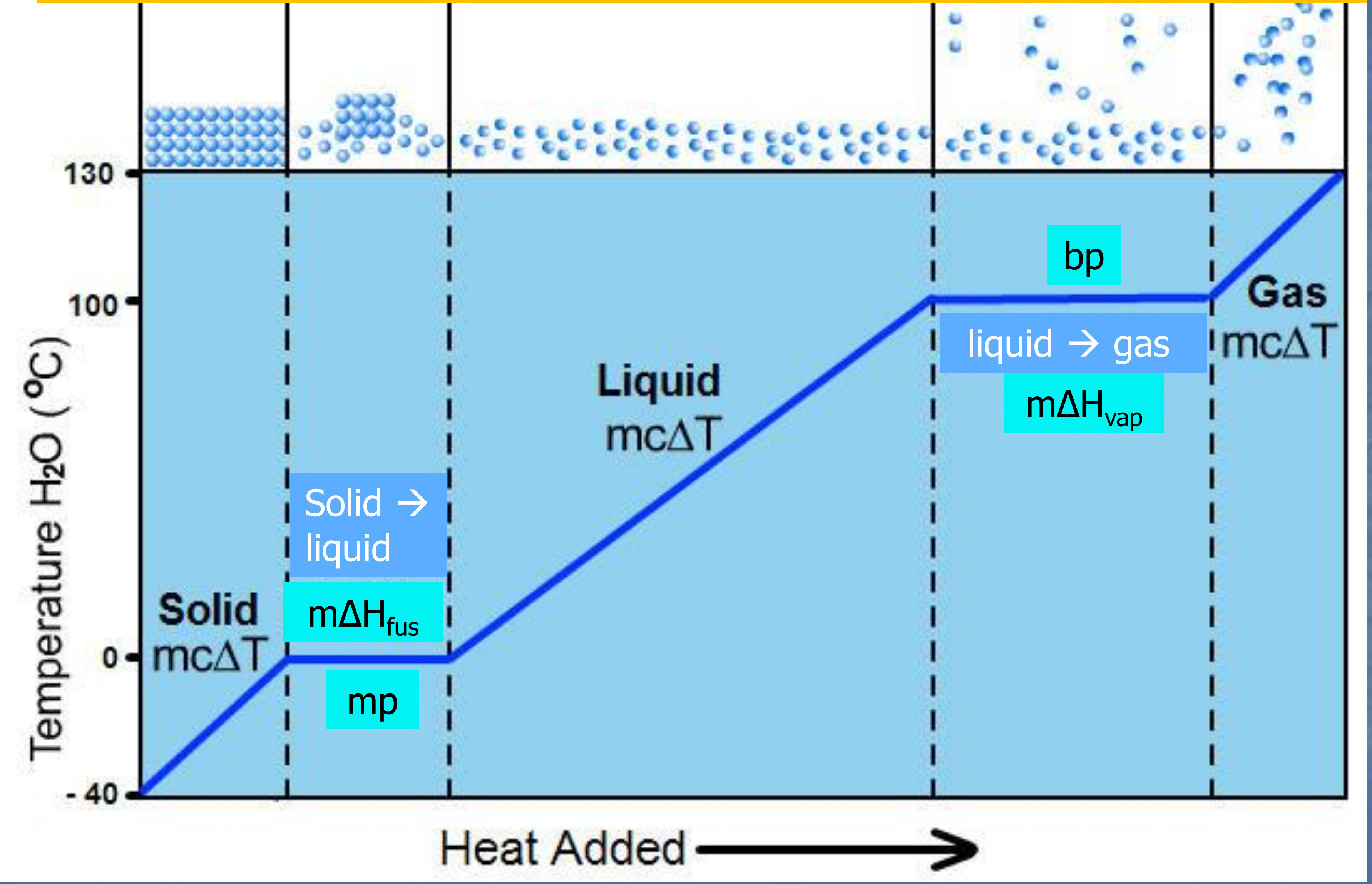
boiling pt	D→E	100	l → g	**constant**	**increase**
	E→F	100 → 110	g	**increase**	**constant**

Notice the "equations" for each change



Phase Change Diagrams

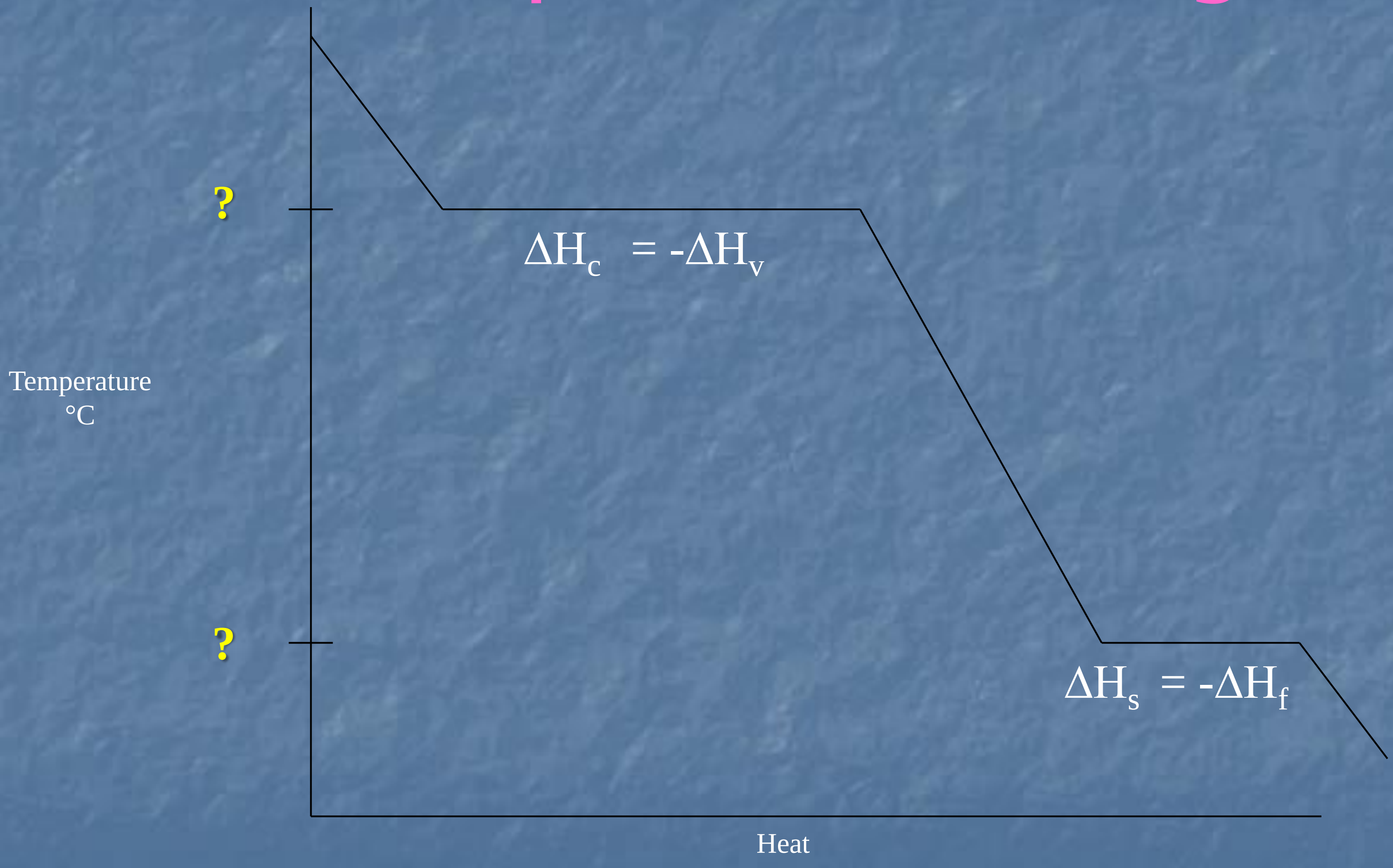
- Endothermic $\Delta H = +$
  - Heating Curve
  - heat flows **into** the system from the surroundings

