Phase Changes and Interactions

Auburn Mountainview

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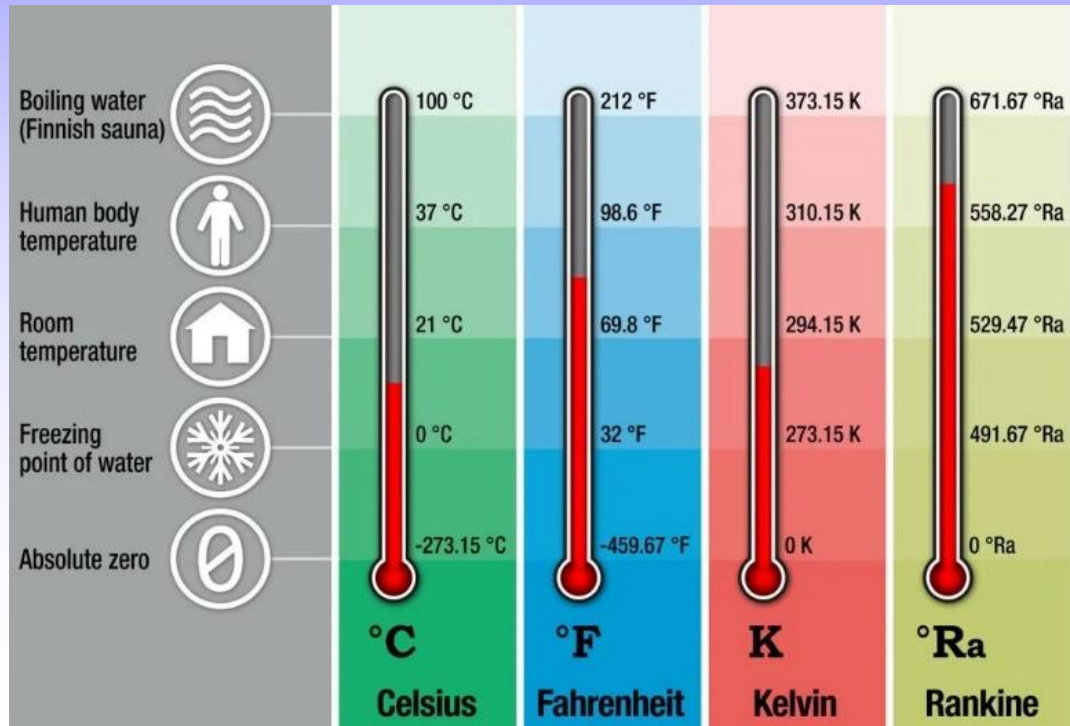
By end of this unit I can…

- A.1 Create and/or evaluate a claim about the relationship between transfer of thermal energy and the temperature change in different samples.
- A.2 Perform calculations using data gathered from a simple constant-pressure calorimetry experiment.
- B.1 Use data to explain the direction of energy flow into or out of a system.
- C.1 Explain the relationship between changes in states of matter and the attractions among particles.
- C.2 Create and/or interpret models representing phase changes.
- D.1 Create and/or interpret heating and cooling curves and/or phase diagrams of pure substances.
- D.2 Calculate the energy transferred when a substance changes state.

