States of Matter

Kevin Pyatt, Ph.D.

AUTHOR
Kevin Pyatt, Ph.D.

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CHAPTER 1 States of Matter

CHAPTER OUTLINE

So far, much of our study of the chemical world has focused on investigating elements, compounds, and reactions. We have assessed the characteristics of elements and studied how they combine through chemical reactions to produce new compounds. Now, we are going to shift our attention away from compounds and reactions and look at what happens to matter when it changes state. We are going to study why substances exist as solids, liquids, or gases. We are also going to look at what takes place on the molecular level when matter changes state from one form to another. As you might already imagine, some substances exist as solids at room temperature, others exist as liquids, and still others exist as gases. Life as we know it would be much different, for instance, if water was solid at room temperature, or if oxygen and carbon dioxide were liquid. This chapter will help us understand the different states of matter.

Mike Cline. commons.wikimedia.org/wiki/File:FireholeRiverOjoCalientiBend.jpg. Public Domain.

1.1 The Kinetic-Molecular Theory of Gases

Lesson Objectives

- Describe kinetic energy as it applies to molecules.
- Explain how kinetic energy is related to the mass and velocity of a particle.
- Describe the nature of elastic collisions and the implications for matter at the molecular level.
- Describe the origins and assumptions of the kinetic-molecular theory, and use this model to describe the nature of matter at the molecular level.
- Describe ideal behavior as it applies to a gas.

Lesson Vocabulary

- **kinetic energy**: The energy associated with motion.
- **elastic collision**: A collision in which momentum is conserved.
- **kinetic-molecular theory**: Describes the molecular behavior of an ideal gas.
- **ideal gas**: Gases that conform to the kinetic-molecular theory.

Check Your Understanding

- What are the similarities and differences between different states of matter?

Introduction

Ice, water, and steam appear quite differently to the eye. If you were to look at these three states of matter on the molecular level, you would see that the arrangement of molecules is very different here as well. However, solids and liquids have definite volumes, unlike gases which tend to take on the shape of their container. In this lesson, you will learn about the unique behavior of gas particles on a molecular level and the basis for kinetic molecular theory.

States of Matter - A Microscopic View

If we see matter at the macroscopic level, we can easily tell whether it is solid, liquid, or gas. In **Figure** 1.1, we see a green chlorine gas-liquid equilibrium. The various states of matter can largely be explained by studying interactions between particles that occur at the microscopic level. **Figure** 1.2 shows how the three states of matter differ on the molecular level.

FIGURE 1.1
Chlorine gas.

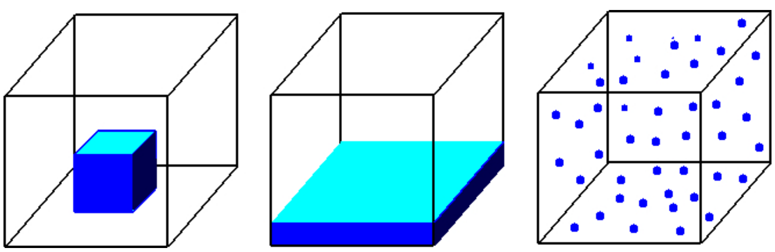


FIGURE 1.2
In this image, we see how atoms of argon interact in the solid, liquid, and gas phases.

As we saw in our chapter on the mole, matter is ultimately composed of particles. We cannot "see" individual molecules, but we can see the effects exerted by the structure of each molecule on the behavior of the bulk material. What accounts for the very different properties exhibited by the same substance when it exists in different phases of matter? The sizes and properties of individual atoms and molecules do not change when a substance changes phase. Rather, it is the interactions between particles that are changing.

Liquids and Solids

As shown in **Figure** 1.2, each state of matter looks quite different at the molecular level. In the case of liquids and solids, the distances between particles are negligible relative to the size of each particle; they are essentially in direct contact with one another. In liquids, particles are free to move and exchange neighbors, resulting in the properties of a fluid. In solids, they are rigidly fixed in space and held tightly to neighboring particles.

