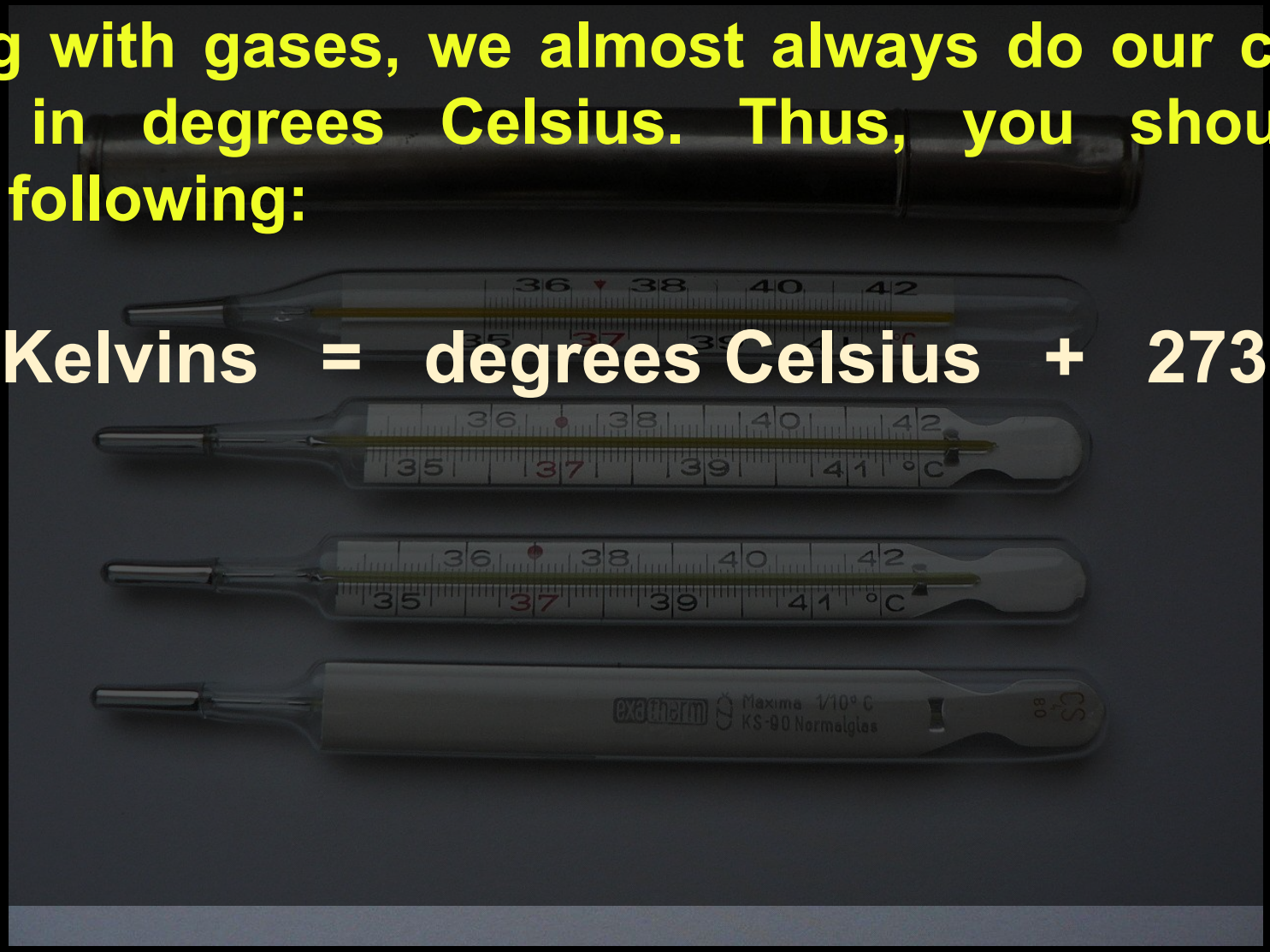

Gases
(Temperature, Volume, and Pressure)

Temperature

When working with gases, we almost always do our calculations in Kelvins, not in degrees Celsius. Thus, you should *definitely* memorize the following:

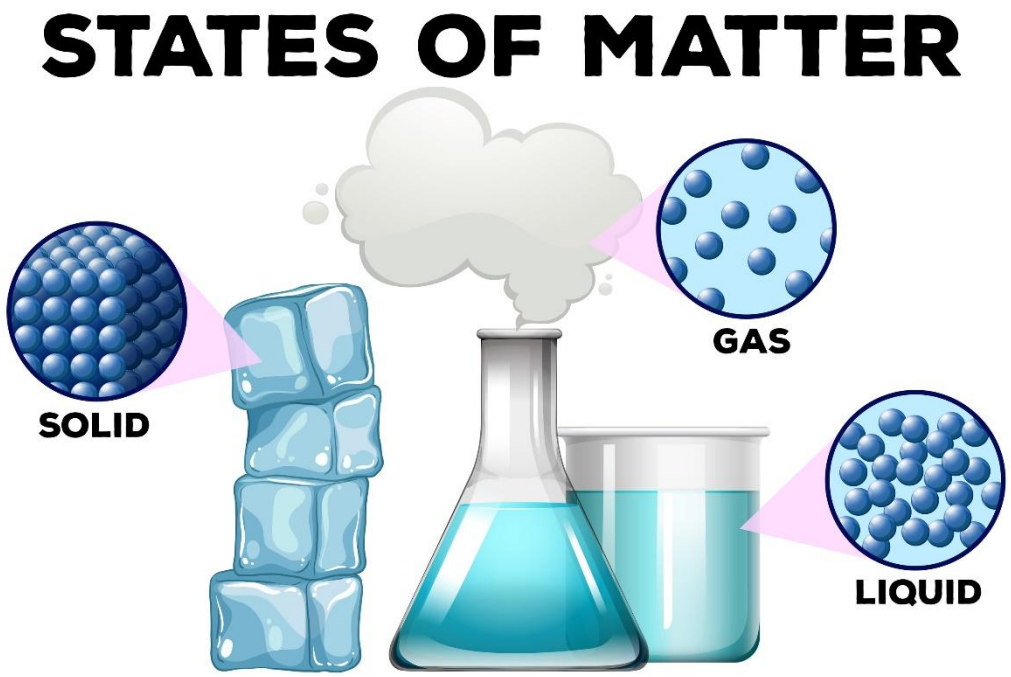


Kelvins = degrees Celsius + 273

Volume

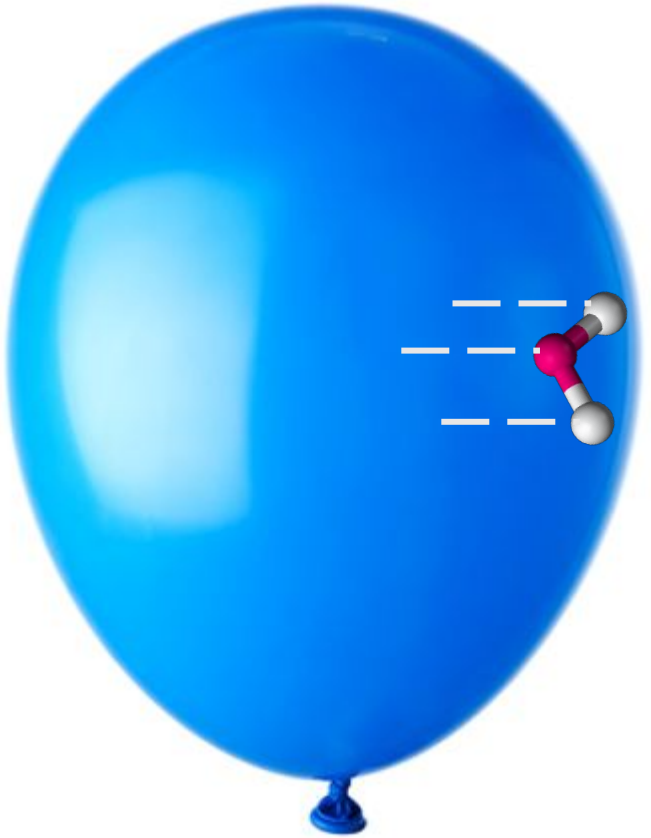
As you likely know, both ***solids*** and ***liquids*** have distinct ***volumes***. Liquids' shapes conform to that of their containers.

Although ***gases*** <u>also</u> conform to the shapes of their containers, gases do NOT have distinct ***volumes***. Instead, gas molecules typically spread out to fill whatever container they are placed in. Thus, their volumes can change considerably. Furthermore, gases can be compressed or expanded to occupy much larger or much smaller volumes.



Volume

For example, the volume (size) of a balloon is caused by gas molecules pressing, colliding, or hitting against the walls of the balloon. This is also true of all expandable containers that are filled with gases.



