Lectures: Physics 3306

Provides an introduction to a wide variety of topics in classical (pre-quantum) physics as a bridge to prepare students for subsequent upper-level courses in physics. The topics covered include thermodynamics, fluid mechanics, mechanical waves, optics, radiation, electromagnetic phenomena, atoms, and laboratory techniques. Prerequisites: C- or better in PHYS 1106; and in PHYS 1304 or PHYS 1308.

Saptaparna Bhattacharya

March 30th, 2026

Based on Simon Dalley's lectures taught in Spring 2025

Labs

Lectures

Schedule

No class

Month	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday	Sunday
January	19	20	21 ✅	22	23 ✅	24	25
	26 ❄️🌨️❄❄❄	27	28 ❄️🌨️❄❄❄	29	30 ✅	31	1
February	2 ✅	3	4 ✅	5	6 ✅	7	8
	9 ✅	10	11 HWB due ✅	12	13 ✅	14	15
	16 ✅	17	18 ✅	19	20 HWC due ✅	21	22
	23 Hegi Center ✅	24	25 HWD due ✅	26	27 ✅	28	1
March	2 ✅	3	4 HWE due	5	6 ✅	7	8
	9 ✅	10	11	12	13 Midterm	14	15
	16	17	18	19	20	21	22
	23 ✅	24	25	26	27 ✅	28	29
April	30 Lecture 11	31	1 HWF due	2	3	4	5

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April	6 Midterm 2	7	8 HWG due	9	10 Lecture 13	11	12
	13 Lecture 14	14	15 HWH due	16	17 Lecture 15	18	19
	20 Lecture 16	21	22 HWI due	23	24 Lecture 17	25	26
May	27 Lecture 18	28	29 HWJ due	30	1 Lecture 19	2	3
	4 Lecture 20	5	6	7	8	9	10

Key concepts: Entropy

- An irreversible process is one that cannot be reversed by means of small changes in the environment
- The direction is which an irreversible process proceeds is set by the change in entropy ΔS of the system undergoing the process
- Entropy S is a state property (or state function) of the system and it depends
  - on the state of the system
    - not in the way in which the system reached that state
  - The entropy postulate states:
    - If an irreversible process occurs in a closed system, the entropy of the system always increases
  - The entropy change ΔS for an *irreversible* process that takes a system from an initial state i to a final state f is exactly equal to the entropy change ΔS for any *reversible* process that takes the system between those same two states
    - Reversible $\Delta S = S_f - S_i = \int_i^f \frac{dQ}{T}$, where Q is the energy transferred

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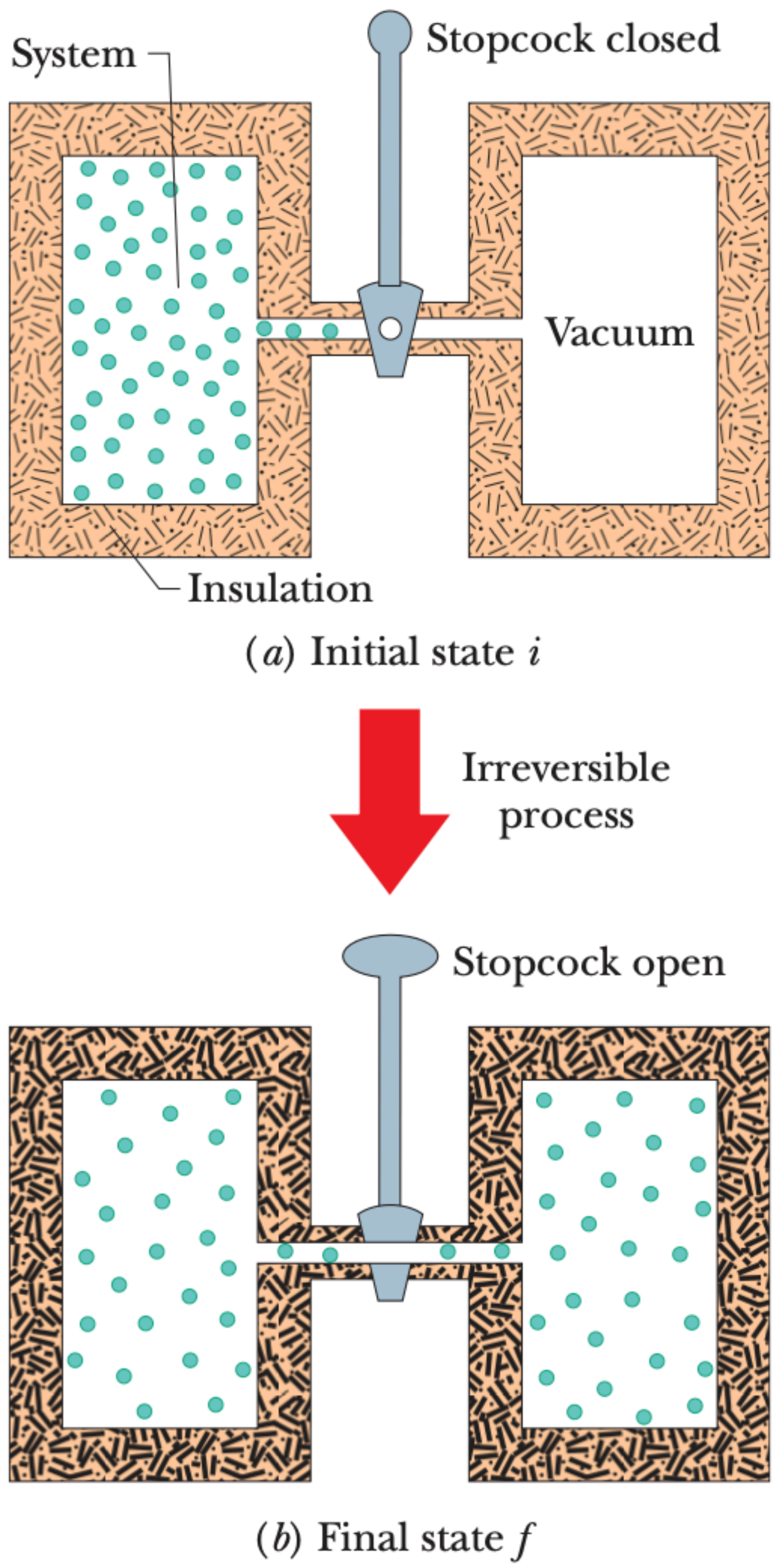
- For a reversible isothermal process, the expression for an entropy change reduces to:
  - $\Delta S = S_f - S_i = \frac{Q}{T}$
- When the temperature change ΔT of a system is small relative to the temperature (in kelvins) before and after the process, the entropy change can be approximated as:
  - $\Delta S = S_f - S_i \approx \frac{Q}{T_{\text{avg}}}$