Gases

The story is quite different for gases. Gases take the shape of their container, and they are relatively easy to compress. There are fewer gas particles per unit volume than for the same substance in the liquid or solid form. In fact, the liquid form of a given material is generally several hundred times more dense than the gas form at normal pressures. Despite the large amounts of empty space, a sample of a gas contains many particles moving around, colliding and imparting force on their surroundings. For example, in a one mole sample of gas at 0°C and 1 atm of pressure, each cubic centimeter contains roughly 2.7×10^{19} molecules. Each molecule participates in several billion collisions every second, moving only about 10-100 nanometers between collisions. Additionally, these gas particles move at very high speeds. For example, at 25°C, the average speed of hydrogen molecules in a sample of hydrogen gas is 1960 m/s.

These are just some of the differences we see when we look at the molecular level and study the different states for a particular chemical species. The following simulation illustrates how particles behave over time in the liquid, solid, and gas states: http://phet.colorado.edu/en/simulation/states-of-matter .

In this simulation, you can watch different types of molecules form a solid, liquid, or gas. Add or remove heat to watch the phase change. Change the temperature or volume of a container and see a pressure-temperature diagram respond in real time.

One of the concepts shown by this animation is that particles are constantly moving and vibrating. This is an important assumption that we make when we study matter at the molecular level. Anything that is moving and has mass also possesses some amount of **kinetic energy**, or the energy of motion. Kinetic energy increases as the molecular mass increases and as the velocity of the particle increases.

Another thing we can see in this animation is that particles are constantly colliding with one another. One assumption that we make when talking about collisions between gas particles is that they are completely elastic collisions. In an **elastic collision**, momentum is conserved, which means that none of the kinetic energy of the colliding particles is lost in some other form (such as the emission of light). This makes sense, because if energy were lost in collisions, the speeds of the particles would gradually decrease over time, and eventually everything would condense down into a solid form.

The Kinetic-Molecular Theory of Gases

Some of the observations and assumptions we just made about particle behavior at the molecular level were proposed in independent works by August Kroning (1856) and Rudolf Clausius's 1857 work titled "the theory of moving molecules." This work became the foundation of the **kinetic-molecular theory** of gases. The kinetic-molecular theory of gases makes the following assumptions:

1. Gases are comprised of large numbers of particles that are in continuous, random motion and travel in straight lines.
2. The volume of gas particles in a sample is extremely small compared to the total volume occupied by the gas.
3. Attractive and repulsive forces between gas molecules are negligible.
4. Energy can be transferred between molecules during collisions. Collisions are completely elastic.
5. The average kinetic energy of the molecules is proportional to the temperature of the sample.
6. Gases that conform to these assumptions are called **ideal gases**.

By applying these principles to gases, it is possible to show that the properties of gases on the macroscopic level are a direct result of the behavior of molecules on the microscopic level.

Lesson Summary

- Differences between solids, liquids, and gases depend upon the interactions between the individual particles.

- Kinetic energy is directly proportional to the mass and velocity of a particle; that is, as mass and/or velocity increase, so does the kinetic energy.
- The kinetic-molecular theory describes the behavior of an ideal gas.
- Assumptions of the kinetic-molecular theory include the following:
  - Gas particles are in constant, random motion.
  - The volume of gas particles is negligible in comparison to the volume of the container.
  - There are no attractive forces between gas particles.
  - Collisions of gas particles are elastic, so no energy is lost.
  - The speed of a gas particle is directly proportional to the temperature of the system.

Lesson Review Questions

1. What do we mean when we say molecular view of matter? Can you draw a diagram to describe what particles might look like at the molecular level for solids, liquids, and gases?
2. What is kinetic energy? Does kinetic energy increase or decrease as particle speed increases?
3. Describe what is meant by an elastic collision. What would happen to particles over time if most collisions were not elastic?
4. Summarize the major points of the kinetic-molecular theory.
5. How are ideal gases and the kinetic-molecular theory related?
6. Determine whether or not the following gases would be ideal; that is, do they fit the points of kinetic-molecular theory?
   a. As a gas is heated, its particles start to move more slowly.
   b. When one gas particle bumps into another, no energy is lost.
   c. The gas particles follow predictable, circular paths within a container.