- Exothermic $\Delta H = -$
  - Cooling Curve
  - Heat flows **from** the system into the surroundings

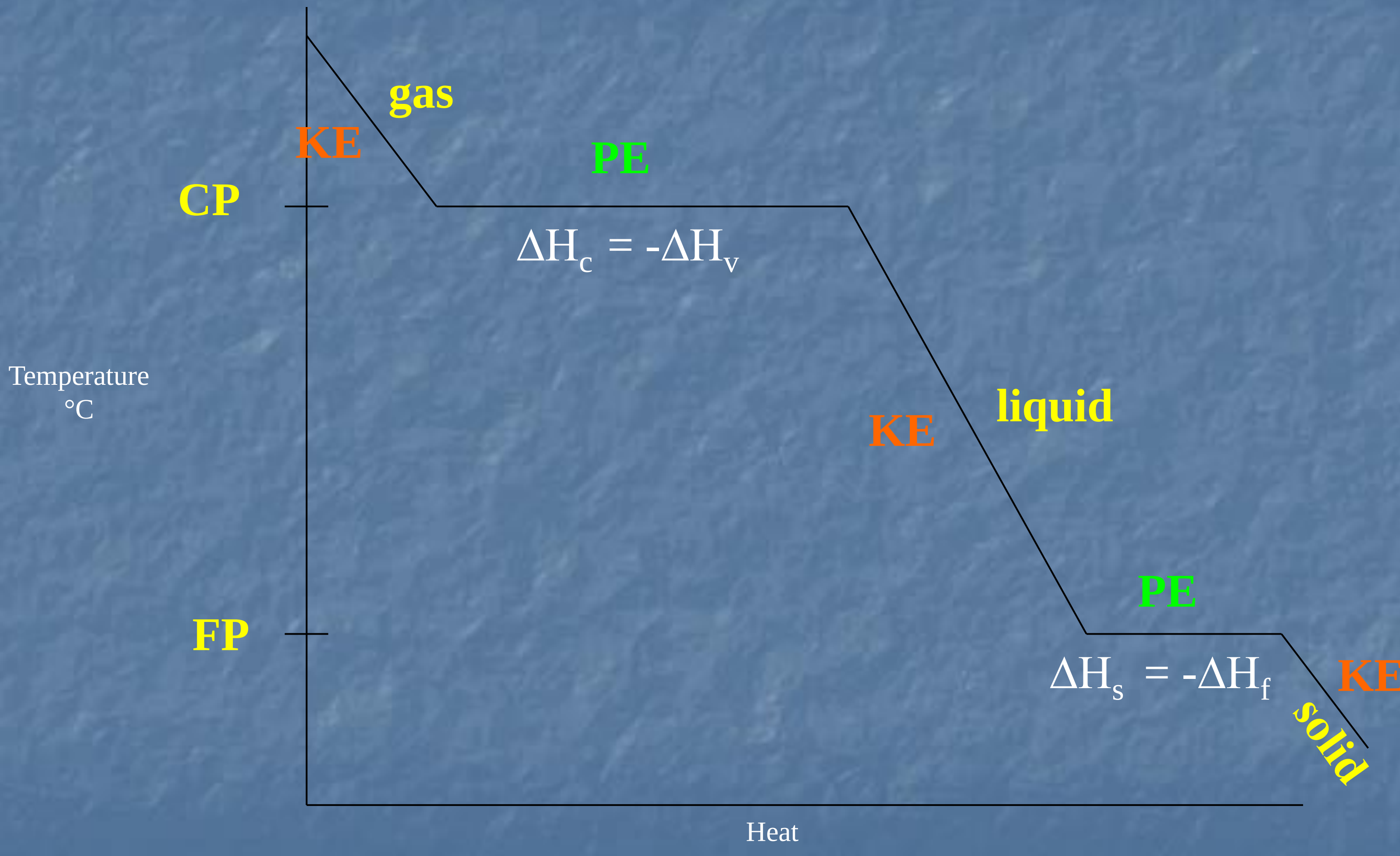
Exothermic (Cooling) Curve

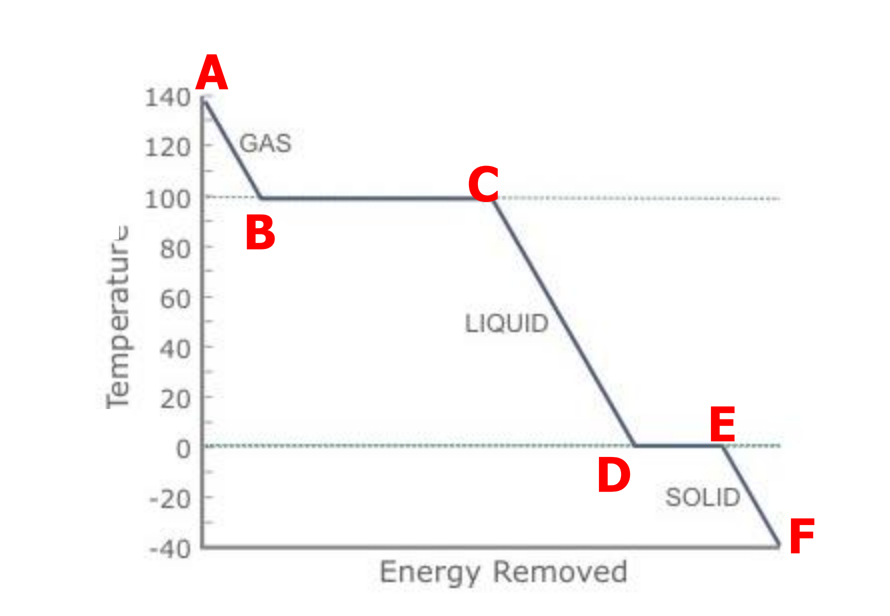
Label the phases & Energies

TRY IT

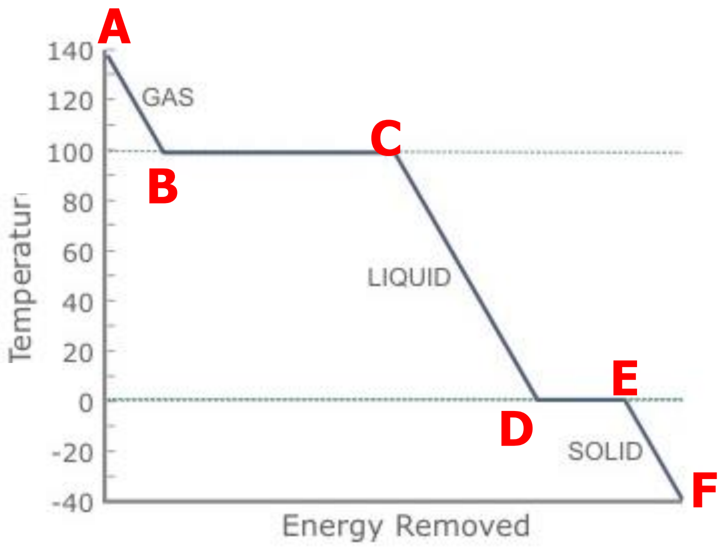


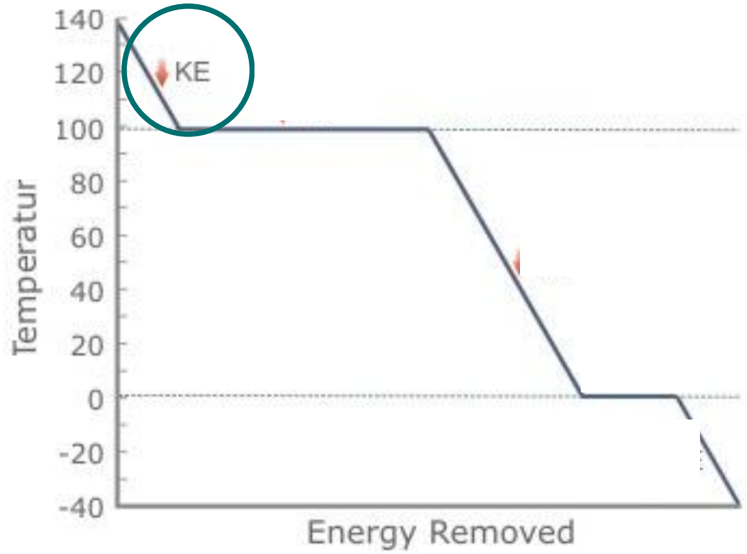
Exothermic (Cooling) Curve





	Graph	Temp	Phase	KE	PE
	A→B	140 → 100	g	?	?
	B→C	100	g → l	?	?
	C→D	100 → 0	l	?	?
	D→E	0	l → s	?	?
	E→F	0 → -40	s	?	?