The S.I. unit for volume is ***liters***. You should definitely memorize the following:

$1\ cm^3 = 1\ mL$

Volume

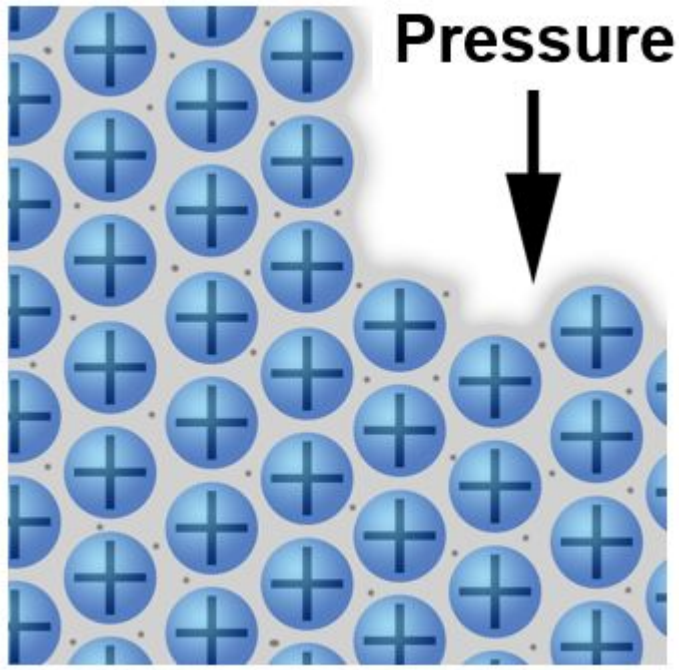
$1\ cm^3 = 1\ mL$

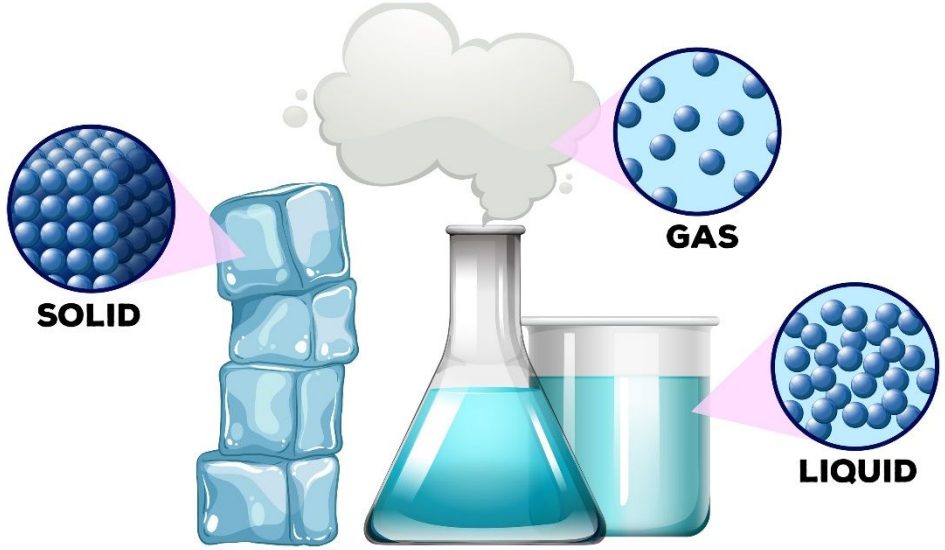
A cubic centimeter is also called a “cc”, which is the unit typically used to convey the volume size of a motorcycle’s engine.

Pressure

As I mentioned a moment ago, ***gases*** are much more compressible than liquids or solids. In other words, if you impose ***pressure*** on a gas, you can decrease its ***volume*** and place it in a much smaller container.

Pressure = Force/Area

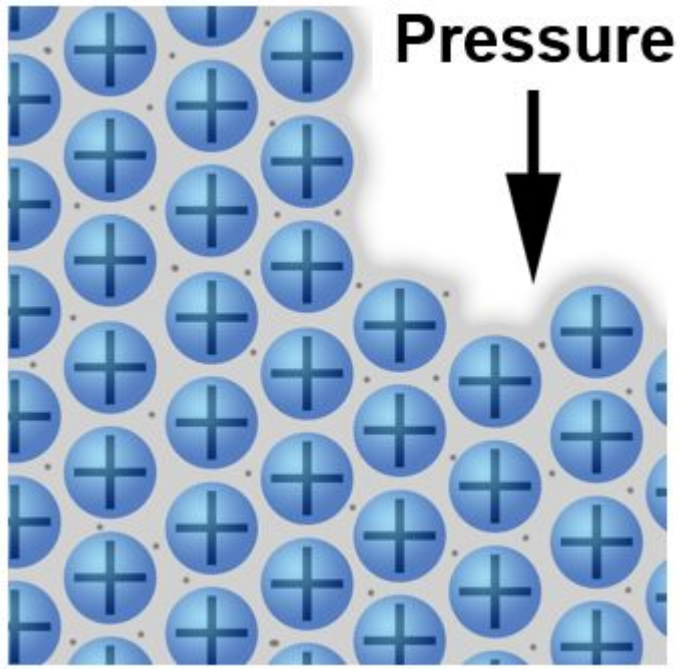




Pressure

To increase the pressure, you can either increase the *force*, or decrease the *area*.