Further Reading / Supplemental Links

- Brenner, H. C. (1992). The kinetic molecular theory and the weighing of gas samples. Journal of Chemical Education, 69(7), 558-null. doi: 10.1021/ed069p558
- Hildebrand, J. H. (1963). An introduction to molecular kinetic theory - Selected topics in modern chemistry The University of California, Berkeley. New York: Reinhold Pub. Corp.
- Timm, J. A. (1935). The kinetic-molecular theory and its relation to heat phenomena. Journal of Chemical Education, 12(1), 31-null. doi: 10.1021/ed012p31
- TedEd "The Invisible Properties of Gases": http://www.youtube.com/watch?v=EHxdVtygP1g

Points to Consider

- One of the assumptions of the kinetic-molecular theory is that collisions between particles are elastic –that momentum is conserved. Can you think of collisions you have witnessed in your everyday life that are completely elastic? Are collisions that you typically see elastic or not? Give examples.

1.2 Changes of State

Lesson Objectives

- Describe how the phase of a material is affected by changes in the temperature.
- Describe how the phase of a material is affected by changes in pressure.
- Draw phase diagrams to relate the pressure, temperature, and the phase of a substance.
- Describe the energy changes associated with changes of state.

Lesson Vocabulary

- **heating curve**: A curve where supplying heat to a solid substance will gradually raise its temperature, and eventually, it will melt.
- **melting point**: The temperature of a point where heat is used to break up the attractive forces holding them rigidly in place.
- **boiling point**: The temperature of a point where particles start to enter the gas phase.
- **phase diagram**: A plot of temperature vs. pressure that indicates the states of matter present at each point.
- **triple point**: A point where all three states can exist simultaneously.

Check Your Understanding

1. Compare and contrast the properties of liquids, solids, and gases.
2. Which of the following statements about solids and liquids are true? (There may be more than one.)
   a. Solids and liquids are virtually incompressible; their volume is constant.
   b. Solids are typically more dense than liquids.
   c. All liquids have the same density.

Introduction

In the last lesson, we studied the characteristics of liquids and solids at a macroscopic level and at the molecular level. Increasing the temperature of a solid transforms the particles from a rigid arrangement to a fluid (a liquid or gas). Conversely, decreasing the temperature of a liquid or gas slows the particles down, going back from free movement to a fixed arrangement. In this section, we will further explore how temperature and pressure affect the characteristics and behavior of matter. Pressure has a larger effect on gases, which are very compressible, than liquids and solids. However, changes in pressure are still relevant to solids and liquids. For example, **Figure** 1.3 shows ice skates on ice. Because your entire weight is all concentrated on a thin blade, ice skates exert quite a bit of pressure on the ice below them. An interesting property of water is that increasing the pressure on its solid form

FIGURE 1.3

(ice) will eventually convert it to liquid water. The pressure exerted by ice skates makes small amounts of liquid water on the surface, allowing the skates to glide smoothly.

Heating Curves

Supplying heat to a solid substance will gradually raise its temperature, and eventually, it will melt. Heating the resulting liquid will cause a further increase in temperature until the liquid begins to boil. If we were to graph the temperature of a substance against the amount of heat added, it would look something like the **Figure** 1.4.

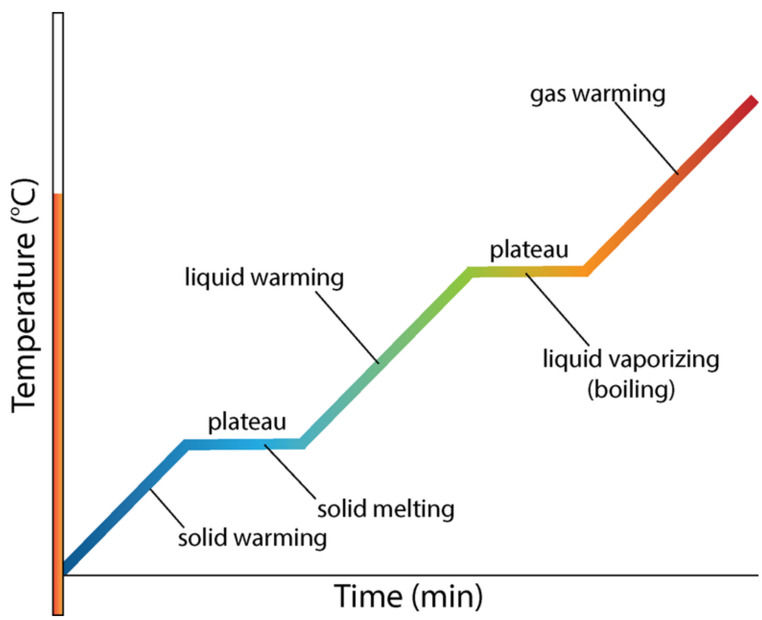


FIGURE 1.4
Heating curve depicting phase changes between the solid, liquid, and gas forms of a given substance.

