Announcements for Wednesday, 11DEC2024

- End-of-Semester Surveys due **Monday, 16DEC2024, at 11:59 PM (EST)**
- Week 14 Homework Assignments available on eLearning
  - Graded and Timed Quizzes 12a & 12b – "Gases – Intermolecular Forces" due **tonight at 6:00 PM (EST)**
- Final Exam Conflict Requests
  - Due by Friday, 13DEC2024, 11:59 PM
  - See Canvas Announcement from 09DEC2024
- Final Exam is Wednesday, 18DEC2024, 4:00-7:00 PM **(EST)**
  - Coverage: Chapters E-11
  - Exam details and locations will be announced on Canvas

ANY GENERAL QUESTIONS? Feel free to see me after class!

Final Exam Format

- ACS Exam ***Only***
  - 70 multiple-choice questions, 110 minutes
  - **YOU MUST NOT WRITE IN THE EXAM BOOKLET!!**
    - 10-point penalty if you do
  - everything will be collected after 110 minutes
  - see Canvas Announcement from 11DEC2024 for more information

Try This On Your Own

Choose the compound that is most likely a gas at room temperature:

BBr_3	CH_3OH	BF_3	$C_{20}H_{40}Cl_2$	CF_3COOH
(ℓ)	(ℓ)	**(g)**	(s)	(ℓ)

The Critical Point

- **critical point of a liquid =** the temperature and pressure at which the liquid and gas of a substance cease to exist as separate phases
  - the gas and liquid phases transitions into a single, homogeneous phase
- supercritical fluid
  - has properties of both gases and liquids
  - can expand to fill container like a gas but has the density more of a liquid
- **critical temperature (T_c) =** the temperature at which a substances becomes a supercritical fluid
  - above this temperature the supercritical fluid cannot condense into a traditional liquid no matter how much pressure is applied
- **critical pressure (P_c) =** the pressure at which a substance becomes a supercritical fluid
- for water, critical point is at 374 °C and 218 atm
- for CO_2, critical point is at 31.1 °C and 72.9 atm

Phase Changes – Melting vs. Freezing

- melting = (s) $\longrightarrow$ (ℓ)
  - also known as "fusion"
  - IMFs within the solid must be overcome
  - endothermic

- freezing = (ℓ) $\longrightarrow$ (s)
  - exothermic process

- strong intermolecular forces = high melting points and high freezing points
  - H_2O(s) melts at 0 °C
  - CH_4(s) melts at –182 °C

Heat of Fusion (ΔH_{fus})

- **ΔH_{fus}** = the amount of energy needed to convert 1 mole of a substance from a solid to a liquid at its melting point

TABLE 11.9 Heats of Fusion of Several Substances

Liquid	Chemical Formula	Melting Point (°C)	ΔH_{fus} (kJ/mol)
Water	H_2O	0.00	6.02
Rubbing alcohol (isopropyl alcohol)	C_3H_8O	−89.5	5.37
Acetone	C_3H_6O	−94.8	5.69
Diethyl ether	$C_4H_{10}O$	−116.3	7.27

- in general lower melting points indicate weaker intermolecular forces
- at 0 °C, $H_2O(s) \longrightarrow H_2O(\ell)$ $\Delta H = +6.02$ kJ/mol
- for liquid water at 0 °C, $\Delta H_{freezing} = -(\Delta H_{fus}) = -6.02$ kJ/mol

Try This On Your Own

How much heat must be removed to freeze 1.25 kg of water at 0 °C given that ΔH_{fus} = +6.02 kJ/mol? **418 kJ**

$\Delta H_{\mathbf{fus}}$ = **+**6.02 kJ/mol, so $\Delta H_{\mathbf{freezing}}$ = **–**6.02 kJ/mol

$$1.25\ \mathrm{kg\ H_2O} \times \frac{1000\ \mathrm{g}}{1\ \mathrm{kg}} \times \frac{1\ \mathrm{mol\ H_2O}}{18.02\ \mathrm{g}} \times \frac{-6.02\ \mathrm{kJ}}{1\ \mathrm{mol\ H_2O}} = -418\ \mathrm{kJ}$$

