SUMMARY OF PAGES

Are you just copying the formulas for Gases down, memorizing them, but not understanding them? Do you think you'll get screwed when the Inquiry questions come along (applying your knowledge)?

Then, this thing is for you...

A formula and graph sheet will be provided on the first two pages. After that its pure explanation of the formulas, as well as how to do each assignment sheet.

Definitions:

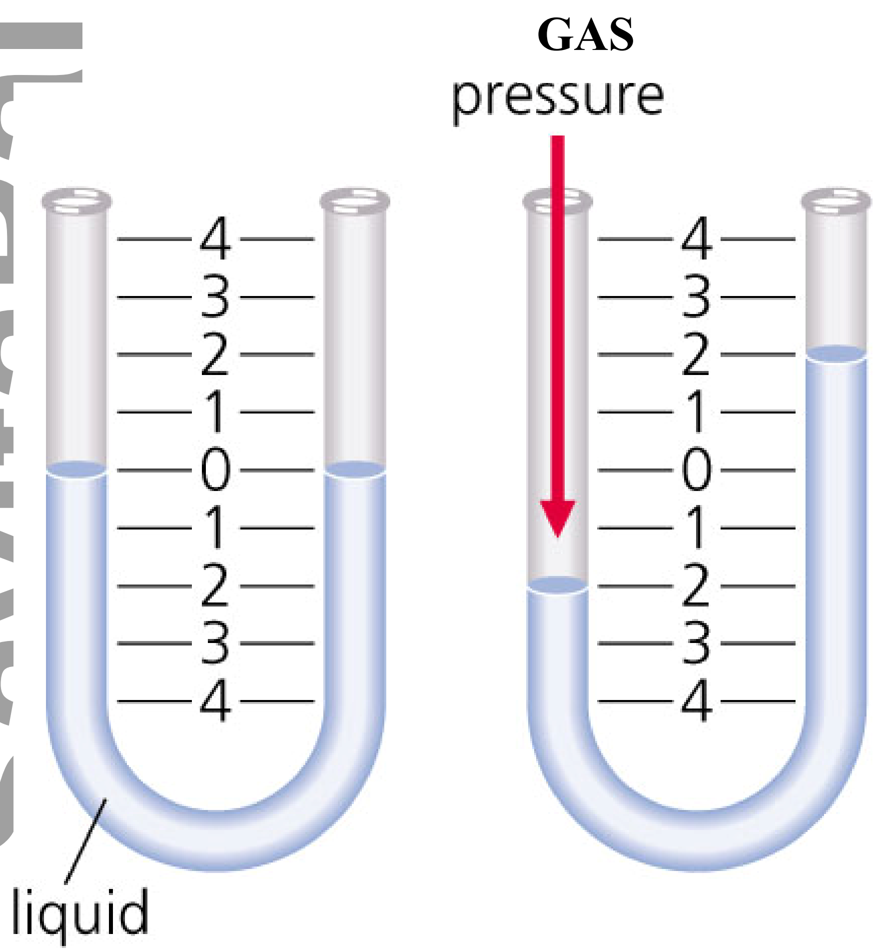
Diffusion:
The intermingling of substances, the random movement of atoms, molecules, or ions from one site in a medium to another, resulting in complete mixing.

Effusion: The flowing of a gas through a small aperture. (hole)

Barometer: Device used to Measure Earth's atmospheric pressure. <u>A type of manometer.</u>

Manometer: A device used to measure the pressure of any gas.

How a barometer/manometer works (see diagram below).
Therefore, if you look at the second diagram of the manometer, the gas makes the liquid go up two notches. To calculate pressure in a U-tube manometer, add the sum of the readings above and below zero. This diagram shows an atmospheric reading of 4.



THE DIFFERENT WAYS TO USE THE WORD PRESSURE.
Well most ppl know that pressure is measured in Pascals. Thats the most common unit – the SI unit (International System of Units **)**

But of course, she has to make it more complicated for us, and use other units as well...