	Graph	Temp	Phase	KE	PE
	A→B	140 → 100	g	**decrease**	**constant**
	B→C	100	g → l		
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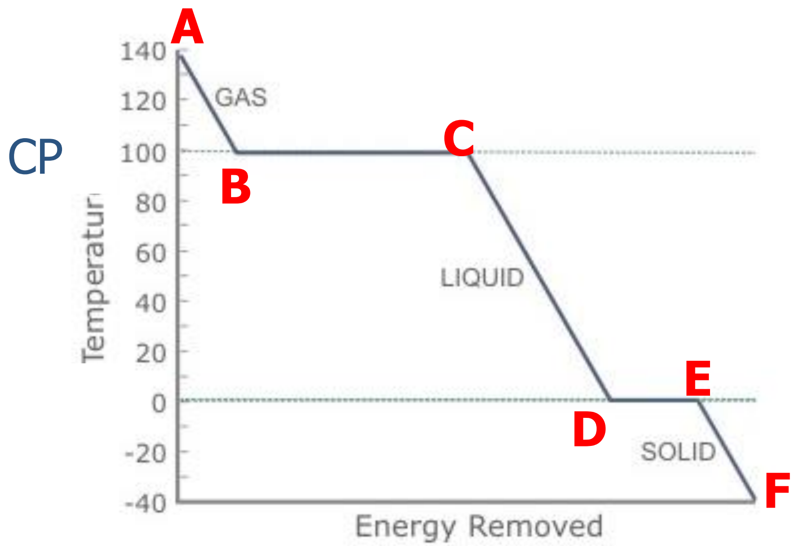

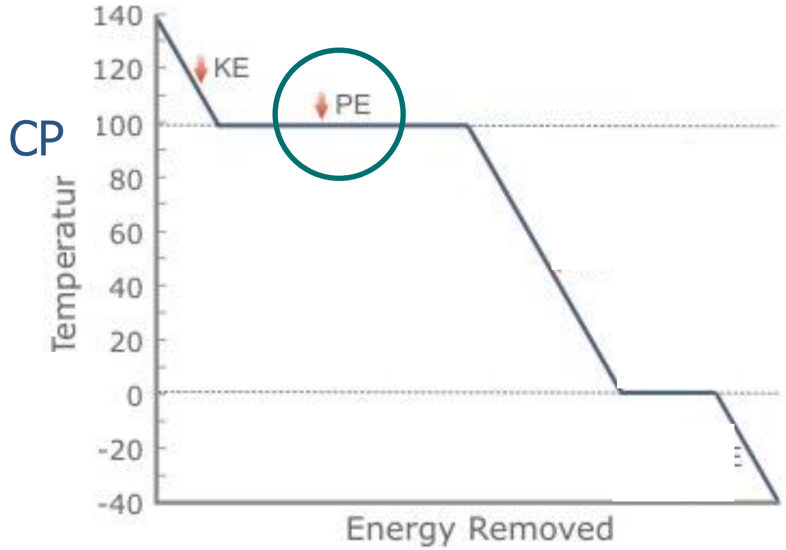

	Graph	Temp	Phase	KE	PE
	A→B	140 → 100	g	**decrease**	**constant**
Condensation pt	B→C	100	g → l	**constant**	**decrease**
	C→D	100 → 0	l		
	D→E	0	l → s		
	E→F	0 → -40	s		

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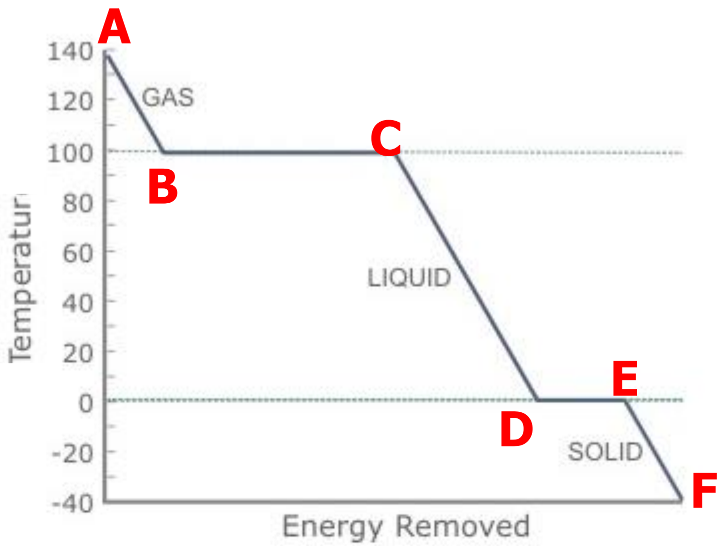