418 kJ must be removed

Phase Changes – Sublimation vs. Deposition

- sublimation = (s) $\longrightarrow$ (g)
  - IMFs within the solid must be overcome
  - endothermic
  - sublimation of CO_2
- deposition = (g) $\longrightarrow$ (s)
  - exothermic process

CO_2(s) (dry ice)

- strong intermolecular forces = high sublimation temperatures

A Summary of Phase Changes

$$\text{liquid } (\ell) \underset{\Delta H_{condensation}\ (-)}{\overset{\Delta H_{vaporization}\ (+)}{\rightleftarrows}} \text{gas (g)}$$

$$\text{solid (s)} \underset{\Delta H_{freezing}\ (-)}{\overset{\Delta H_{fusion}\ (+)}{\rightleftarrows}} \text{liquid } (\ell)$$

$$\text{solid (s)} \underset{\Delta H_{deposition}\ (-)}{\overset{\Delta H_{sublimation}\ (+)}{\rightleftarrows}} \text{gas (g)}$$

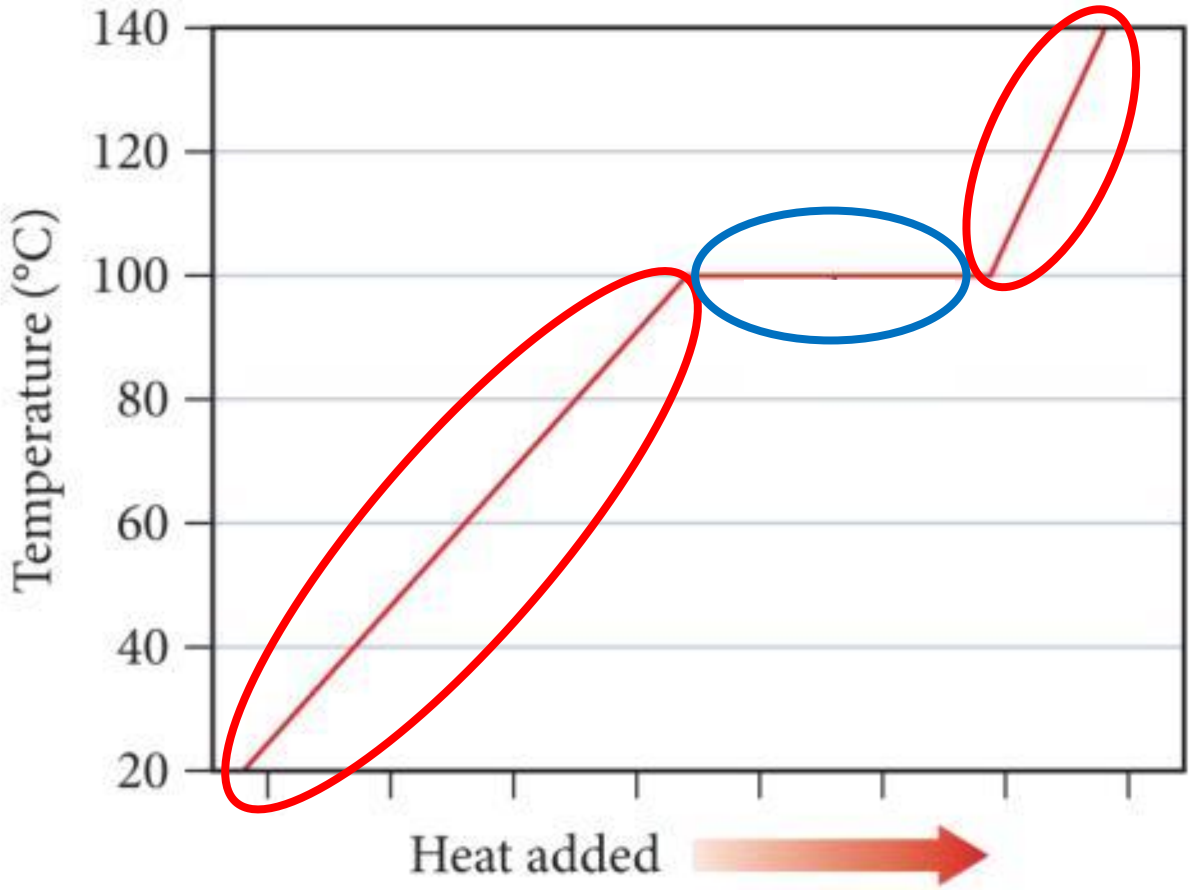
in general when undergoing phase changes, substances exhibiting relatively strong intermolecular forces **require** greater amounts energy (or **release** greater amounts of energy) than substances with weaker intermolecular forces

