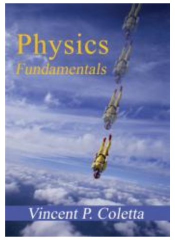
Medical Physics Class
Temperature Part-1



Physics Fundamentals by Vincent P. Coletta

Learning Goals

Looking forward at …

- **Temperature**
- **Temperature Measurement**
- **Ideal Gas Law**

Why?

Concept	Relevance in Pharma
Temperature	Affects drug stability and storage, reactions, solubility
Kinetic Energy	Influences reaction rates and molecular behavior
Applications	Formulation, drug design, storage, and release

1. Many drugs are sensitive to temperature. Higher temperatures increase the kinetic energy of molecules, which can speed up degradation reactions.
2. The rate of chemical reactions, including synthesis of active pharmaceutical ingredients, depends on the kinetic energy of molecules.
3. Temperature affects solubility by increasing molecular motion, which can help dissolve solid drugs in solvents.
4. Temperature-sensitive drug delivery systems use kinetic energy to control the release of a drug.
5. Molecular kinetic energy is a factor in **molecular dynamics simulations**, which help predict how a drug molecule interacts with its target.

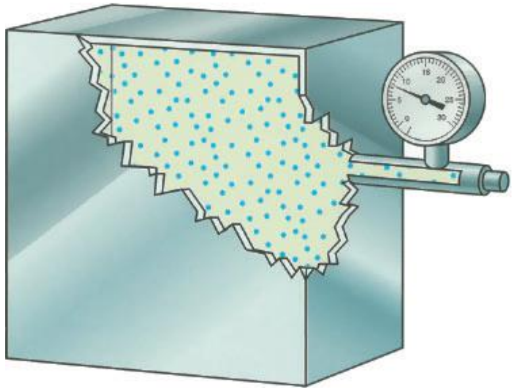
Temperature

- We all know what it feels like when it's really hot in the summer or freezing cold in the winter.
- Our bodies notice even small changes in temperature and react by sweating when it's hot or shivering when it's cold. These reactions help keep our body temperature steady.
- This natural sensitivity to heat and cold is where our understanding of temperature comes from.
- In science, **temperature** means how hot or cold something is, and we can measure it using different types of thermometers.
- In this class, we'll learn how temperature is measured and how it's actually related to how fast molecules are moving — which is called **molecular kinetic energy**.

Temperature Measurement

Thermometers

- Temperature is measured with a thermometer.
- Galileo invented the first thermometer, which made use of air’s property of expanding as it is heated. The air’s volume indicated the temperature.
- Today there are various kinds of thermometers, each appropriate for the range of temperatures and the system to be measured.
- For example, in addition to the common mercury thermometer, there are different kinds of thermometers.



Temperature Measurement

Thermometers

- Each thermometer depends on the existence of some thermometric property of matter.
- For example, the expansion or contraction of mercury in a fever thermometer correlates with the body’s sensation of hot and cold.
- A person who has a high fever both is hot to the touch and will register a higher than normal temperature on a mercury thermometer.
- The length of the mercury column in the glass stem of the thermometer gives us a quantitative measure of temperature.

Temperature Measurement

Temperature Scales

- To assign a numerical value to the temperature of a body, we need a temperature scale.
- The two most common scales in everyday use are the Celsius scale (formerly known as the centigrade scale) and the Fahrenheit scale.
- In establishing a temperature scale, one could use any kind of thermometer and any thermometric property.
- For example, the Celsius scale was based on the expansion of a column of liquid such as mercury in a thin glass tube.
- The **absolute**, or **Kelvin**, scale is now the standard in terms of which all other scales, such as Celsius and Fahrenheit, are defined.

Temperature Measurement

Temperature Scales

- The Kelvin scale is chosen as the standard for important reasons.
- First, various laws of physics are most simply expressed in terms of this scale. We shall see an example of this in the ideal gas law, described in the next section.
- Second, zero on the absolute scale has fundamental significance. It is the lowest possible temperature a body can approach.

Temperature Measurement

🌡️ **Temperature Scales – Celsius**

- The **Celsius scale** is one of the most common ways to measure temperature.
- When it was first made, scientists used the **freezing point** of water as **0°C** (zero degrees Celsius).
- The **boiling point** of water was marked as **100°C**.
- They used a **mercury thermometer** and marked the levels at these two points.
- Then, they divided the space between those two marks into **100 equal parts** — each part is **1 degree Celsius (1°C)**.

Temperature Measurement

Today, the Celsius scale is based on another scale called the **Kelvin scale**, which is used in science.

•The temperature in Celsius (T_C) is related to Kelvin (T) by this simple formula:

$$T_C = T - 273.15$$

So if you know the temperature in Kelvin, subtract 273.15 to get Celsius.