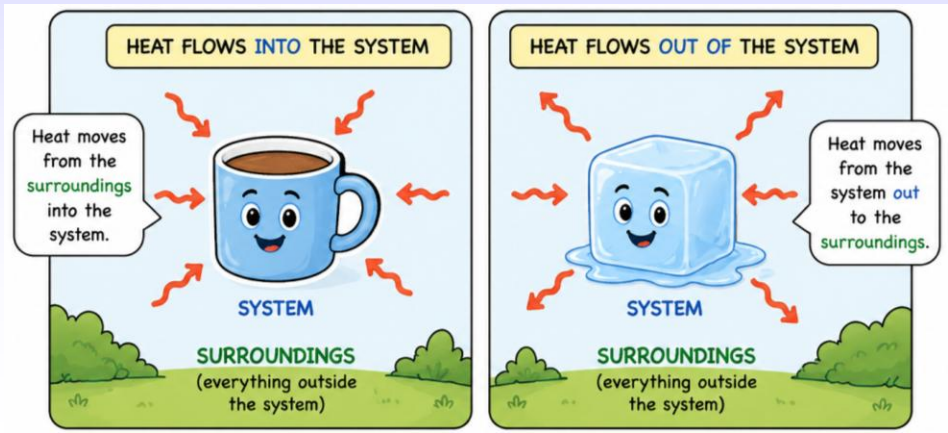

Temperature

- **Temperature: Quantifiable measure of the average Kinetic Energy of the particles in a substance.**
- **Two common scales with an absolute:**
  - **Centigrade-°C (H_2O: 0→100)**
    - **Kelvin-K (273.15/373.15)**
    - **C→K: $T_K = T_C + 273.15$**
  - **Fahrenheit-°F (Brine: 0→ Human 100)**
    - **Recalibrated for H_2O (32→212: 180°)**
    - **Rankine-°Ra (491/671)**
    - **F→C: $T_F = (9/5)T_C + 32$**



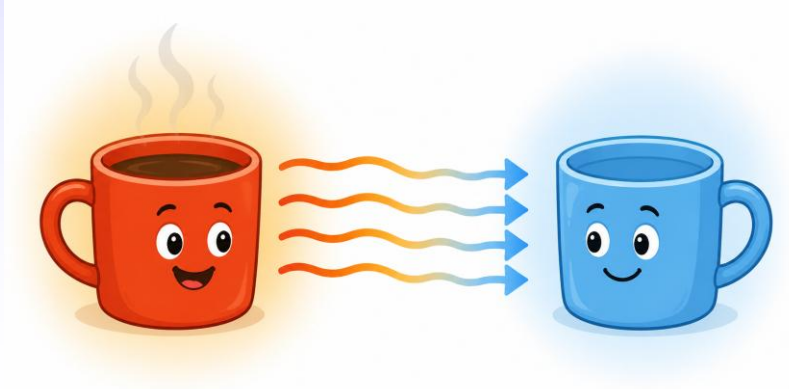
Relationships

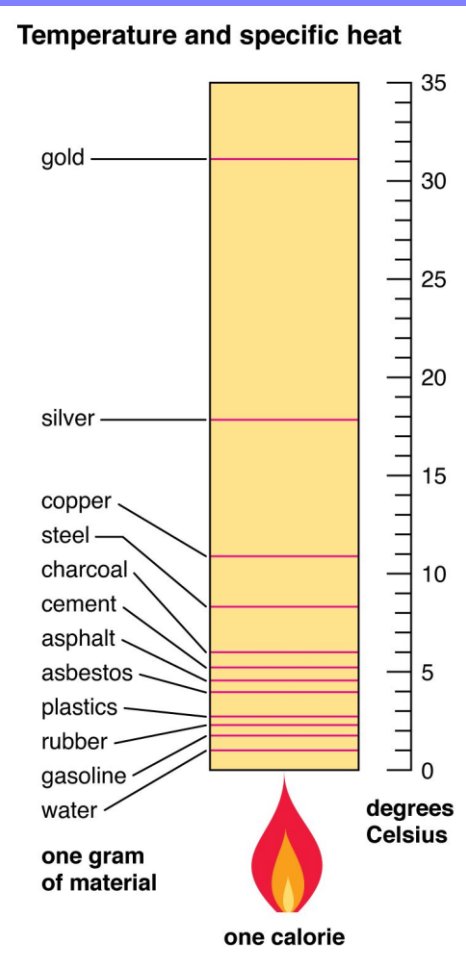
- **Energy: The capacity to supply heat or do work: Joule**
- **Heat: The transfer of thermal energy between a system and its surroundings.**
  - **System: The reaction or process being studied.**
  - **Surrounding: Everything else not related to the system.**
  - **Heat flows always from high temperatures to low ones.**
- **Quantified Relationship: $q = cm\Delta T$**