Heating Curves

- **heating curve =** a plot that shows how the temperature of a substance changes as it is heated up at a constant rate

- as heat is added to a substance, it will be used to do two things

1. increase the temperature of a substance
   - slope (rate of temperature change) depends on the substance's specific heat
   - $q = m\ C_s\ \Delta T$ (Chapter 9)
2. cause phase transitions at specific temperatures
   - for H_2O, melting point at 0 °C and boiling point at 100 °C
   - during phase transitions, the temperature of the substance does **NOT** change as heat is added



Heating Curve of Water

How much heat (kJ) is required to warm 1.00 mole of ice at –50.0 °C to liquid water at 95.0 °C? **15.04 kJ**

For **ice**: C_s = 2.09 J/g·°C, melting point = 0 °C, and ΔH_{fus} = 6.02 kJ/mol

For **liquid water**: C_s = 4.18 J/g·°C

1. warm ice from –50.0 °C to 0 °C (ΔT = 50.0 °C)

   q = (18.0 g)(2.09 J/g·°C)(50.0 °C) = **1.88 kJ**

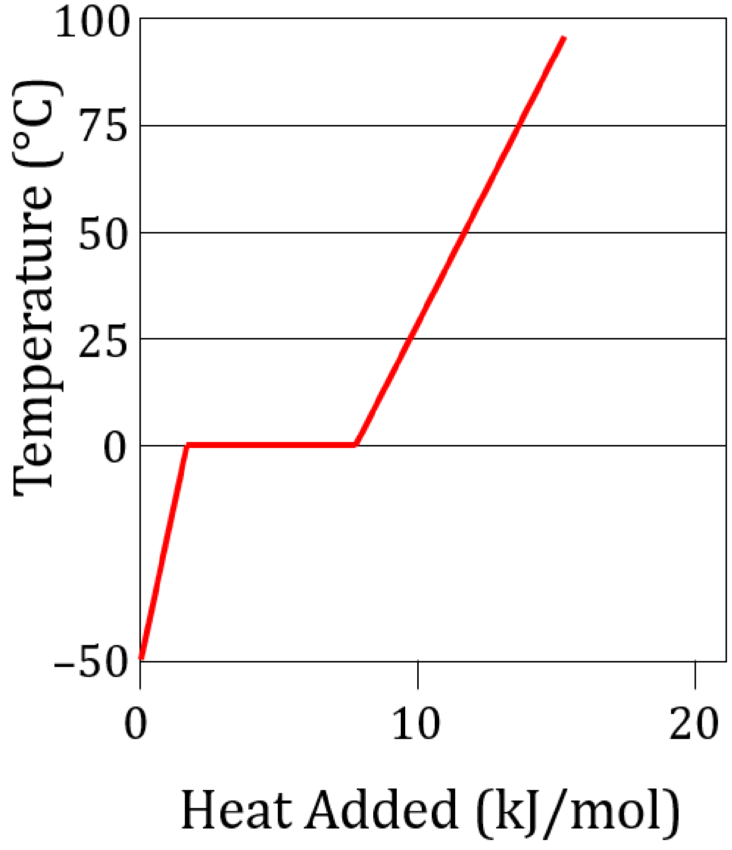
2. convert ice to liquid ($q = n\,\Delta H_{fus}$)

   q = (1.00 mole)(6.02 kJ/mol) = **6.02 kJ**

3. warm liquid from 0 °C to 95.0 °C (ΔT = 95.0 °C)