Temperature Measurement

Temperature Scales

$$T_C = T - 273.15$$

- The normal freezing point of water open to the air at one atmosphere of pressure is $273.15K$, or $0.00°\text{C}$.
- The normal boiling point of water is $373.15K$, or $100.00°\text{C}$.
- This definition of the Celsius scale conforms to the earlier definition based on the freezing and boiling points of water.
- Temperature intervals on the Celsius and Kelvin scales are the same.
- For example, the difference in temperature between the boiling point of water and its freezing point is 100 Celsius degrees (°C), or 100 kelvins.

Temperature Measurement

Temperature Scales

- Fahrenheit temperature T_F, measured in degrees Fahrenheit (°F), is defined relative to the Celsius temperature T_C by the

$$T_F = 32 + \frac{9}{5} T_C$$

- The normal freezing and boiling points of water on the Fahrenheit scale are 32 °F and 212 °F respectively. The interval between these points is 180 °F. There are only 100 °C between these same points.

- Thus Celsius degrees are bigger than Fahrenheit degrees: 1°C is $\frac{180}{100}$, or $\frac{9}{5}$, times 1°F .

Temperature Measurement

Measuring a Fever on the Celsius Scale

Normal internal body temperature is 98.6°F. A temperature of 106°F is considered a high fever. Find the corresponding temperatures on the Celsius scale.

$$T_C = \tfrac{5}{9}(T_F - 32.0)$$

A temperature of 98.6° F corresponds to

$$T_C = \tfrac{5}{9}(98.6 - 32.0) = 37.0° \text{ C}$$

and a temperature of 106° F corresponds to

$$T_C = \tfrac{5}{9}(106 - 32.0) = 41.1° \text{ C}$$

Ideal Gas Law

Introduction to Gas Laws

- Gases are made of tiny particles moving around quickly. When we **heat** a gas, the particles move **faster** — this means the **temperature goes up**.
- The **Gas Laws** help us understand how a gas will behave when we change:
  - **Temperature** (how hot it is)
  - **Volume** (how much space it takes up)
  - **Pressure** (how hard it pushes on the container)

Ideal Gas Law

Heat and Gases

- When you **add heat**, gas particles gain **energy** and move faster.
- This can cause the gas to:
  - **Expand** (take up more space),
  - Or **increase pressure** if the space stays the same.

Ideal Gas Law

- The pressure of a gas can be changed in several ways:

1. One way to increase the pressure of a gas confined to a fixed volume is to increase the number of gas molecules in the volume. You do this, for example, when you pump air into a gas container (gas tank, bicycle tire, or an automobile tire).
2. Another way to change the pressure of a gas is to change its temperature. For example, when the air in a gas container heats up, its pressure increases significantly.
3. A third way to change gas pressure is to change the volume containing the gas; decreasing volume causes an increase in pressure.

Ideal Gas Law

- For low-density gases, there is a simple, universal relationship between the gas $pressure\ P$, $volume\ V$, $Kelvin\ temperature\ T$, and $number\ of\ gas\ molecules\ N$. The product of P and V is proportional to the product of N and T:

$$PV = NkT$$

- This equation is called the **ideal gas law.** The constant k is known as “Boltzmann’s constant” and is found from e $k = 1.380 \times 10^{-23}\ \mathrm{J/K}$ e value

  - In applying the ideal gas law, temperature must be expressed in kelvins, not in °C or °F.

Ideal Gas Law

- Special cases of the gas law are found when one considers the variation of two of the variables P, V, N, and T, while the other two variables are held constant.
- For example, if N and T are fixed, the ideal gas law implies that the product PV is constant:

$$\boxed{PV = \text{constant}}$$

- This result is known as Boyle's law.
- Boyle's law implies that if the volume of a gas is reduced to half its original value the pressure of the gas is doubled.

Ideal Gas Law

- If P and N are fixed, the ideal gas law implies that the volume of the gas is directly proportional to its temperature:

$$\boxed{V \propto T} \qquad \text{(for constant } N \text{ and } P\text{)}$$

- If V and N are fixed, the ideal gas law implies that

$$\boxed{P \propto T} \qquad \text{(for constant } N \text{ and } V\text{)}$$

- The very definition of temperature on the Kelvin scale requires that this relationship be satisfied, at least in the limit of a very low-density gas.

Ideal Gas Law

EXAMPLE 2 The Temperature of an Ideal Gas After Compression

An ideal gas initially has a volume of 1.0 liter (L), a pressure of 1.0 atmosphere (atm), and a temperature of 27° C. The pressure is raised to 2.0 atm, compressing the volume of the gas to 0.60 L. Find the final temperature of the gas.

