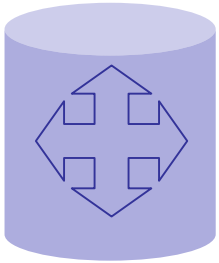
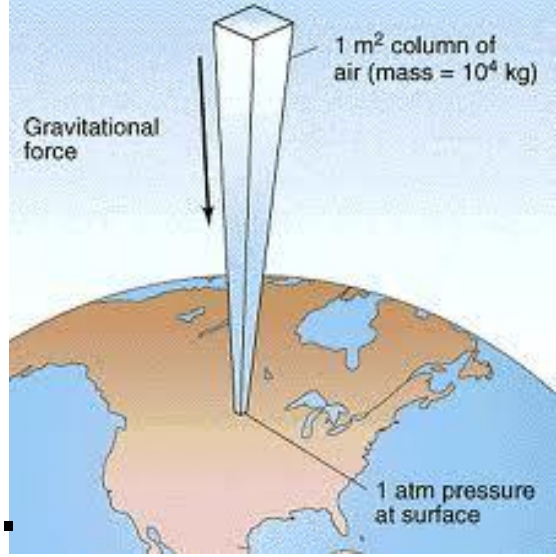


Pressure

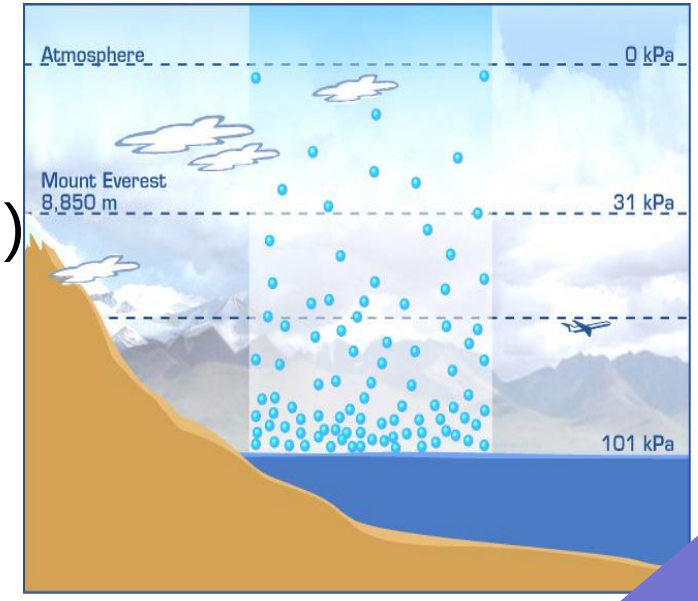
- Pressure is the force exerted by a gas on a surface.
  - The surface that we measure the pressure on is usually the inside of the gas's container.
- Pressure and the Kinetic Theory
  - Gas pressure is caused by billions of particles moving randomly, and striking the sides of the container.
- Pressure Formula: $$P = \frac{F}{A}$$

Pressure = force divided by area

Atmospheric Pressure



- This is the force of a 100 km high column of air pushing down on us.
- Standard atmospheric pressure is
  - 1.00 atm (atmosphere), or
  - 101.3 kPa (kilopascals), or
  - 760 Torr (mmHg), or
  - 14.7 psi (pounds per square inch)
- Pressure varies with:
  - Altitude. (lower at high altitude)
  - Weather conditions. (lower on cloudy days)



Write this!

Atmospheric pressure

- At sea level the atmospheric pressure is set to 1.00 atm

1.00 atm = 101.3 kPa = 760 mm Hg (Torr)

- Standard Temperature & Pressure (STP)

0°C = 273 K & 101.3 kPa

- Standard Ambient Temperature & Pressure (SATP)

25°C = 298 K & 101.3 kPa

Pressure conversions

SP
1.00 atm
760 mmHg
760 Torr
101.3 kPa
14.7 psi

Example 1: convert 540 mmHg to kilopascals

$$\frac{540\,mmHg}{760\,mmHg} \quad \frac{P}{101.3\,kPa} = 72.0\text{ kPa}$$

Multiply

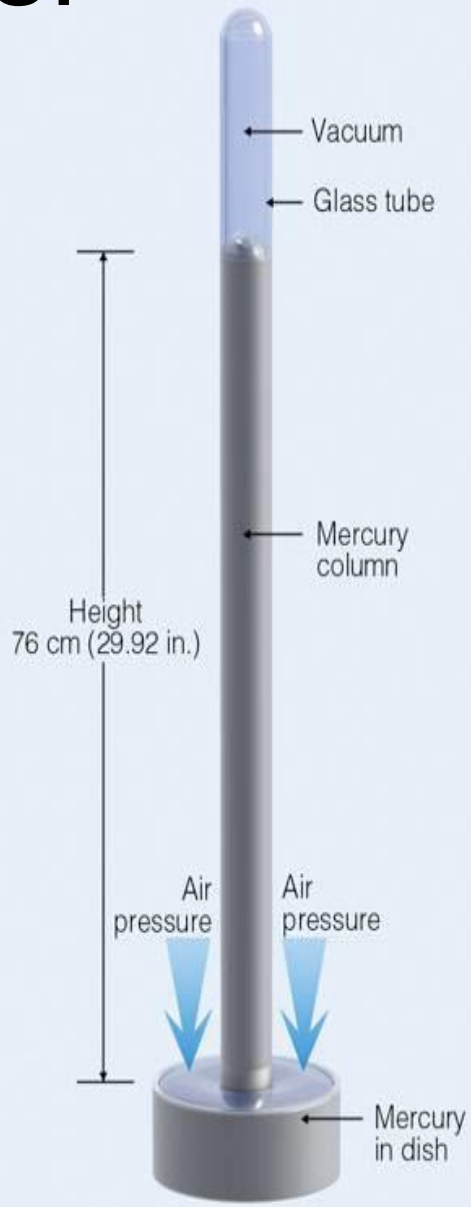
Divide

Example 2: convert 155 kPa to atmospheres

$$\frac{155\,kPa}{101.3\,kPa} = \frac{P}{1.00\,atm} = 1.53\text{ atm}$$

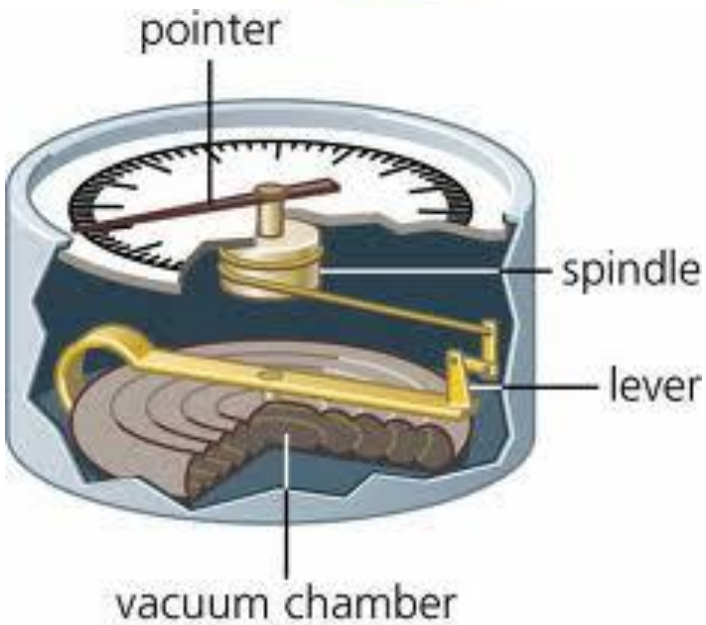
the Mercury Barometer

- A tube at least 800 mm long is filled with mercury (the densest liquid) and inverted over a dish that contains mercury.
- The mercury column will fall until the air pressure can support the mercury.
- On a sunny day at sea level, the air pressure will support a column of mercury 760 mm high.
- The column will rise and fall slightly as the weather changes.
- Mercury barometers are very accurate, but have lost popularity due to the toxicity of mercury.