Key concepts: Entropy

- When an ideal gas changes reversibly from
  - an initial state with temperature T_i and volume V_i
  - an initial state with temperature T_f and volume V_f
  - Change in entropy:
  - $$\Delta S = S_f - S_i = nR \ln \frac{V_f}{V_i} + nC_V \ln \frac{T_f}{T_i}$$
- The second law of thermodynamics, which is an extension of the entropy postulate, states: If a process occurs in a closed system, the entropy of the system increases for irreversible processes and remains constant for reversible processes. It never decreases
  - $\Delta S \geq 0$

Key concepts: Change in entropy

- Change in entropy:
  - $\Delta S = S_f - S_i = \int_i^f \frac{dQ}{T}$
- For isothermal expansion:
  - $\Delta S = S_f - S_i = \frac{1}{T} \int_i^f dQ$
  - $\int_i^f dQ = Q$
  - $\Delta S = S_f - S_i = \frac{Q}{T}$
- To find the entropy change for an irreversible process, replace that process with any reversible process that connects the same initial and final states. Calculate the entropy change for this reversible process



(*a*) Initial state *i*

(*b*) Final state *f*

Figure 20-1 The free expansion of an ideal gas. (*a*) The gas is confined to the left half of an insulated container by a closed stopcock. (*b*) When the stopcock is opened, the gas rushes to fill the entire container. This process is irreversible; that is, it does not occur in reverse, with the gas spontaneously collecting itself in the left half of the container.

Which of the following has the most entropy?

a) An empty sports stadium

b) A flock of birds flying in V formation

c) A crowd of people on the streets of New York

d) A marble inside a box

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a) An empty sports stadium: This is a highly ordered state. There is no movement or chaos because there is nothing there to move around. (Low entropy)

b) A flock of birds flying in V formation: While the birds are moving, they are flying in a highly structured, predictable, and synchronized pattern. (Low entropy)

c) A crowd of people on the streets of New York: This represents chaos. People are walking in different directions, at different speeds, stopping, crossing streets, and interacting randomly. There is a massive amount of unpredictability and disorder. (Highest entropy)

d) A marble inside a box: This is a very simple system with just one object. Even if it's rolling around, it has far fewer possible random arrangements than a massive crowd of people. (Low entropy)

What is the equation for entropy of a system if two parts 1 and 2 having entropies S1 and S2 are considered in equilibrium?

a. $S = S_1 – S_2$

b. $S = S_1 + S_2$

c. $S = (S_1 + S_2) / 2$

d. $S = \sqrt{}S_1\ S_2$

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Entropy (S) is an extensive property.
Because it measures the total disorder or number of possible microstates in a system, adding more matter or combining systems means you are adding their individual amounts of disorder together. Therefore, the total entropy of a composite system is simply the sum of the entropies of its macroscopic parts.

Key concepts: Entropy as a state function

- First law of thermodynamics:
  - $dE_{\text{int}} = dQ - dW$
  - $dQ = pdV + nC_V dT$
  - $\dfrac{dQ}{T} = nR\dfrac{dV}{V} + nC_V\dfrac{dT}{T}$
  - $\displaystyle\int_i^f \frac{dQ}{T} = \int_i^f nR\frac{dV}{V} + \int_i^f nC_V\frac{dT}{T}$
  - $\Delta S = S_f - S_i = nR\ln\dfrac{V_f}{V_i} + nC_V\ln\dfrac{T_f}{T_i}$

Which thermodynamic process can always be plotted accurately on thermodynamic coordinates?

a. reversible process

b. irreversible process

c. open system

d. closed system

e. none of the above

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To plot a process accurately on thermodynamic coordinates (like a Pressure-Volume or P-V diagram), you must know the exact state of the system—its pressure, temperature, and volume—at every single moment during the change.