   q = (18.0 g)(4.18 J/g·°C)(95.0 °C) = **7.14 kJ**

q_{TOTAL} = 1.88 + 6.02 + 7.14 = **15.04 kJ**



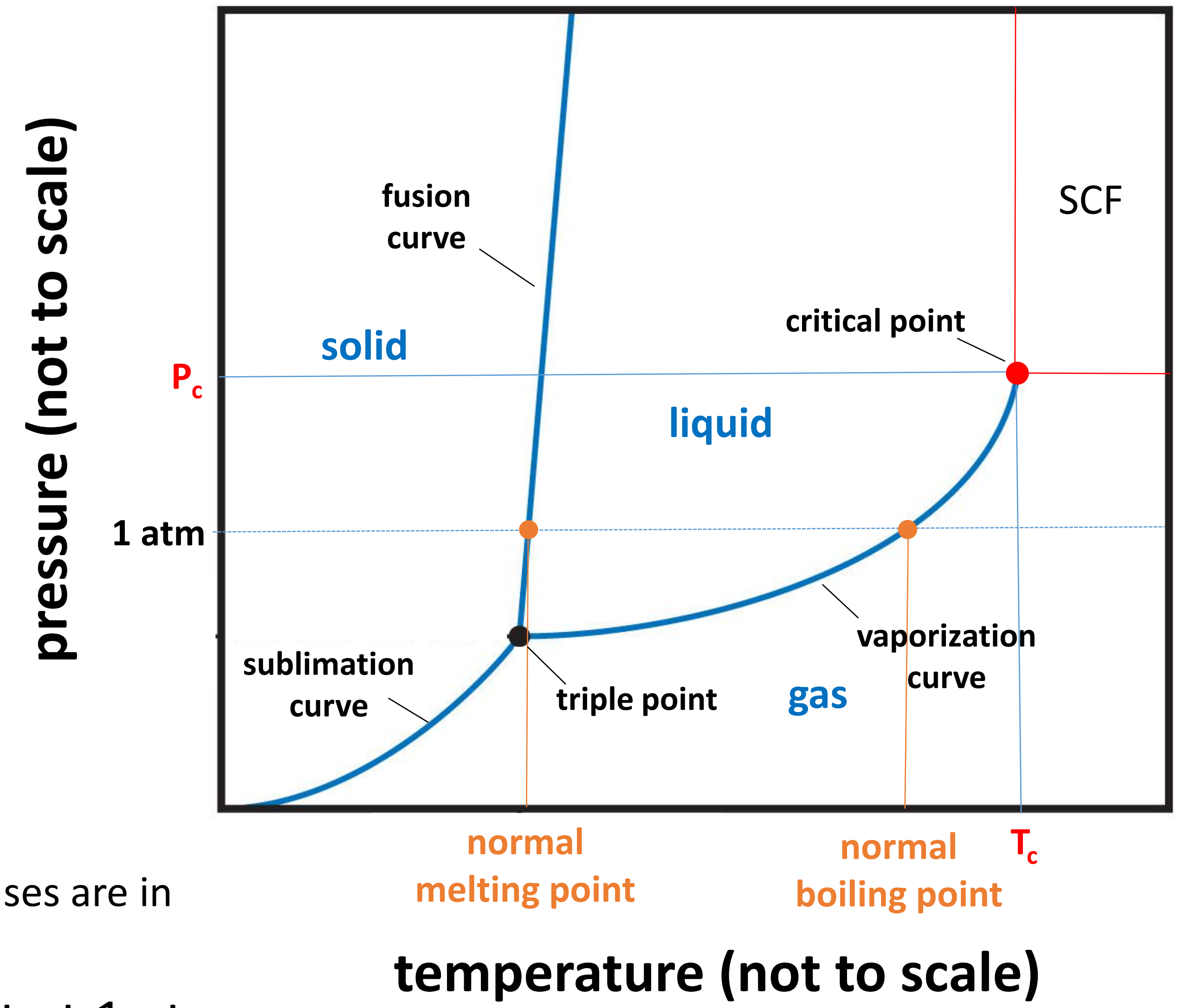
Phase Diagrams

- **phase diagram =** a map of the state of a substance as a function of temperature and pressure
  - we can predict the state of a substance under a given set of conditions
  - we can predict whether a change of state will occur when conditions are changed

