Critical Temperature Explained

https://gemini.google.com/app/12555f47806972e1

Response: Based on the principles of phase equilibria and the properties of real gases, here is the breakdown of the statements:

Correct Answer

(4) It is the highest temperature at which the gas phase and the liquid phase are in equilibrium.

Detailed Explanation

- **Critical Temperature (T_c):** The critical temperature of a substance is the highest temperature at which a gas can be liquefied by applying pressure. Above T_c, the kinetic energy of the molecules is so high that no amount of pressure can force the gas into a liquid state (it becomes a supercritical fluid). Therefore, it marks the upper boundary where distinct gas and liquid phases can coexist in equilibrium.

Analysis of Incorrect Options:

- **(1)** The temperature at which intermolecular forces balance out and real gas behavior closely mimics ideal gas behavior is known as the **Boyle temperature**, not the critical temperature.
- **(2)** T_c corresponds to the minimum pressure required for liquefaction at that maximum temperature (known as the critical pressure, P_c), but it is not defined as the temperature for the overall lowest pressure of liquefaction.
- **(3)** The temperature and pressure at which gas, liquid, and solid are all in equilibrium is called the **triple point**.
- **(5)** The van der Waals equation can be used to derive T_c ($T_c = \frac{8a}{27Rb}$), but critical temperature is a specific unique constant for a gas, not a temperature calculated "at any pressure."
