Temperature

Temperature is very difficult physical quantity to define, but we have a pretty good idea of what it means:

*Stick your hand in a bucket of ice water, and then a pot of boiling water – you’ll experience different temperatures!*

But, what physical quantity does temperature actually measure?

We’ll find this out in Chap 18; for now, we’ll just say that for two systems:

If: $T_1 \neq T_2$

Then: thermal energy flows between the systems.
# Temperature Scales

We will use several different temperature scales that you should have seen:

**TABLE 16.4** Temperatures measured with different scales

<table>
<thead>
<tr>
<th>Temperature</th>
<th>Celsius</th>
<th>( T (^\circ C) )</th>
<th>Kelvin (^*)</th>
<th>Fahrenheit</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting point of iron</td>
<td>1538</td>
<td>1811</td>
<td>2800</td>
<td></td>
</tr>
<tr>
<td>Boiling point of water</td>
<td>100</td>
<td>373</td>
<td>212</td>
<td></td>
</tr>
<tr>
<td>Normal body temperature</td>
<td>37</td>
<td>310</td>
<td>99</td>
<td></td>
</tr>
<tr>
<td>Room temperature</td>
<td>20</td>
<td>293</td>
<td>68</td>
<td></td>
</tr>
<tr>
<td>Freezing point of water</td>
<td>0</td>
<td>273</td>
<td>32</td>
<td></td>
</tr>
<tr>
<td>Boiling point of nitrogen</td>
<td>(-196)</td>
<td>77</td>
<td>(-321)</td>
<td></td>
</tr>
<tr>
<td>Absolute zero</td>
<td>(-273)</td>
<td>0</td>
<td>(-460)</td>
<td></td>
</tr>
</tbody>
</table>

**Conversions:**

\[
T_K = T_C + 273 \\
T_F = \frac{9}{5} T_C + 32
\]

*Note: the Kelvin scale is an absolute scale which means that the zero is at absolute zero.*
Whiteboard Problem 16-4

The lowest and highest natural temperatures ever recorded on Earth are \(-127^\circ F\) in Antarctica and \(136^\circ F\) in Libya.

a) What is \(-127^\circ F\) in Celsius?

b) What is \(136^\circ F\) in Kelvin?

Conversions:

\[
T_K = T_C + 273 \\
T_F = \frac{9}{5} T_C + 32
\]
Pressure

Pressure is an important property of fluids (gases and liquids) that will be used in all of thermodynamics. It is introduced in Chap 15 (sections 15.2 and 15.3). There are two types of pressure that we will see:

1. Gas Pressure

   When the particles collide with the wall, they exert a force
   
   \[ P = \frac{\text{force on the wall}}{\text{Area}} \]
   
   MKS Units: \( 1 \frac{N}{m^2} \equiv 1 \text{ Pascal (Pa)} \)

   e.g. Atmospheric Pressure: \( P_{\text{atm}} = 1.013 \times 10^5 \text{ Pa} = 14.7 \text{ pounds/in}^2 \)

   The atmosphere is another common unit of Pressure:

   \( P_{\text{atm}} = 1.0 \text{ atmospheres} \)

   **Gauge Pressure** is the pressure relative to the atmospheric pressure  
   (because that’s what a pressure gauge measures)

   \[ P_{\text{gauge}} = P - P_{\text{atm}} \]

   *in most of our calculations, absolute pressure must be used.*
Pressure

2. **Hydrostatic Pressure** is a contribution to the pressure in a fluid in a gravitational field. It is caused by the weight of the fluid above that point.

(If you have been to the bottom of the dive well at the Rec Center, you’ve felt hydrostatic pressure.)

e.g. in a liquid:

![Diagram of liquid with pressure at depth formula](image)

- \( P_0 \): pressure at the surface (atmospheric if the container is open to the atmosphere)
- \( \rho \): liquid density
- \( d \): depth

Pressure at depth \( d \) is:

\[
P(\text{at } d) = P_0 + \rho gd
\]
Whiteboard Problem 16-5

The deepest point in the ocean is 11 km below sea level, deeper than Mt. Everest is tall. What is the pressure in atmospheres at this depth?
The Ideal Gas Law

Consider a system that is a gas in a container:

During the 18th century it was discovered experimentally that the state variables of a gas are related to each other.