The **Figure** 1.4 is referred to as a **heating curve**. The most notable feature is that the temperature rise is not steady;

there are plateaus during which heat is being added, but the temperature is not increasing. What is happening here? Let's start with the first slope on the left. At this point, the substance is in its solid form. Adding heat causes the particles to move faster. Faster particles means more kinetic energy, which also means a higher temperature. At the first plateau, the vibrations of the particles become energetic enough to break free of the rigid solid form, and the substance converts to a liquid. Heat continues to be added, but instead of increasing the kinetic energy of the particles, it is used to break up the attractive forces holding them rigidly in place. This process is known as melting, and the temperature at which it occurs is the **melting point**. Because the added heat is used to break up attractive forces instead of adding to the kinetic energy of the particles, the temperature of the material stays constant until the phase change is completed.

Further heating then adds energy to the liquid particles, increasing their speed, kinetic energy, and temperature. This is the second slope on the curve. Once the particles are energetic enough to completely break free of each other, they start to enter the gas phase. Boiling occurs at the second plateau of this curve, and the temperature at this point is referred to as the **boiling point**. Again, the added heat is being used to break up the interactions between particles instead of increasing their kinetic energy, so no temperature increase is observed until all particles are in the gas phase. Finally, adding even more energy will further speed up the gas particles, increasing the kinetic energy and temperature of the substance.

The reverse process can also be diagrammed, where we start with a gas and gradually remove heat until it condenses to a liquid and then freezes into a solid. For a given amount of a certain substance at a given pressure, heating and cooling curves should be mirror images. The melting point will be equal to the freezing point, and the boiling point will be equal to the condensation point. Additionally, the amount of heat added to completely melt the sample is the same as the amount that must be removed to completely freeze it.

Two other changes of state can occur under at appropriate pressures. The direct conversion of a solid to a gas without becoming a liquid is called sublimation. The reverse process (gas to solid) is known as deposition. Depositions of hot metal vapors are often used in the electronics industry to produce thin films of metal on solid bases. Most substances require reduced pressures (less than one atmosphere) for these processes to occur. At higher pressures, substances would transition through the liquid phase. However, some materials, such as carbon dioxide, will sublime even at standard pressure. Iodine and naphthalene (found in mothballs) are other substances that can sublime at only slightly reduced pressures.

The **Figure** 1.5 summarizes the different processes involved in phase changes.

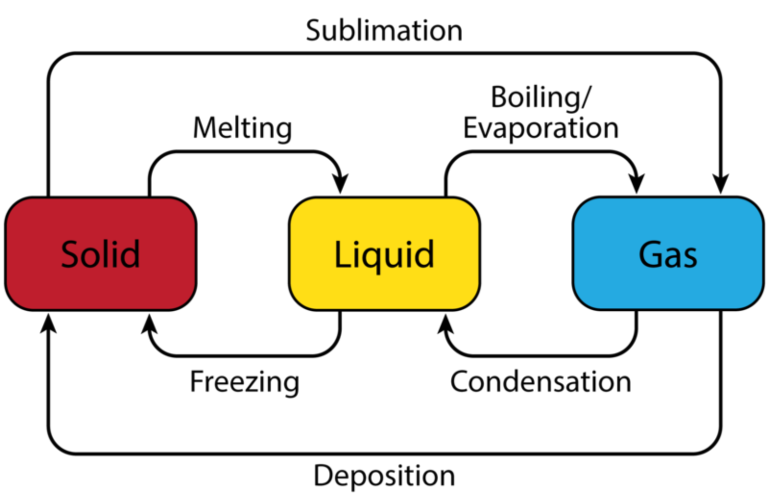


FIGURE 1.5
Summary of phase changes.

Phase Diagrams

Both temperature and pressure have an effect on the phase in which a given substance exists. A plot of temperature vs. pressure that indicates the states of matter present at each point is known as a **phase diagram**. **Figure** 1.6 shows the phase diagram for water.