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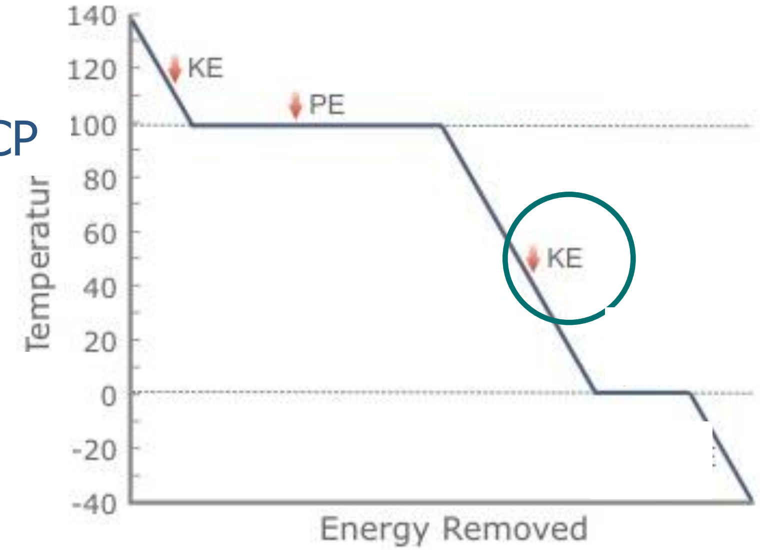

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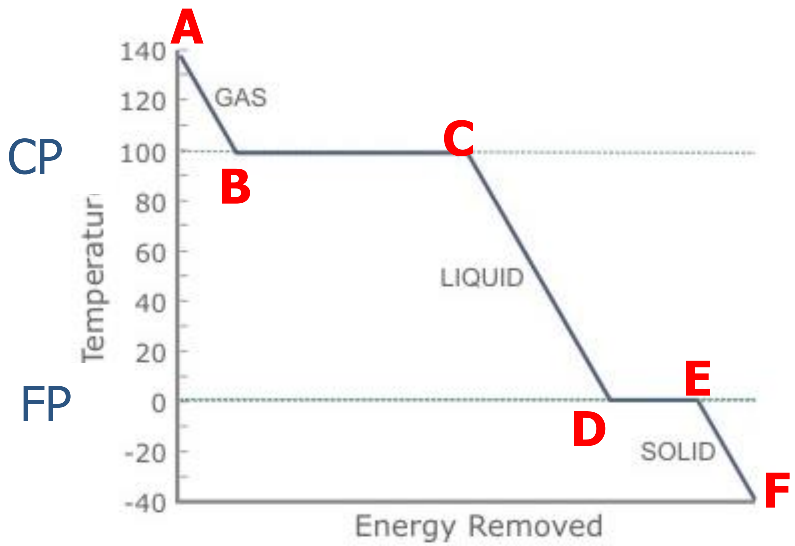

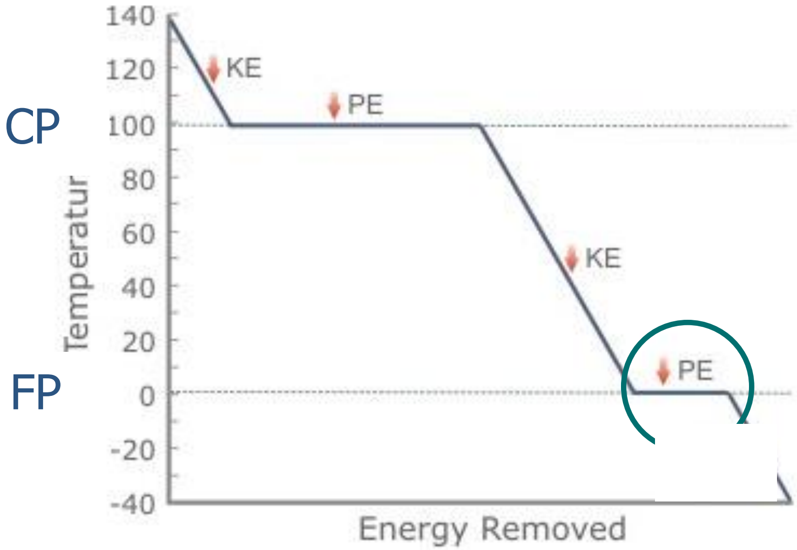

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Freezing pt	D→E	0	l → s	**constant**	**decrease**
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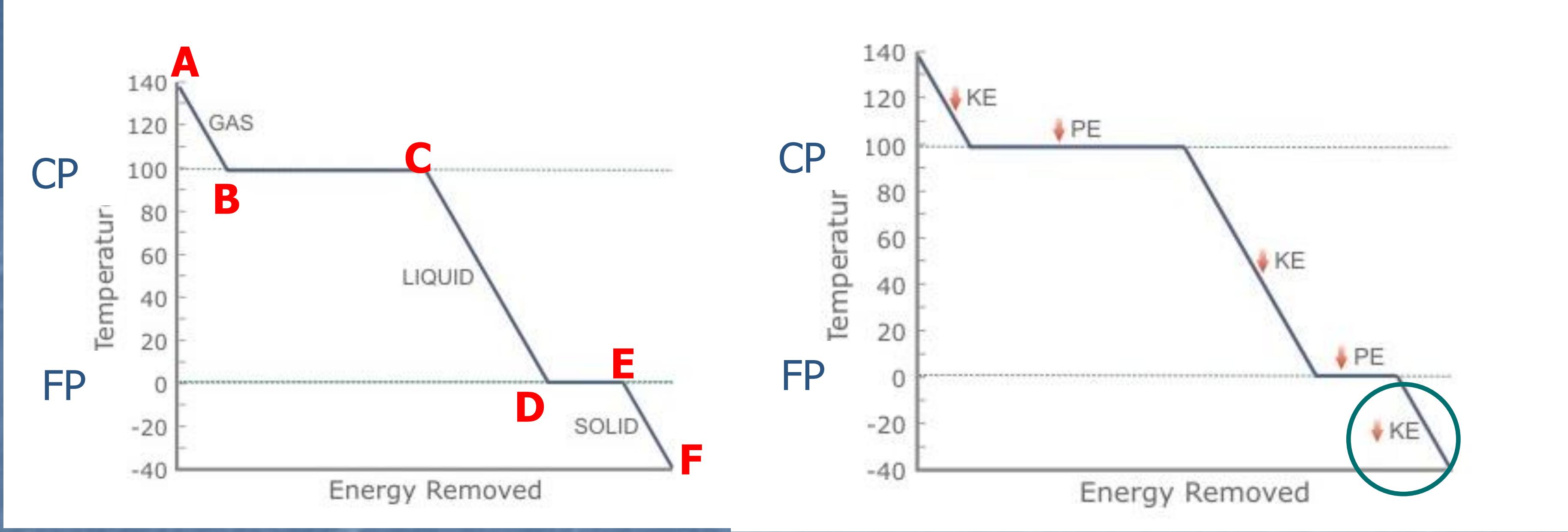