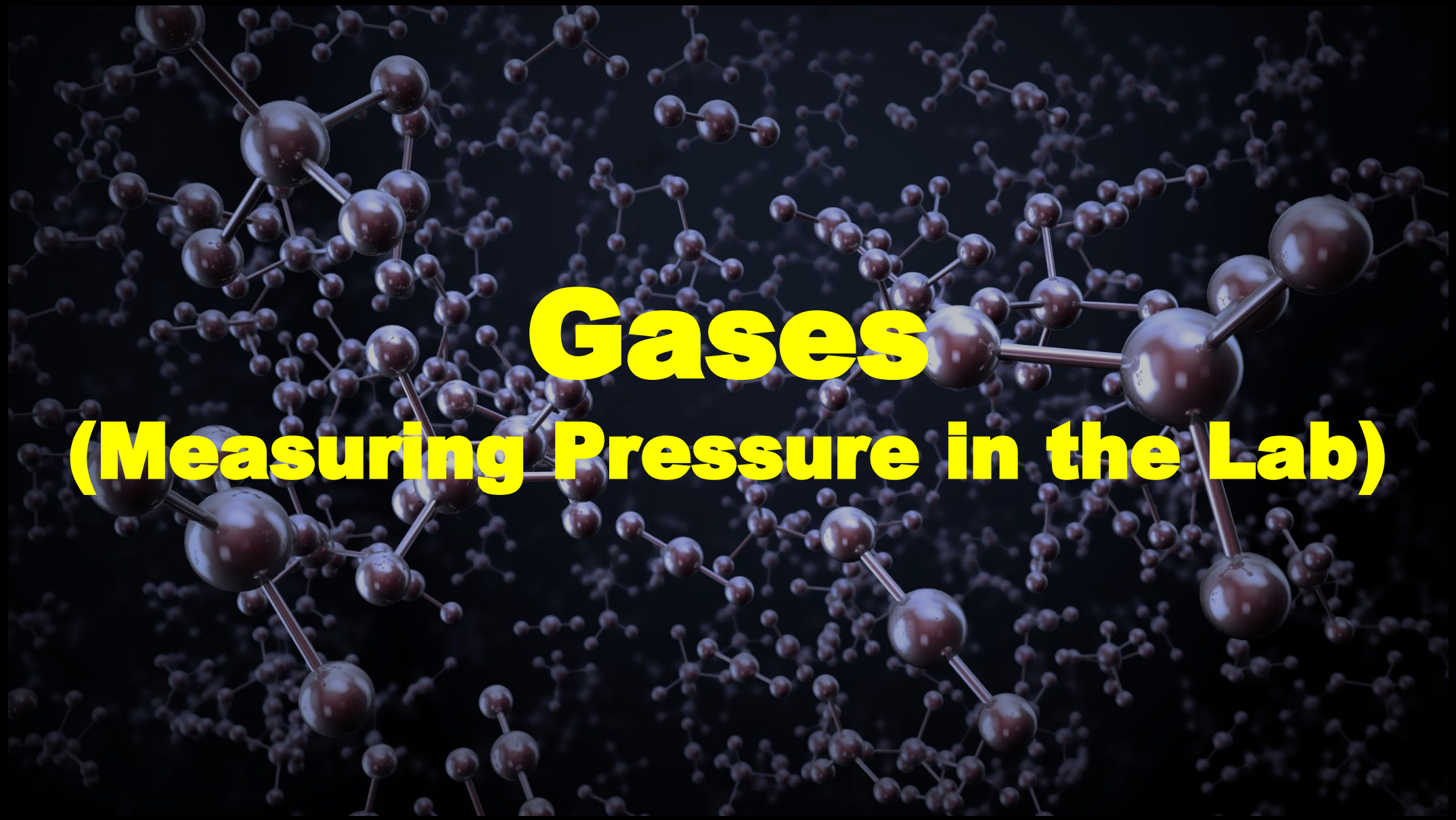
Pressure = Force/Area



Pressure Units

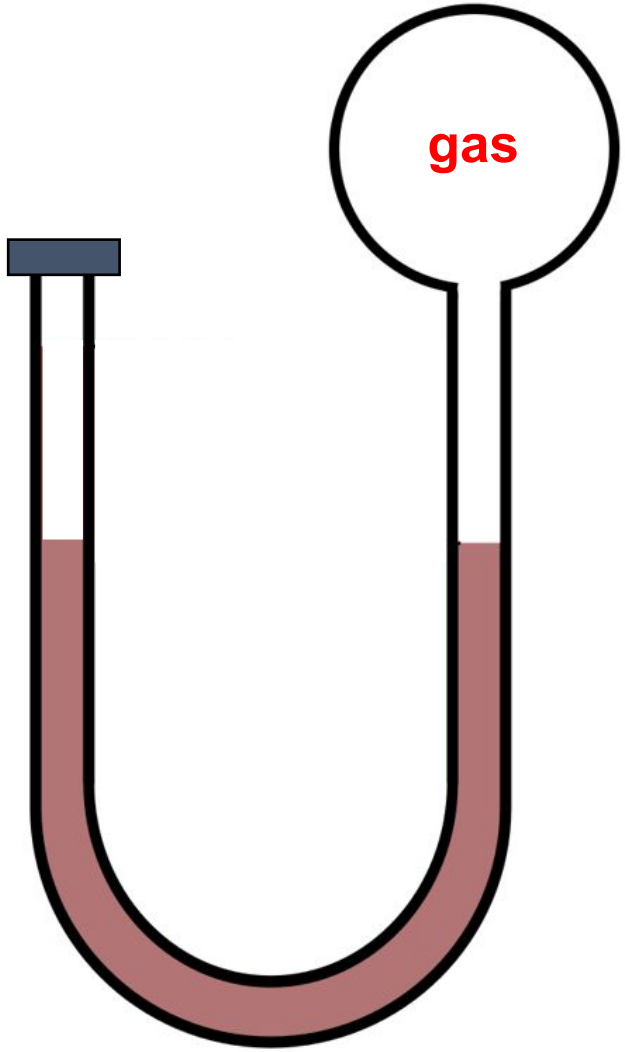
I finish this section by telling you that you should definitely memorize the following:

1 atm = 760 torr = 760 mmHg


Gases
(Measuring Pressure in the Lab)

Using a Hg Manometer

We often measure a gas's pressure by using a ***mercury*** (**Hg**) ***manometer***.

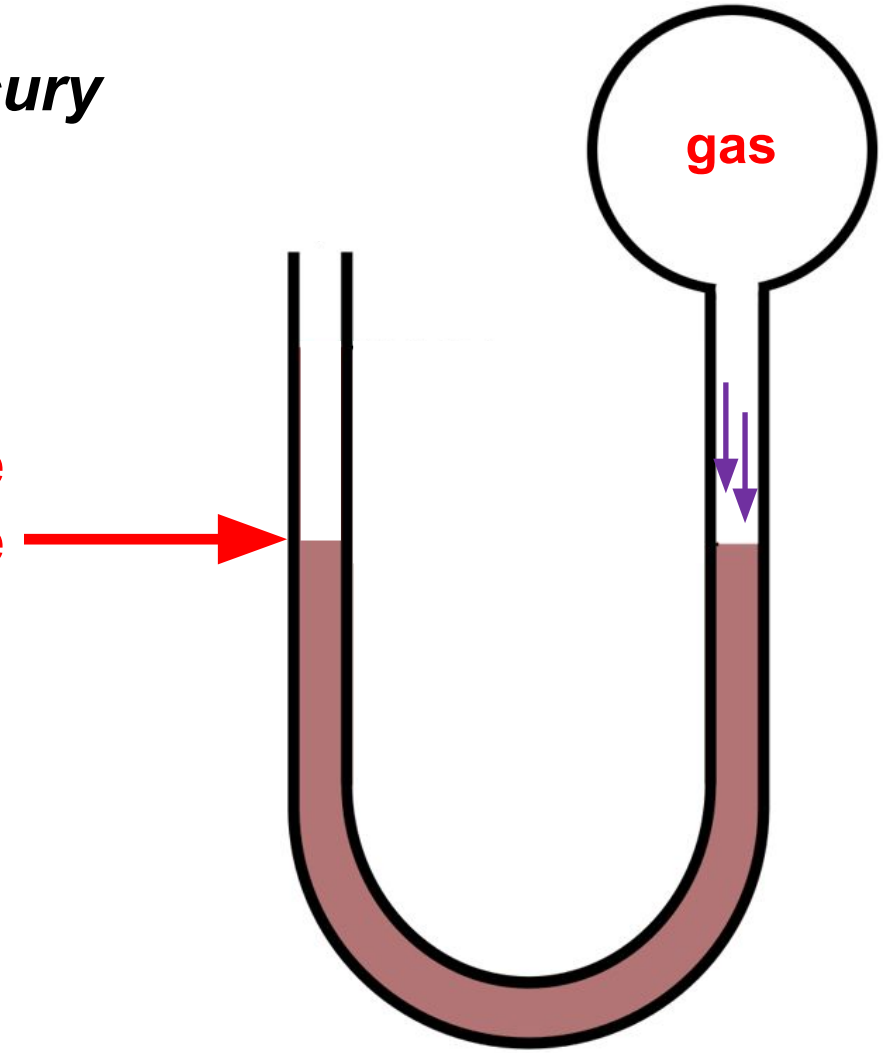


mercury (Hg) manometer

Using a Hg Manometer

We often measure a gas's pressure by using a ***mercury*** (**Hg**) ***manometer***.

If the pressure of the gas and the atmosphere are equal (1 atm), then the Hg height will be the same in both tubes.



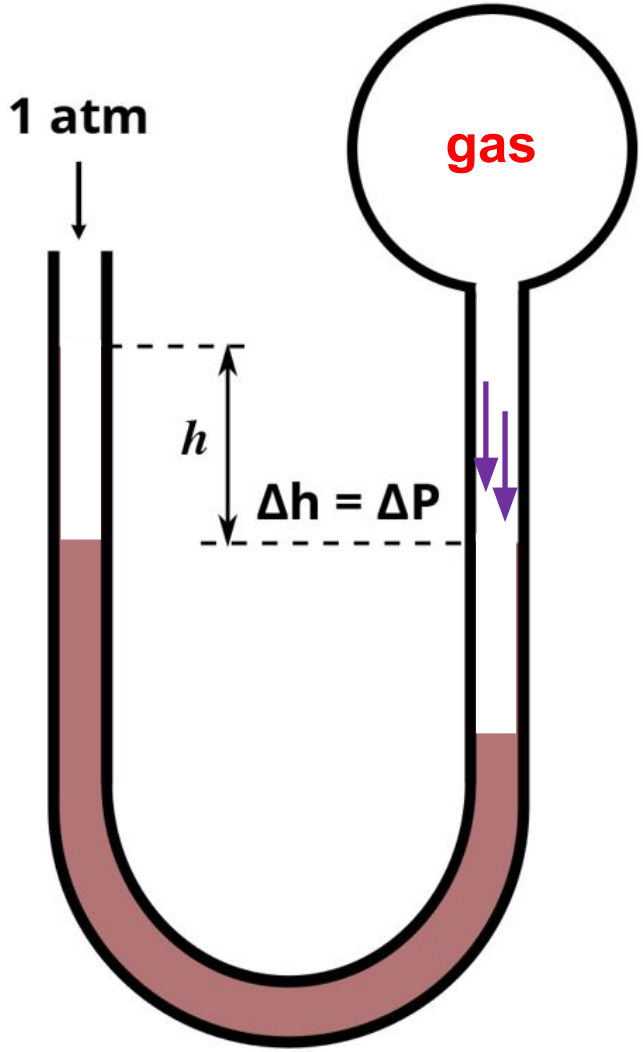
mercury (Hg) manometer

Using a Hg Manometer

We often measure a gas's pressure by using a ***mercury*** (**Hg**) ***manometer***.

If the pressure of the unknown gas is greater than the atmosphere's, then the gas will push down the tube of mercury.

We can use this information to help us identify the gas.

