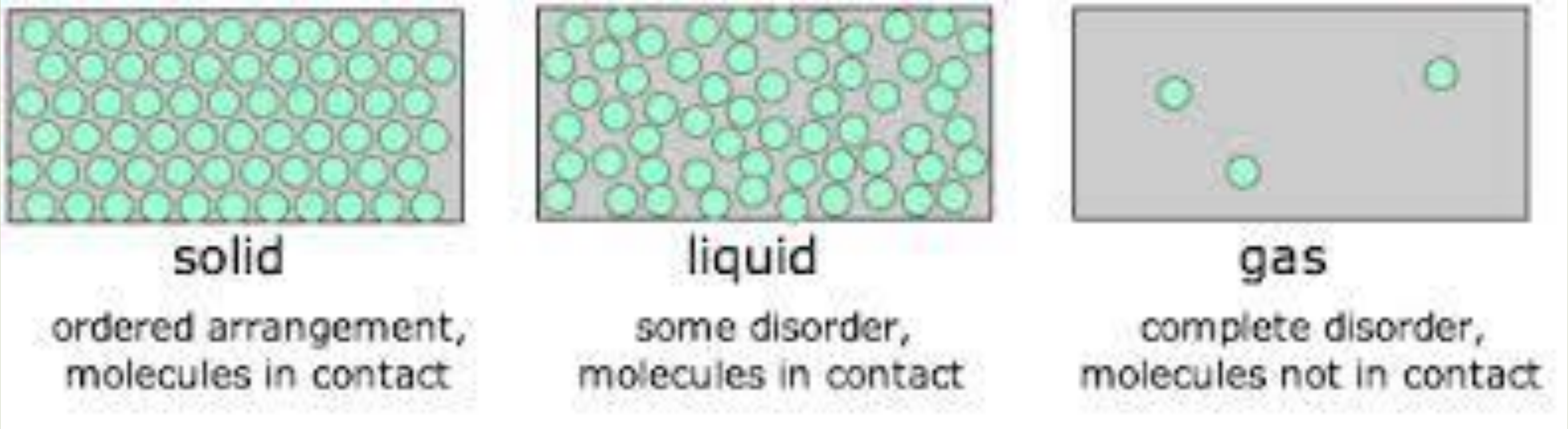
CHAPTER 14: LIQUIDS AND SOLIDS

14-1 CONDENSED STATES OF MATTER

- Condensed State- substances in these states have much higher densities than they do in the gaseous state.

PHYSICAL PROPERTIES OF THE STATES OF MATTER



Occupy their own shape and volume	Occupy the shape of their container but have their own volume	Occupy the shape and volume of their container

KINETIC-MOLECULAR THEORY APPLIES TO LIQUIDS AND SOLIDS ALSO

- According to the kinetic-molecular theory, the state of a substance at room temperature depends on the strength of the attractions between its particles.
- Attractive forces between solids are the strongest

ATTRACTIONS AND PHYSICAL STATE

- Water below $0^{o}C$ is a solid…why?
  - Kinetic energy of the water molecules is too low to overcome the strong attractions between the water molecules.
  - Above $0^{o}C$, molecules have enough kinetic energy to get away from each other and flow.
  - At $100^{o}C$, kinetic energy of the molecules is so high they can escape the container.

*INTRA*MOLECULAR FORCES

- 3 types of chemical bonds.
  - Ionic – metal + nonmetal transfer electrons (all ionic compounds are solids at room temp)
  - Metallic – share electrons (sea of electrons) most are solid at room temp.
  - Covalent – sharing of electrons.

*INTER*MOLECULAR FORCES

- Intermolecular forces are the attractions that affect physical states
- 3 types of intermolecular forces
  - Dispersion
  - Dipole-Dipole
  - Hydrogen bond

DISPERSION (AKA…LONDON FORCES)

- Only type of intermolecular attraction between NONPOLAR molecules…which includes noble gases, BOFINCH molecules, and other nonmetallic elements
- There are no permanent dipoles but there are attractions between ***temporary*** dipoles
- The heavier the molecules the STRONGER the dispersion forces
- The stronger the dispersion force, the higher the boiling point, melting point, evaporation point…etc.