- **Reversible Processes:** These are "quasi-static," meaning they happen infinitely slowly. Because they happen so slowly, the system is always in a state of thermodynamic equilibrium. This allows you to know its exact properties at every point and plot a solid, continuous line.
- **Irreversible Processes:** These happen rapidly or involve things like friction or turbulence. Because the change is sudden or chaotic, the system is *not* in equilibrium during the transition. The properties (like pressure or temperature) might be different in different parts of the system, meaning you cannot define a single state to plot accurately. (These are usually just represented by dashed lines connecting the known start and end points).
- **Open/Closed Systems:** These terms describe whether mass can cross the system's boundaries, not how the process itself unfolds.

Which of the following are true for thermodynamics?

a) for an isolated (closed) system, $dS \geq 0$

b) for a reversible process in any system, $dS = 0$

c) for an irreversible process in any system, $dS > 0$

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Here is a breakdown of why option A is correct and the others are false:

- **a) for an isolated system, $dS \geq 0$ (TRUE):** This is the fundamental definition of the Second Law of Thermodynamics. Because an isolated system cannot exchange heat or matter with its surroundings, its total entropy can never decrease. It will either remain constant (if perfectly reversible, $dS=0$) or increase (if irreversible, $dS>0$). The universe itself is often considered an isolated system, which is why we say the entropy of the universe is always increasing.
- **b) for a reversible process in any system, $dS=0$ (FALSE):** The entropy of the *universe* (system + surroundings) is zero during a reversible process. However, the entropy of the *system itself* can absolutely change. For example, if you reversibly melt ice by adding heat infinitesimally slowly, the entropy of the ice (the system) increases, while the entropy of the surroundings decreases by the exact same amount.
- **c) for an irreversible process in any system, $dS>0$ (FALSE):** Similar to the previous point, in an irreversible process, the entropy of the *universe* must strictly increase ($dS_{universe} > 0$). But the specific *system* you are looking at can lose entropy. Think of a hot cup of coffee cooling down in a room: the coffee (the system) loses heat and its entropy goes down, but it heats up the air around it, increasing the room's entropy by an even larger amount.

Rank the entropy changes of water heated on a stove as its temperature rises

1)From 20 ºC to 25 ºC
2)From 30 ºC to 35 ºC
3)From 80 ºC to 85 ºC

a) 1 > 2 > 3

b) 2 > 3 > 1

c) 1 > 3 > 2

d) 3 > 2 > 1

e) 3 > 1 > 2

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d) 3 > 2 > 1

e) 3 > 1 > 2

The correct answer is **a) 1 > 2 > 3**. Here is a breakdown of why:

The fundamental equation for a change in entropy (dS) when heat (dQ) is added to a system is:

$dS=TdQ$
(Where T is the absolute temperature in Kelvin).

Notice that temperature (T) is in the denominator. This tells us a crucial rule of thermodynamics: **adding a specific amount of heat to a cold system causes a larger increase in entropy than adding that exact same amount of heat to a hot system.** Let's look at your specific scenario:

- In all three cases, you are heating the water by exactly 5◦C (or 5 Kelvin).
- Because the temperature change is identical, the amount of heat energy (dQ) you are adding is basically the same for each step.
- Since dQ is constant, the only thing changing is the temperature (T) at which that heat is added.
- Because the water in scenario 1 (20 °C to 25 °C) is at the lowest starting temperature, adding that heat creates the most significant relative increase in disorder. In scenario 3 (80 °C to 85 °C), the water is already very hot and highly disordered, so adding that same amount of heat makes a smaller relative impact.

The Mathematical Proof

If you prefer to see the math, the formula for macroscopic entropy change when heating a liquid is:

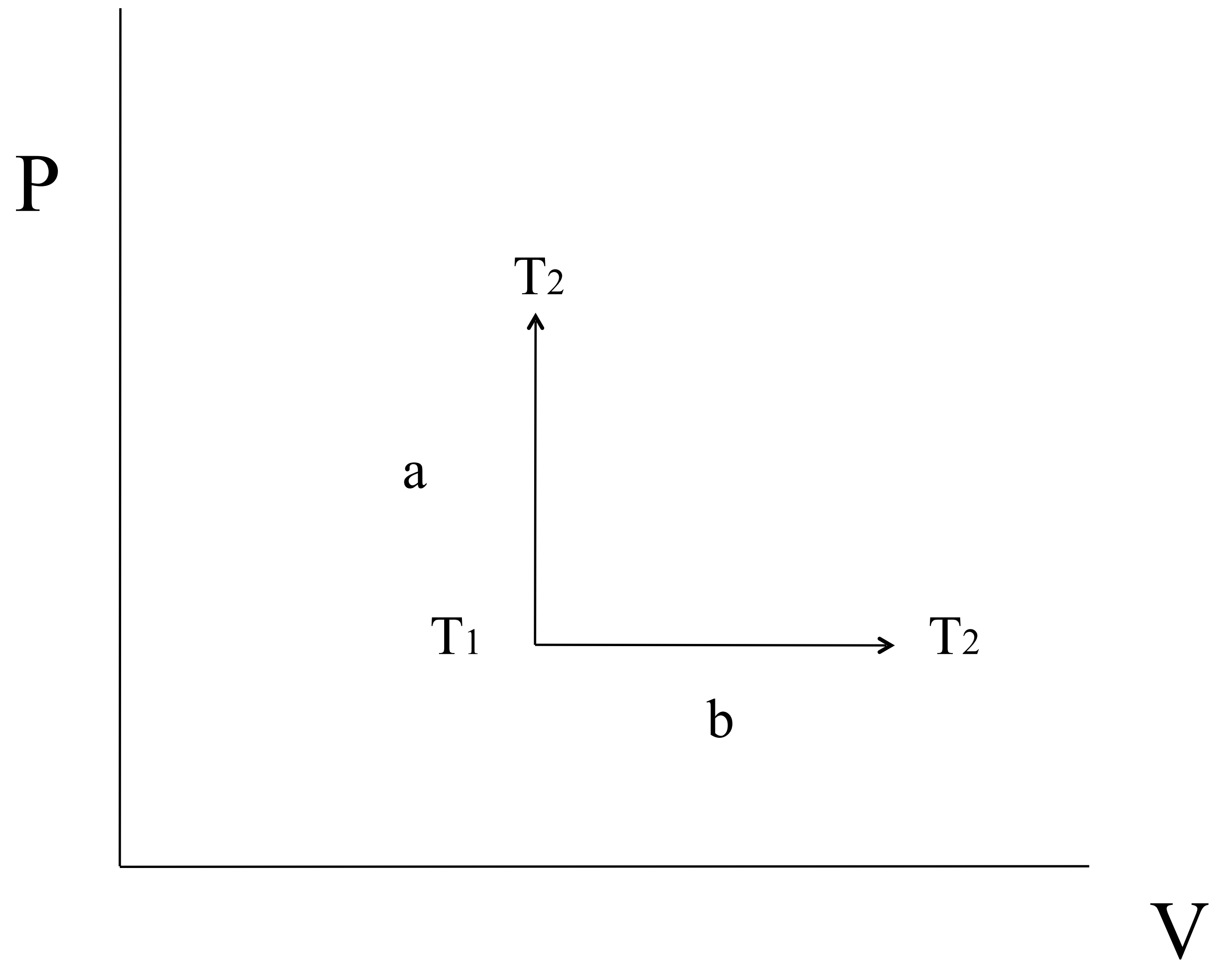
$\Delta S=mCln(T1 /T2)$
(Note: Temperatures must be in Kelvin)

If we look at the ratio of T1/ T2 for each:

1. **20 to 25 °C** (293.15 K to 298.15 K): 293.15/298.15 =1.017
2. **30 to 35 °C** (303.15 K to 308.15 K): 303.15/308.15 =1.016
3. **80 to 85 °C** (353.15 K to 358.15 K): 353.15/358.15 =1.014

Because the natural log (ln) increases as the ratio increases, the largest ratio gives the largest entropy change. Therefore, 1>2>3.

The P-V diagram shows two paths for an ideal gas, a and b, from the temperature T_1 to T_2. Rank the paths according to the entropy change of the gas, greatest first.