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Condensation pt	B→C	100	g → l	**constant**	**decrease**
	C→D	100 → 0	l	**decrease**	**constant**
Freezing pt	D→E	0	l → s	**constant**	**decrease**
	E→F	0 → -40	s	**decrease**	**constant**

General Heating Curve

Fixed points (based on PE) … temp is "fixed"

- Melting point ⟷ Freezing point (s ↔ l)
- Boiling point ⟷ Condensation pt (l ↔ g)
- Sublimation ⟷ Deposition (s ↔ g)

Fixed points for water

- Melting point ⟷ Freezing point

  0° C

- Boiling point ⟷ Condensation point

  100° C

Quantitative Heat Measurements

Heat Equation: $q = m\, c\, \Delta T$

KE

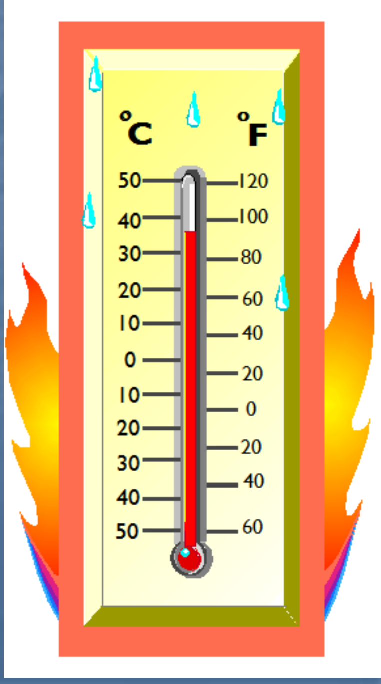
- **q: change in heat energy for system or surroundings**
- **m: mass of substance**
- **c_p: specific heat of substance**

c_p (H20): 1.0 cal/g·⁰C = 4.18 j/g·⁰C
c (ice): 0.5cal/g·⁰C = 2.09 j/g·⁰C
c (steam): 0.5cal/g·⁰C = 2.09 j/g·⁰C

Δ T: change in temperature (closed system)

Δ T = final Temp – initial Temp

Δ T = Tf – Ti



Heat of Fusion / Solidification

- The amount of energy gained or released when a substance melts (s → l) or freezes (l → s)

PE

Melting (s → l): endothermic $\Delta Hf = +$

Freezing (l → s): exothermic $\Delta Hf = -$

ΔH_s solidification

For water: $\Delta H_f = +\underline{80.0\ cal/g} = -\Delta H_s$

80.0 cal/g x 4.18 j/cal = <u>334.4 j/g</u>

Heat of Vaporization / Condensation

- The amount of energy gained or released when a substance vaporizes (l → g) or condenses (g → l)