You may have seen these relations in a chemistry class – remember Charles’ Law, Boyle’s Law, and Gay Lussac’s laws?

All of these laws can be combined together into **The Ideal Gas Law***:

\[
PV = nRT
\]

Where \( R = \text{Universal Gas Constant} \)

\[
R = 8.314 \frac{J}{mol \cdot K} \quad (\text{MKS units})
\]

*Note: you may be used to seeing \( R \) as a different number. That’s a different set of units. We’ll use MKS.

*sometimes you see this referred to as an “Equation of State” since it relates the state variables of the system.
Some Important Things about the Ideal Gas Law (IGL)

Anytime you use the IGL, both the temperature and the pressure must be absolute temperature and pressure!

This means that when doing problems, we want to make sure that the temperature is in Kelvins and the pressure is the absolute pressure, not the gauge pressure.

The IGL is not valid for all gases in all situations. It is reasonably valid if:

1. The gas density is low (how low is low? We’ll see in chap 18)
2. The temperature is large compared to the condensation temperature
How do we use the IGL?

Of course, we can use the IGL to find an unknown state variable if we know all of the other ones.

In most problems that we’ll do, the state of the system will be changing from some initial state to some final state. Most of the time the quantity of gas will be fixed \((n = \text{constant})\), and we can use the IGL to find how the other state variables change:

\[
\begin{align*}
\text{Initial state} & \quad n, P_1, V_1, T_1 \\
\text{Final state} & \quad n, P_2, V_2, T_2
\end{align*}
\]

e.g. suppose temperature is constant:

\[
PV = nRT = \text{constant} \Rightarrow P_1 V_1 = P_2 V_2
\]

or, suppose pressure is constant:

\[
PV = nRT \Rightarrow \frac{V}{T} = \frac{nR}{P} = \text{constant} \Rightarrow \frac{V_1}{T_1} = \frac{V_2}{T_2}
\]

You can work out the others for constant volume or when only \(n\) is constant.
A compressed-air cylinder is known to fail if the pressure exceeds 110 atm. A cylinder that was filled to 25 atm at 20ºC is stored in a warehouse. Unfortunately, the warehouse catches fire and the temperature reaches 950ºC. **Does the cylinder blow up?**
For a fixed amount of gas ($n = \text{constant}$), we can represent the state of the system in PV phase space:

If the system changes slowly from state 1 to state 2, it follows a path in PV space.

This is really in a three dimensional PVT space, but that gets a little hard to draw!
Special Process or Paths on the PV Diagram

Constant Pressure
“Isobaric”

\[ P_1 = P_2 \]

\[ V_1 \quad V_2 \]

Constant Volume
“Isochoric”

\[ P \]

\[ P_1 \quad P_2 \]

\[ V_1 = V_2 \]

Note: for \( T = \text{constant} \):

\[ PV = nRT = \text{constant} \]

\[ \Rightarrow P = \frac{\text{constant}}{V} \propto \frac{1}{V} \]

What shape is this? It’s a hyperbola!

There’s one more special process that we’ll see in chapter 17, the adiabatic process.
0.020 moles of a gas undergoes the isothermal process shown below.

a) What is the final temperature, $T_2$, in °C?
b) What is the final volume, $V_2$?
Now for some fun*: Whiteboard Problem 16-8

A 50 kg lead piston shown in the figure floats on 0.12 mol of compressed air. (Assume that the piston is circular, as most pistons are.)

a) What is the piston height $h$ if the temperature is 30°C?
b) How far does the piston move if the temperature is increased by 100°C?

**Hint for part a:** Find the pressure first

**Hint for part b:** Does the pressure change?

*wherein we use some the great stuff that we learned in PHY191!
Homework Hints: Problem 16-72

72. The cylinder in FIGURE CP16.72 has a moveable piston attached to a spring. The cylinder’s cross-section area is 10 cm$^2$, it contains 0.0040 mol of gas, and the spring constant is 1500 N/m. At 20°C the spring is neither compressed nor stretched. How far is the spring compressed if the gas temperature is raised to 100°C?

This is a fairly tough problem. It is similar to WB16-8, but there are some important differences:

Instead of gravity pushing back on the pressure, it’s the spring force which is not constant.

Draw good initial and final sketches, and don’t be surprised if you have to solve a quadratic somewhere in your solution.