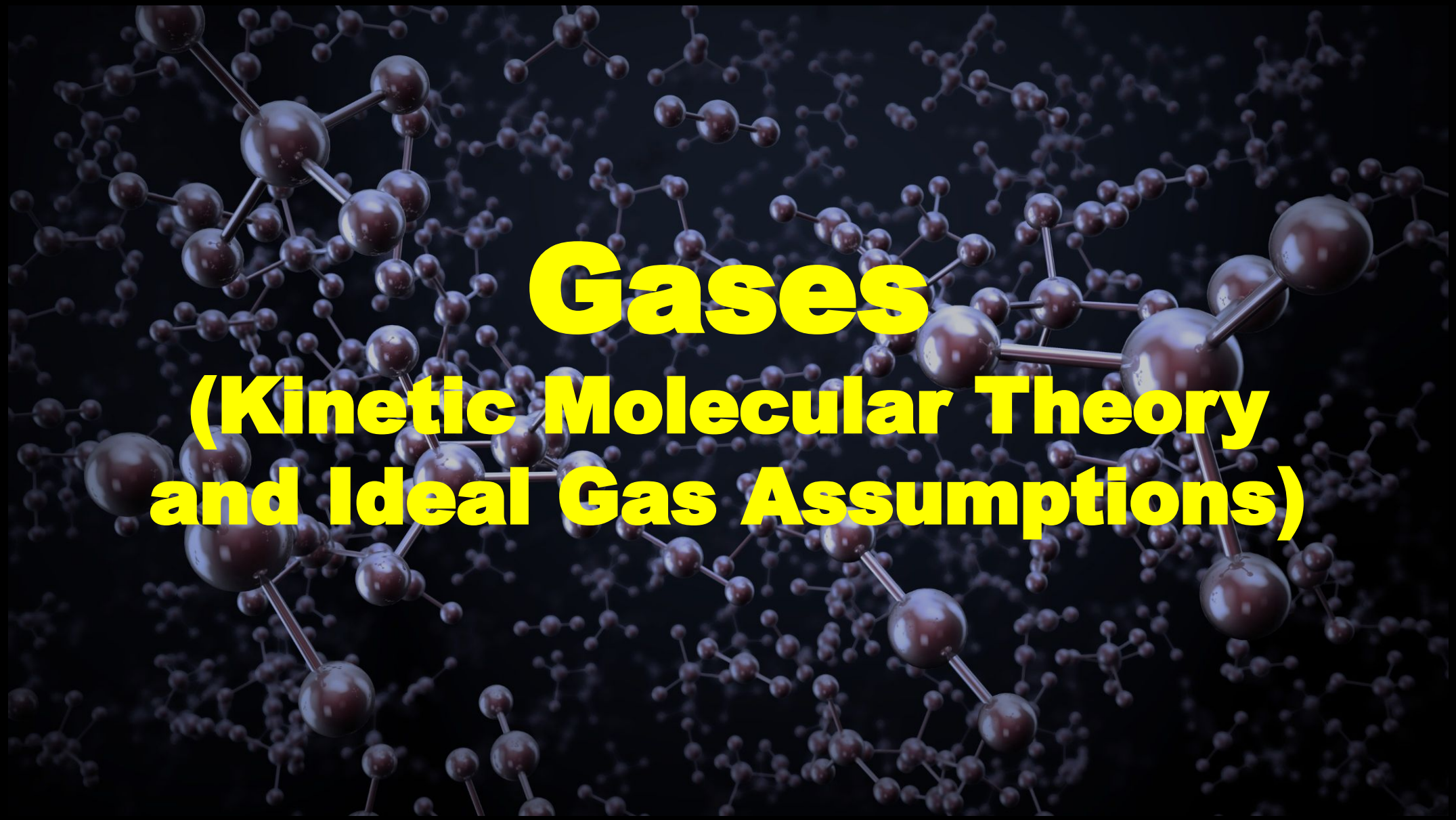


mercury (Hg) manometer

Using a Hg Manometer

Say the unknown gas pushes down the tube of mercury by 40mm. What is the pressure of the gas?

The change in height is equal to 40mm, and if we're using mercury, we also know 760 mmHg = 1 atm. So the pressure of the gas is equal to whatever atmospheric pressure is, in this example 760 mmHg, plus the extra 40mm, so the pressure of the unknown gas is 800 mmHg


Gases
(Kinetic Molecular Theory
and Ideal Gas Assumptions)

What is an Ideal Gas?

An ***ideal gas*** is a theoretical gas whose behavior can be perfectly predicted by the ***ideal-gas law*** (which I'll explain in a later video).

In real life, most gases behave pretty much like an ***ideal gas*** (they behave "ideally") within a certain range of temperatures and pressures. Gases begin to behave non-ideally under extreme temperatures and pressures.

In order for a gas to behave ***ideally*** (that is, like an ***ideal gas***), there are three things that it must do. We call these the three "assumptions" of an ***ideal gas.***

Ideal Gas Assumptions

These assumptions simplify gases' real-life behaviors in order to make easier predictions about them. In other words, we assume that:

1. Gas particles' volumes are insignificant compared to the massive space between them.
2. Gas particles experience zero intermolecular forces with each other and their containers.
3. Gas particles are always in continuous, random motion
4. Gas molecules' collisions with each other are perfectly elastic, so there is zero loss or gain of kinetic energy when they collide with each other.
5. The average ***kinetic energy*** (E_K) of a gas depends only on the system's ***temperature***.

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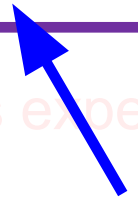


Exam Tip: *This* assumption is most accurate at LOW pressures.

2. Gas particles experience zero intermolecular forces with each other and their containers.

3. Gas particles are in constant, random motion.

4. Gas molecules' collisions with each other are perfectly elastic, so there is zero loss or gain of kinetic energy when they collide with each other.

5. The average ***kinetic energy*** (E_K) of a gas depends only on the system's ***temperature***.

Ideal Gas Assumptions

These assumptions simplify gases' real-life behaviors in order to make easier predictions about them. In other words, we assume that:

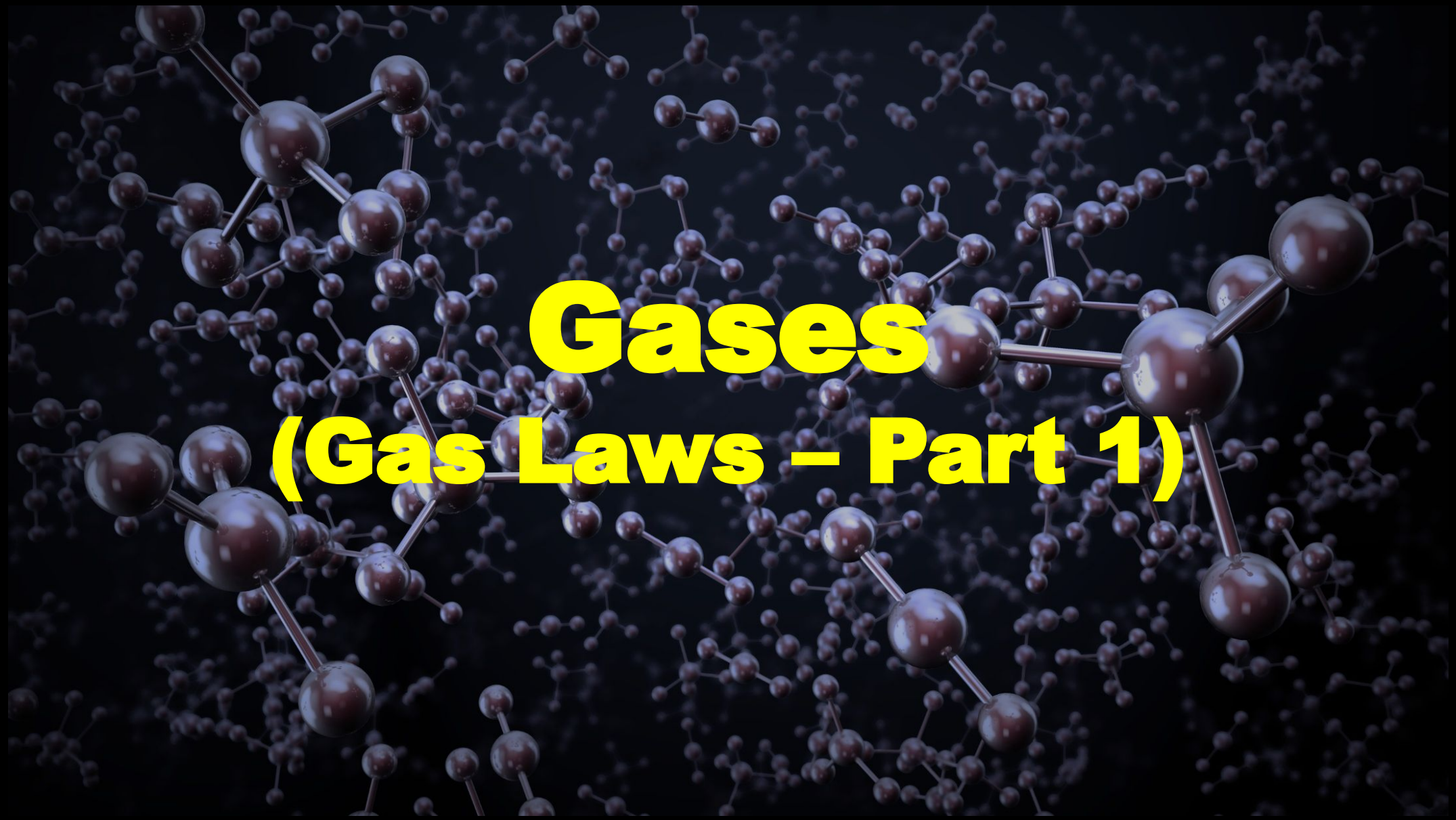
<u>Exam Tip</u>: *This* assumption is most accurate at high temperatures and for gases with weak intermolecular forces

1. Gas particles' volumes are insignificant compared to the massive space between them.
2. Gas particles experience zero intermolecular forces with each other and their containers.
3. Gas particles are always in continuous, random motion
4. Gas molecules' collisions with each other are perfectly elastic, so there is zero loss or gain of kinetic energy when they collide with each other.
5. The average ***kinetic energy*** (E_K) of a gas depends only on the system's ***temperature***.