$q = cm\Delta T$

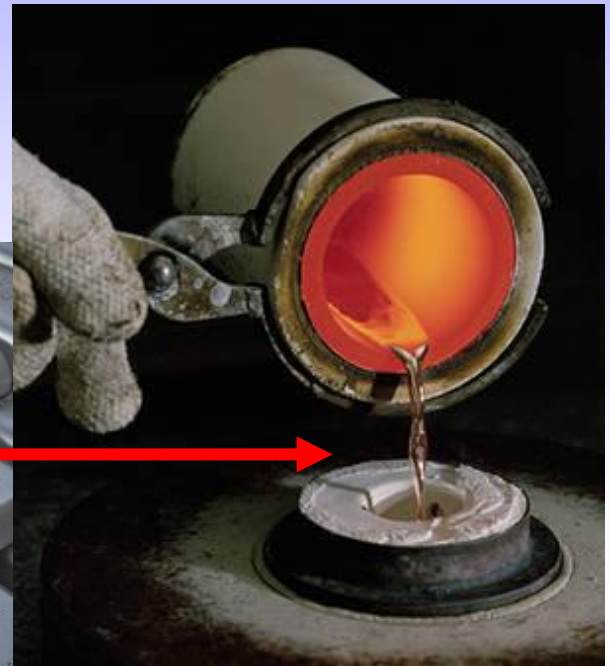
- **q: heat added or removed from the system: Joules or calories**
  - **q>0: heat added (exothermic)**
  - **q<0: heat removed (endothermic)**
- **c: specific heat capacity is the heat required to change the temperature: Joules/grams·°Centigrade**
  - **For 1 gram H_2O: 1-cal/g°C or 4.184-J/g°C**
- **m: mass of the object (system)**
- **ΔT: change in temperature**
  - **Δ: final minus initial**





Example 1

How much heat is needed to raise 20.0-g of silver (c= .235-J/g°) to 1200.0-ºC? Assume the silver starts at room temperature (20.0-ºC).



Example 1

$$q = cm\Delta T$$

$$q = .235 - \frac{J}{g°C} \cdot 20.0 - g \cdot 1180 - °C$$

$$q = 5546 - J$$

q =
c = .235-J/g•°C
m = 20-g
ΔT = (1200-20)-°C
ΔT = 1180-°C

$$\boxed{q = 5550 - J}$$

Example 2

A 1.3-kg watermelon (assume same c as water) is left out on a hot (35-°C) day. To cool it down it is placed in a cooler containing 1500.0-g of ethanol at -10.0-°C. Equilibrium in the cooler occurs and the temperature is found to be 15-°C. What is the specific heat of ethanol?

Assume there is no loss of heat outside the cooler, so this is a closed system (q=0)

Example 2

$$q_{system} = q_{watermellon} + q_{ethanol}$$

$$0 = q_{watermellon} + q_{ethanol}$$

$$-q_{watermellon} = q_{ethanol} \qquad q = cm\Delta T$$

$$-cm\Delta T = cm\Delta T$$

$$-4.184 - \frac{J}{g℃} \cdot 1300 - g \cdot -20 - ℃ = c \cdot 1500 - g \cdot 35 - ℃$$

$$108784 - J = 52500 - g \cdot ℃$$

$$q = 2.07 - \frac{J}{g℃}$$

c = 4.184-J/g•°C
m = 1300-g
ΔT = (15-35)-°C
ΔT = -20-°C

c =
m = 1500-g
ΔT = (15-(-20))-°C
ΔT = 35-°C

$$q = 2.1 - \frac{J}{g℃}$$

Phase Changes

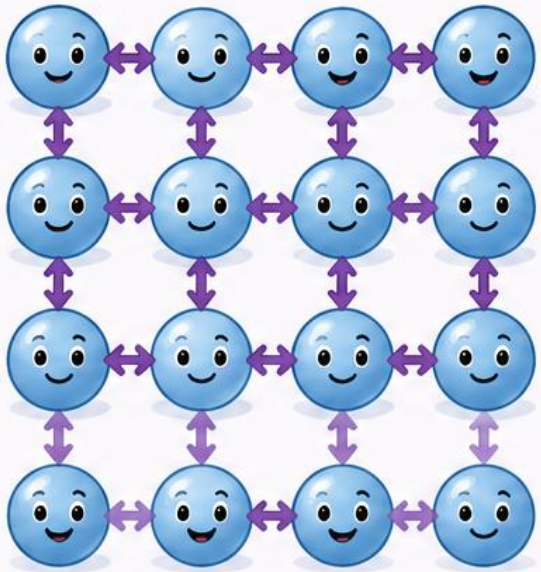
- **Intermolecular Forces: Attractions between molecules.**
  - **London, Dipole, Hydrogen: electron based (future topic)**
  - **Metals: Not IMF, delocalized electrons (future topic)**
- **Kinetic Energy tries to pull molecules apart.**
- **Adding heat to a solid/liquid/gas increases Kinetic Energy. This can cause KE to overpower the IMF and change the molecules phase.**
  - **Can also be reversed (remove heat: cooling)**