The Aneroid Barometer

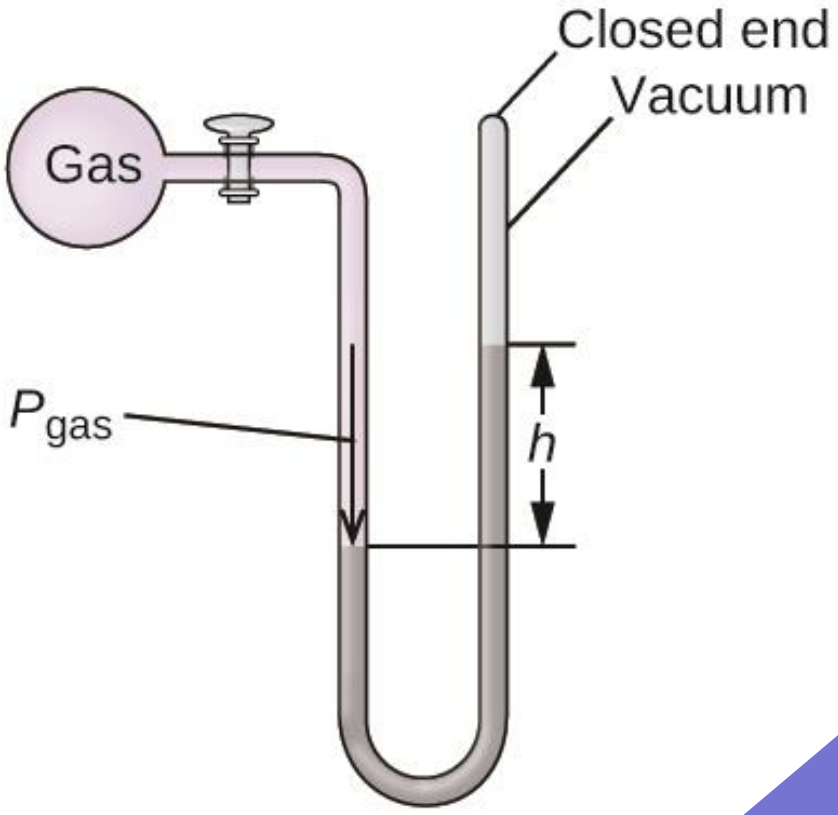
- In an aneroid barometer, a chamber containing a partial vacuum will expand and contract in response to changes in air pressure
- A system of levers and springs converts this into the movement of a dial.



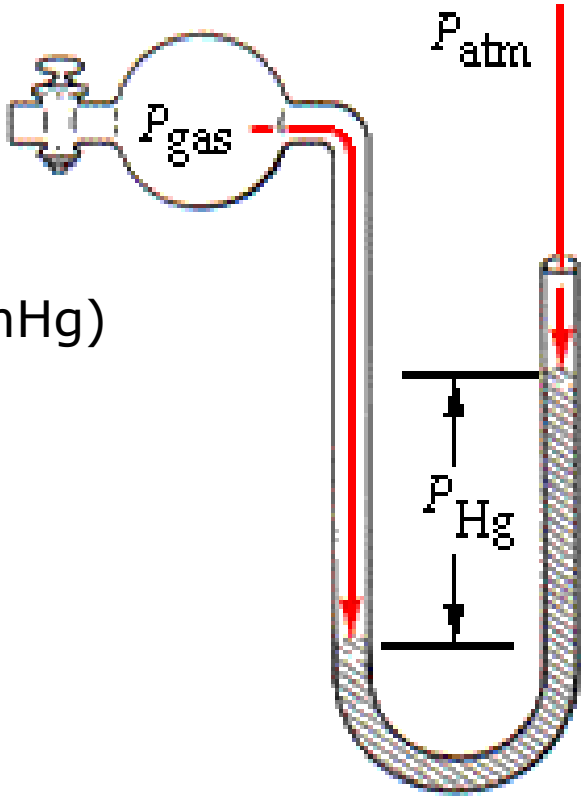
Write this!

- A manometer is a pressure gauge that measures the pressure difference between the inside and outside of a container.
- 2 types
  1. Closed ended manometer

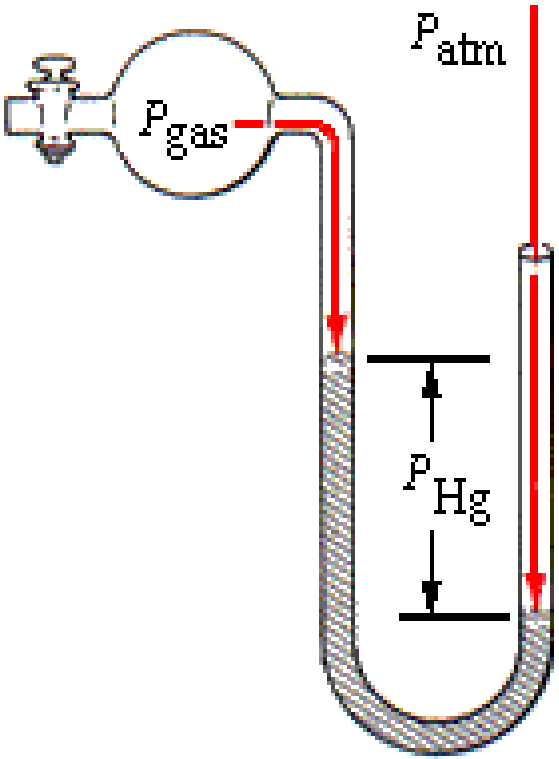
$$P_{gas(mm\ Hg)} = h_{\ (mm\ Hg)}$$



2. Open ended manometer



$$P_{gas(mmHg)} = P_{atm(mmHg)} + h_{(mmHg)}$$



$$P_{gas(mmHg)} = P_{atm(mmHg)} - h_{(mmHg)}$$

Diagrams on p72

Manometer Examples

on a day when the air pressure is 763mmHg (101.7 kPa)

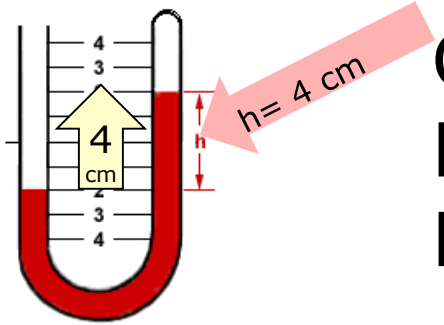


Fig. 2-1

Closed tube: $P_{gas(mm\ Hg)} = h_{(mm\ Hg)}$

$P_{gas} = h = 4\ cm = 40\ mm\ Hg$

$$P_{gas} = \frac{40 mm Hg}{760 mm Hg} \times 101.3 kPa = 5.3 kPa$$

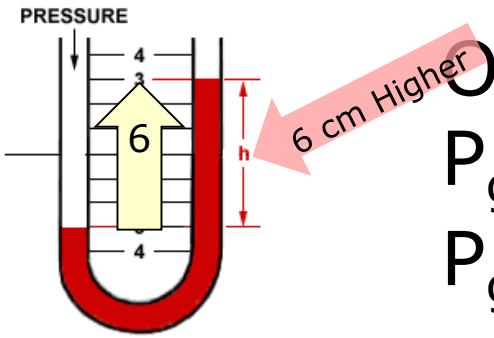


Fig. 2-2

Open: $P_{gas(mmHg)} = P_{atm(mmHg)} + h_{(mmHg)}$

$P_{gas} = 763 + 60_{mm\ Hg} = 823_{mm\ Hg}$

$$P_{gas} = \frac{823 mm Hg}{760 mm Hg} \times 101.3 kPa = 109.7 kPa$$

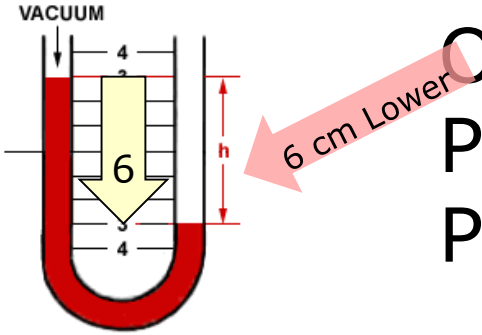


Fig. 2-3

Open: $P_{gas(mmHg)} = P_{atm(mmHg)} - h_{(mmHg)}$

$P_{gas} = 763 - 60_{mm\ Hg} = 703_{mm\ Hg}$

$$P_{gas} = \frac{703 mmHg}{760 mmHg} \times 101.3 kPa = 93.7 kPa$$

- Four factors affecting gases:
  - Pressure (P)
  - Volume (V)
  - Temperature (T)
  - # of moles (n)

