CLIMATE CHANGE TOOLKIT
How Many Molecules make a Trace Gas?

Standards
NGSS PS1.A: Structure & Properties of Matter
NGSS ESS3.D: Global Climate Change
Grade Level: Middle & High School

Equipment:
Calculator

Question: How many molecules of CO_2 are in a one-liter bottle of air at standard pressure and temperature?

The answer to this question has two parts:
(1) the Ideal Gas Law tells us how many molecules of air are in one liter
(2) we need to know the current concentration of CO_2 in air

Overview
Trace gases are gases whose volume makes up less than 1% of the Earth's atmosphere. But just because trace gases occupy a small percentage of the atmosphere's volume doesn't mean they're not important. Some trace gases—such as water vapor, carbon dioxide, and methane—play a big role in the Earth's climate system because of their influence on temperature.

Also, the name "trace gases" suggests that they're barely present in the atmosphere, as if they hardly leave a trace. But really there are vastly more trace gas molecules in a liter of air than there are people on Earth or stars in our galaxy. Because gas molecules are so tiny, it can be hard to visualize how many are present in even a small volume. This activity runs through a calculation to show how many carbon dioxide molecules are present in a liter of air.

Calculation
(1) Start with the Ideal Gas Law:

$$PV = NRT$$

P = pressure
V = volume
N = number of moles of air in the volume V
R = 8.314 J/(mole * K) (*the ideal gas constant*)
T = temperature

Let n_{air} = number density of air (number of molecules of air per unit volume)

$$n_{air} = \frac{AN}{V}$$

where A = Avogadro's number = 6.022 x 10^{23} molecules/mole

Substitute for V from the Ideal Gas Law:

$$n_{air} = AN * \frac{P}{NRT} = \frac{AP}{RT}$$

Using the values above and values for standard pressure (10^5 Pa) and temperature (273 K),

$$n_{air} = \frac{\left(6.022 * \ 10^{23} \frac{molecules}{mole}\right)(10^5\ Pa)}{8.31 \frac{J}{mole * K} * 273\ K} = \ 2.69 * 10^{25}\ \frac{molecules}{m^3}.$$

(2) We can look up the current value of atmospheric CO_2 concentration measured at the NOAA Mauna Loa Atmospheric Observatory:
(https://scripps.ucsd.edu/programs/keelingcurve/)

$$C\ = 417\ ppmv\ (parts\ per\ million\ by\ volume).$$

So the number density of CO_2 in air is:

$$n_{CO2} = \ Cn_{air} = (417 * \ 10^{-6})\left(2.69 * 10^{25}\ \frac{molecules}{m^3}\right) = 1.12 * \ 10^{22}\ \frac{molecules}{m^3}.$$

Finally, to answer our original question, adjust the volume from m^3 to 1 liter.
A one-liter bottle has a volume of 0.001 m^3, so the number of molecules of CO_2 in this bottle is:

$$0.001\ m^3 * 1.12 * 10^{22} \frac{molecules}{m^3} = \mathbf{1.12 * \ 10^{19}\ \boldsymbol{molecules}}.$$

That's about ten-billion-billion molecules!

STUDENT DATA TABLE: *fill in green cells*

Parameter	***Value***	***Units***
P = pressure	10^5	Pa = N/m^2
V = volume = 1 liter	0.001	m^3
N = number of moles of air in the volume *V*		
R = ideal gas constant	8.314	J/(mole*K)
T = temperature	273	K
A = Avogadro's number	6.022 x 10^{23}	molecules/mole
n_{air} = "number density" = the answer we are looking for in part (1); how many molecules of air per m^3		molecules/m^3
C = concentration of CO_2	[look this up]	ppmv
n_{CO2} = number density of CO2 molecules per m^3 = the answer to part (2)		molecules/m^3
Part (1): calculate n_{air}		
n_{air} = AP/RT		
Part (2): calculate n_{CO2}		
n_{CO2} = C*n_{air}		
Final answer: convert volume from m^3 to 1-liter		
# molecules CO_2 in a 1-liter bottle at STP		