- regions vs. lines vs. points on a phase diagram
  - **regions =** areas that represent conditions where a particular state is stable
  - **lines =** sets of temperatures and pressures at which the substance is in equilibrium between two states
  - **points =** unique sets of temperatures and pressures at which specific phenomena occur

Phase Diagram – General Features

- gas phase
  - favored at high temp and low pressure
- liquid phase
- solid phase
  - favored at low temp and high pressure
- sublimation curve
  - solid and gas in equilibrium
- fusion curve
  - solid and liquid in equilibrium
- vaporization curve
  - liquid and gas in equilibrium
- triple point
  - temperature and pressure at which three phases are in equilibrium
- normal melting point and boiling point at 1 atm

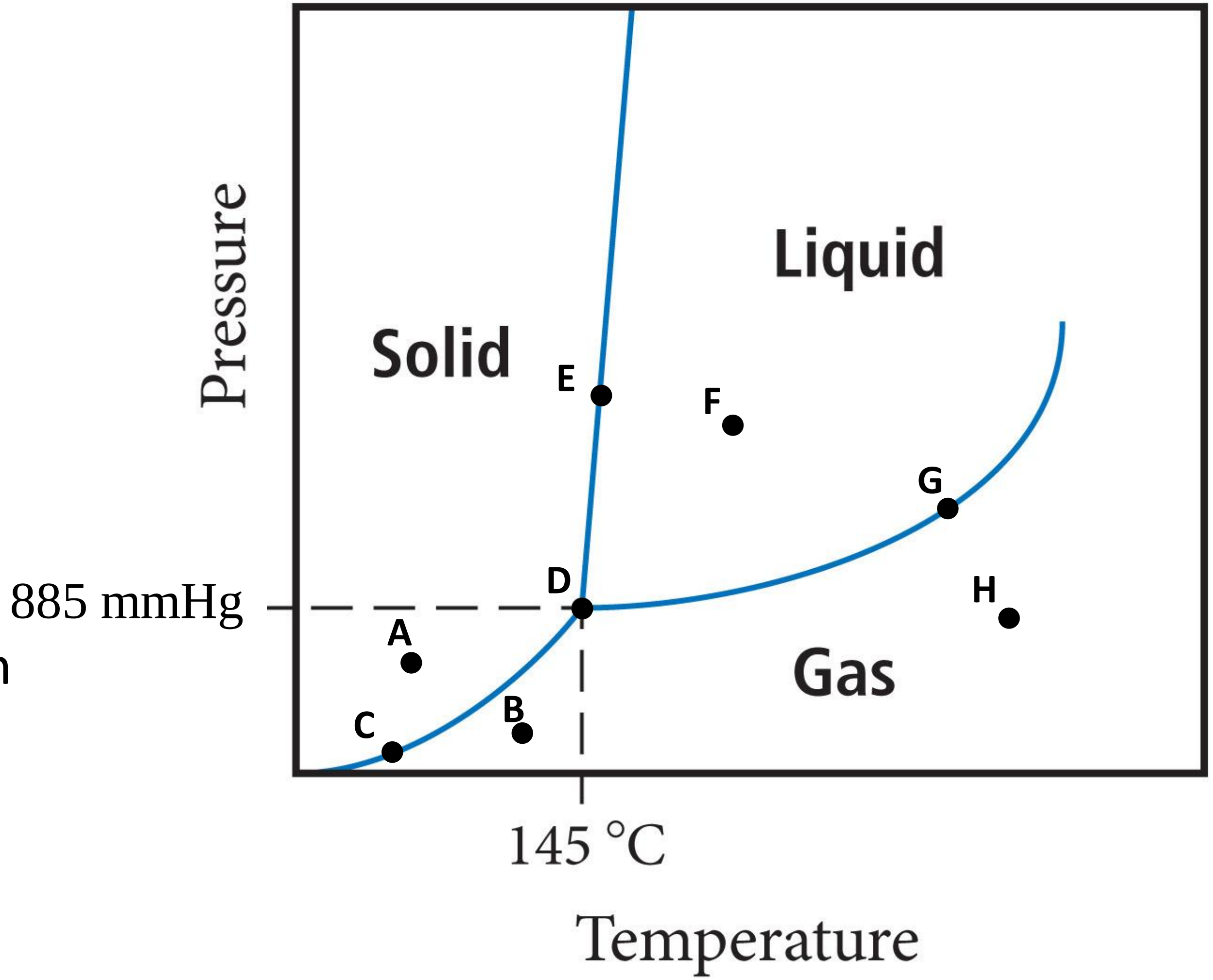


Navigating Within a Phase Diagram

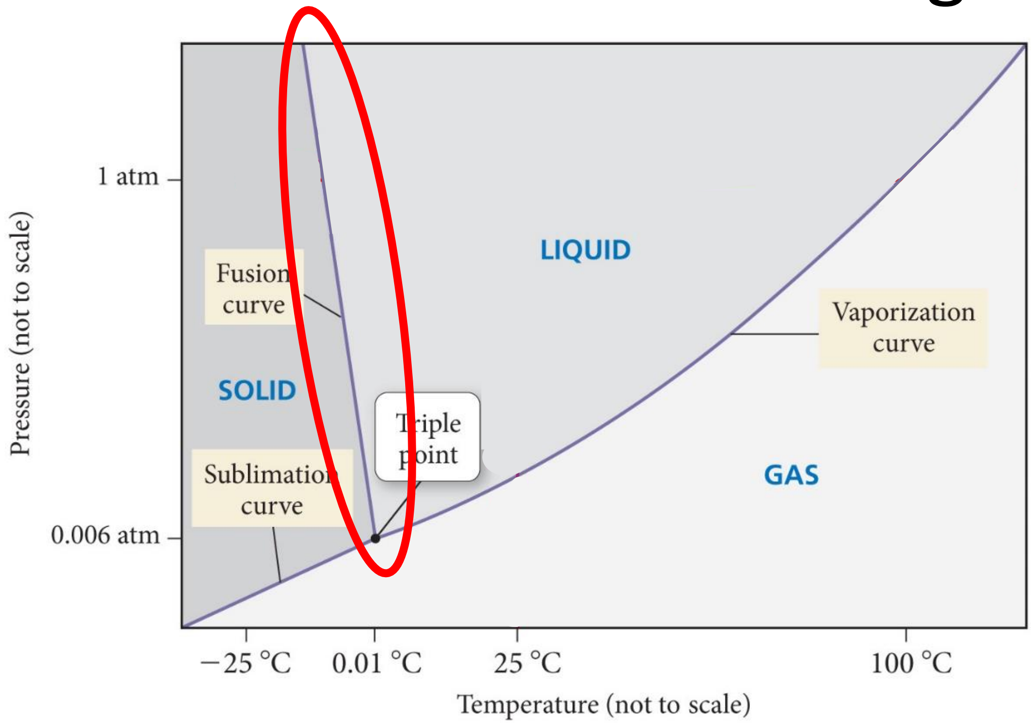
- changing temperature = horizontal movement
- changing pressure = vertical movement

consider the phase diagram shown

- What phase is present at 140 °C and 900 mmHg? **solid**
- What happens when the substance initially at 160 °C and 800 mmHg is put under increased pressure at constant temperature? **condensation**
- Which point has solid in equilibrium with gas? **point C**
- How many points have at least two states in equilibrium? **four**
- Does the substance have a normal melting point or boiling point? **no**
- What is the densest phase for this substance? **solid**