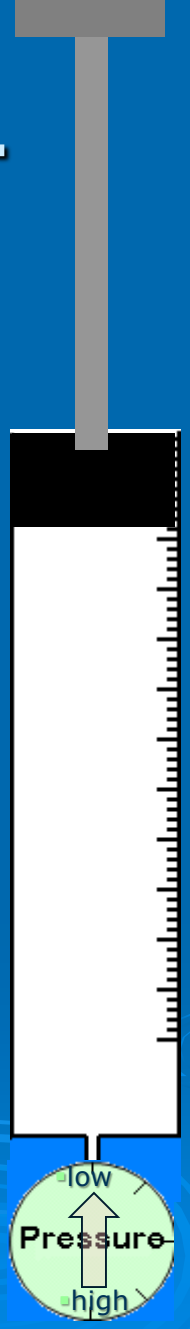
- The Simple Gas Laws
  - Boyle's Law — Relates volume & pressure
  - Charles' Law — Relates volume & temperature
  - Gay-Lussac's Law — Relates pressure & temperature
  - Avogadro's Law — Relates to the number of moles

Robert Boyle

- Born: 25 January 1627 Ireland
- Died 31 December 1691 (64)
- Very rich and influential.
- Fields: Physics, chemistry;
- Considered to be the founder of modern chemistry
- **But…**
- **H. Power** & **R. Towneley** did the actual experiments.
- Boyle was just the one who published the results.
- The law was also discovered by French chemist **Edme Mariotte.**

Consider the air in a syringe…

- Assumptions:
  - No gas enters or leaves (n is constant)
  - Temperature is constant
- The harder you press, the smaller the volume of air becomes.
  - ↑ pressure = ↓ volume
- As the volume of a contained gas increases, the pressure decreases.
  - ↑ volume = ↓ pressure



Boyle's Law (PV relationship)

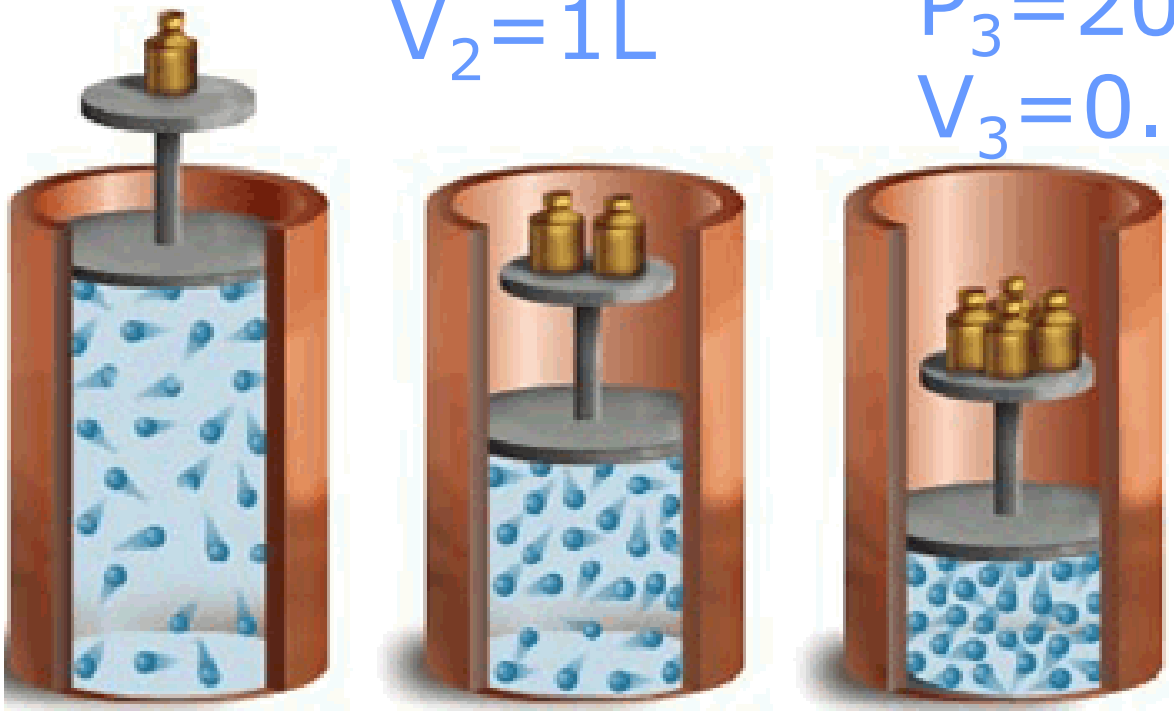
Write this!

- Experimented with manometers
  - Concluded that ↑ pressure = ↓ volume
- Consider

$P_1 = 50kPa$
$V_1 = 2L$

$P_2 = 100kPa$
$V_2 = 1L$

$P_3 = 200kPa$
$V_3 = 0.5L$



- What would the P-V graph look like?
  - On blackboard …. **Copy graphs!**
- Conclusion from P vs V graph: **Write this!**
  - At constant temperature, the volume occupied by a given quantity of gas is inversely proportional to the pressure of the gas

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

- Conclusion of P vs $\frac{1}{\ }$:

$$P = k^a \times \frac{1}{V} \rightarrow PV = k^a \rightarrow P_1V_1 = P_2V_2$$

Write this!

$$P_1V_1 = P_2V_2$$

Where:

P_1 is pressure of the gas before the container changes shape.*

P_2 is the pressure after (using the same units as P_1).

V_1 is the volume of the gas before the container changes (L or mL)

V_2 is the volume of the gas after (same units as V_1)

*appropriate pressure units include:

kPa & mmHg & atm.

Do these question:

- P74 2,3
- P97 1-4

- You have ~ 15 min

Example 1

- You have 30 mL of air in a syringe at 100 kPa. If you squeeze the syringe so that the air occupies only 10 mL, what will the pressure inside the syringe be?
- $P_1 \times V_1 = P_2 \times V_2$, so..
- 100 kPa × 30 mL = **?** kPa × 10 mL
- 3000 mL·kPa ÷ 10 mL = 300 kPa
- The pressure inside the syringe will be 300 kPa

Graph of Boyle's Law

The Pressure-Volume Relationship

Boyle's Law produces an inverse relationship graph.

P(kpa) x V(L)