INTERMOLECULAR FORCES (IMFs)

Forces of attraction between particles.

SOLID

Particles are close together in a fixed position.

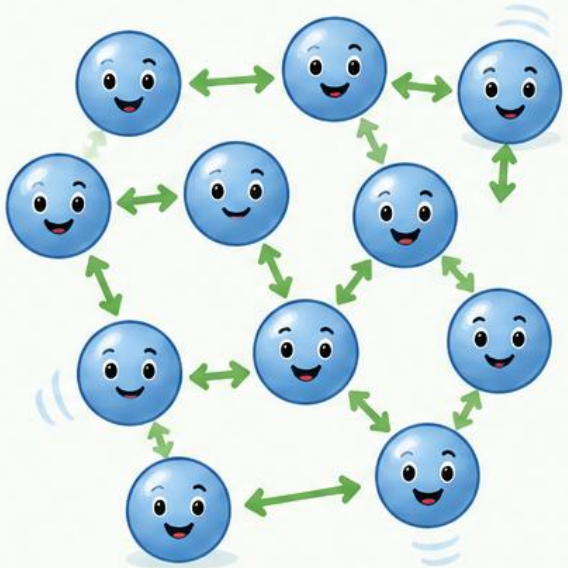


IMFs are **STRONG**

Particles vibrate in place but do not move past each other.

LIQUID

Particles are close together but can move around.

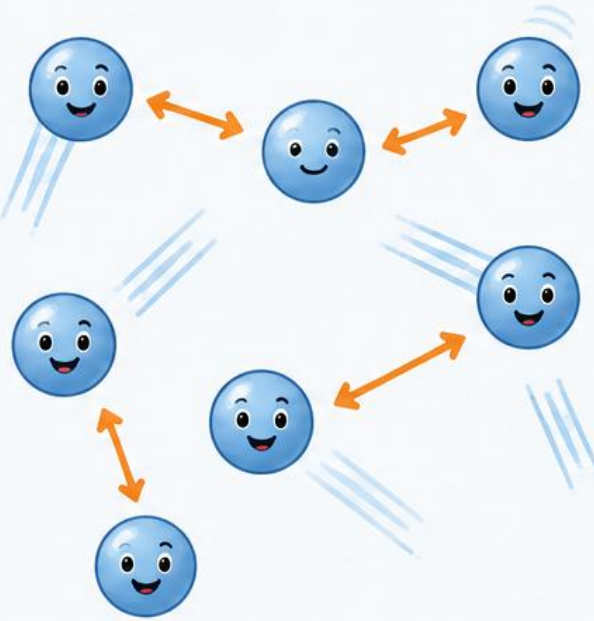


IMFs are **MODERATE**

Particles move around and slide past each other.

GAS

Particles are far apart and move freely.

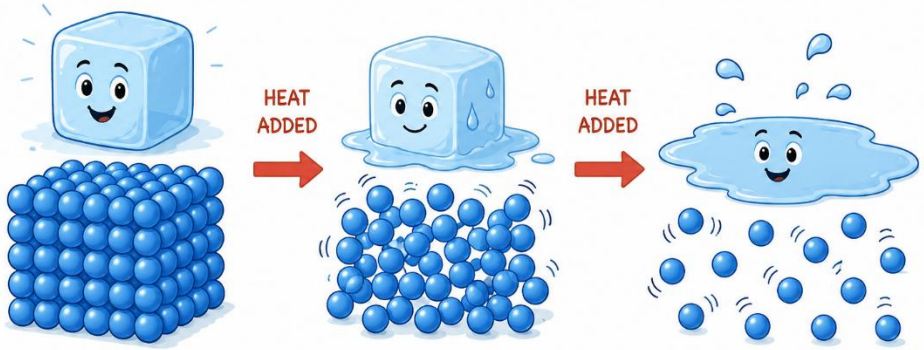


IMFs are **WEAK**

Particles move quickly in random directions and rarely attract each other.