DIPOLE-DIPOLE

- Dipole-Dipole forces – attractions between the permanent dipoles in polar molecules.
- A permanent dipole is a result of a covalent bond between 2 atoms with a > 0.4 difference in electronegativity.
- Since they have different electronegativities, the atom with the higher EN has a $\delta-$ and the lower EN has a $\delta+$
- Opposites attract creating a dipole-dipole attraction

HYDROGEN BOND

- Hydrogen Bonding- Hydrogen has a low EN and F, N, and O have high EN. Because of the large difference in EN, the bonds are *very* polar and there is an extremely strong dipole-dipole force. It's so strong, it's called a H bond.

- Only in molecules with H-F, H-N, or H-O bonds

14-2 PROPERTIES OF LIQUIDS

- The physical properties of liquids are determined by the type and strength of the *intermolecular* forces between their molecules.
- **Viscosity** – resistance to flow. Tells you how easily a liquid will pour.
  - The stronger the attraction between the molecules, the greater the resistance to flow…the higher the viscosity.
  - Viscosity increases as temperature decreases.
- **Surface Tension** –molecules at the surface of a liquid only attract down and sideways they act like there is a film on top of the liquid.

- Water…Why it's so good…
  - Unusually high bp (HF and NH_3 are corrosive gases at room temp)
  - it can absorb or release heat without large changes in temperature (high specific heat)
  - density of ice is less than liquid water
  - high surface tension which allows plants to carry water up from root to tip
  - has a high heat of vaporization (allows cooling with sweating and evaporation)
  - is the universal solvent because it's polar

14-3 THE NATURE OF SOLIDS

- Crystalline solids- a solid with a highly ordered, repeating pattern.
- Unit cells – the smallest possible repeating pattern of a crystal.
- Crystal Formation and Water of hydration
  - Many chemical reactions for ionic substances occur in water. If water gets trapped in the bonds, it is called a hydrate and the water it has in it is known as *water of hydration*

- Bonding in solids – the physical properties of solids, such as hardness, electrical conductivity, and melting point, depend on the kind of particles that make up the solid and the strength of the attractive forces between them.

TYPES OF SOLIDS

- *Metals* – sea of electrons, good conductors because their valence electrons are free to move. (wires) Malleable (sheets) Ductile (wires) electrons can just change shape to adjust
- *Molecular solids* – low mp because only the weak intermolecular forces have to be broken. Do not conduct b/c no free electrons
- *Ionic solids* – make a crystal to maximize attraction between + and -'s and minimize repulsion.

- *Covalent-Network solids* – Atoms are bonded together with strong covalent atoms without forming molecules. Instead they form a network throughout the crystal.

- *Amorphous solids*- appears to be a solid but does not behave like a solid. Glass, rubber, and plastic. Think of them as a super-cooled liquid where the viscosity has become really high.