Unit	**Equivalent measurements, comments**
Pascal (Pa)	1 pascal = a force of 1 Newton per square meter 1 Pa = N/m^2
Atmosphere (atm) *Or called* **Bar**	Normal atmospheric pressure is defined as 1 atmosphere. 1 atm = 101.325 kPa = 1 bar
Torr/ (torr) Or **mm Hg**	Based on the original Torricelli barometer design, one atmosphere of pressure will force the column of mercury (Hg) in a mercury **barometer** to a height of 760 millimeters. A pressure that causes the Hg column to rise 1 millimeter is called a torr (you may still see the term **1 mm Hg** used; this has been replaced by the torr). 1 atm = 760 mm Hg
Millibar (mb or mbar)	*There are 1,000 millibar in one bar. This unit is used by meteorologists who find it easier to refer to atmospheric pressures without using decimals. One millibar = 0.001 bar = 0.750 torr = 100 Pa.*
Kilopascal (kPa)	*The prefix "kilo" means "1,000", so one kilopascal = 1,000 Pa. Therefore, 101.325 kPa = 1 atm = 760 torr and 100 kPa = 1 bar = 750 torr.*
Megapascal (MPa)	*The prefix "mega" means "1,000,000", so one megapascal = 1,000 kPa = 1,000,000 Pa. Such high pressures are rarely encounterd.*
Gigapascal (GPa)	*The prefix "giga" means "1,000,000,000", so one gigapascal = 1,000 MPa = 1,000,000 kPa = 1,000,000,000 Pa = 9,870 atm = 10,000 bar. Pressures of several gigapascals can convert graphite to diamond or make hydrogen a metallic conductor!*

OK NEXT we have Brett Anderson's list of equations. This should be referred as a 'summary sheet' since all I'll be doing for the next lot of pages is explain them. :D. How -- interesting.

Equations

$1\text{Pa} = \frac{N}{m^2} = \text{kg}\,m^{-1}\,s^{-1}$

Standard Atmospheric Pressure (STA) $= 101.325\text{kPa} = 760\text{mm Hg}$

Boyles Law: $P_1 v_1 = p_2 v_2$

Charles' Law: $\frac{v_1}{T_1} = \frac{v_2}{T_2}$

Gay - Lussac's Law: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Gay - Lussac's Law: $\frac{v_1}{n_1} = \frac{v_2}{n_2}$

General Gas Equation: $\frac{P_1 v_1}{T_1} = \frac{P_2 v_2}{T_2}$

Ideal Gas Equation: $Pv = nRT$

Variations:

$n = \frac{m}{m_R} \Rightarrow Pv = \frac{m}{m_R}RT$

$\rho = \frac{m}{v} \Rightarrow P = \frac{\rho RT}{m_R}$

Van - der - Waal's Equation: $\left(P + \frac{a}{v^2}\right)(v - b) = nRT$

Kinetic Energy: $E_K = \frac{1}{2}mv^2$

$\frac{\text{volume}}{\text{molar volume}} = \frac{\text{mass}}{\text{molar mass}} = \frac{\#\text{ molecules}}{\text{Avogadro's Number}} = \text{moles}$

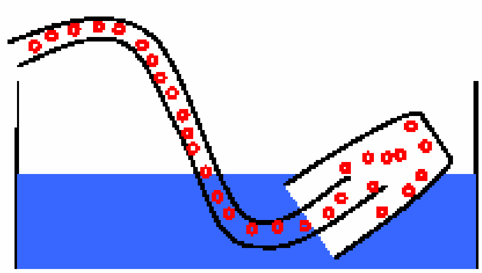
Molar volume at STP $= 22.4\text{L/mol}$

Partial Pressures: $P_{Total} = P_1 + P_2 + P_3 + \ldots$

Partial Pressures: $\frac{n_1}{n_{TOT}} = \frac{P_{g1}}{P_{TOT}}$

Collecting Water of a Gas: $P_T = P_{(H_2O)} + P_{Gas}$

Graham's Law of Diffusion: $\frac{v_1}{v_2} = \sqrt{\frac{m_{R_2}}{m_{R_1}}}$



Boyle's Law

There. A big colourful word – describing a chemistry law. Colours make chemistry laws so much more amusing to look at.