⟷ Arrows represent the strength of intermolecular forces. Purple = strong Green = moderate Orange = weak

Melting: A Solid Changing to a Liquid

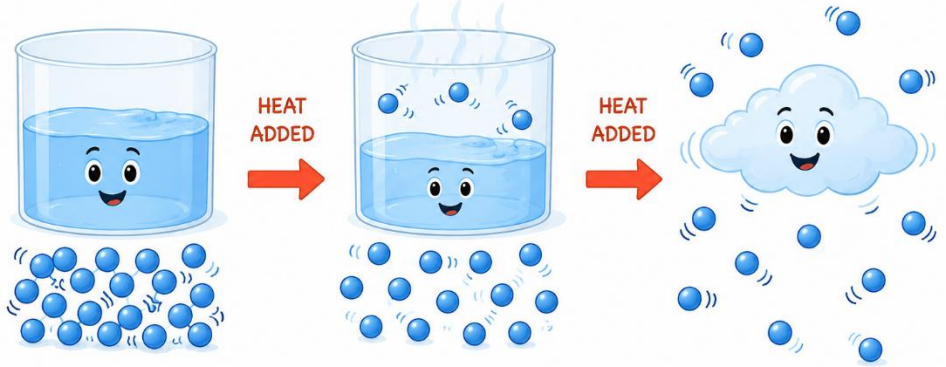


SOLID
Particles are close together in a fixed, ordered arrangement.

MELTING
Particles gain energy. They move more and break away from their fixed positions.

LIQUID
Particles are farther apart and move freely around each other.

Evaporation: A Liquid Changing to a Gas



LIQUID
Particles are close together and move around each other.

EVAPORATING
Particles gain energy. Some move faster and escape from the surface into the air.

GAS
Particles are far apart and move quickly and freely in all directions.