1.a > b

2.b > a

3.a = b

4.Not enough information

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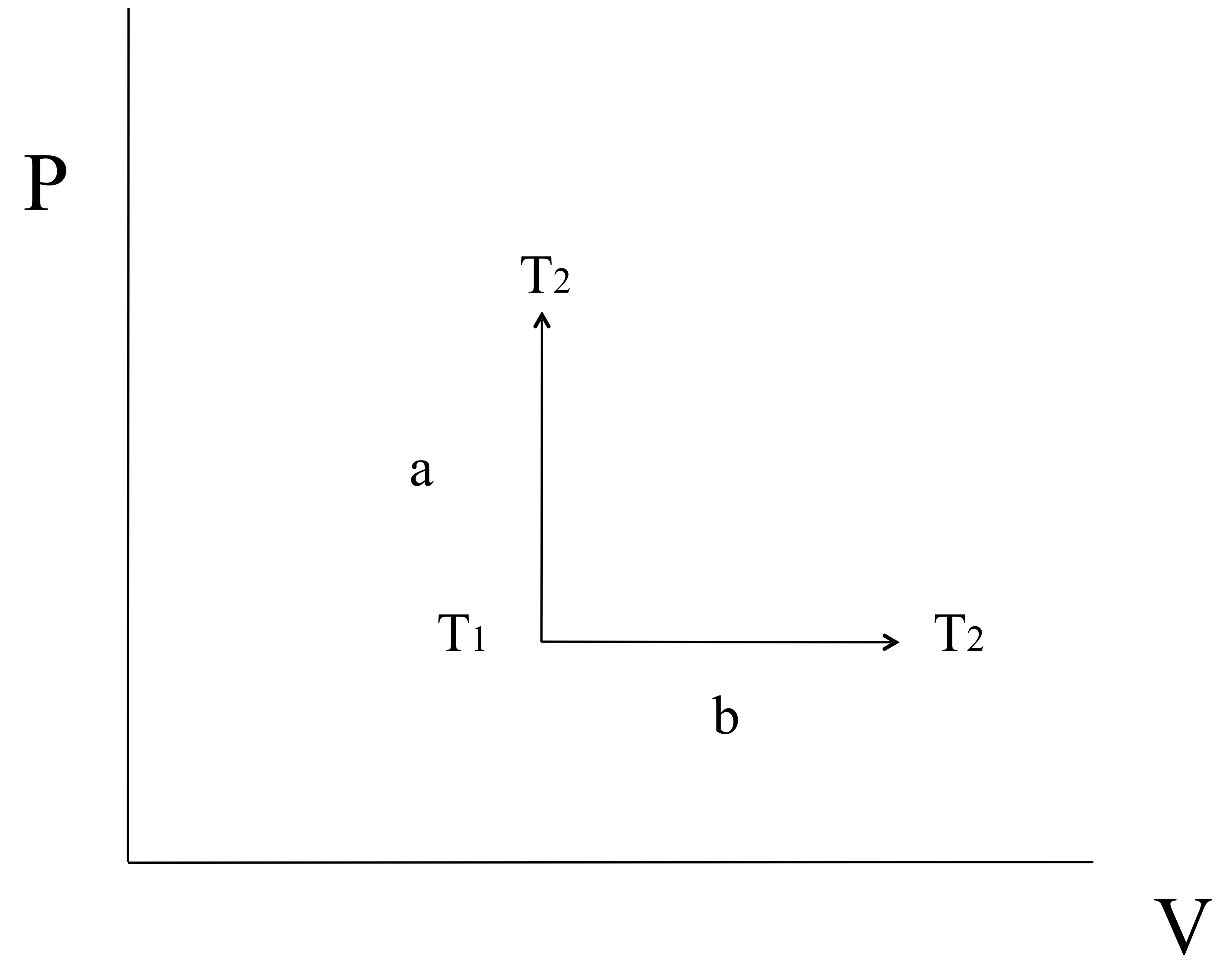
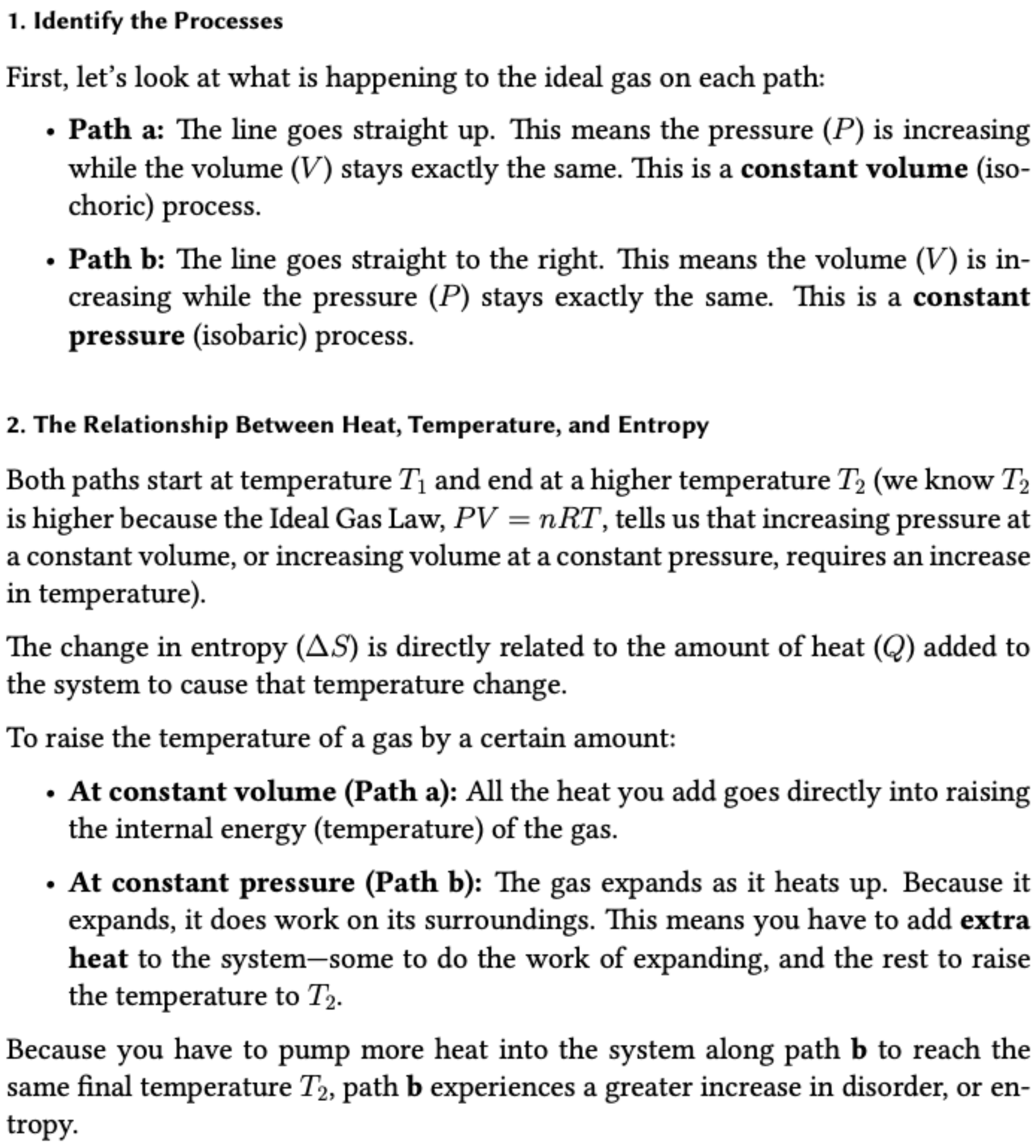


1. a > b
2. b > a
3. a = b
4. Not enough information



1. Identify the Processes

First, let's look at what is happening to the ideal gas on each path:

- **Path a:** The line goes straight up. This means the pressure (P) is increasing while the volume (V) stays exactly the same. This is a **constant volume** (isochoric) process.
- **Path b:** The line goes straight to the right. This means the volume (V) is increasing while the pressure (P) stays exactly the same. This is a **constant pressure** (isobaric) process.