Ok, by now you know that

$$P_1V_1 = P_2V_2$$

Thats pretty easy to explain, once you think about it.
Just think, that when PRESSURE INCREASES, VOLUME DECREASES.

So if I try to squish some gas in a space with a lot of pressure, the space occupying the gas will obviously be *smaller* than before.

So what she's trying to say is that the ratio of Pressure of Gas 1 to Volume of Gas 1, is the same as the ratio between Pressure of Gas 2 to Volume of Gas 2

But the main point is: PRESSURE IS INVERSELEY PROPORTIONAL TO VOLUME. Or in other words, as PRESSURE INCREASES, VOLUME DECREASES

Then she says that n and T are constant, and you're like WHAT? You think you're lost.

But it really isn't that bad. Look at the boxed sentence with a bit of blue text a few lines up. She's saying that this ratio thing only works if:

1) The temperature is the equal in Gas 1 and Gas 2
2) The # of moles of the gas is equal in Gas 1 and Gas 2.

Coz if they weren't, you'd have different variables in gas 1 and gas 2, that would affect your law, so obviously if they weren't equal to each other, than this ratio thing wouldn't work. So we just say that it only works if n (# of moles) and T (temperature) is constant.

Okay just when we're about to rejoice that she's not using so many units for the same thing, we find a few units describing the same factor of TEMPERATURE

UNIT	EXPLANATION
^{0}C (Celsius)	Okay. Why would you want an explanation for the unit of Celsius? I assure you, if you do not know what this unit means, you need more help than ISU's FOR DUMMIES CAN GIVE YOU. You are basically what is known as an 'idiot'.
K (Kalvin)	Okay, the unit 'K' is the symbol for the word Kelvin. Basically it measures from the coldest possible temperature *(absolute zero)* and then goes up. So, 0K = 273 degrees Celsius. There is no temperature lower than 0K. That's why its called ABSOLUTE ZERO!
Fahrenheit (F)	We don't use Fahrenheit in the ISU. If we don't use it, I'm not explaining it. :P

(we will be using K and Celsius in the next law)

Charles's Law

Next we have CHARLES'S LAW.

Charles's Law

$V \alpha T(K)$ where n, P are constant.

$273^0 C = 0K$

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ T is in Kelvin, $2 \times T = 2 \times V$

Ta-dah. Okay now for the explanation.

You should know that α means 'proportional to' by now.

So Volume is proportional to Temperature.

Well this part – you really learned in grade eight, its just that it looks a little complicated.

The basic law is that when Temperature increases, Volume also increases?

Remember – you learned that when you heat something, it rises or expands? Well that's all this law is trying to tell you.

The second part is that the ratio between One Gas's Volume and Temperature is the same between another gas's Volume and Temperature.

This is assuming the # of moles, and the pressure is the same on BOTH GASES. If not, then this ratio isn't true.

By now, you should be getting the hang of it. You can probably understand the laws without the next few pages – but I'm going to explain them anyways, just in case.

Gay-Lussac's Law

Gay-Lussac's Law

$P \alpha T$ where n, V are constant.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$
$$2T = 2P$$

Pressure is proportional to Temperature.

This is just common sense. If I try to squeeze something – er sorry – 'exert pressure on a substance', it's going to get hotter than before.

BUT

Think of this: If this law is true, then technically, **if I get a temperature of ABSOLUTE ZERO** (273 degrees celscius for people who thought they were so smart and didn't read page 7)...

(technically, in the PERFECT WORLD, and in the IDEAL (meaning perfect) GAS)

I SHOULD HAVE NO PRESSURE ON THE GAS.

But zero pressure is impossible in a real gas. You have to take in consideration, that in a gas, although the molecules are far away from each other, they still are going to bang into each other, or into another substance's molecules. THEREFORE, that itself is a tiny, tiny, pressure.