100 x 8 = 800

200 x 4 = 800

300 x $2.6\overline{6}$ = 800

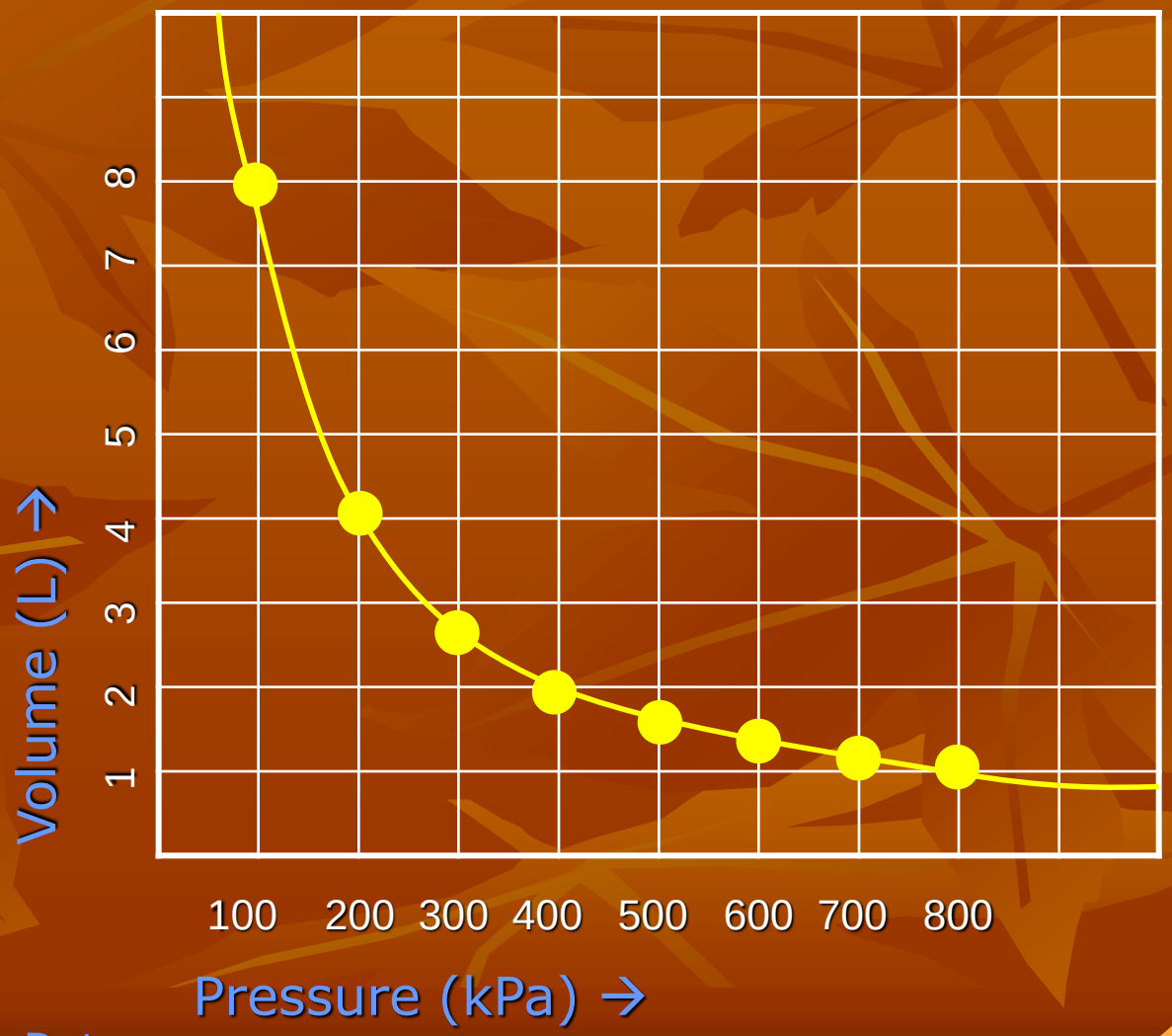
400 x 2 = 800

500 x 1.6 = 800

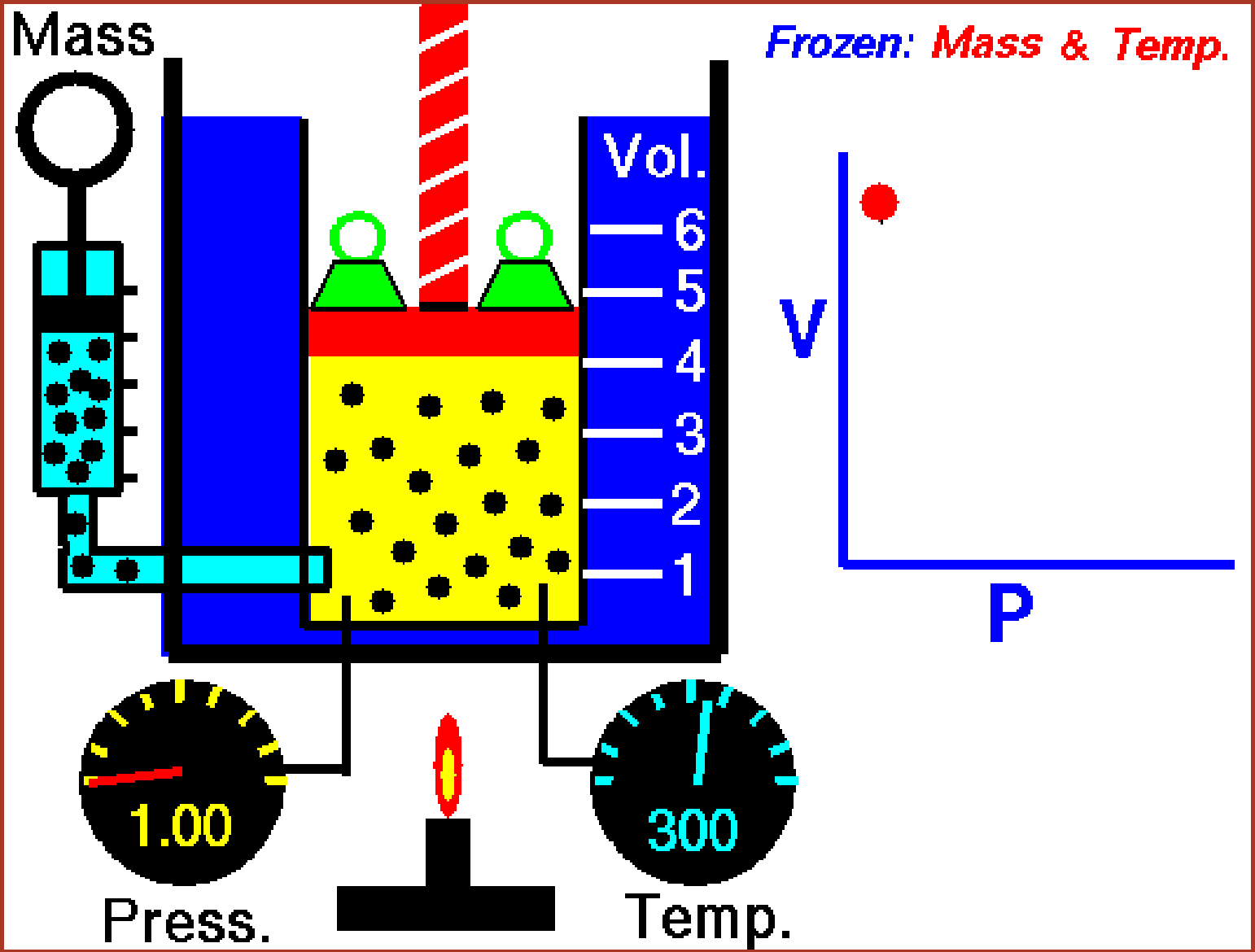
600 x $1.3\overline{3}$ = 800

700 x 1.14 = 800

800 x 1 = 800



Next slide: Real Life Data


Mass
Frozen: Mass & Temp.
Vol.
6
5
4
3
2
1
V
P
1.00
Press.
300
Temp.

Assignments on Boyle’s Law

- Read pages 75 to 79
- Do questions 1 to 10 on page 97

	Bromide Br^-	Carbonate CO_3^{2-}	Chloride Cl^-	Chlorates ClO_3^-	Hydroxide OH^-	Nitrate NO_3^-	Oxide O^{2-}	Phosphate PO_4^{3-}	Sulfate SO_4^{2-}	Dichromate $Cr_2O_7^{2-}$
Aluminium Al^{3+}	S	X	S	S	I	S	I	I	S	I
Ammonium NH_4^+	S	S	S	S	S	S	X	S	S	S
Calcium Ca^{2+}	S	I	S	S	sS	S	sS	I	sS	I
Copper(II) Cu^{2+}	S	I	S	S	I	S	I	I	S	I
Iron(II) Fe^{2+}	S	I	S	S	I	S	I	I	S	I
Iron(III) Fe^{3+}	S	X	S	S	I	S	I	I	sS	I
Magnesium Mg^{2+}	S	I	S	S	I	S	I	I	S	I
Potassium K^+	S	S	S	S	S	S	S	S	S	S
Silver Ag^+	I	I	I	S	X	S	I	I	sS	I
Sodium Na^+	S	S	S	S	S	S	S	S	S	S
Zinc Zn^{2+}	S	I	S	S	I	S	I	I	S	I
	Bromide Br^-	Carbonate CO_3^{2-}	Chloride Cl^-	Chlorates ClO_3^-	Hydroxide OH^-	Nitrate NO_3^-	Oxide O^{2-}	Phosphate PO_4^{3-}	Sulfate SO_4^{2-}	Dichromate $Cr_2O_7^{2-}$

Solubility Chart

I Insoluble, S Soluble, SS slightly soluble

	Ag	Al	Ba	Bi	Ca	Cd	Co	Cr	Cu	Fe	H	Hg	K	Mg	Mn	Na	NH_4	Ni	Pb	Sn	Sr	Zn
Acetate	I	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S
Bromide	I	S	S		S	S	S	S	S	S	S	I	S	S	S	S	S	S	I	S	S	S
Carbonate	I	I	I	I	I	I	I	I	I	I	S	I	S	I	I	S	S	I	I	I	I	I
Chlorate	S	S	S		S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S
Chloride	I	S	S	S	S	S	S	S	S	S	S	I	S	S	S	S	S	S	I	S	S	S
Chromate	I	S	I	S	S	S	S	S	S	S	S	I	S	S	S	S	S	S	I	S	I	S
Cyanide	I		S	I	I	S	I	I	I	I	S	S	S	S		S	S	I	I		S	I
Fluoride	S	S	I	S	I	S	S	S	I	I	S		S	I	I	S	S	I	I	S	I	I
Hydroxide	I	I	S	I	SS	I	I	I	I	I	S	I	S	I	I	S	S	I	I	I	SS	I
Iodide	I	S	S	S	S	S	S	S	S	S	S	I	S	S	S	S	S	S	I	S	S	S
Nitrate	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S		S	S
Oxide	I	I	S	I	I	I	I	I	I	I	S	I	S	I	I			I	I	I	S	I
Phosphate	I	I	I	I	I	I	I	I	I	I	S	I	S	I	I	S	S	I	I	I	I	I
Silicate	I	I	I	I	I	I	I	I	I	I	S	I	S	I	I	S	S	I	I	I	I	I
Sulfate	SS	S	I	S	SS	S	S	S	S	S	S	I	S	S	S	S	S	S	I	S	I	S
Sulfide	I	I	I	I	I	I	I	I	I	I	S	I	S	I	I	S	S	I	I	I	I	I
Sulfite	I	I	I	I	I	I	I	I	I	I	S	I	S	I	I	S	S	I	I	I	I	I

Charles' Law
(Lesson 2.4.2 p80)

The Relationship between Temperature and Volume.