PE

vaporizing (l → g): endothermic $\Delta Hv = +$

condensing (g → l): exothermic $\Delta Hv = -$

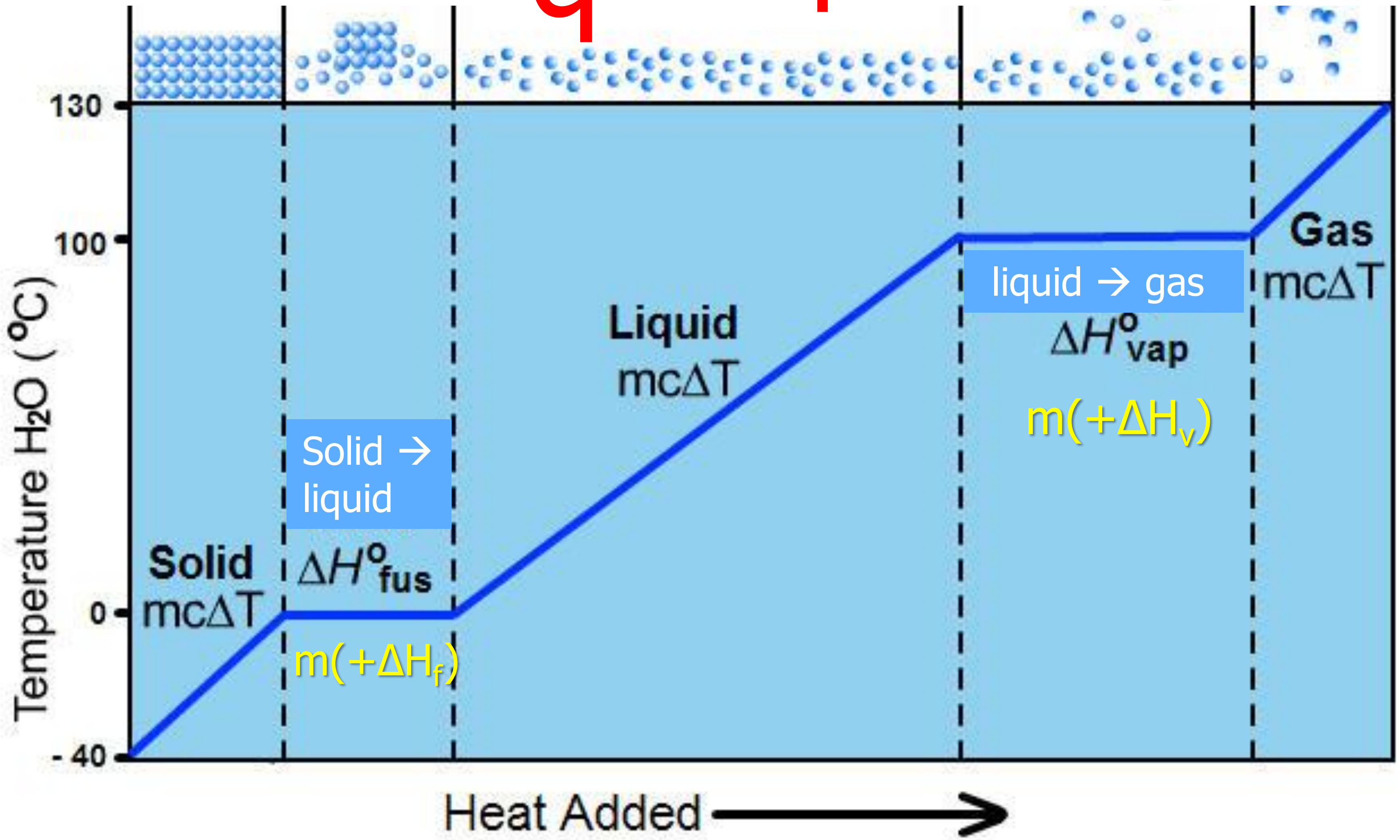
ΔH_c = condensation

For water: $\Delta H_v = +$<u>540. cal/g</u> = $-\Delta H_c$

540. cal/g x 4.18 j/cal = <u>2.26 kJ/g</u>

Notice the "equations" for each change

Endothermic $q = +$ Heating Curve



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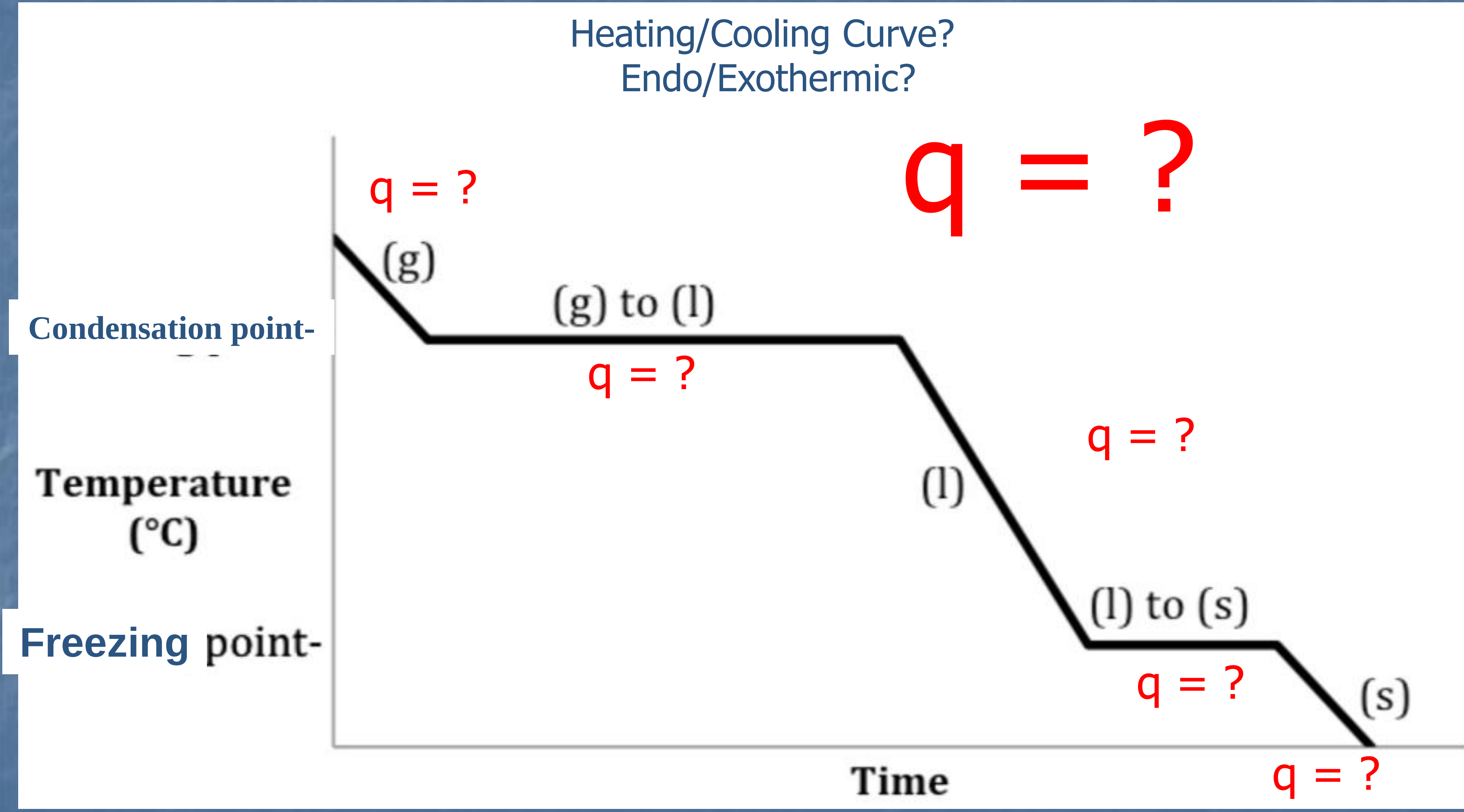

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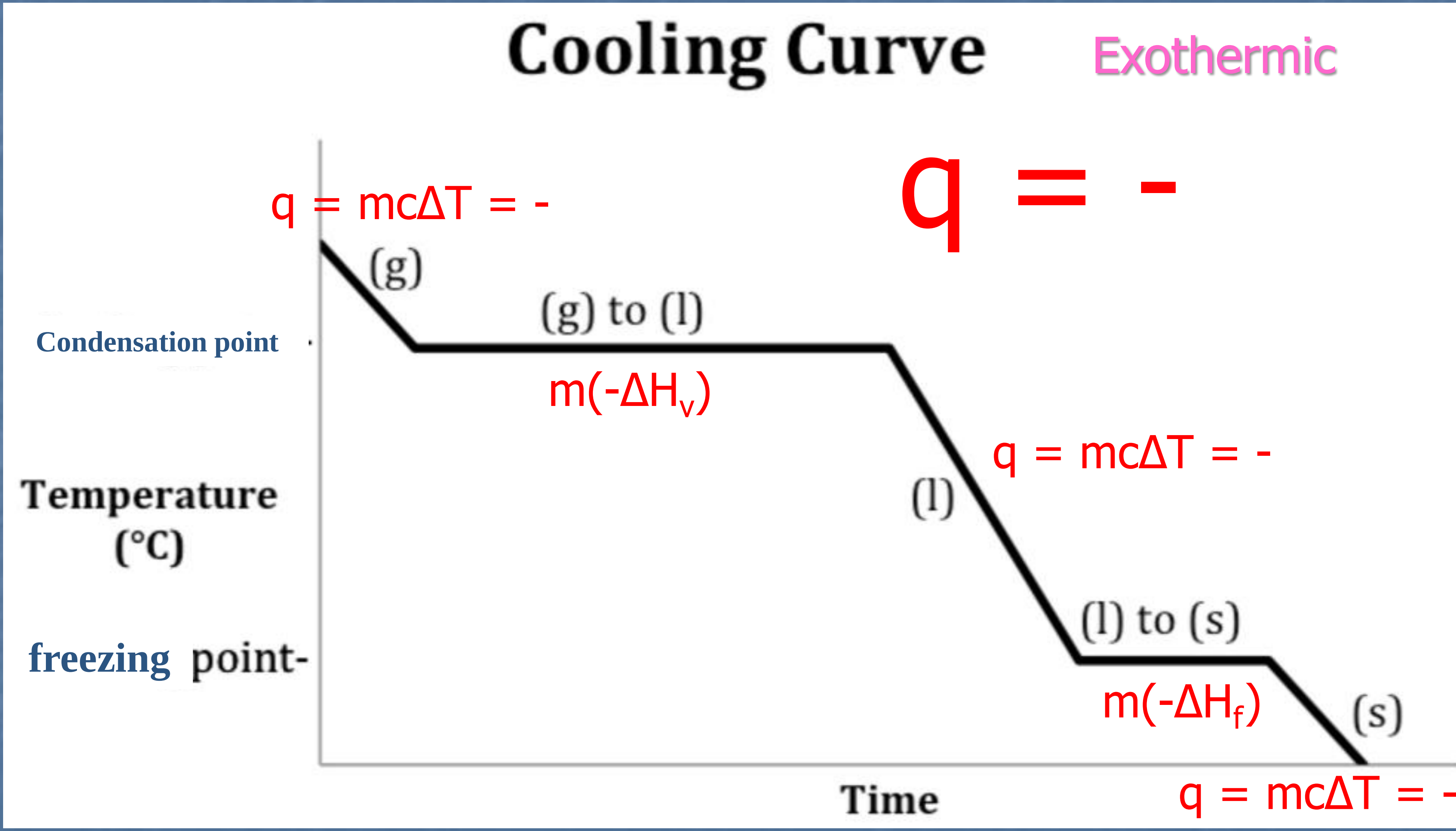


Quantitative Heat Measurements

How much heat flows when 20.0 g of water at 10.0 °C is heated to 40.0 °C? (Endo/Exothermic?)