<u>SO NO GAS IN THE REAL WORLD IS PERFECT OR 'IDEAL'.</u>

<u>So if pressure can't be zero in a real gas, then it follows that ABSOLUTE ZERO TEMPERATURE cannot be achieved. Scientists have come really close in getting absolute zero in labs, but its never hit it on the spot. Go figure.</u>

Dalton's Law of PARTIAL PRESSURE

Okay the next law is called Dalton's Law of Partial Pressure (as you can see from the beautiful title). It states that:

$P_{total} = P_1 + P_2 + P_3 + \ldots$

$$\frac{\text{Moles of gas 1}}{\text{Total Moles}} = \frac{\text{Partial Pressure of Gas 1}}{\text{Total Pressure}}$$

This basically means that you have a space right? And you obviously have a lot of gases… So you want to measure the total pressure of the gases, so you take each gas, and find its pressure. Then you add them all together.

Okay this part is really important:

Partial pressure of a gas is independent of the type of gas itself, but is dependent on the number of moles of gas particles present.

Soooo.. for those of you who didn't understand/get that, she's saying that when you're calculating pressure of individual gases, the type of gas is irrelevant – only the number of moles of the gas is relevant (will change the amount of pressure).

This is a kinda weird theory, but if you say the above statement over and over again, you'll understand it SOMEDAY. Hopefully it's a day before the test. ☺

The next formula:

$$\frac{\text{Moles of gas 1}}{\text{Total Moles}} = \frac{\text{Partial Pressure of Gas 1}}{\text{Total Pressure}}$$

just states a formula for the statement in bold. So the ratio between the moles of a gas and the total # of moles, is the same ratio as the pressure of one gas, to the total pressure. Which makes sense. Get it?

Moving on…

Okay this is the formula that the author's tutor said that you'll be using the most.
So make sure you memorize it!!!!!!!1

$P_tV=n_tRT$

And then, you're like WHOA! When did we learn THAT? But really, this formula is just a combination of ALL the previous formulas. Don't ask why we use it, we just do. ☺
Oh yeah, I forgot to describe each unit in that equation: ***$P_tV=n_tRT$***

Pt is obviously total pressure
V is obviously volume
Nt is the total number of mols in all the gases in your space.
T is the temperature.

What's R? R is the "Ideal (Universal) Gas Constant" and it's the same for all gases. The value for R is always 8.31. Remember THAT!

Gay Lussac's Law and Avagadro's hypothesis

Well didn't Gay Lussac already have a law? Well he made another one. :)

Here's his other one

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

So, volume is PROPORTIONAL to the # of moles.

Now if you don't get this, it's okay. I'll explain it.

Consider this example:

This guy's really mad because his wife found out about this girl he was cheating with. And the girl hadn't known that he'd had a wife, so she was pretty mad too. They were both mad at him, so he's like what-the-heck, and makes up a plan so that he could have a day of peace. He finds this potion/gas that makes you faint for a day. So he releases a little bit of gas into his old girlfriend's house, and a little bit of gas, into his wife's new mansion. The gas spreads all over both of the homes. Now it just happens to be that his girlfriend was living on welfare, so she has a small townhouse. Obviously, the gas covering the entire interior of the townhouse, will be a smaller amount of gas than the gas covering the entire mansion. This is because the VOLUME of the mansion is bigger than the volume of the townhouse.

(And yes, the example is quite stupid, but I like it all the same)

So in conclusion, if you have a bigger space to fit the gas in, more gas will fit in, therefore more mols of the gas will fit in.

Why is this section labelled Avagadro's Hypothesis as well? Well Avagadro says that if you have two equal volume boxes, you can fit as much mols of gas in each space. So therefore, you use avagadro's number, and you then have equal number of molecules in each space as well.

A REALLY IMPORTANT FACT:

When you have a standard temperature (o degrees Celsius) and you have a standard pressure (101.3 kpa), then you have a standard volume – which is *22.4L*
This number is always called the – STANDARD MOLAR VOLUME

This will help you do the assignment with the chart referring to molar volume.

(duh)

Grahm's Law of Diffusion

Grahm's Law of Diffusion:

She says diffusion is the process whereby a substance spreads from a region of high concentration to one of lower concentration. For those of you in Biology, you'll remember diffusion across a membrane. I'll use that as my example.

If you have a cell membrane, and you have molecules diffusing into it. (moving inside the cell). The fat molecules will have a hard time getting in, where the smaller molecules, like water, are easy to get in. The fat molecules are also really really slow moving, where as the small molecules are faster.