"Volume varies directly with Temperature"

$$V \propto T$$

Jacques Charles (1787)

- 1746 – 1823
- Nationality: France
- Fields: physics, mathematics, hot air ballooning

"The volume of a fixed mass of gas is directly proportional to its temperature (in kelvins) if the pressure on the gas is kept constant"

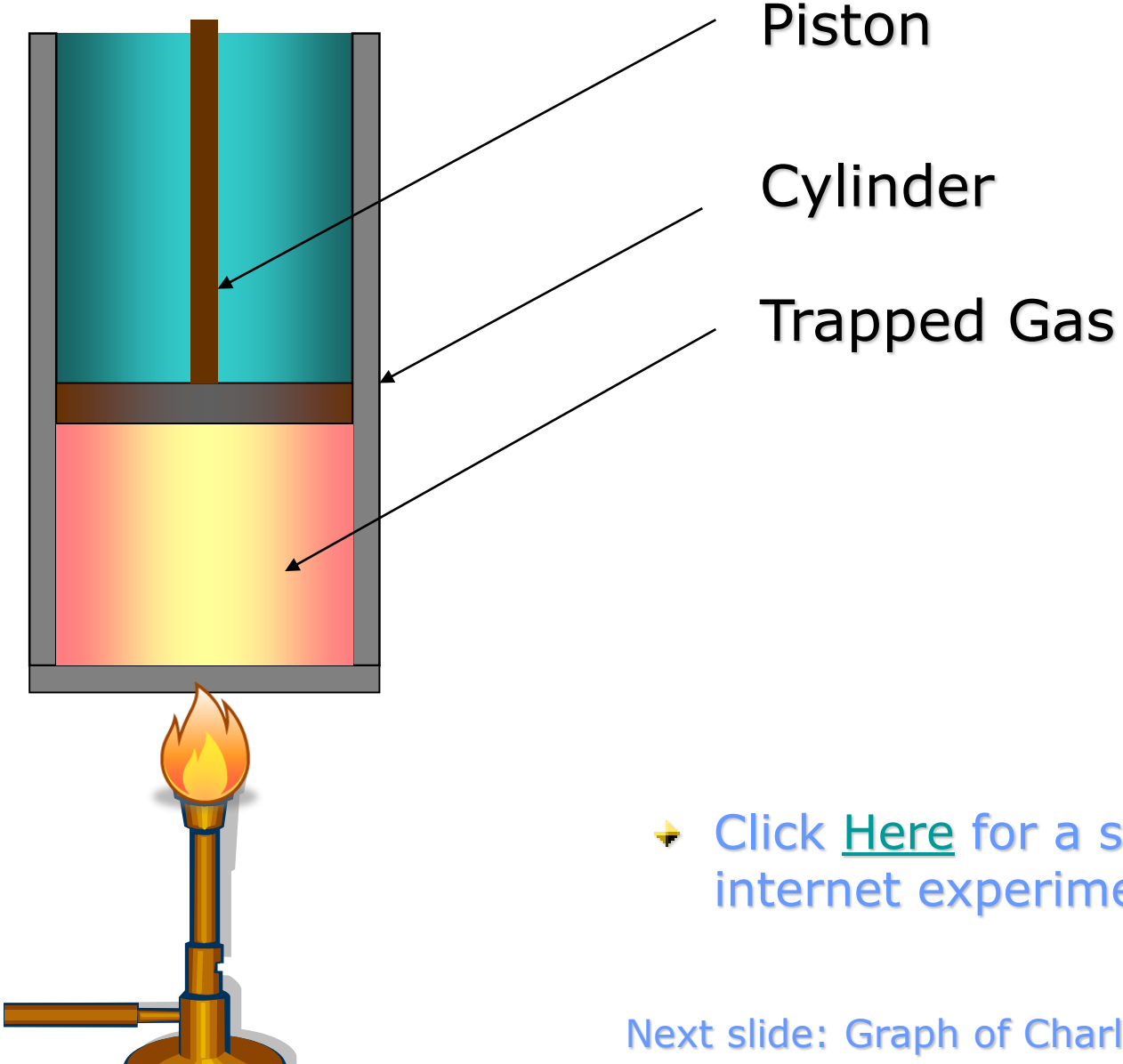
- This assumes that the container can expand, so that the pressure of the gas will not rise.

Charles Law Evidence

- Charles used cylinders and pistons to study and graph the expansion of gases in response to heat.

- Lord Kelvin (William Thompson) used one of Charles' graphs to discover the value of absolute zero.