2. The Relationship Between Heat, Temperature, and Entropy

Both paths start at temperature T_1 and end at a higher temperature T_2 (we know T_2 is higher because the Ideal Gas Law, $PV = nRT$, tells us that increasing pressure at a constant volume, or increasing volume at a constant pressure, requires an increase in temperature).

The change in entropy (ΔS) is directly related to the amount of heat (Q) added to the system to cause that temperature change.

To raise the temperature of a gas by a certain amount:

- **At constant volume (Path a):** All the heat you add goes directly into raising the internal energy (temperature) of the gas.
- **At constant pressure (Path b):** The gas expands as it heats up. Because it expands, it does work on its surroundings. This means you have to add **extra heat** to the system—some to do the work of expanding, and the rest to raise the temperature to T_2.

Because you have to pump more heat into the system along path **b** to reach the same final temperature T_2, path **b** experiences a greater increase in disorder, or entropy.

If you prefer to see this using thermodynamics equations, we can look at the formulas for entropy change where n is the number of moles:

- **For Path a (constant volume):**

$$\Delta S_a = nC_v \ln\left(\frac{T_2}{T_1}\right)$$

(Where C_v is the molar heat capacity at constant volume)

- **For Path b (constant pressure):**

$$\Delta S_b = nC_p \ln\left(\frac{T_2}{T_1}\right)$$

(Where C_p is the molar heat capacity at constant pressure)

For any ideal gas, it is a fundamental rule that $C_p > C_v$ (specifically, $C_p = C_v + R$).

Since both equations share the exact same $n \ln\left(\frac{T_2}{T_1}\right)$ term, the equation multiplied by the larger heat capacity (C_p) will result in the larger entropy change. Therefore, $\Delta S_b > \Delta S_a$.

Rank these in terms of multiplicity W

1) Getting 10 heads total when I throw 10 coins
2) Getting 6 dots total when I throw 6 dice.
3) Getting 2 tails total when I throw 5 coins