Gas

Heat Removed

Condensation

Vaporization

Deposition

Sublimation

Heat Added

Liquid

Freezing

Melting

Solid

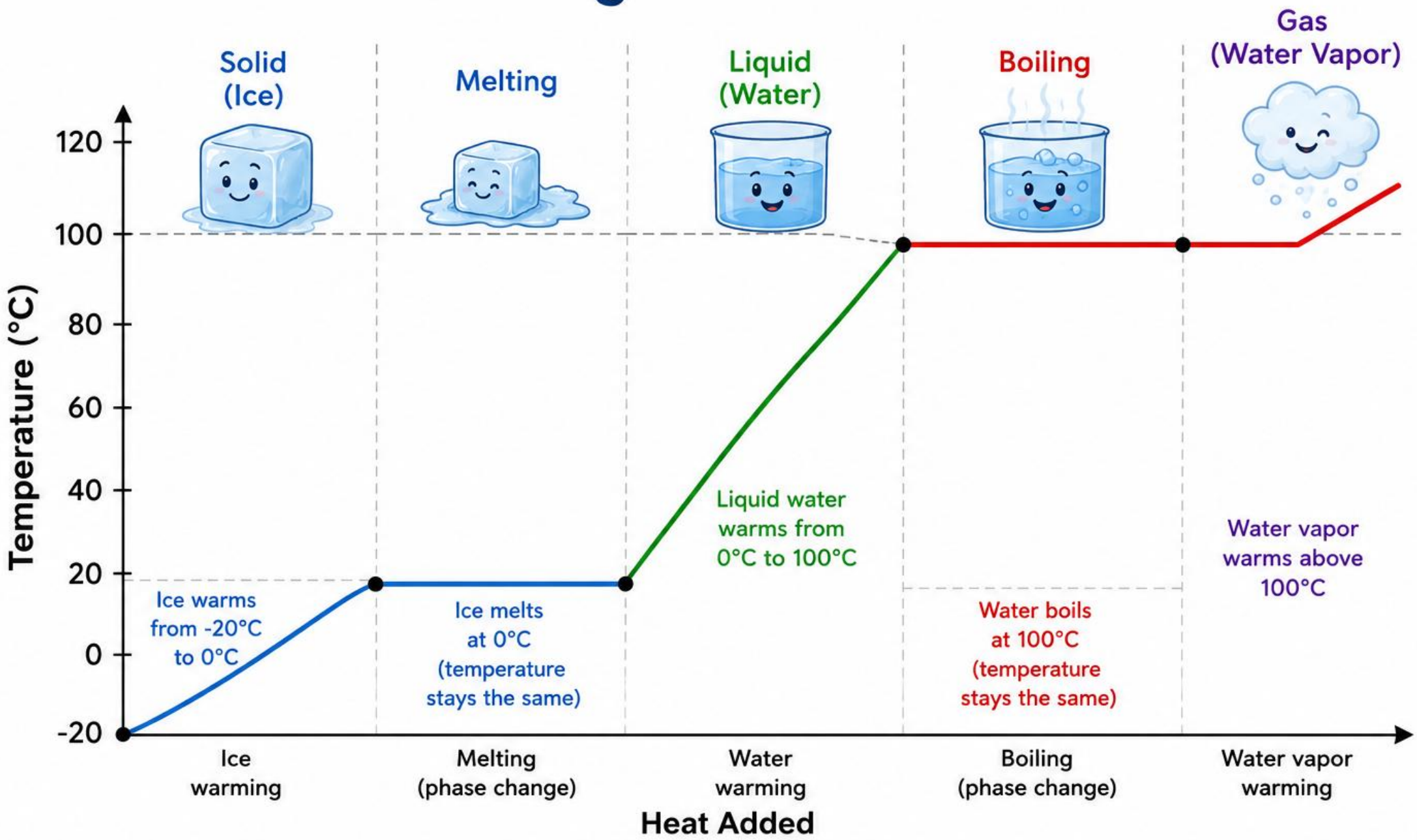
Phase Change Transitions

- **For a Solid/Liquid/Gas increasing heat increases the KE which can be seen as a temperature increase.**
  - **At some temperature, the IMF can't hold it together.**
    - **Temperature is different for every molecule or metal.**
    - **Causes a transition between phases.**
- **During this transition, all heat is now used to break the IMF's, until the phase change is complete.**
  - **There is no change in temperature during this process.**

Phase Change Transitions

- **A Heating Curve can show this visually.**
  - **y-axis: Temperature, x-axis: heat or time heating**
- **Solid/Liquid/Gas: Slope represents specific heat of material. ($q=cm\Delta T$)**
- **Plateau: Transition, breaking of IMF. ($q=mL_x$)**
  - **Heat needed is based on a constant called Latent Heat.**
    - **L_F (Fusion) Solid/Liquid or L_V (Vaporization) Liquid/Gas**
- **Cooling curve shows the reverse (same values just -)**

Heating Curve for Water



Example 1: Revisited

How much heat is needed to raise 20.0-g of silver (c= .235-J/g°) to 1200.0-ºC? Assume the silver starts at room temperature (20.0-ºC).

Silver melts at 961.8°C. The heat needed to break the metallic bonds (not IMF) was not accounted for.