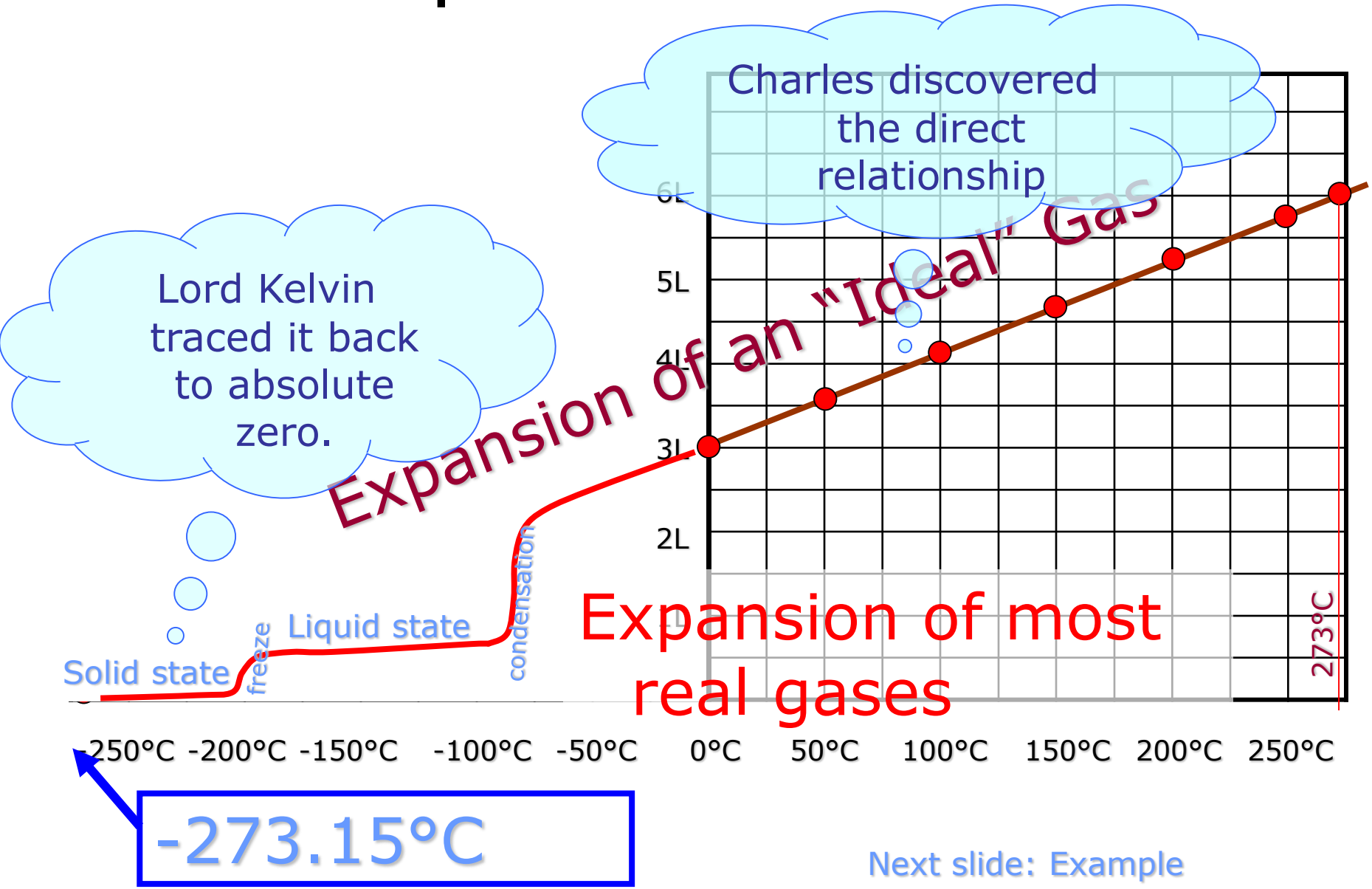
Charles Law Example



- Click Here for a simulated internet experiment

Next slide: Graph of Charles' Law

Graph of Charles Law

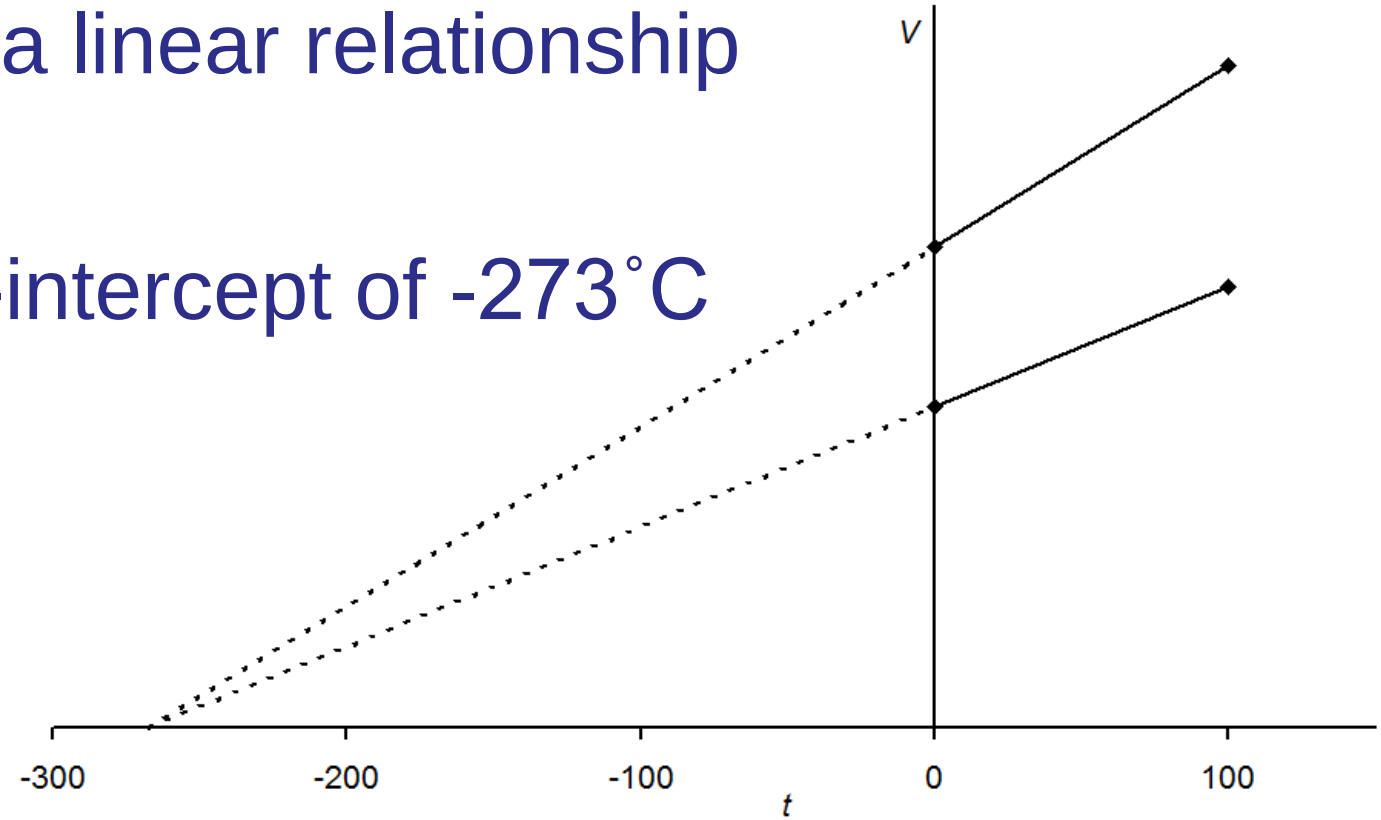


Next slide: Example

Charles' Law (VT relationship)

Write this!

- Observed that the volume of a gas increased by 1/273 of its initial value for each ˚C.
- Noticed a linear relationship
- Same x-intercept of -273˚C



ABSOLUTE ZERO BY LORD KELVIN

-273.15°C is called **<u>absolute zero</u>**. It is the coldest possible temperature.

At absolute zero, molecules stop moving and even vibrating.

Since temperature is based on the average kinetic energy of molecules, temperature cannot be said to exist if there is no kinetic energy (movement)

ABSOLUTE ZERO IS REALY COOL!

Kelvin's Scale

In 1848 Lord Kelvin suggested using a temperature scale based on absolute zero, but with divisions exactly the same as the Celsius scale.

To convert from Celsius to Kelvin, simply add **273** to the Celsius temperature. To convert back, subtract 273

Note: Temperature readings are always assumed to have at least 3 significant digits. For example, 6°C is assumed to mean 279 K with 3 sig.fig., even though the data only showed 1 sig.fig.

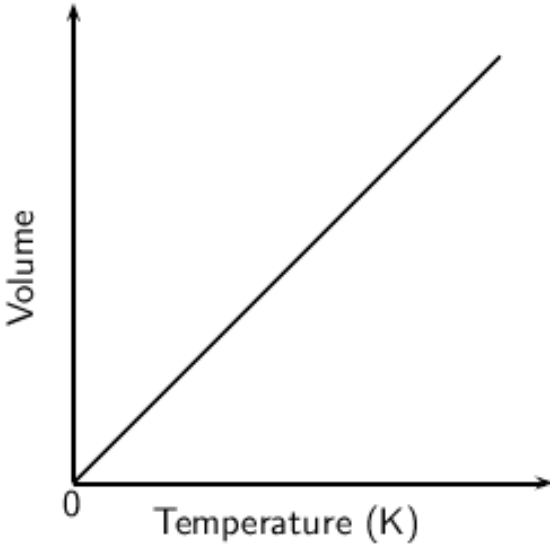
- Conclusion of V vs T graph:
  - At constant pressure, the volume occupied by a given quantity of gas is directly proportional to the absolute temperature of the gas.

Write this!

$$V \propto T$$

$$V =$$

$$\frac{V}{T} =$$



- Charles' Law: $$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Write this!

T_1 Temperature before

T_2 Temperature after

V_1 Volume before (L or mL)

V_2 Volume after

- Temperature in kelvin (T=˚C + 273)
- Volume in L or ml

Example

- If 2 Litres of gas at 27°C are heated in a cylinder, and the piston is allowed to rise so that pressure is kept constant, how much space will the gas take up at 327°C?
- Convert temperatures to kelvins: 27°C =300k, 327°C = 600k
- Use Charles' Law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{2\,Litres}{300\,K} = \frac{x\,Litres}{600\,K}$$

- Answer: 4 Litres

Charles' Law Assignments

- Read pages 80 to 84
- Do questions 11 to 21 on pages 97 and 98

Charles' Law Practice

1. The temperature inside my fridge is about 4˚C, If I place a balloon in my fridge that initially has a temperature of 22˚C and a volume of 0.50 litres, what will be the volume of the balloon when it is fully cooled? (for simplicity, we will assume the pressure in the balloon remains the same)

Data:

T_1=22˚C =295K

T_2=4˚C =277K

V_1=0.50 L

To find:

V_2= *unknown*

Temperatures must be converted to kelvin

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

multiply

divide

So:

$V_2 = V_1 \times T_2 \div T_1$

$V_2 = \frac{0.5\,L \times 277\,K}{295\,K}$

V_2=0.469 L

The balloon will have a volume of 0.47 litres

2. A man heats a balloon in the oven. If the balloon has an initial volume of 0.40 L and a temperature of 20.0°C, what will the volume of the balloon be if he heats it to 250°C.