a)1 = 2 = 3

b)1 > 2 > 3

c)2 > 3 > 1

d)3 > 2 > 1

e)3 > 2 = 1

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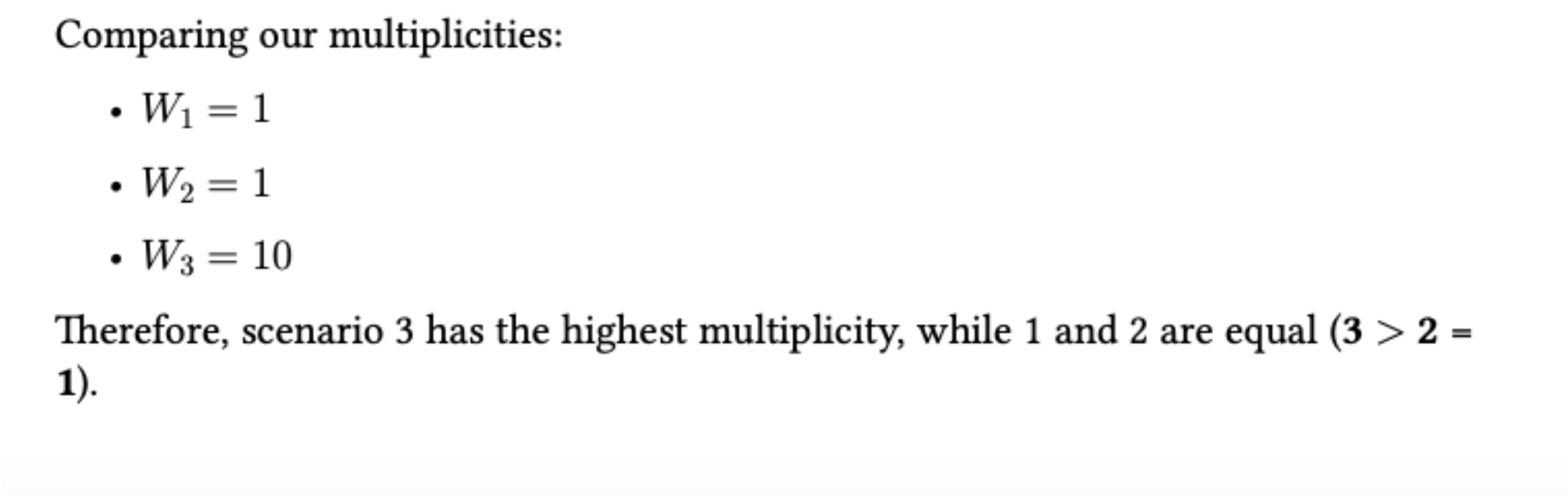
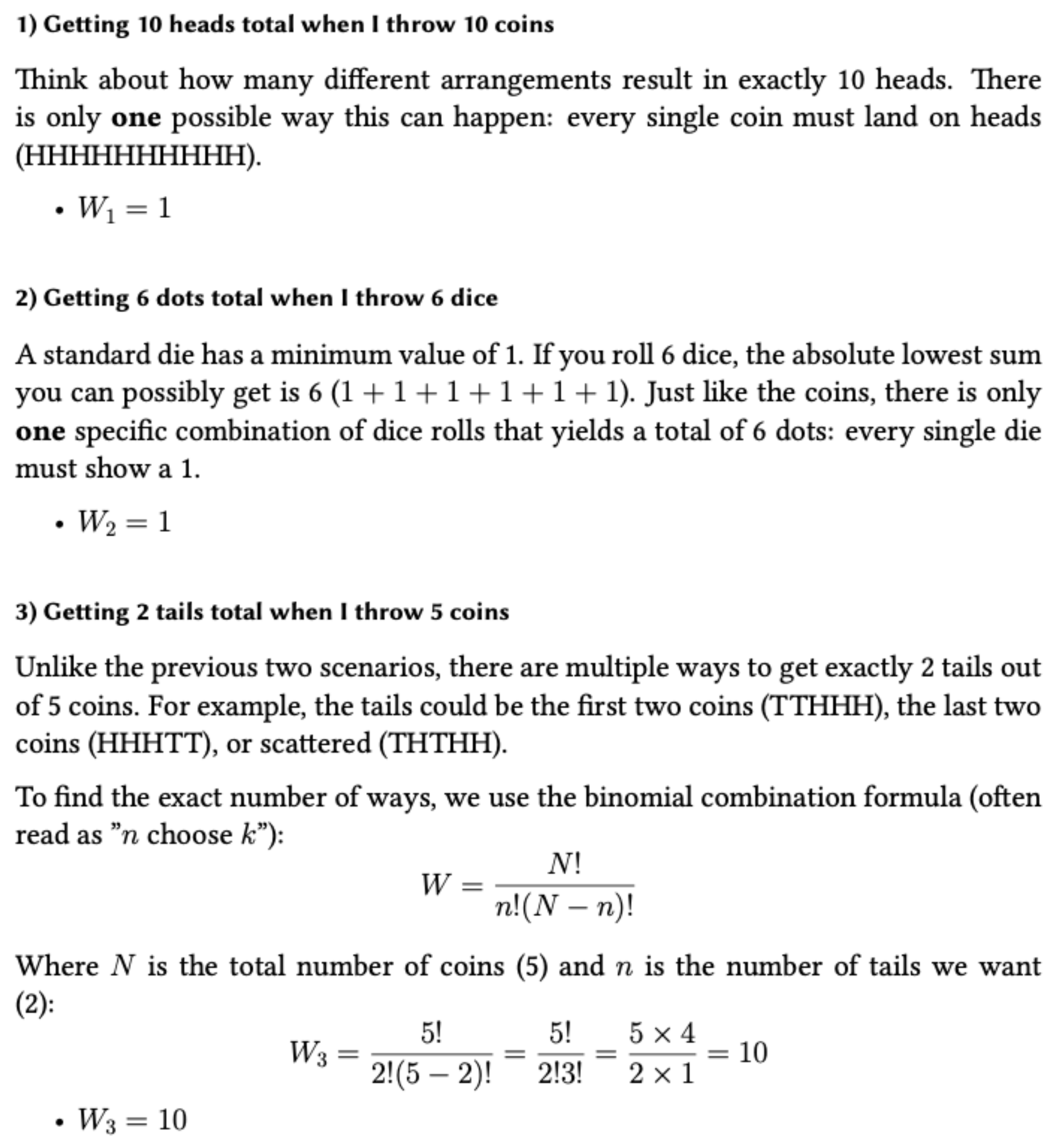


Comparing our multiplicities:

- $W_1 = 1$
- $W_2 = 1$
- $W_3 = 10$

Therefore, scenario 3 has the highest multiplicity, while 1 and 2 are equal (**$3 > 2$ = 1**).



1) Getting 10 heads total when I throw 10 coins

Think about how many different arrangements result in exactly 10 heads. There is only **one** possible way this can happen: every single coin must land on heads (HHHHHHHHHH).

- $W_1 = 1$

2) Getting 6 dots total when I throw 6 dice

A standard die has a minimum value of 1. If you roll 6 dice, the absolute lowest sum you can possibly get is 6 $(1 + 1 + 1 + 1 + 1 + 1)$. Just like the coins, there is only **one** specific combination of dice rolls that yields a total of 6 dots: every single die must show a 1.

- $W_2 = 1$

3) Getting 2 tails total when I throw 5 coins

Unlike the previous two scenarios, there are multiple ways to get exactly 2 tails out of 5 coins. For example, the tails could be the first two coins (TTHHH), the last two coins (HHHTT), or scattered (THTHH).