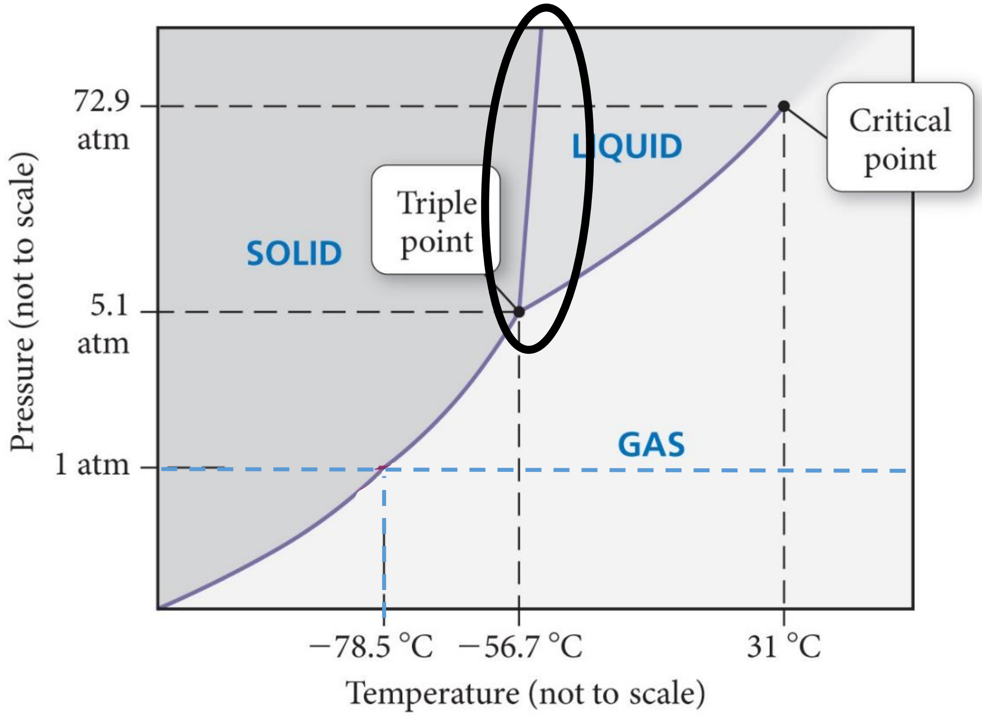


Phase Diagrams – Water vs. CO_2



water

- unusual negative slope of fusion curve
  - the liquid is more dense than the solid
- triple point at 0.01 °C and 0.006 atm
- liquid exists at normal atmospheric pressure



carbon dioxide

- positive slope of fusion curve
  - solid is the most dense phase
- triple point at –56.7 °C and 5.1 atm
- liquid cannot exist at normal atmospheric conditions
  - solid sublimes directly into gas
- critical point at 31 °C and 72.9 atm

Try This On Your Own

$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g) \quad \Delta H_{rxn} = -2044\ kJ$

20.0 g $C_3H_8(g)$ is burned in excess oxygen and all the heat generated is absorbed by a 2.00-kg block of ice at –80 °C? Assuming no heat is lost to the surroundings, determine the temperature and physical state of H_2O that results.

For **<u>ice</u>**: C_s = 2.09 J/g·°C, melting point = 0 °C, and ΔH_{fus} = 6.02 kJ/mol

For **<u>liquid water</u>**: C_s = 4.18 J/g·°C, boiling point = 100 °C, and ΔH_{vap} = 40.7 kJ/mol

H_2O (s) AND $H_2O(\ell)$ at 0 °C

heat released by combustion:

$$20.0\ \mathrm{g\ C_3H_8} \times \frac{1\ \mathrm{mol\ C_3H_8}}{44.09\ \mathrm{g}} \times \frac{-2044\ \mathrm{kJ}}{1\ \mathrm{mol\ C_3H_8}} = 927\ \mathrm{kJ\ released\ by\ combustion\ into\ the\ ice}$$

heat needed to raise temperature of ice from –80 °C to melting point (0 °C):

$$\mathrm{q} = \mathrm{m\ C_s\ \Delta T} = (2000\ \mathrm{g})(2.09\ \mathrm{J/g \cdot {}^\circ C})(0\ {}^\circ\mathrm{C} - (-80\ {}^\circ\mathrm{C})) = 3.34 \times 10^5\ \mathrm{J} = 334\ \mathrm{kJ}$$

heat needed to <u>completely</u> melt ice at 0 °C:

$$\mathrm{q} = \mathrm{n}\ \Delta\mathrm{H_{fus}} = \left(\frac{2000\ \mathrm{g}}{18.02\ \mathrm{g/mol}}\right)(6.02\ \mathrm{kJ/mol}) = 668\ \mathrm{kJ}$$

$$\mathrm{total\ heat\ required} = 334\ \mathrm{kJ} + 668\ \mathrm{kJ} = 1002\ \mathrm{kJ;\ but\ only\ 927\ kJ\ available}$$

there is enough heat to bring ice to 0 °C but not enough to completely convert all ice into water at 0 °C