Data

V_1= 0.40L

T_1= 20°C =293 K

T_2= 250°C=523 K

V_2= 0.7139L

Convert temperatures to kelvin

20+273= 293K, 250+273=523k

Use Charles' Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \cdots \frac{0.40L}{293K} = \frac{V_2}{523K}$$

0.40L x 523 K ÷ 293 K = 0.7139L

Answer: The balloon's volume will be 0.71 litres

3. On hot days you may have noticed that potato chip bags seem to inflate. If I have a 250 mL bag at a temperature of 19.0°C and I leave it in my car at a temperature of 60.0°C, what will the new volume of the bag be?

(assume that most of the bag is filled with gas, that the chips are negligible volume)

Data:

V_1=250 mL

T_1= 19.0°C=292 K

T_2=60.0°C =333 K

V_2= 285.10 mL

Convert temperatures to kelvin

19+273= 292K, 60+273=333K

Use Charles' Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \ldots \frac{250mL}{292K} = \frac{V_2}{333K}$$

250mL x 333 K ÷ 292 K = 285.10mL

Answer: The bag will have a volume of 285mL

Although only the answers are shown here, in order to get full marks you need to show all steps of the solution!

4. The volume of air in my lungs will be 2.35 litres

Be sure to show your known information

Change the temperature to Kelvins and show them.

Show the formula you used and your calculations

State the answer clearly.

5.

6. **The temperature is 279.7 K, which corresponds to 6.7^0 C. A jacket or sweater would be appropriate clothing for this weather.**

Lesson 2.4.3

Gay-Lussac's Law

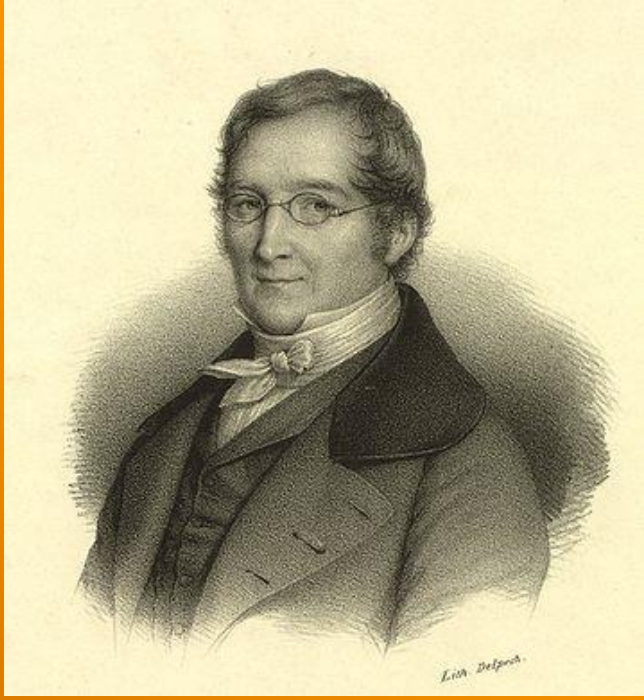
For Temperature-Pressure changes.

"Pressure varies directly with Temperature"

$$P \propto T$$

Next slide:'

Joseph Gay-Lussac

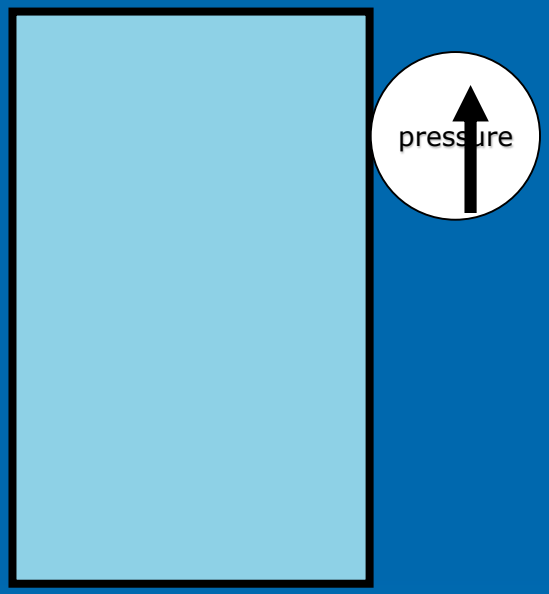


- 1778 - 1850
- **Nationality:** French
- **Fields:** Chemistry
- **Known for** Gay-Lussac's law

- **"The pressure of a gas is directly proportional to the temperature (in kelvins) if the volume is kept constant."**

Gay-Lussac's Law

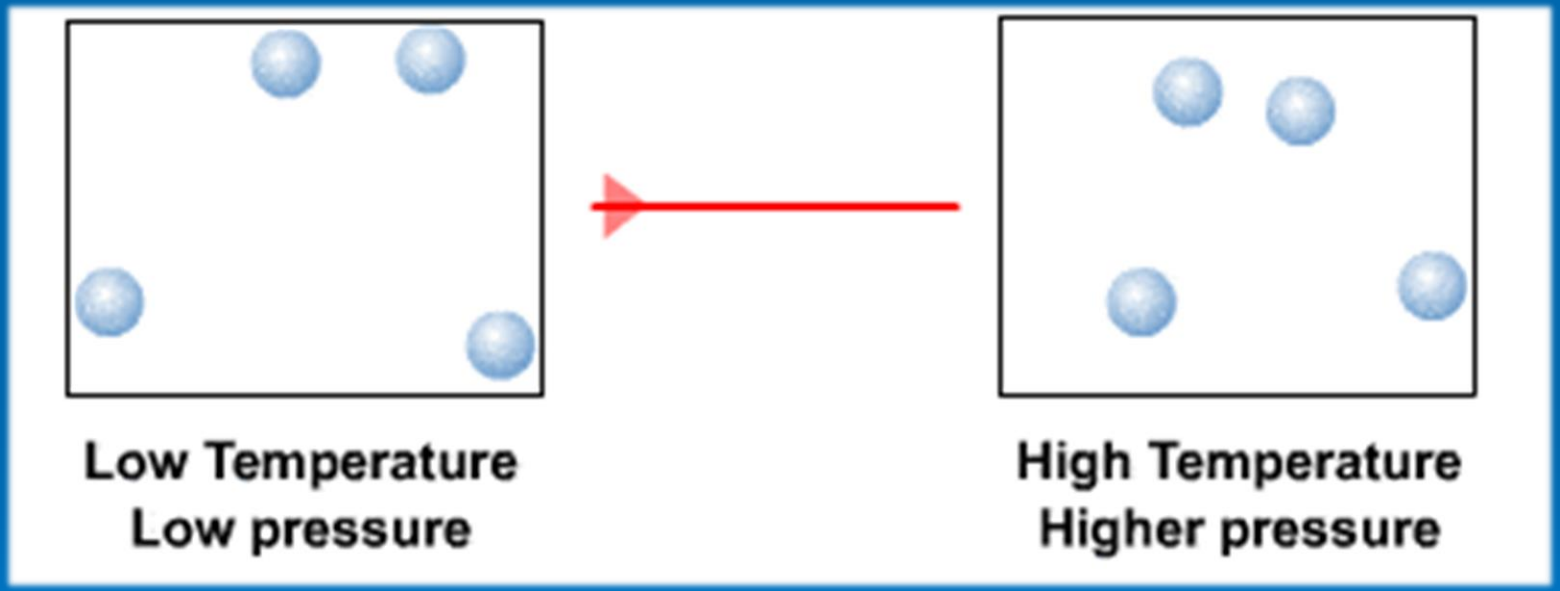
- As the gas in a sealed container that cannot expand is heated, the pressure increases.