SOLUTION We are given the following initial and final values of P, V, and T:

$$P_i = 1.0\text{ atm} \qquad P_f = 2.0\text{ atm}$$

$$V_i = 1.0\text{ L} \qquad V_f = 0.60\text{ L}$$

$$T_{Ci} = 27^\circ\text{ C} \qquad T_{Cf} = ?$$

The number of molecules, N, is constant. The problem is to find the final temperature T_{Cf}. We can do this simply by first writing the ideal gas law for the initial state of the gas and again for the final state and then taking the ratio of the two expressions:

$$P_iV_i = NkT_i$$

$$P_fV_f = NkT_f$$

$$\frac{T_f}{T_i} = \frac{P_fV_f}{P_iV_i}$$

Since this equation involves ratios of pressures and volumes, we may insert these quantities in the units in which they are given, that is, atmospheres and liters, rather than converting to standard units of Pa and m^3. The conversion to standard units would simply introduce identical factors for both initial and final values, and these factors would cancel. We must be careful, however, to convert temperature from degrees Celsius to kelvins, even when a ratio is used, as it is here, since this change in units involves an additive term, rather than a multiplicative factor. Thus we must use $T_i = 27 + 273 = 300$ K. Substituting values into the preceding equation, we obtain

$$\frac{T_f}{300\text{ K}} = \frac{(2.0\text{ atm})(0.60\text{ L})}{(1.0\text{ atm})(1.0\text{ L})} = 1.2$$

or

$$T_f = (1.2)(300\text{ K}) = 360\text{ K}$$

This corresponds to a final Celsius temperature T_{Cf} of

$$T_{Cf} = T_f - 273 = 360 - 273$$

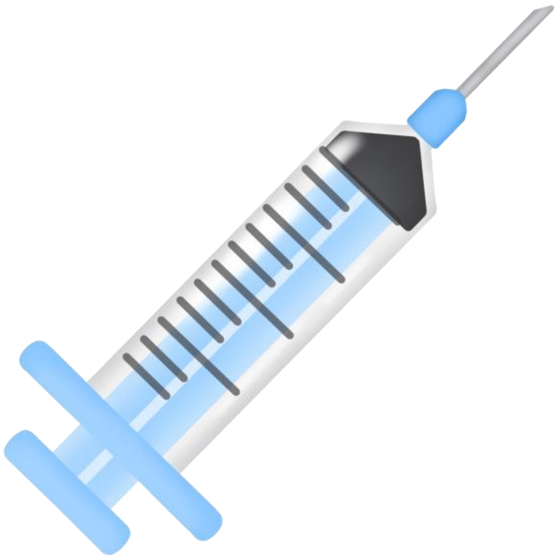
$$= 87^\circ\text{ C}$$

Ideal Gas Law

A pressurized inhaler contains a gas-based medication. The temperature inside the inhaler is **37.0°C**, while the temperature of the surrounding room air is **25.0°C**. Assuming the **pressure and volume inside and outside are the same**, calculate the **ratio of the number of gas molecules** in the inhaler at **temperature (37.0°C)** to the number at **room temperature (25.0°C)**.

Ideal Gas Law

A pharmacist prepares a dose of a gas-based medication in a **sealed syringe**. At **room temperature (25°C)**, the gas fills **10.0 mL** of volume. If the syringe is later **warmed to body temperature (37°C)** while keeping the **pressure constant**, what will be the **new volume of the gas**?



Ideal Gas Law

It is often convenient to express the ideal gas law in a slightly different form, known as the “molar form.” To accomplish this, we first define

1. atomic mass
2. the mole.

Ideal Gas Law

Atomic mass tells us how heavy an atom is compared to other atoms. Scientists use a scale where the most common carbon atom is given a mass of exactly **12**.

- A **hydrogen atom** is about **1/12** the mass of a carbon atom, so its atomic mass is about **1**.
- A **helium atom** is about **4/12** the mass of a carbon atom, so its atomic mass is about **4**.

You can find atomic masses on the **periodic table**.

- But! These numbers are usually **averages**, because elements come in slightly different types (called **isotopes**).
- For example, carbon's atomic mass is listed as **12.01** instead of exactly 12, because a small percent of carbon atoms are a bit heavier (mass of 13).

Ideal Gas Law

Molecular mass is the total mass of all the atoms in a molecule.

For example, in a water molecule (**H_2O**):

- Oxygen has a mass of about **16**
- Hydrogen has a mass of about **1**, and there are **two** of them

So, the molecular mass of H_2O is:

$$16+2(1)=18$$

The unit used for these tiny masses is called the **atomic mass unit**, written as **u**.