14-4 CHANGES OF STATE

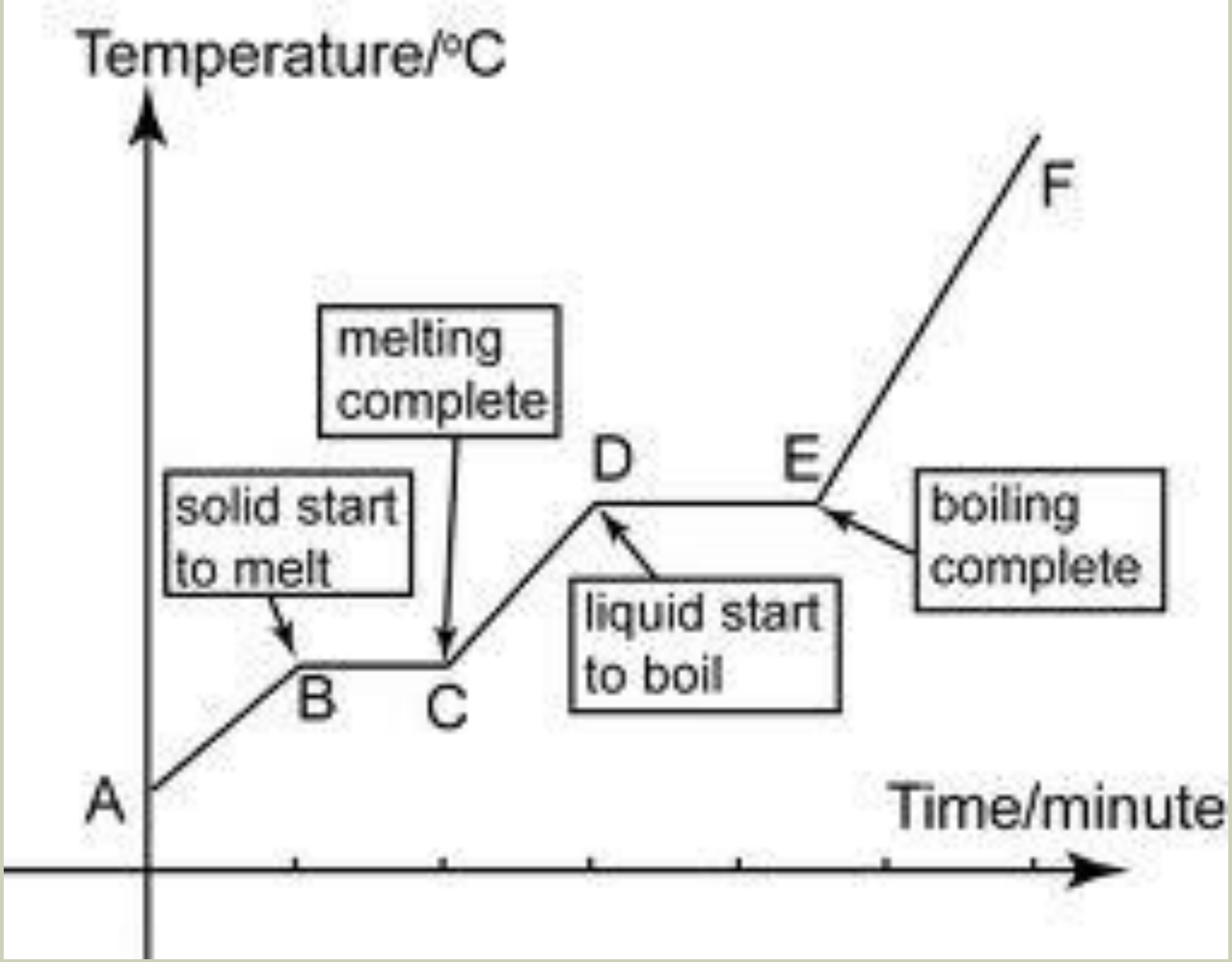
- Conversion of a substance from one of the three physical states of matter to another.

- It **ALWAYS** involves a change of energy.
  - Kinetic energy – movement of particles (temperature)
  - Potential energy – stored energy in bonds or attractions between molecules

ENERGY AND CHANGE OF STATE

- When a substance goes from solid to liquid or liquid to gas, it must overcome the attractive forces holding it together in the more condensed state
  - It must break free!
- When a substance goes from liquid to solid or gas to liquid, it must find attractions to hold it together in a more condensed state

HEATING/COOLING CURVES



VAPORIZATION AND CONDENSATION

- Vaporization = Liquid to gas
- Condensation = gas to liquid

- Evaporation = going from a liquid to a gas at *surface only*
- Vaporization = evaporation *anywhere* in sample aka…boiling

EVAPORATION

- When *surface molecules* have enough energy to fly out of their containers. The higher the temperature, the more evaporation
  - Volatile liquid – one that evaporates easily because the molecules are not strongly attracted to each other (rubbing alcohol)
- Evaporative cooling – molecules that evaporate absorb energy from their surroundings (your skin) and this decreases the energy (and temperature) of the surroundings

LIQUID-VAPOR EQUILIBRIUM

- If you have a closed container with water in it, the water will evaporate at a constant rate, then it will condense and eventually will be in a dynamic equilibrium (a continually changing balance where evaporation and condensation are happening at the same speed)
- Equilibrium Vapor Pressure – As you increase the temp, there will be a greater number of vapor molecules so the pressure will increase.