Recap

These assumptions simplify gases' real-life behaviors in order to make easier predictions about them. In other words, we assume that:

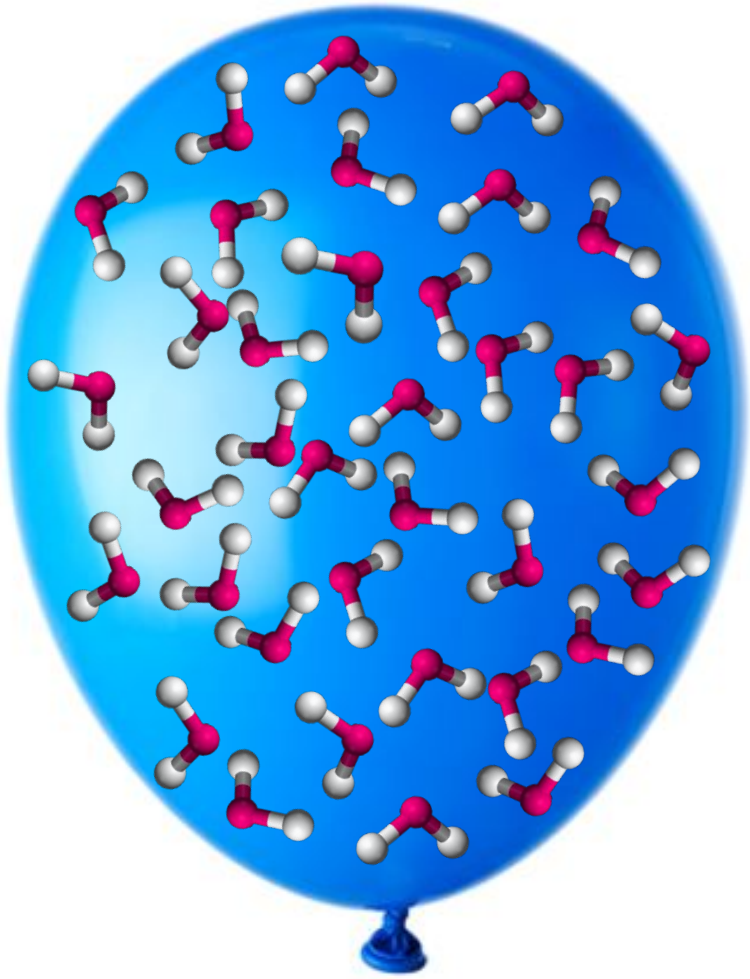
1. Gas particles' volumes are insignificant compared to the massive space between them.
2. Gas particles experience zero intermolecular forces with each other and their containers.
3. Gas particles are always in continuous, random motion
4. Gas molecules' collisions with each other are perfectly elastic, so there is zero loss or gain of kinetic energy when they collide with each other.
5. The average ***kinetic energy*** (E_K) of a gas depends only on the system's ***temperature***.


Gases
(Gas Laws – Part 1)

Boyle's Law

Imagine that you have a balloon filled with gas molecules. Those gas molecules will, of course, exert a pressure against the walls of the balloon.

Now, imagine that the volume of the balloon suddenly decreases.

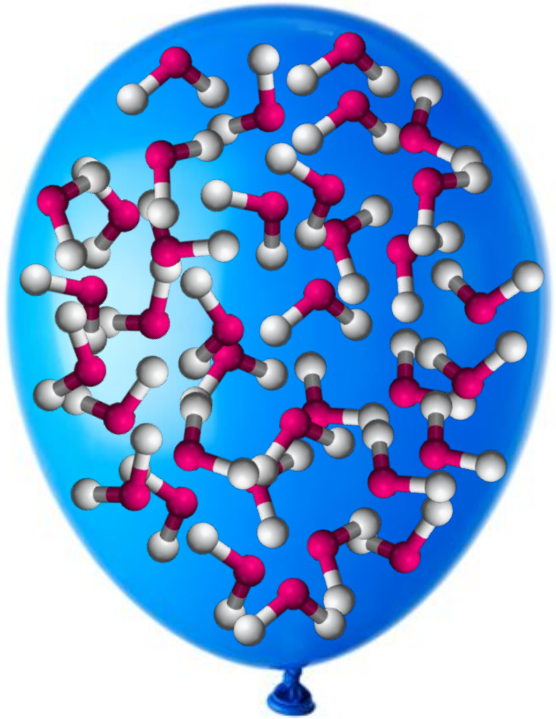


Boyle's Law

Imagine that you have a balloon filled with gas molecules. Those gas molecules will, of course, exert a pressure against the walls of the balloon.

Now, imagine that the volume of the balloon suddenly decreases.

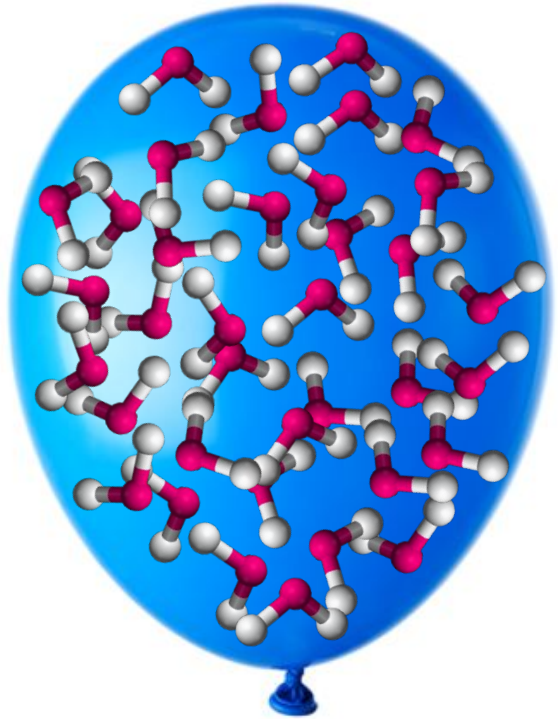
What will happen to the pressure inside the balloon?



Boyle's Law

It's going to increase! We see, then, that as long as everything else remains constant:

$$\downarrow V \propto \uparrow P$$

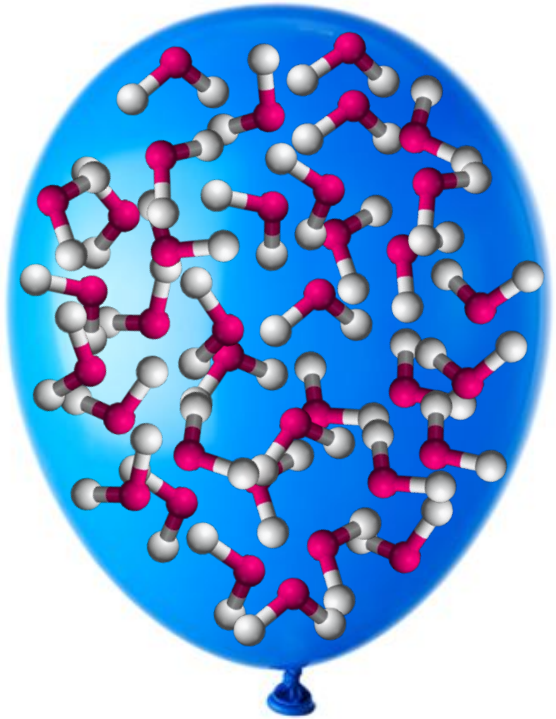


Boyle's Law

It's going to increase! We see, then, that as long as everything else remains constant:

$$V \propto \frac{1}{P}$$

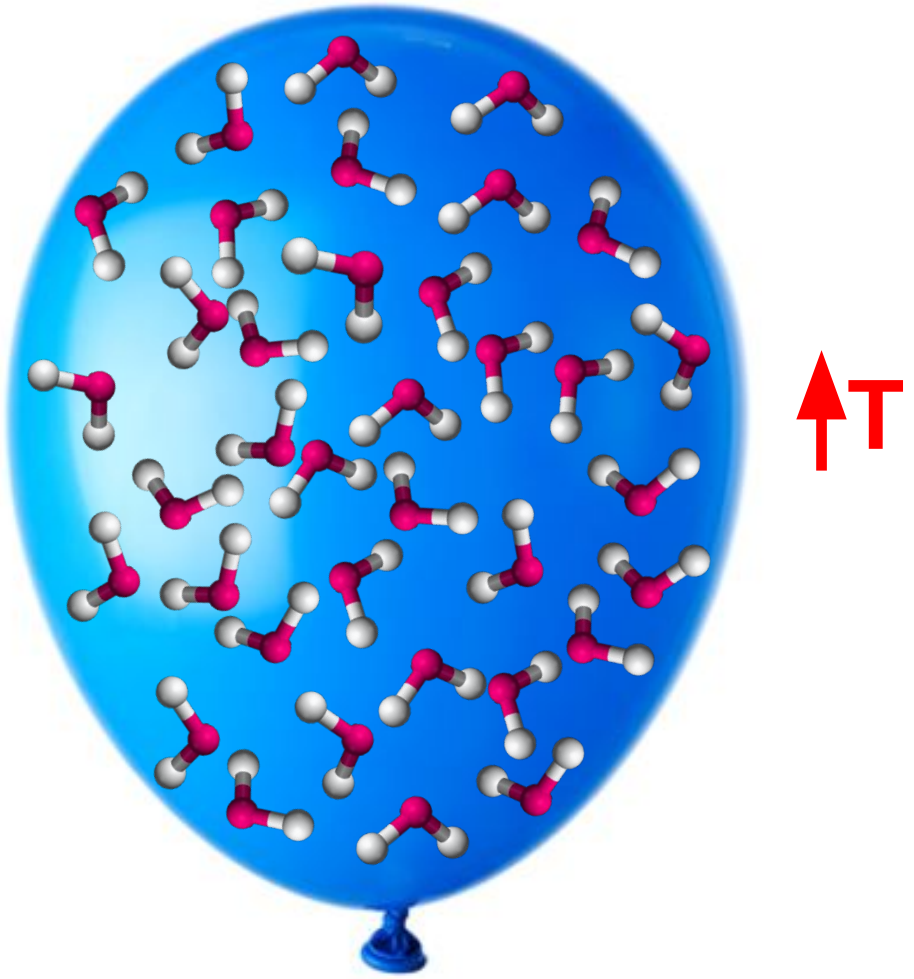
This is called ***Boyle's Law***.