So therefore, if you have a SMALLER molar mass of a molecule, the molecule will diffuse FAST across the membrane.

If you have a LARGER molar mass of a molecule, the molecule will diffuse SLOW across the membrane. Pretty simple.

Okay the next thing you have to learn is about density and our old $P_tV=n_tRT$ formula.

DENSITY = MASS/VOLUME

Its pretty hard to explain, so just memorize that small formula.

But the description of how this formula turns into the main formula ($P_tV=n_tRT$) is what really counts.

This is the explanation for anybody who didn't get hers.

$PV = nRT$ That's your formula that you used before,

$n = \frac{m}{MM}$ # mols = mass/molar mass

$$PV = \frac{m}{MM} \times RT$$

Substitute mass/molar mass (m/MM) for n in your PV=nRT formula.

$$P = \frac{mRT}{MM\,V}$$

$$P = \frac{mRT}{MM\left(\frac{m}{d}\right)}$$

$$P = \frac{\cancel{m}RT\,d}{MM\,\cancel{m}}$$

$$P = \frac{d \cdot RT}{MM}$$

Kinetic Energy

Okay Kinetic energy, is the energy possessed by molecules as they move through space…

That's not hard to understand…

Okay the main formula for Kinetic energy is:

Ke = 1/2mv2

Where M is the mass of the molecules (molar mass) and v is the velocity of the molecules.

$$K_E \alpha T$$

This is supposedly, the hardest thing to understand.

What's easy is that Kinetic Energy depends on temperature completely.
(Kinetic energy is PROPORTIONAL) to temperature

What's hard, or hard to remember at least, is that kinetic energy DOES NOT RELY IN VOLUME, PRESSURE, AND THE # OF MOLS OF THE GAS

Okay so now remember our cheating guy. His wife has a mansion. His now former girlfriend has a small townhouse. But if he releases the gas in BOTH houses, and BOTH houses as the same temperature – than the gas kinetic energy is the same in each of the houses.
So the # of mols of the gas, or the volume of the house doesn't matter.

And pressure, well if the girlfriend was a scientific nerd who, in order for the gas not to get to her, used a lot of pressure on the gas so it would fit into a small box…. Well her kinetic energy is still the same (if the temperature is the same).

But KINETIC ENERGY DOES RELY on TEMPERATURE… This is quite easy to understand – if you heat something, the molecules are obviously going to get very ‘excited’ and move faster.

BUT WAIT!!!!!!!!!!!!!!!!!!!

We have to remember that term again → IDEAL GAS!

In a perfect gas this formula would be true ALL THE TIME. But you see, IT ISN’T because no gas is ever IDEAL/PERFECT

Because, well usually when you heat something, the velocity of the molecules increase rite? But lets say you LOWER the temperature of a gas like WATER VAPOUR. The water vapour, is obviously going to turn into a LIQUID (water) and later into a SOLID (ice). And you know that in a solid, molecules are frozen into one structure, and can only vibrate, not move around.

SO THEREFORE, KINETIC ENERGY DOES NOT DEPEND ON THE $K_E \cdot T$ FORMULA IF THE TEMPERATURE IS LOWERED SO MUCH THAT THE GAS CHANGES STATE.!!!!!!

Okay, its important that you understand that in the Ke=1/2mv2 formula, the m is talking about MOLECULAR MASS, and not the whole gas’s VOLUME.

Its cause volume DOESN’T AFFECT THE KINETIC ENERGY and molecular mass DOES.

But then you’re like: well then how can kinetic energy be the same, if the molar masses of two gases are different? After all Ke does equal half of (the molar mass times volume squared).

Well that question is solved when she says:

“Therefore two gases (e.g; H_2 + O2) at the same temperature will have the same KE but differing molecular velocities (velocity of H2 molecules is greater than those of O2 molecules”

So suddenly, everything seems so EASY!

If you have lets say Hydrogen, with a SMALL molar mass, its going to have a LARGE molecular velocity! (speed of molecules used in $Ke=1/2mv^2$ formula)

If you have Oxygen, it has a larger molar mass, but it has a smaller molecular velocity!