BOILING AND EQUILIBRIUM VAPOR PRESSURE

- Boiling defined

  Vapor pressure = Atmospheric Pressure

- Why????
  - as water is heated, bubbles of vapor are formed within the liquid
  - they will collapse as long as the atmospheric pressure is greater than the vapor pressure (THEY GET SQUASHED!!)
- Boiling point – the temperature where vapor pressure = atmospheric pressure
- The higher/lower the atmospheric pressure, the higher/lower the boiling point

SUBLIMATION AND DEPOSITION

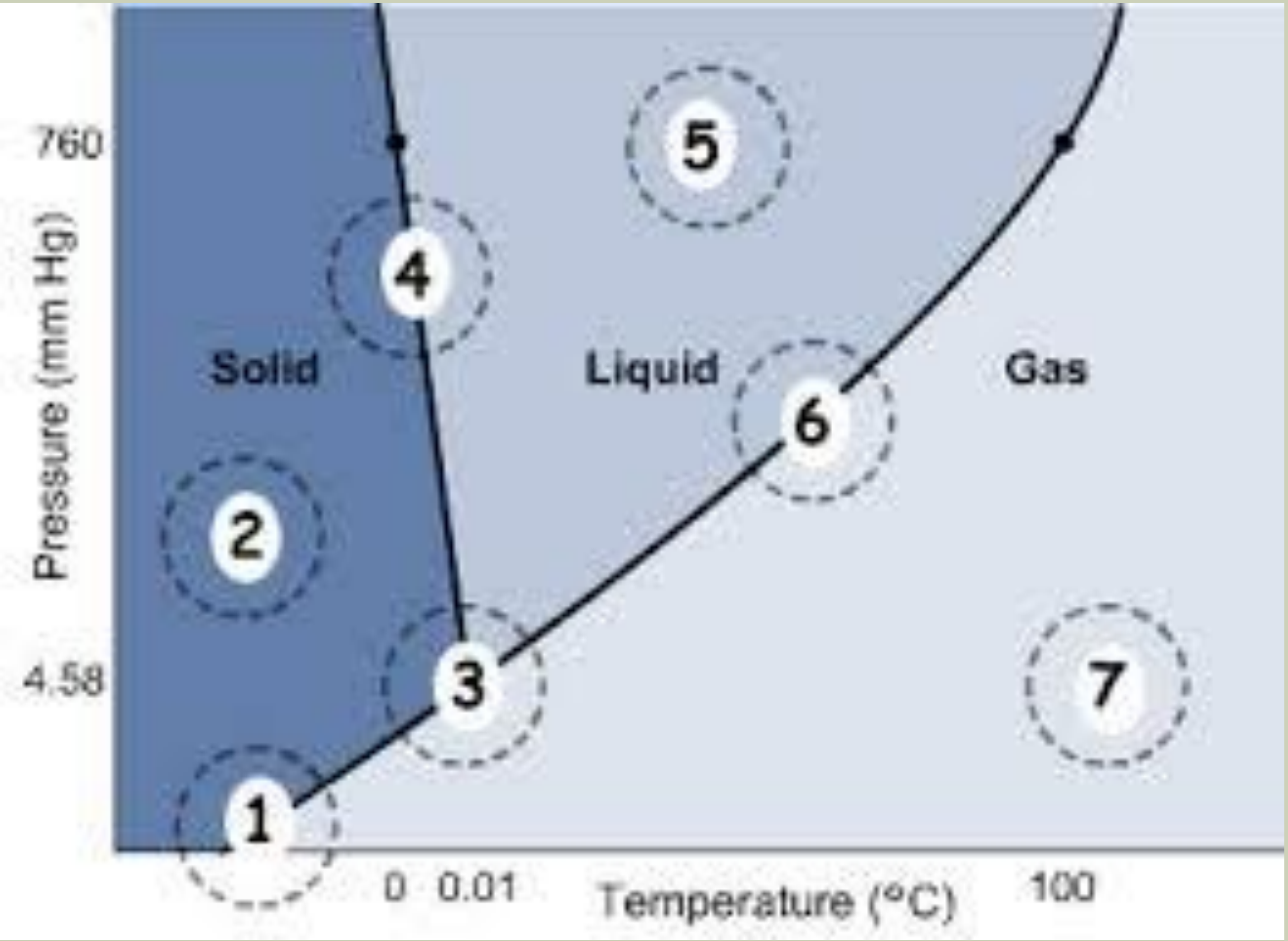
- Solid → gas … sublimation
- Gas → solid … deposition

- These changes never go through the liquid phase!