Scientists have also figured out how to relate this tiny unit to grams using experiments.

$$1\ \text{u} = 1.6606 \times 10^{-24}\ \text{g}$$

Ideal Gas Law

Even a tiny amount of a substance contains **a huge number of atoms or molecules**.

To make this easier to work with, scientists use a big counting unit called a **mole**.

A **mole** means a specific number of particles — this number is called **Avogadro's number**, written as N_a.

$$1 \text{ mole} = 6.02 \times 10^{23} \text{ atoms or molecules}$$

What's cool is:

- The **mass of 1 mole** (in grams) is the **same number** as the atomic or molecular mass. For example:
    - 1 mole of **carbon-12** atoms weighs **12 grams**
    - 1 mole of **water (H_2O)** molecules weighs **18 grams**

Ideal Gas Law

Mole and Number of Molecules – Simple Explanation

We can figure out **Avogadro's number** by dividing:

- The mass of **1 mole** of carbon-12 (which is **12 grams**) by The mass of **one carbon-12 atom** (which is **12 atomic mass units**) This gives us how many atoms are in a mole — about **6.02 × 10²³**.

$$N_A = \frac{12\ \text{g}}{12\ \text{u}} = \frac{12\ \text{g}}{12(1.6606 \times 10^{-24}\ \text{g})}$$

$$\boxed{N_A = 6.022 \times 10^{23}}$$

- We may express the number of molecules, N, of a substance as the product of Avogadro's number, N_A, and the number of moles, denoted by n:

$$N = nN_A$$

Ideal Gas Law

EXAMPLE 5 Number of Atoms in a Nail

Find the number of atoms in an iron nail of mass 5.00 g.

SOLUTION First we inspect the periodic table (shown on the inside back cover) and find that the atomic mass of iron (Fe) is 55.847. This means that 1 mole of naturally occurring iron has a mass of 55.847 g. We can now calculate the number of moles of iron in the nail, which we denote by n:

$$n = (5.00\ \text{g})\left(\frac{1\ \text{mole}}{55.847\ \text{g}}\right) = 8.95 \times 10^{-2}\ \text{mole}$$

Since 1 mole contains Avogadro's number of atoms, the nail contains a number of atoms equal to the number of moles times N_A:

$$N = nN_A = (8.95 \times 10^{-2}\ \text{mole})\left(\frac{6.022 \times 10^{23}\ \text{atoms}}{1\ \text{mole}}\right)$$
$$= 5.39 \times 10^{22}\ \text{atoms}$$

Ideal Gas Law

- **Molar Form of the Ideal Gas Law**

- To obtain the molar form of the ideal gas law, we substitute $N = nN_A$ into our original form of the gas law

$$PV = NkT = nN_AkT$$

- The product N_Ak is called the ideal gas constant, denoted by R.

$$R = N_Ak$$

$$= (6.022 \times 10^{23})(1.380 \times 10^{-23}\ \mathrm{J/K})$$

$$\boxed{R = 8.31\ \mathrm{J/K}}$$

- Substituting R for N_Ak in the ideal gas law, we obtain the molar form of the gas law.

$$\boxed{PV = nRT}$$

Ideal Gas Law

EXAMPLE 6 Finding the Mass of a Volume of Air

Find the mass of air in a room with dimensions 5.00 m $\times$ 4.00 m $\times$ 3.00 m, if the air pressure is 1.00 atm and the temperature is 27.0° C.

SOLUTION First we apply the ideal gas law (Eq. 12–13) to find the number of moles, using the Kelvin temperature ($T = 273 + 27.0 = 300$ K) and expressing pressure in Pa (Eq. 11–5: 1.00 atm $= 1.01 \times 10^5$ Pa).

$$n = \frac{PV}{RT}$$

$$= \frac{(1.00\,\text{atm})(1.01 \times 10^5\,\text{Pa/atm})(5.00\,\text{m} \times 4.00\,\text{m} \times 3.00\,\text{m})}{(8.31\,\text{J/K})(300\,\text{K})}$$

$$= 2.43 \times 10^3\ \text{moles}$$

Since 1 mole has a mass equal to the molecular mass in grams, the mass of air equals the product of the number of moles times the molecular mass in grams. A nitrogen molecule N_2 has a molecular mass of $2(14) = 28$, and an oxygen molecule O_2 has a molecular mass of $2(16) = 32$. Air consists of approximately 80% nitrogen and 20% oxygen, and so the average molecular mass is $0.8 \times 28 + 0.2 \times 32 = 28.8$ Thus

$$m = (2.43 \times 10^3\ \text{moles})(28.8\ \text{g/mole}) = 7.00 \times 10^4\ \text{g}$$

$$= 70.0\ \text{kg}$$

The End Part-1