Quantitative Heat Measurements

How much heat flows when 20.0 g of water at 10.0 °C is heated to 40.0 °C? (Endo/Exothermic?)

heating water → $q_{liquid} = m\,c\,\Delta T$

$$q_{water} = (20.0\ g)(1.00\ cal/g\cdot°C)(40.0°\ C - 10.0°C)$$

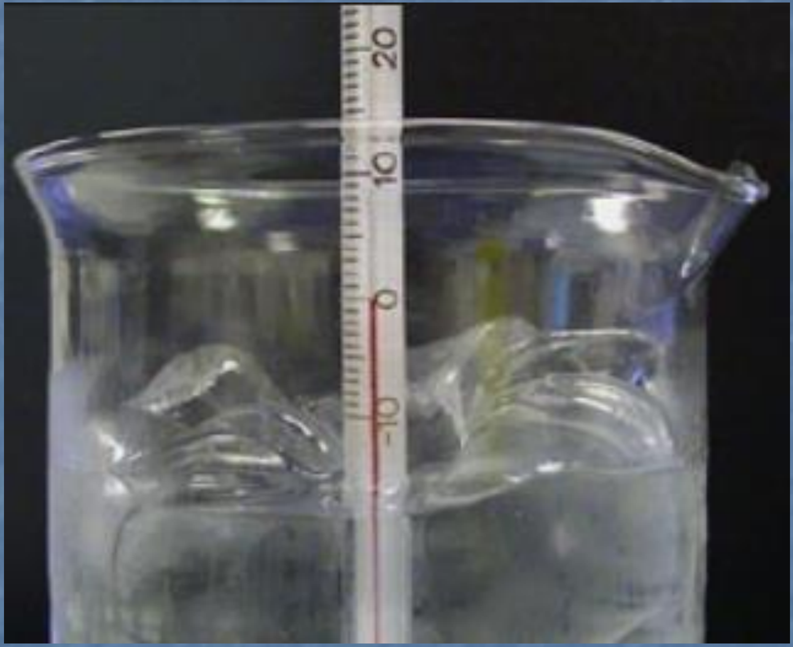
$$q_{water} = (20.0\ g)(1.00\ cal/g\cdot°C)(30.0°C)$$

$$q_{water} = +600\ cal \times 4.18\ J/cal = +2.51\ kJ$$

Endothermic (heat is added or absorbed)

Quantitative Heat Measurements

How much heat flows when 20.0 g of water freezes at 0.0 °C? (Endo/Exothermic?)



Quantitative Heat Measurements

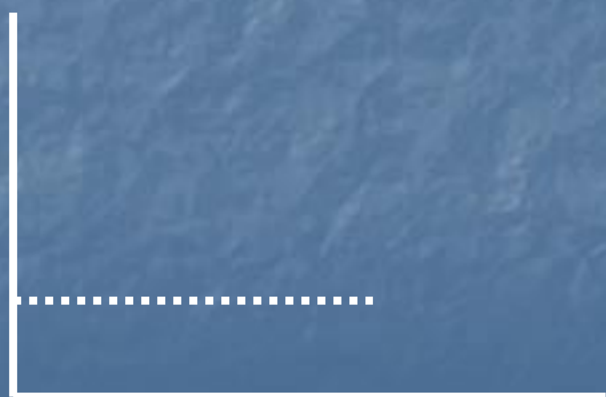
How much heat flows when 20.0 g of water freezes at 0.0 °C? (Endo/Exothermic?)

freezing water → $q_{l \rightarrow s} = m\,\Delta H_f$

$$q_{ice} = (20.0\ g)(80.0\ cal/g)$$

$$q_{ice} = -1{,}600\ cal \times 4.18\ J/cal = -12.8\ kJ$$

Exothermic (heat is "lost" or released)



Specific Heat of Water

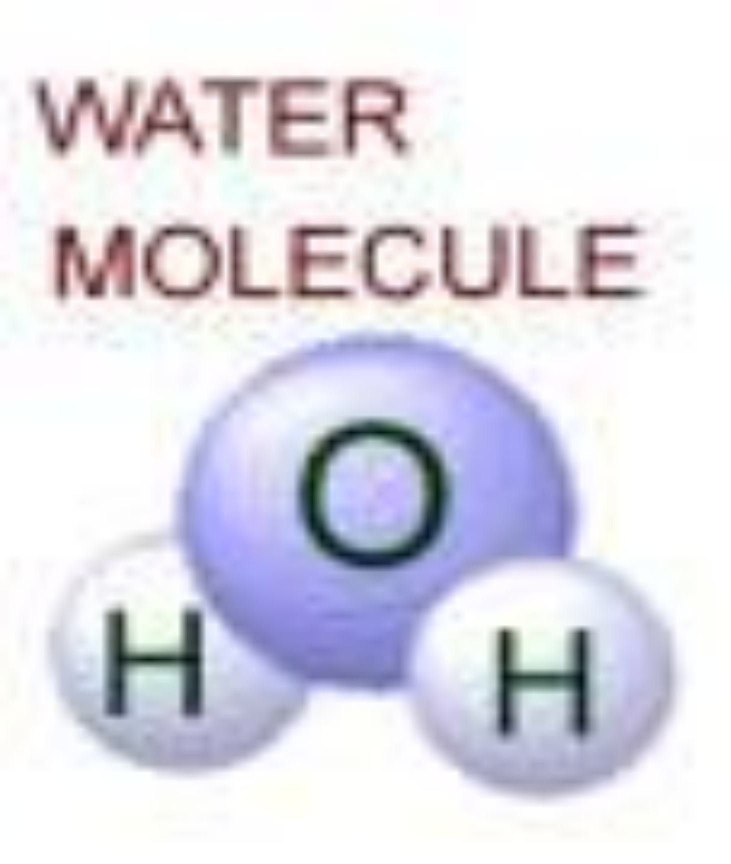
The **specific heat capacity**, or simply the specific heat (C) of a substance, is the amount of heat it takes to raise the temperature of 1 g of the substance 1º C.

- Water has the second highest specific heat of all liquids.
- Metals generally have low specific heats.