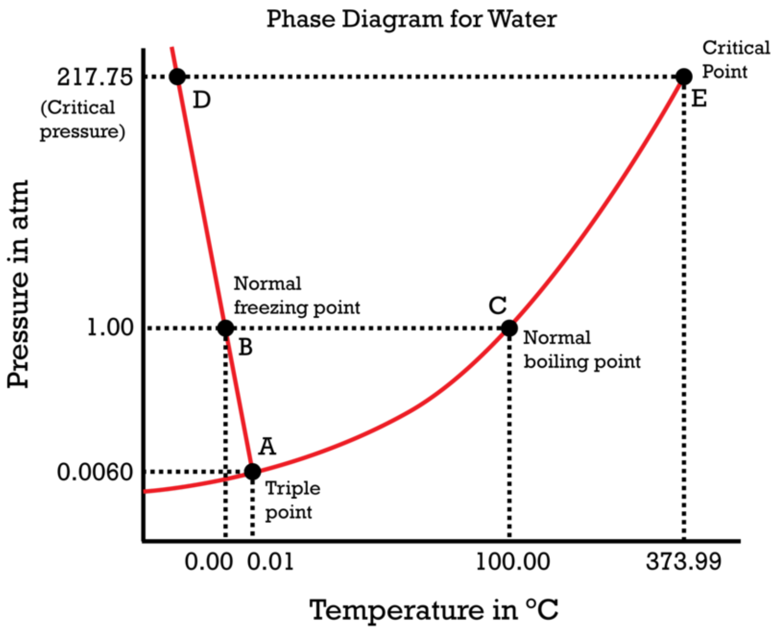


FIGURE 1.6

The lines on this diagram show the boundaries between the three states of matter for water. Notice that there is one point where all three states can exist simultaneously; this is called the **triple point**. Also look at the slope of the solid-liquid boundary. For most substances, this line has a positive slope, which means that increasing the pressure of a liquid will eventually form a solid. Water is unusual in that it has a negative slope for this boundary. This has to do with the fact that the liquid form of water is more dense than the solid form (ice). As a result, putting large amounts of pressure on ice at a given temperature will cause it to melt. We already discussed this phenomenon at the beginning of the lesson in the context of ice skating. The phase diagram gives us a more detailed way to look at this occurrence.

Example 13.1

A pressure cooker operates by keeping water in its liquid form at temperatures above its normal boiling point. Can you use the phase diagram to explain this behavior?

Answer:

The pressure cooker maintains a pressure that is above one atmosphere. At this higher pressure, the boundary between the liquid and gaseous forms of water occurs at a higher temperature.

If we look at the phase diagram for carbon dioxide, we can see that its triple point occurs at approximately 5 atm and -56°C. Below this pressure, liquid CO_2 cannot exist. At one atmosphere of pressure, only the solid and gas forms are possible, which explains why solid CO_2 sublimes instead of melting at standard pressures. Also note that the solid-liquid boundary has a positive slope, which is normal for most substances.

Example 13.2

Why is solid CO_2 is called dry ice? Use the phase diagram to explain.

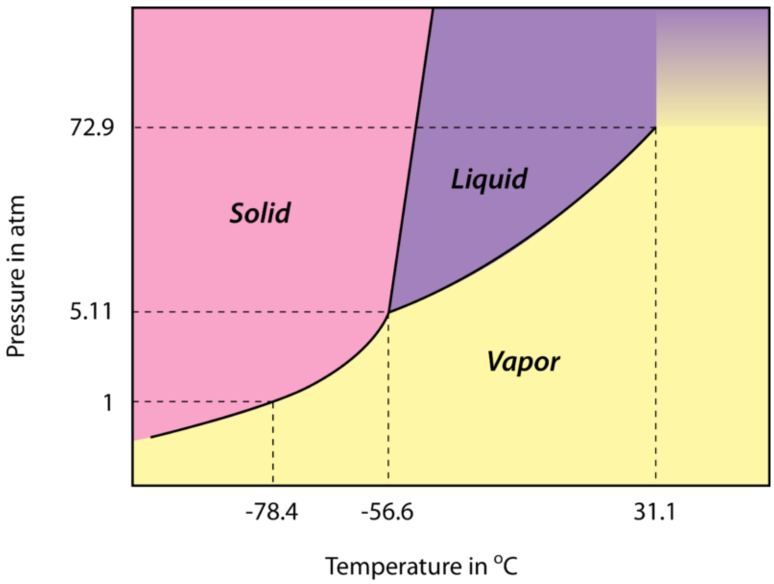


FIGURE 1.7

Answer:

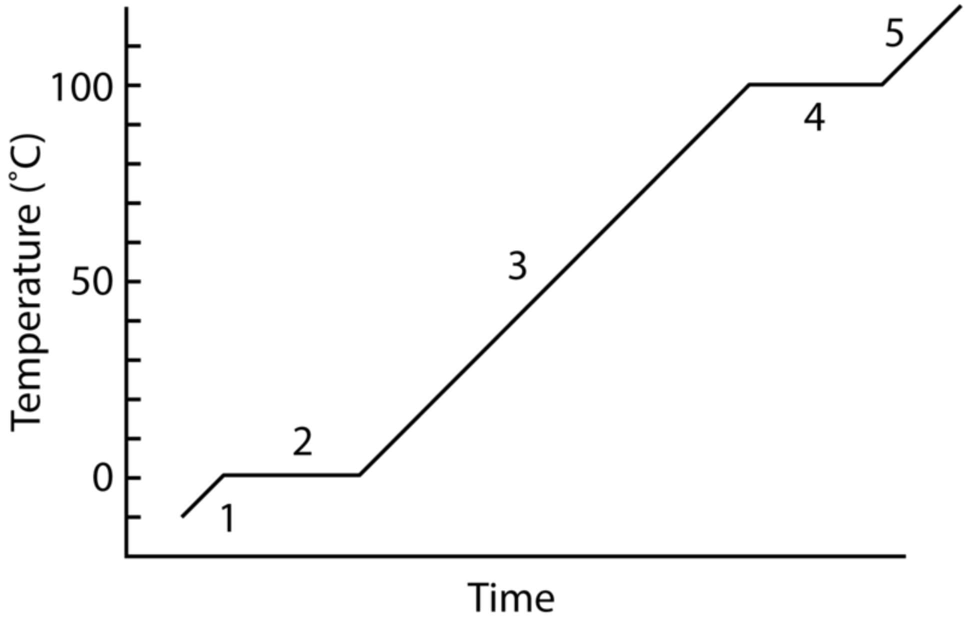
Solid CO_2 exists at temperatures below -78°C at the standard pressure of 1 atm. When solid CO_2 is allowed to warm to room temperature at this pressure, the solid changes to a gas without going through the liquid form. Because it does not become a liquid, it is considered "dry."

Lesson Summary

- Phase transitions occur as heat is added or removed from a substance.
- The solid to liquid transition is called melting.
- The liquid to gas transition is called boiling.
- Sublimation and deposition involve direct transitions between solid and gas without going through the liquid state.
- A phase diagram gives information about the conditions under which a material can exist as a solid, a liquid, or a gas.

Lesson Review Questions

1. The following heating curve of an unknown substance shows several phase changes that take place as heat is added. Label each section indicated by a number.



2. What would happen if you tried to make an ice rink using dry ice? Would you be able to skate? Compare the phase diagrams of water and dry ice to justify your answer.
3. What would you see if you were to observe water at its triple point?
4. Referring to the phase diagram of water, what phases can exist at a temperature of 0°C and a pressure of 0.006 atm?
5. Referring to the phase diagram for CO_2, at what temperature must liquid CO_2 exist if the pressure is 100 atm?
6. What temperature must carbon dioxide be in order to remain solid at 7 atmospheres of pressure? Would it melt or sublime at this pressure?

Further Reading / Supplemental Links

- Observe the triple point of an organic compound, tert-Butyl alcohol: http://www.youtube.com/watch?v=BLRqpJN9zeA
- Materials and their properties: http://www.abpischools.org.uk/resources/solids-liquids-gases/index.asp
- Review on matter and change of state: http://www.chem4kids.com/files/matter_intro.html

Points to Consider

- Can you think of some substances commonly encountered in their gaseous state?
- How might you check for the presence of a gas?

1.3 References

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2. Courtesy of NASA. http://exploration.grc.nasa.gov/education/rocket/state.html . Public Domain
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