Charles' Law

Charles' Law states that as long as everything else remains constant, volume (**V**) is directly proportional to temperature (**T**):

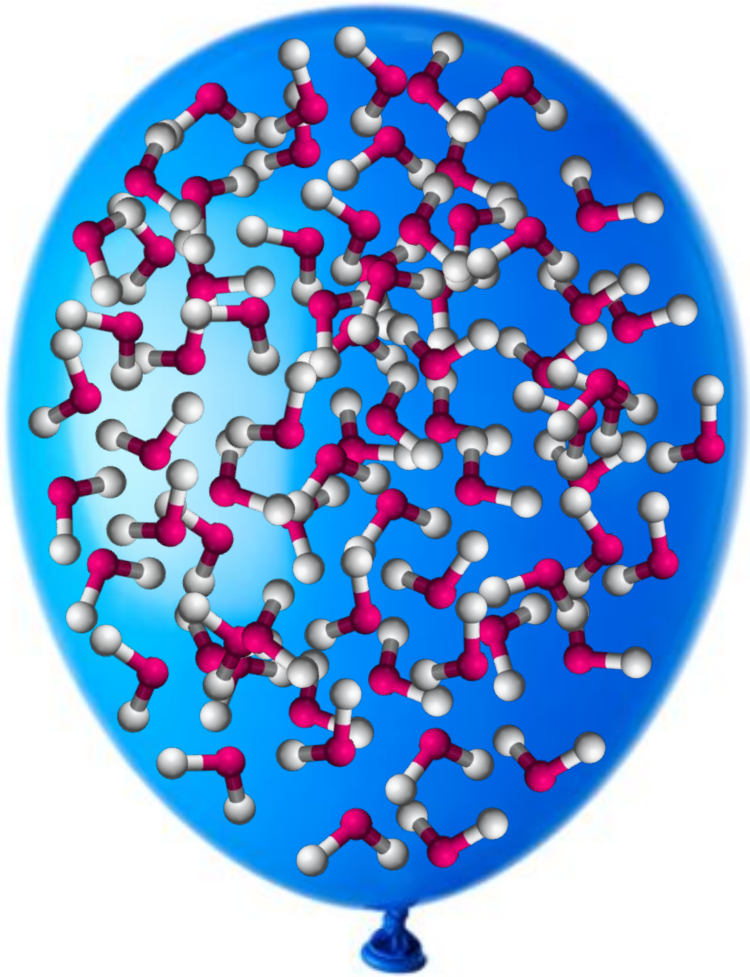
$$\uparrow V \propto \uparrow T$$



Avogadro's Law

Avogadro's Law states that as long as everything else remains constant, volume (**V**) is directly proportional to the number of gas moles (**n**):

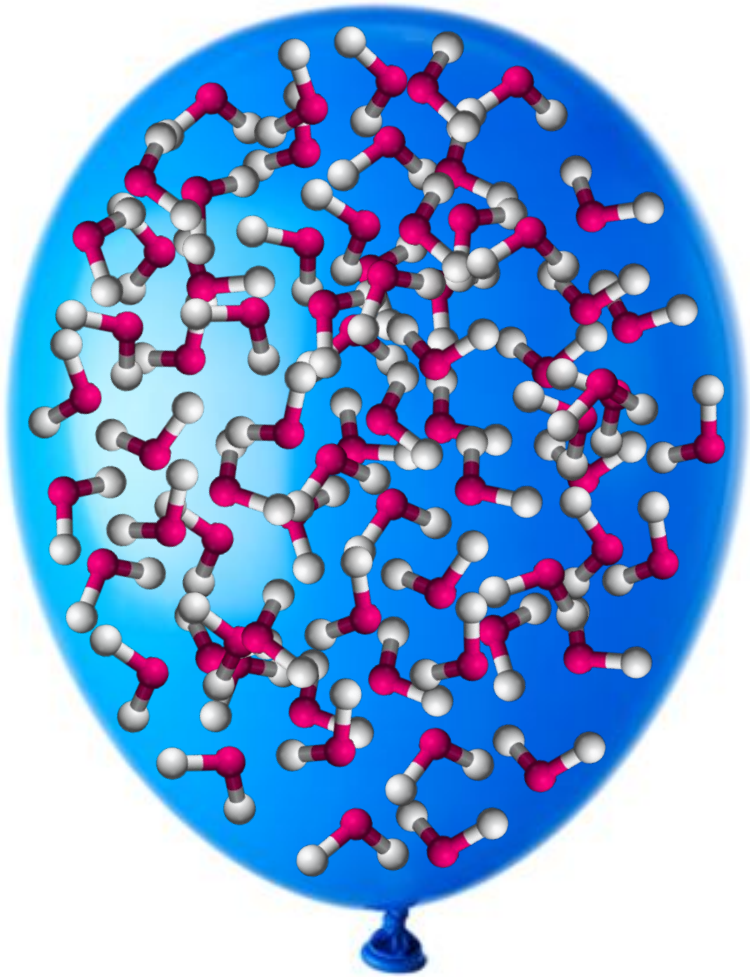
$$\uparrow V \propto \uparrow n$$



Combined Gas Law

The ***Combined Gas Law*** puts pressure (**P**), volume (**V**), temperature (**T**), and number of moles (**n**) together:

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$



If we change ***P***, ***V***, ***n***, or ***T***, this law lets us calculate what the new ***P***, ***V***, ***n***, or ***T*** will be.

Ideal Gas Law

If we know all but one of the following: ***P***, ***V***, ***n***, or ***T***, then the ***Ideal Gas Law*** enables us to calculate the one that we don't know.

$$PV = nRT$$

Where ***P*** is the gas's pressure, ***V*** is the gas's volume, ***n*** is the number of moles of the gas, ***T*** is the gas's temperature, and ***R*** is the ***ideal gas constant***.

The accuracy of the ***ideal gas law*** increases as a gas behaves more "ideally." Gases behave more "ideally" at high temperatures and low pressure. Gases behave less ideally under extreme temperatures or pressures.

Things You Should Memorize

- **Ideal gas constant: *R* = 0.0821 L·atm/mol·K**

  You should definitely memorize these units! In any ideal gas question, you will want to make sure to convert to *these units* (**L·atm/mol·K**).

- ***STP*** stands for "Standard Temperature and Pressure." ***STP* is 0° Celsius (273 K) and 1 atm (760 mmHg or 760 Torr)**

- At ***STP***, **1 mole of any gas = 22.4 L**

Question

A flask contains 2.0 mol of He gas at 25°C and 1.00 atm. How much He gas, in grams, must be added to increase the pressure to 2.00 atm at constant temperature and volume?

A. 4.0 g

B. 8.0 g

C. 12

D. 16 g

E. 20 g

Question

A 0.02 mol sample of an ideal gas in a sealed 1.0 L container was heated from 27 °C to 227°C. What is the pressure of the gas at this new temperature?