Specific Heats of Some Common Substances		
Substance	Specific heat (c_p)	
	J/(g·ºC)	cal/(g·ºC)
Liquid water	4.18	1.00
Ethanol	2.4	0.58
Ice	2.1	0.50
Steam	1.9	0.45
Chloroform	0.96	0.23
Aluminum	0.90	0.21
Iron	0.46	0.11
Silver	0.24	0.057

Specific Heat of Water

- Water has the second highest specific heat of all liquids.



Specific Heat (C)

Some things heat up or cool down faster than others.

e.g. Michigan stays warmer in the winter due to the heat from the lakes (*e.g. hot vegetables stay hot a long time because of their high water content*)

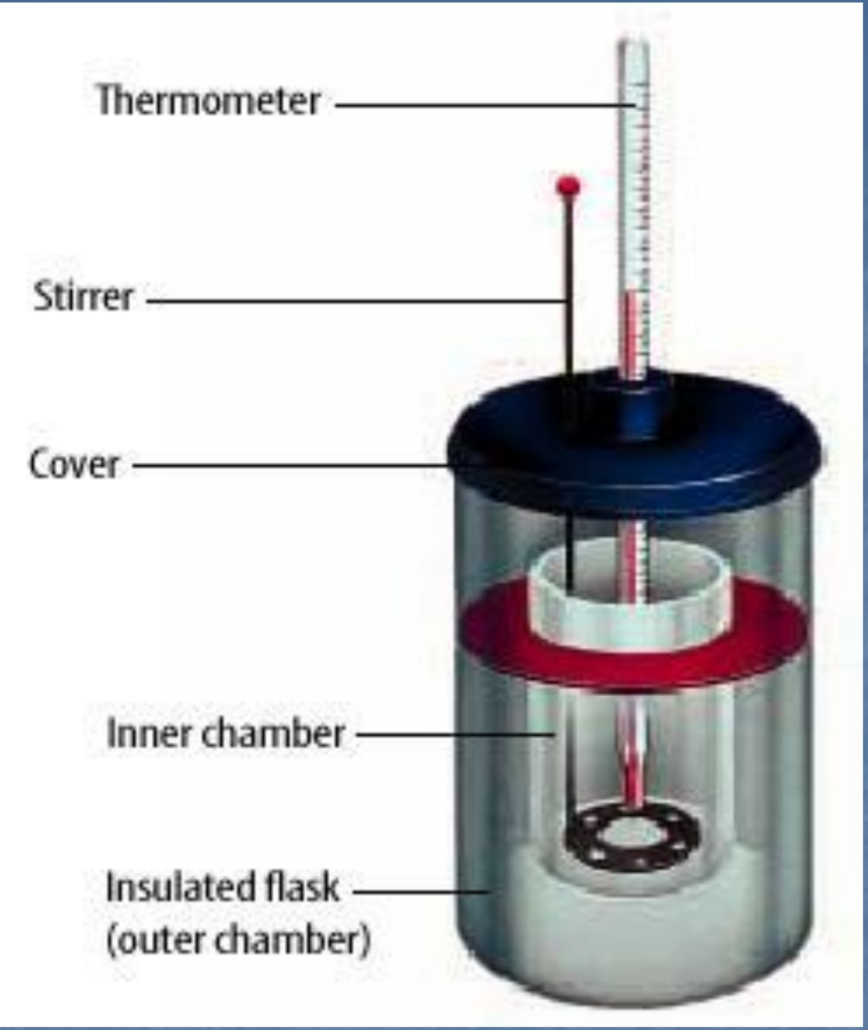
Specific heat is ability of a substance to “hold” heat.

C_{water} = 4.184 J / g°C

C_{sand} = 0.664 J / g°C

This is why land heats up quickly during the day and cools quickly at night and why water takes longer.

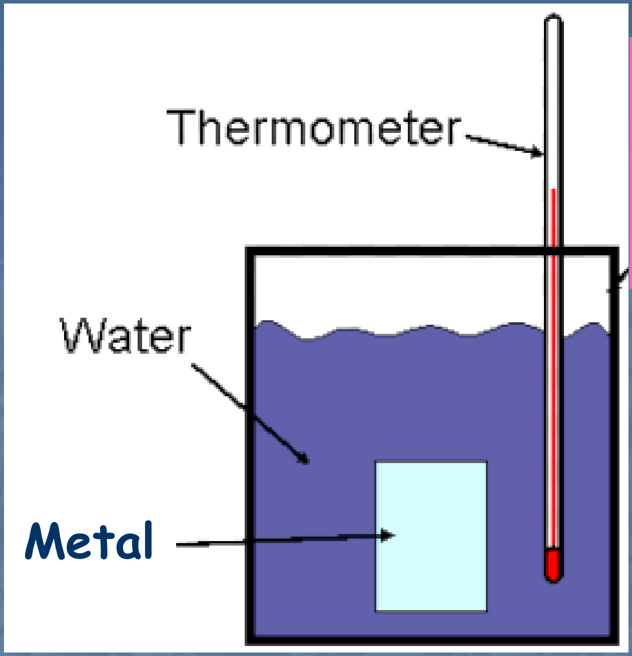
A calorimeter is used to measure the specific heat of a substance.



Rearrange q = mcΔT to solve for c:

$$c = \frac{q}{m \times \Delta T} = \frac{\text{heat (J or cal)}}{\text{mass (g)} \times \text{change in temperature (°C)}}$$

Place a hot metal (~ 75 °C) in water (~25 °C)



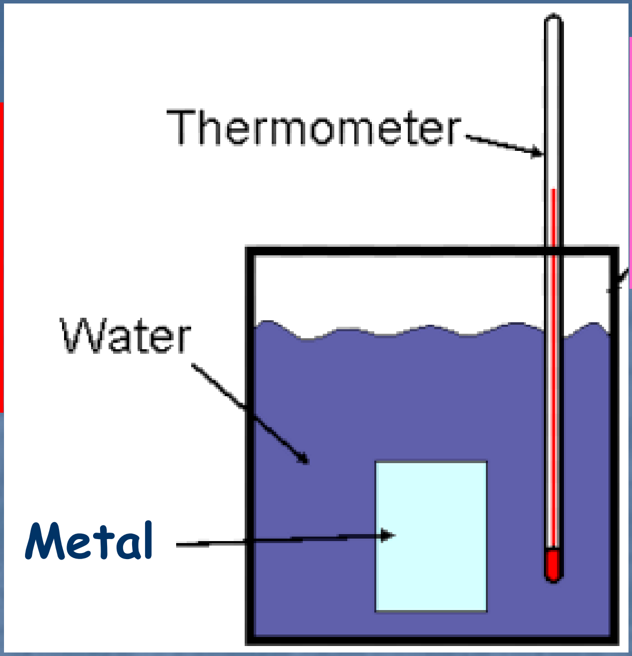
Closed system

What happens to the metal?

What happens to the water?

How does the heat compare?

Place a hot metal (~ 75 C) in water (~25° C)



Closed system

The metal cools off (loses heat); q = -

The water gains heat from the metal; q = +

$$-q_{metal} = q_{water}$$

Exothermic = Endothermic