Example 1: Revisited

$\mathbf{q = mL_F}$

$q = 20 - g \cdot 105 - \frac{J}{g}$

$q = 2100 - J$

q =
m = 20-g
L_F= 105-J/g

Added to the heat needed to change the temperature…

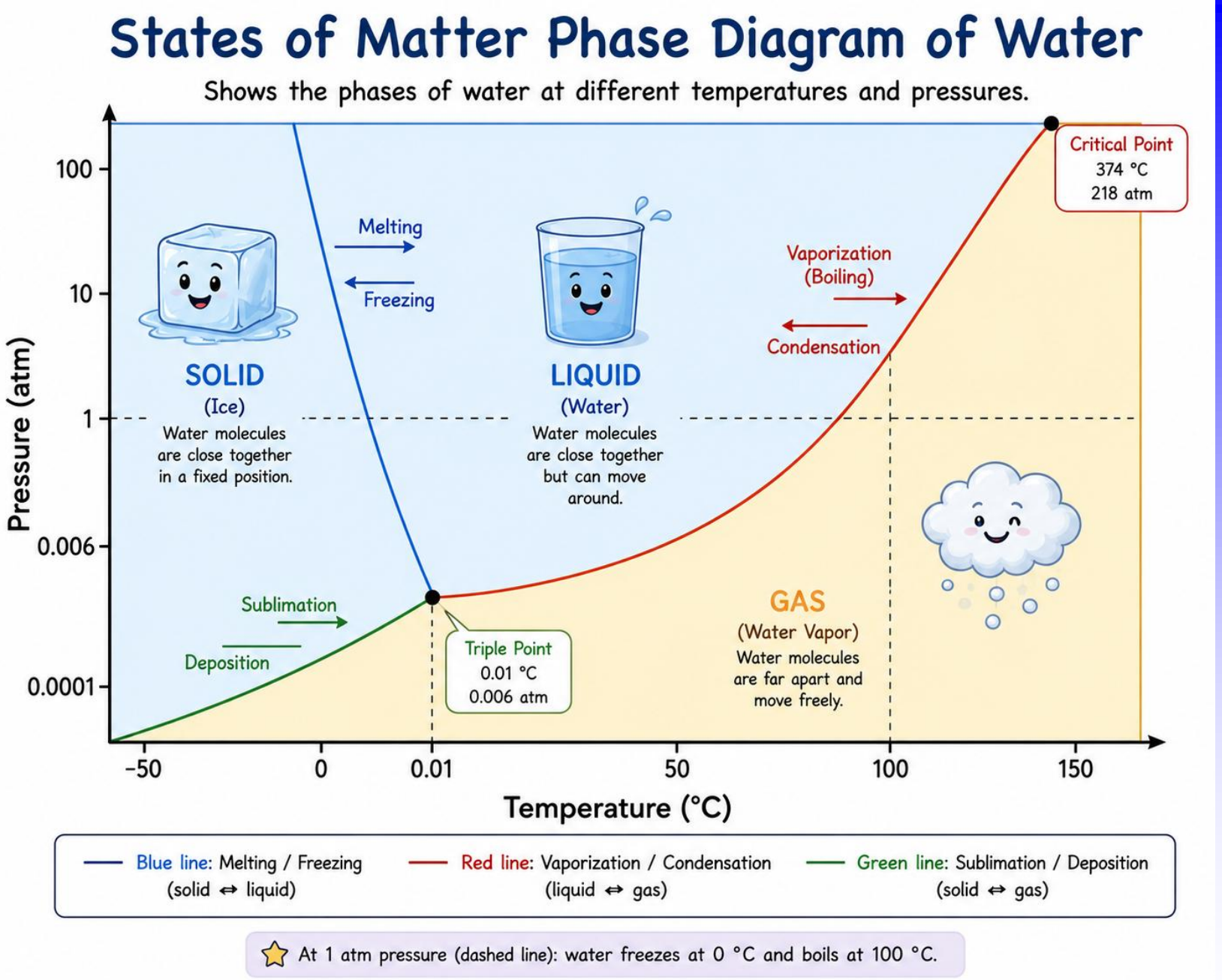
$q = 2100 - J + 5546 - J$

$q = 7646 - J$

$\boxed{q = 7650 - J}$

Phase Diagrams

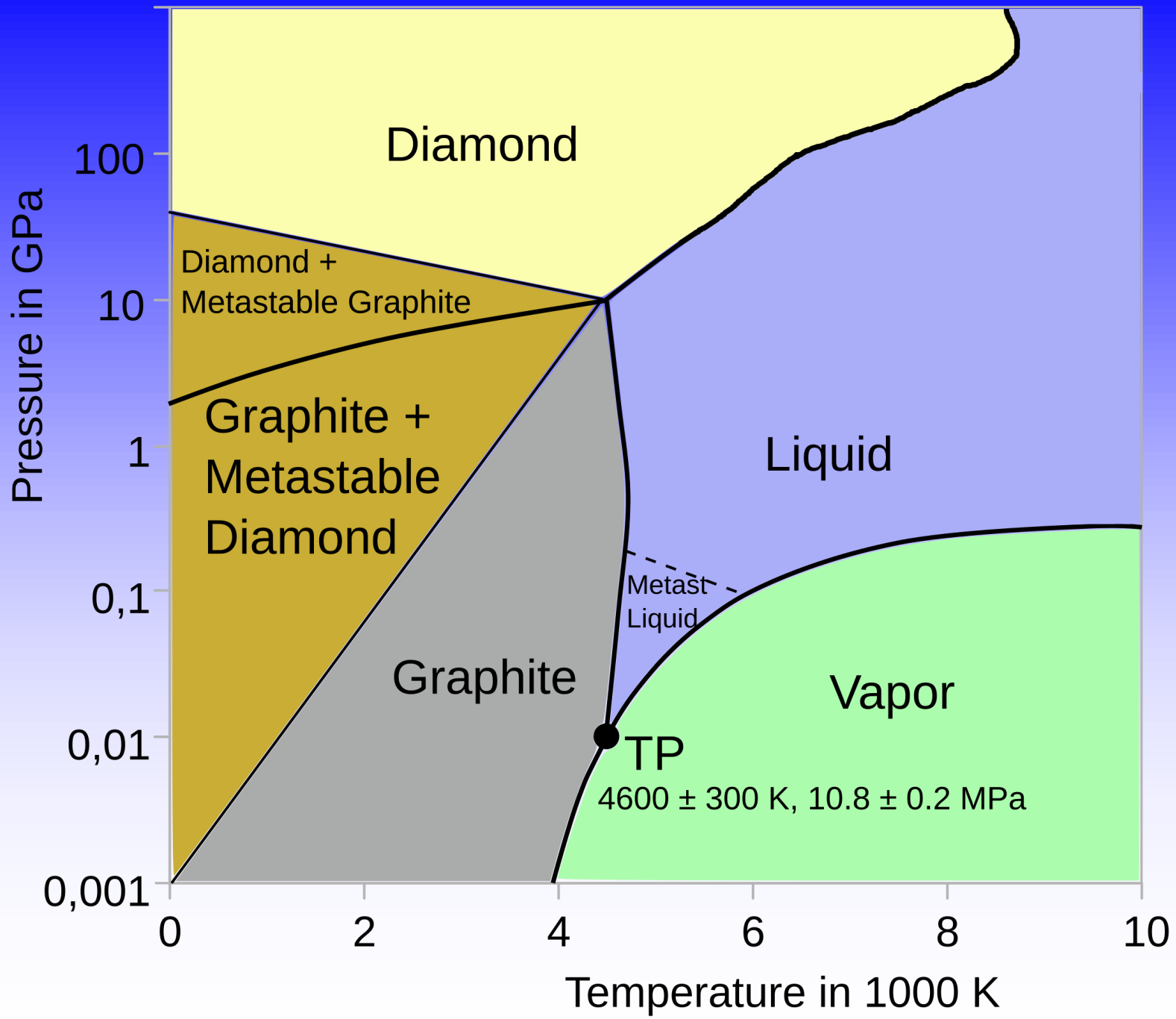
- **Pressure: A force applied over an area.**
- **Kinetic Energy tries to break apart IMF/Metallic bonds.**
- **Pressure can help IMFs**
  - **This can keep phases from changing.**
- **Joseph Louis Gay-Lussac published the work of Jacques Charles and developed a visual look at phases which included the effects of pressure.**



Triple Point: All three phases in equilibrium.

Critical Point: Physical properties of liquid and gas break down.

- **Supercritical Fluid**
- **No Liquefaction**
- **Vanishing Meniscus**



Some materials only have phases in extreme conditions.

Diamond: Extreme pressure/temp (Stable)

Liquid Carbon: Extreme pressure/temp (unstable)

Carbon Vapor: At 3900-K solid carbon sublimates.

Even H_2O has exotic forms!