To find the exact number of ways, we use the binomial combination formula (often read as "n choose k"):

$$W = \frac{N!}{n!(N-n)!}$$

Where N is the total number of coins (5) and n is the number of tails we want (2):

$$W_3 = \frac{5!}{2!(5-2)!} = \frac{5!}{2!3!} = \frac{5 \times 4}{2 \times 1} = 10$$

- $W_3 = 10$

If the two parts A and B in a system are in equilibrium and having thermodynamic probabilities P_A and P_B, what will be the thermodynamic probability of the system?

a. $P = P_A \cdot P_B$

b. $P = P_A + P_B$

c. $P = (P_A + P_B) / 2$

d. $P = \sqrt{}(P_A - P_B)$

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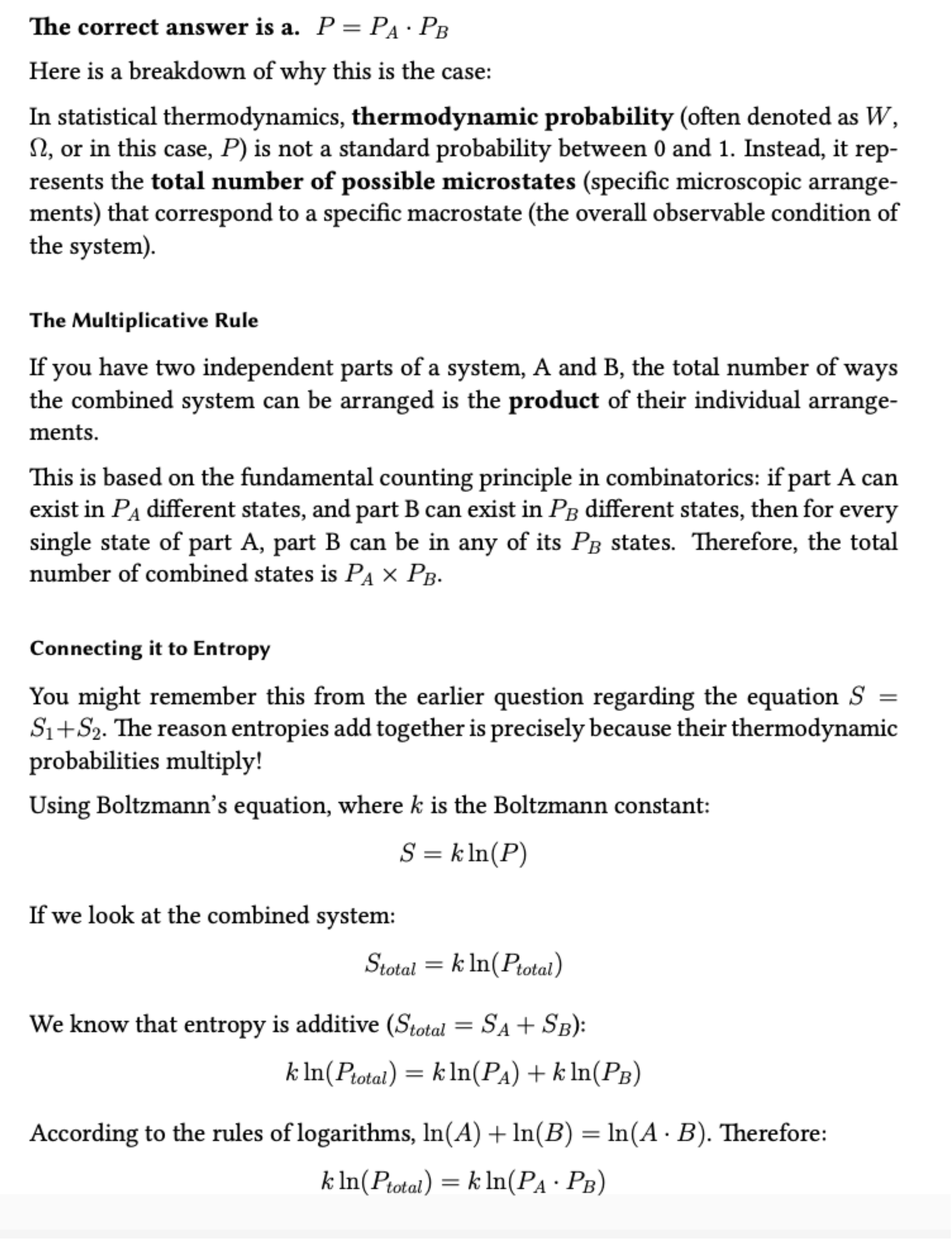

The correct answer is a. $P = P_A \cdot P_B$

Here is a breakdown of why this is the case:

In statistical thermodynamics, **thermodynamic probability** (often denoted as W, Ω, or in this case, P) is not a standard probability between 0 and 1. Instead, it represents the **total number of possible microstates** (specific microscopic arrangements) that correspond to a specific macrostate (the overall observable condition of the system).

The Multiplicative Rule

If you have two independent parts of a system, A and B, the total number of ways the combined system can be arranged is the **product** of their individual arrangements.

This is based on the fundamental counting principle in combinatorics: if part A can exist in P_A different states, and part B can exist in P_B different states, then for every single state of part A, part B can be in any of its P_B states. Therefore, the total number of combined states is $P_A \times P_B$.

Connecting it to Entropy

You might remember this from the earlier question regarding the equation $S = S_1 + S_2$. The reason entropies add together is precisely because their thermodynamic probabilities multiply!

Using Boltzmann's equation, where k is the Boltzmann constant:

$$S = k \ln(P)$$

If we look at the combined system:

$$S_{total} = k \ln(P_{total})$$

We know that entropy is additive ($S_{total} = S_A + S_B$):

$$k \ln(P_{total}) = k \ln(P_A) + k \ln(P_B)$$

According to the rules of logarithms, $\ln(A) + \ln(B) = \ln(A \cdot B)$. Therefore:

$$k \ln(P_{total}) = k \ln(P_A \cdot P_B)$$