So if you find the kinetic energy for both, and the temperature is the same in both cases, they’ll both EQUAL EACH OTHER!

:O (big gasp)

MOST IMPORTANT

Okay if you don't know the $P_tV=n_tRT$ formula by now…

memorize it!

REMEMBER: R is always the same thing: 8.31

Because now, you'll be deriving different things from that formula.

Often times, you'll get a question that has beginning and end variables.

For example:

The pressure on 220 cm3 of a gas in 110 kPa. What will be the volume, if the pressure is changed to 55.0 kPa (keeping temp constant)?

Okay this might seem easy. And actually, it is.

You could use Boyle's Law. I'm not going to restate Boyle's law, just look over the last few sheets.

But what I want to teach you, is something you can use for ALL the problems. And, you can't use separate laws for all the problems.

Lets say you are given beginning and end volume, beginning and end pressure, and beginning temperature, and you're asked to find end temperature.

So basically:
V_1 = given
V_2 = Given
P_1 = Given
P_2 = Given
T_1 = Given

$T_2 = ?$
Where n(number of moles) is constant, R is constant
Okay, so what do you do? Use one of the separate laws? Well that won't work. Actually m only use the separate laws when you're using 2 variables (pressure, temp, volume, etc.) If you have three or four variables, like the question above, use the following formula

$P_tV=n_tRT$

Kind of repetitive formula isn't it? Well this is a combination of all your formulas. But then you might argue that this formula doesn't have 'beginning' and 'end' factors (i.e. P1 and P2) in it, it just has solid variables.

So lets take this example for instance.
You need to use the Volume, Pressure, and Temperature in this solution. You're assuming n(number of moles) and R is constant, since it wasn't mentioned in the question.

So follow these steps:
$P_tV = nRT$

$$\frac{P_tV}{T} = nR$$

Therefore,
$$\frac{P_1V_1}{T_1} = nR$$

Since n and R are constant in this case (they do not move)

$$\frac{P_2V_2}{T_2} = nR$$

SOOO....

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Then substitute all your known values in, and you can get T_2.

....

So, in any problem follow these steps. If u get lost at any point, refer back to the example.

1) DETERMINE HOW MANY VARIABLES (I.E. PRESSURE, TEMP, VOLUME) IN YOUR EQUATION.
2) IF TWO, USE ONE OF THE SEPARATE LAWS (I.E. BOYLE'S LAW)

3) IF THREE, LOOK AT $P_tV=n_tRT$ EQUATION. THEN LOOK AT THE PROBLEM, AND DETERMINE IF THE BEGINNING-END VARIABLES ARE GIVEN, OR JUST STRAIGHT OUT VARIABLES (I.E. P1, P2 V.S. P)
4) IF THEY ARE JUST STRAIGHT OUT VARIABLES, SUBSTITUTE IN EQUATION, AND SOLVE FOR MISSING VARIABLE.
5) IF THEY ARE BEGINNING-END VARIABLES, PERFORM THE FOLLOWING STEPS
   - MAKE ALL THE VARIABLES MENTIONED IN YOUR PROBLEM ON ONE SIDE, AND ALL THE NOW-CONSTANTS, ON THE OTHER.
   - CHANGE ALL THE VARIABLES TO 'BEGINNING-VARIABLES' (I.E. CHANGE ALL THE P'S TO P1S)
   - THIS WILL BE EQUAL TO YOUR CONSTANTS.
   - THEN, ON THE NEXT LINE REWRITE THE SAME THING, EXCEPT WITH END VARIABLES (I.E.: P2S)
   - SET THEM EQUAL TO EACH OTHER, AND SOLVE!

Constants you should know:
8.31: Always Replaces R in your most important formula.
24.32: This constant is used for something called MOLAR VOLUME. If you're asked to find the molar volume of a substance, take the # of mols, and times it by this number.
101.31: (kPa) This is Standard Atmospheric Pressure, chances are you won't be asked to use it.

**CONGRATULATIONS!!!!!!!!!!!!!!!!!!!
YOU'VE UNDERSTOOD (WELL I HOPE YOU HAVE) THE WHOLE UNIT!!!!!!!!**