A. 0.044 atm

B. 0.37 atm

C. 0.40 atm

D. 0.82 atm

E. 1.3 atm

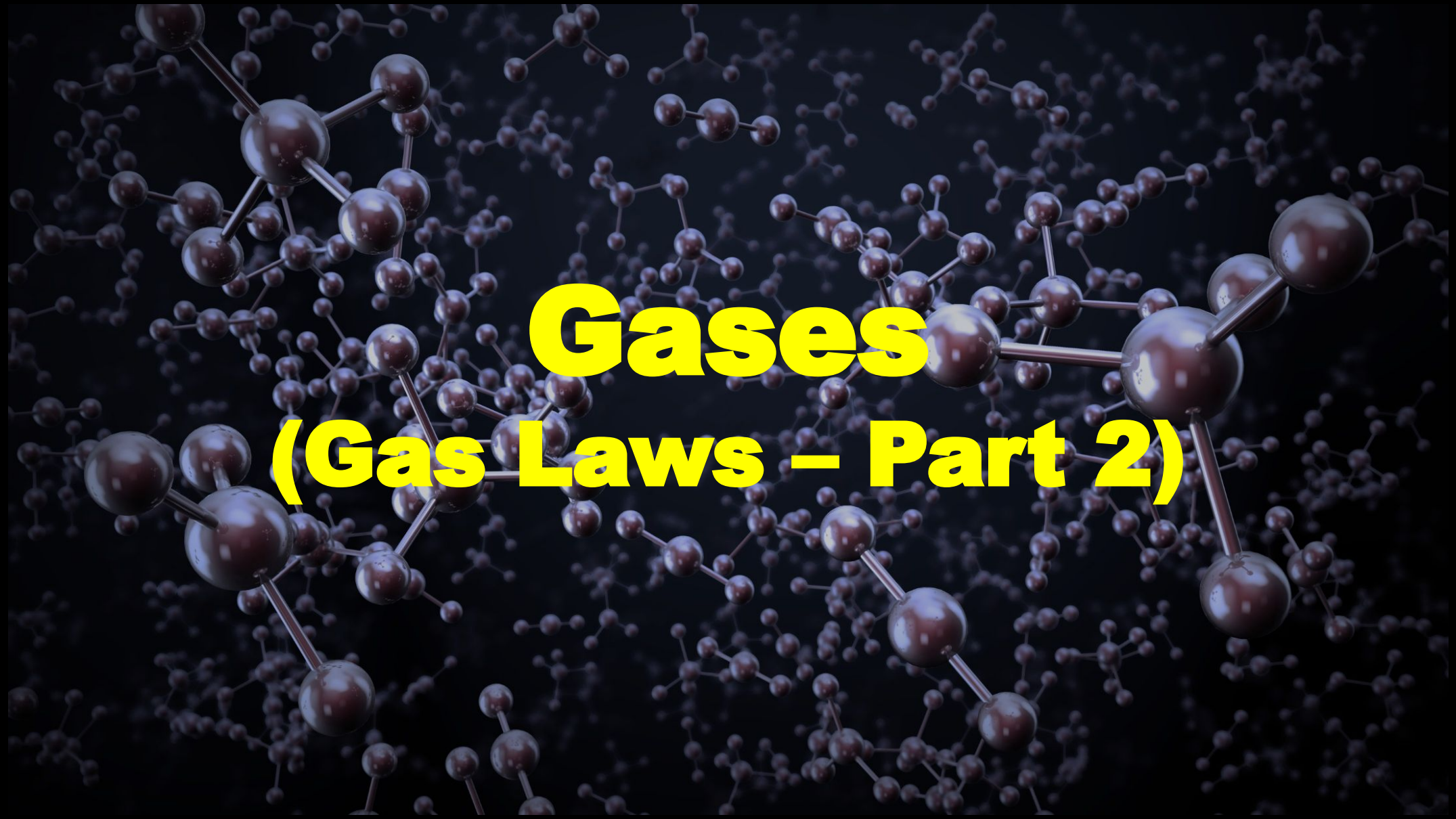
Question

Consider the gaseous reaction

$$A_{(g)} + B_{(g)} \rightarrow AB_{(g)}$$

A mixture of 0.04 mole of gas A and 0.1 mole of gas B was allowed to react at 300 K in a 1 L sealed container. At the end of the reaction, what is the total pressure of the mixture? (The temperature and volume remain constant during the reaction.)

A. 0.99 atm
B. 1.3 atm
C. 1.5 atm
D. 2.5 atm
E. 3.4 atm


Gases
(Gas Laws – Part 2)

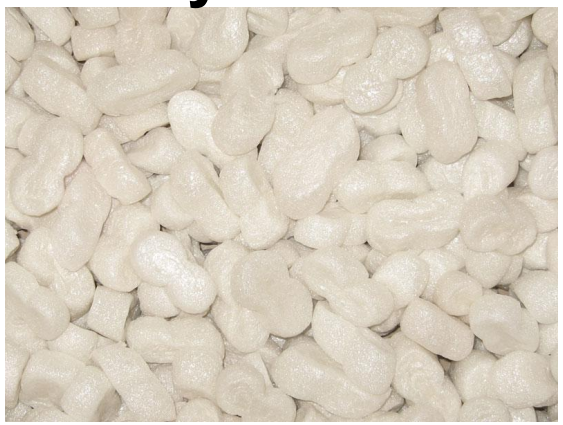
A Gas's Density

Generally speaking, ***density*** is how heavy a certain amount of a substance is. To be more technical:

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

gold

Styrofoam



A Gas's Density

A gas's ***density*** is technically the same thing: its mass divided by its volume. However, with gases, there is *another* way to calculate ***density:***

$$\text{density} = \frac{P \times \text{molar mass}}{R \times T}$$

Where ***P*** is the gas's pressure, ***R*** is the ***ideal gas constant*** (**0.0821 L·atm/mol·K**), and ***T*** is the gas's temperature.

A Gas's Density

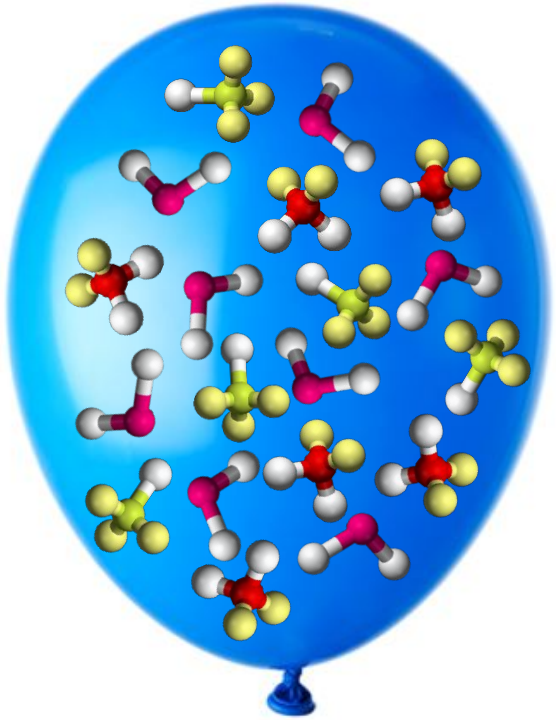
A gas's ***density*** is technically the same thing: its mass divided by its volume. However, with gases, there is *another* way to calculate ***density:***

$$\text{density} = \frac{P \times \text{molar mass}}{R \times T}$$

This equation is helpful if a problem gives you ***pressure***, ***temperature***, and ***gas density***, and then asks you to identify which gas it is. Here, you would simply use these values to calculate the ***molar mass*** and then identify the gas.

Dalton's Law of Partial Pressures

Dalton's Law of Partial Pressures says that the total pressure inside a container filled with multiple gases is equal to the sum of all of those gases' *individual* pressures.



Pressure of gas A = 1 atm

Pressure of gas B = 0.5 atm

Pressure of gas C = 0.25 atm

Total Pressure = 1 + 0.5 + 0.25
= 1.75 atm

Dalton's Law of Partial Pressures

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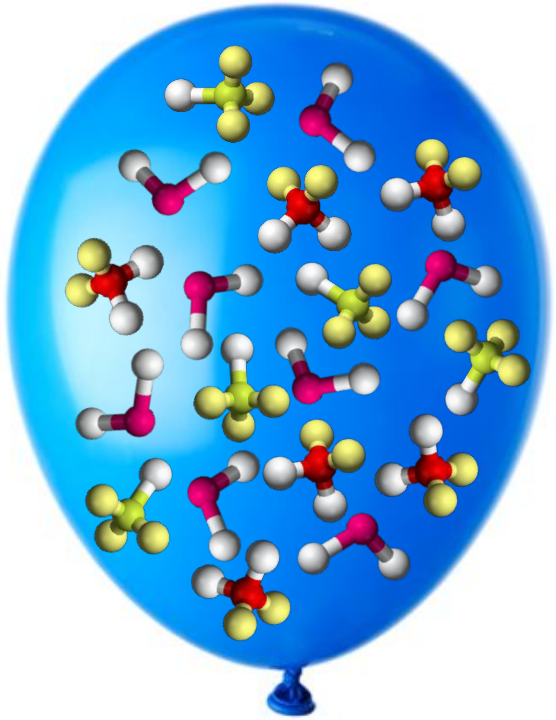
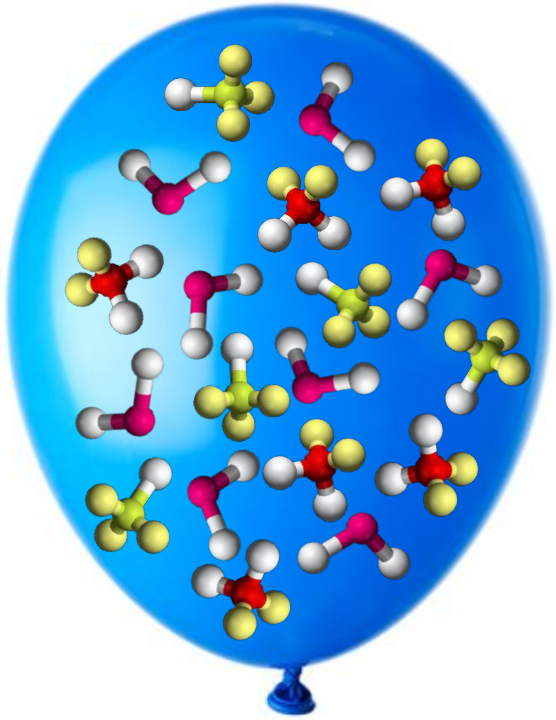


$$P_{total} = P_A + P_B + P_C \ldots$$

Dalton's Law of Partial Pressures

Alternatively, you can calculate P_A, P_B, P_C, etc., as follows:



$$P_A = (\chi_A) \times P_{total}$$

Let's pretend that ¼ (25%) of all the gas moles in this balloon are gas ***A***. Because of this, the pressure of gas ***A*** (P_A) equals: $(0.25) \times P_{total}$

Example Question

You have an expandable container whose total pressure (P_{total}) is 100 atm. Contributing to that system are the following gases, with the given numbers of moles (n):

- n_{H2O} = 5 mol
- n_{N2} = 10 mol
- n_{CO2} = 15 mol

Total moles = 5 + 10 + 15 = 30 moles

% CO_2 = 15 / 30 = 50%

What is the partial pressure of CO_2 (P_{CO2})?