SUMMARY OF PHASE CHANGES

- Endothermic changes (those that require energy)
  - Melting (solid to liquid)
  - Vaporizing (liquid to gas)
  - Subliming (solid to gas)
- Exothermic changes (those that release energy)
  - Freezing (liquid to solid)
  - Condensing (gas to liquid)
  - Depositing (gas to solid)

PHASE CHANGE DIAGRAMS



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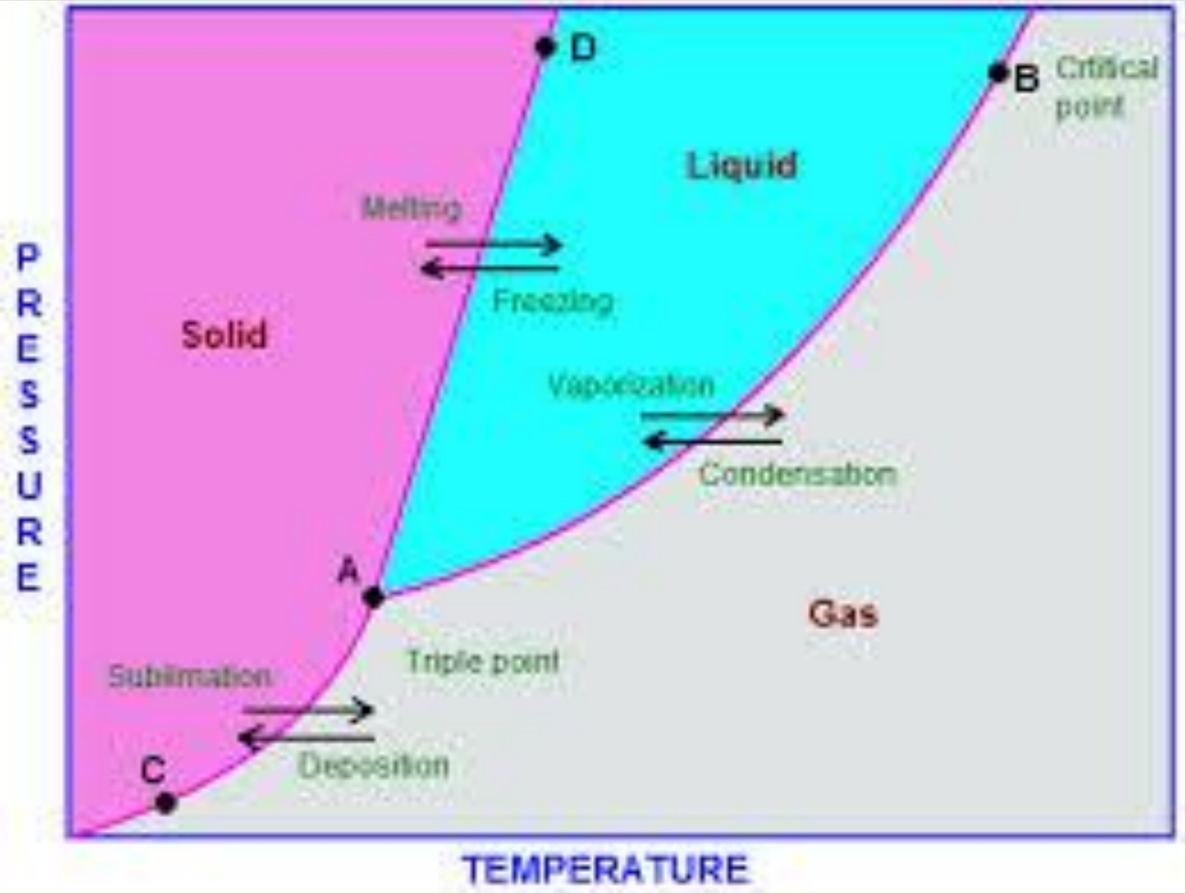

D
B
Critical point
Liquid
Melting
Freezing
Solid
PRESSURE
Vaporization
Condensation
A
Gas
Triple point
Sublimation
Deposition
C
TEMPERATURE