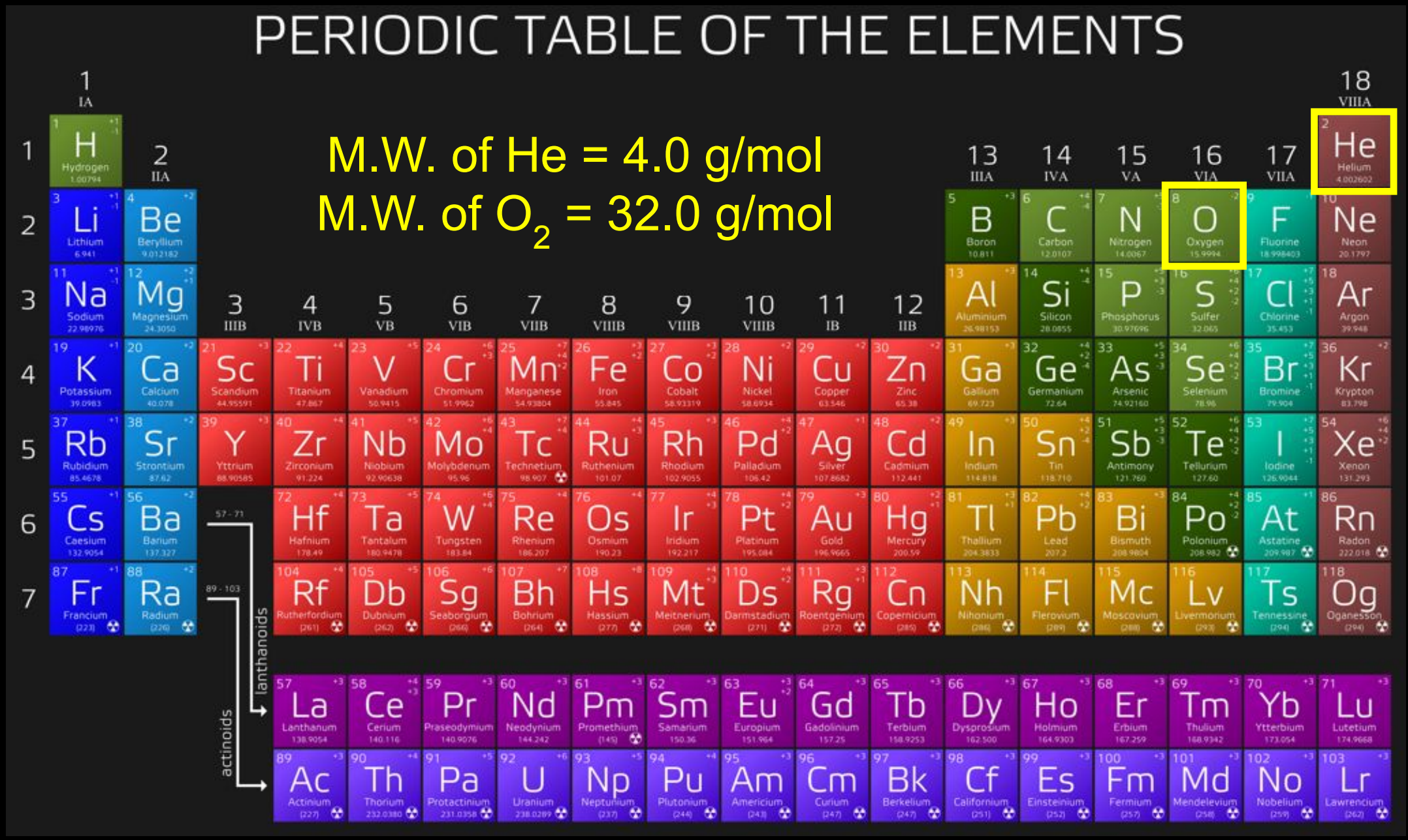
$$P_{CO2} = \text{mole fraction} \times P_{total}$$

$$= 0.50 \times 100 \text{ atm} = 50 \text{ atm}$$

Graham's Law of Effusion

When gases escape through a narrow slit, the lighter gases (gases with lower molecular weights) escape (or "effuse") more quickly. Heavier gases (gases with higher molecular weights) ***effuse*** more slowly. This is called ***Graham's Law of Effusion***.

If a balloon is filled with O_2 and He, which gas will effuse first?


PERIODIC TABLE OF THE ELEMENTS
M.W. of He = 4.0 g/mol
M.W. of O_2 = 32.0 g/mol
1 IA
2 IIA
3 IIIB
4 IVB
5 VB
6 VIB
7 VIIB
8 VIIIB
9 VIIIB
10 VIIIB
11 IB
12 IIB
13 IIIA
14 IVA
15 VA
16 VIA
17 VIIA
18 VIIIA
H Hydrogen 1.00794
He Helium 4.002602
Li Lithium 6.941
Be Beryllium 9.012182
B Boron 10.811
C Carbon 12.0107
N Nitrogen 14.0067
O Oxygen 15.9994
F Fluorine 18.998403
Ne Neon 20.1797
Na Sodium 22.98976
Mg Magnesium 24.3050
Al Aluminium 26.98153
Si Silicon 28.0855
P Phosphorus 30.97696
S Sulfer 32.065
Cl Chlorine 35.453
Ar Argon 39.948
K Potassium 39.0983
Ca Calcium 40.078
Sc Scandium 44.95591
Ti Titanium 47.867
V Vanadium 50.9415
Cr Chromium 51.9962
Mn Manganese 54.93804
Fe Iron 55.845
Co Cobalt 58.93319
Ni Nickel 58.6934
Cu Copper 63.546
Zn Zinc 65.38
Ga Gallium 69.723
Ge Germanium 72.64
As Arsenic 74.92160
Se Selenium 78.96
Br Bromine 79.904
Kr Krypton 83.798
Rb Rubidium 85.4678
Sr Strontium 87.62
Y Yttrium 88.90585
Zr Zirconium 91.224
Nb Niobium 92.90638
Mo Molybdenum 95.96
Tc Technetium 98.907
Ru Ruthenium 101.07
Rh Rhodium 102.9055
Pd Palladium 106.42
Ag Silver 107.8682
Cd Cadmium 112.441
In Indium 114.818
Sn Tin 118.710
Sb Antimony 121.760
Te Tellurium 127.60
I Iodine 126.9044
Xe Xenon 131.293
Cs Caesium 132.9054
Ba Barium 137.327
Hf Hafnium 178.49
Ta Tantalum 180.9478
W Tungsten 183.84
Re Rhenium 186.207
Os Osmium 190.23
Ir Iridium 192.217
Pt Platinum 195.084
Au Gold 196.9665
Hg Mercury 200.59
Tl Thallium 204.3833
Pb Lead 207.2
Bi Bismuth 208.9804
Po Polonium 208.982
At Astatine 209.987
Rn Radon 222.018
Fr Francium (223)
Ra Radium (226)
Rf Rutherfordium (261)
Db Dubnium (262)
Sg Seaborgium (266)
Bh Bohrium (264)
Hs Hassium (277)
Mt Meitnerium (268)
Ds Darmstadium (271)
Rg Roentgenium (272)
Cn Copernicium (285)
Nh Nihonium (286)
Fl Flerovium (289)
Mc Moscovium (288)
Lv Livermorium (293)
Ts Tennessine (294)
Og Oganesson (294)
57 - 71
89 - 103
lanthanoids
actinoids
La Lanthanum 138.9054
Ce Cerium 140.116
Pr Praseodymium 140.9076
Nd Neodymium 144.242
Pm Promethium (145)
Sm Samarium 150.36
Eu Europium 151.964
Gd Gadolinium 157.25
Tb Terbium 158.9253
Dy Dysprosium 162.500
Ho Holmium 164.9303
Er Erbium 167.259
Tm Thulium 168.9342
Yb Ytterbium 173.054
Lu Lutetium 174.9668
Ac Actinium (227)
Th Thorium 232.0380
Pa Protactinium 231.0358
U Uranium 238.0289
Np Neptunium (237)
Pu Plutonium (244)
Am Americium (243)
Cm Curium (247)
Bk Berkelium (247)
Cf Californium (251)
Es Einsteinium (252)
Fm Fermium (257)
Md Mendelevium (258)
No Nobelium (259)
Lr Lawrencium (262)

Graham's Law of Effusion

When gases escape through a narrow slit, the lighter gases (gases with lower molecular weights) escape (or "effuse") more quickly. Heavier gases (gases with higher molecular weights) ***effuse*** more slowly. This is called ***Graham's Law of Effusion***.

If a balloon is filled with O_2 and He, which gas will effuse first?

The lighter one! (The one with the lower molecular weight) **He**

Graham's Law of Effusion (Takeaways)

- Both gases (in this case, O_2 and He) have the same average kinetic energy. Recall average kinetic energy depends only on temperature. both gases are at the same temperature, then they have the same ***kinetic energy***.

- In order to have the same average kinetic energy, the lighter molecules have to move faster.

Kinetic energy = ½ mass × (velocity)^2

Graham's Law of Effusion

Mathematically, ***Graham's Law*** is:

$$\frac{\text{effusion rate of gas 1}}{\text{effusion rate of gas 2}} = \sqrt{\frac{\text{molecular weight 2}}{\text{molecular weight 1}}}$$

$$\frac{\text{Rate 1}}{\text{Rate 2}} = \sqrt{\frac{M_2}{M_1}}$$

Graham's Law of Effusion

Gas A has a molar mass of 50 g/mol and Gas B has a molar mass of 2 g/mol. Which gas escapes faster and how much faster does it effuse than the other gas at the same temperature?