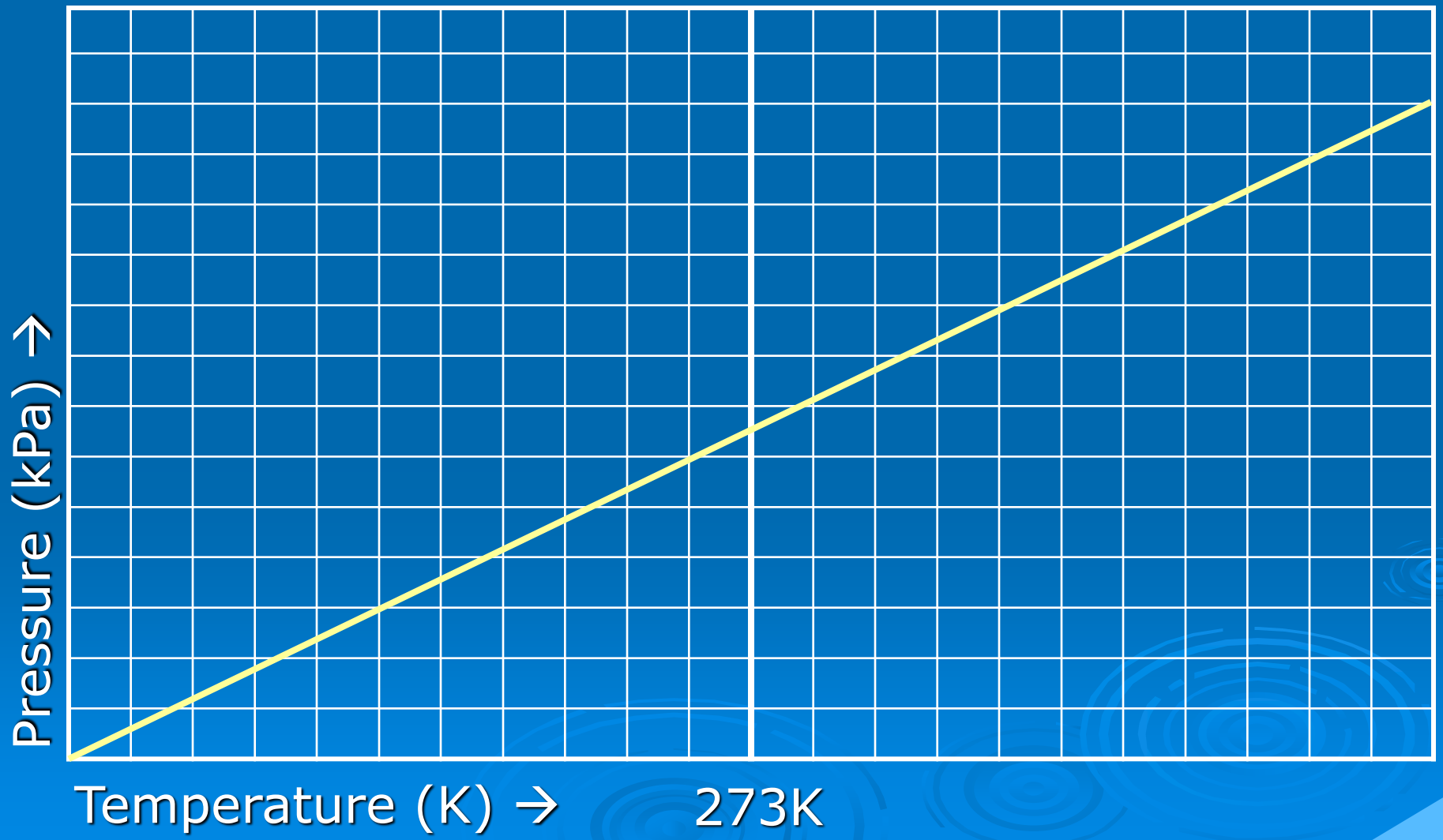
- For calculations, you must use Kelvin temperatures:
  - K=°C+273

Let's make a P vs T graph!

- Do experiment together



Graph of Pressure-Temperature Relationship (Gay-Lussac's Law)



Write this!

Gay-Lussac's Law (PT relationship)

- At constant volume, the pressure of a given quantity of gas is directly proportional to the absolute temperature of the gas.

$$P \propto T$$

$$P =$$

$$\frac{P}{T} =$$

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

Write this!

Where: P_1 pressure before (mm hg, kPa or atm)

P_2 pressure after

T_1 temperature before (in K)

T_2 temperature after

Do these mixed questions.

1. If 2 Litres of gas at 27°C are heated in a cylinder, and the piston is allowed to rise so that pressure is kept constant, how much space will the gas take up at 327°C?

   4 L

2. On hot days you may have noticed that potato chip bags seem to inflate. If I have a 250 mL bag at a temperature of 19.0°C and I leave it in my car at a temperature of 60.0°C, what will the new volume of the bag be?

   (assume that most of the bag is filled with gas, that the chips are negligible volume)

   285 mL

3. A sealed can contains 310 mL of air at room temperature (20°C) and an internal pressure of 100 kPa. If the can is heated to 606 °C what will the internal pressure be?

   3.00×10^2 kPa

Example

- **A sealed can contains 310 mL of air at room temperature (20°C) and an internal pressure of 100 kPa. If the can is heated to 606 °C what will the internal pressure be?**

Remove irrelevant fact

Data:

P_1= 100kPa

~~V_1=310 mL~~

T_1=20°C

P_2=unknown

T_2=606°C

Celsius must be converted to kelvins

20°C =293K 606°C =879K

Formula:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{100kPa}{293K} = \frac{x}{879K}$$

multiply

divide

$x = 87900 \div 293$

$x = 300$

Answer: the pressure will be 3.00×10^2 kPa

Next slide: T vs P graph

Assignment on Gay-Lussac's Law

- Read pages 85 to 87
- Answer questions #22 to 30 on page 98

Lesson 2.4.4

Avogadro's Laws

For amount of gas.

"The volume or pressure of a gas is directly related to the number of moles of gas"

$$V \propto n$$

$$P \propto n$$

Lorenzo Romano Amedeo Carlo Avogadro di Quaregna

- 1776 - 1856
- Field: Physics
- Known for:
  - Avogadro's hypothesis
  - Avogadro's number.
- "Equal volumes of gas at the same temperature and pressure contain the same number of moles of particles."

- You already know that:
  - That a mole contains a certain number of particles (6.02×10^{23})
  - So 1 mole of any gas will occupy the same volume at a given T & P!

- New! How will changing the amount of gas present affect pressure or volume?

Inc. # of moles = inc. volume (if it <u>can</u> expand),

Inc. # of moles = inc. pressure (if it <u>is unable</u> to expand).

It’s mostly common sense…

- If you pump more gas into a balloon, and allow it to expand freely, the volume of the balloon will increase.

- If you pump more gas into a container that can’t expand, then the pressure inside the container will increase.

Avogadro's Law

Write this!

- The volume of a fixed amount of gas is directly proportional to the # of moles.

  (if P & T are constant)

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

- The pressure of a fixed amount of gas is directly proportional to the # of moles.

  (if V & T are constant)

$$\frac{P_1}{n_1} = \frac{P_2}{n_2}$$

n = # of moles

- The volume of 1 mole of an ideal gas depends on the conditions:
  - At STP one mole of an ideal gas has a volume of **22.4 litres**
  - AT SATP one mole of an ideal gas has a volume of **24.5 litres**
- Since all common gases are very near ideal at these temperatures, we can use these as standard molar volumes for ANY common gas.

Write this!

Comparison of Conditions

STANDARD	Standard Temperature & Pressure (STP)	Ambient Temperature & Pressure (SATP)
Pressure	101.3 kPa	101.3 kPa
Temperature °C	0 °C	25 °C
Temperature K	273.15 K	298.15 K
Molar Volume	22.4141 L/mol	24.4714 L/mol
# moles	1.00 mol	1.00 mol

Today:

- Hand-in lab by the end of lunch.
- Return & go over test.
- Textbook pages 98-99 questions

  31, 32, 34, 37, 38, 39, 49, 50

Go over if time / finish for homework if necessary.

Simple gas Laws: Summary

Boyle's Law: $P \propto \frac{1}{V}$ $P_1V_1 = P_2V_2$

Charles' Law: $V \propto T$ $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Gay-Lussac's Law: $P \propto T$ $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Avogadro's Laws: $P \propto n$ $\frac{P_1}{n_1} = \frac{P_2}{n_2}$

$V \propto n$ $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

How can we combine these?

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

$$V \propto T$$

$$P \propto T$$

$$P \propto n$$

$$V \propto n$$

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

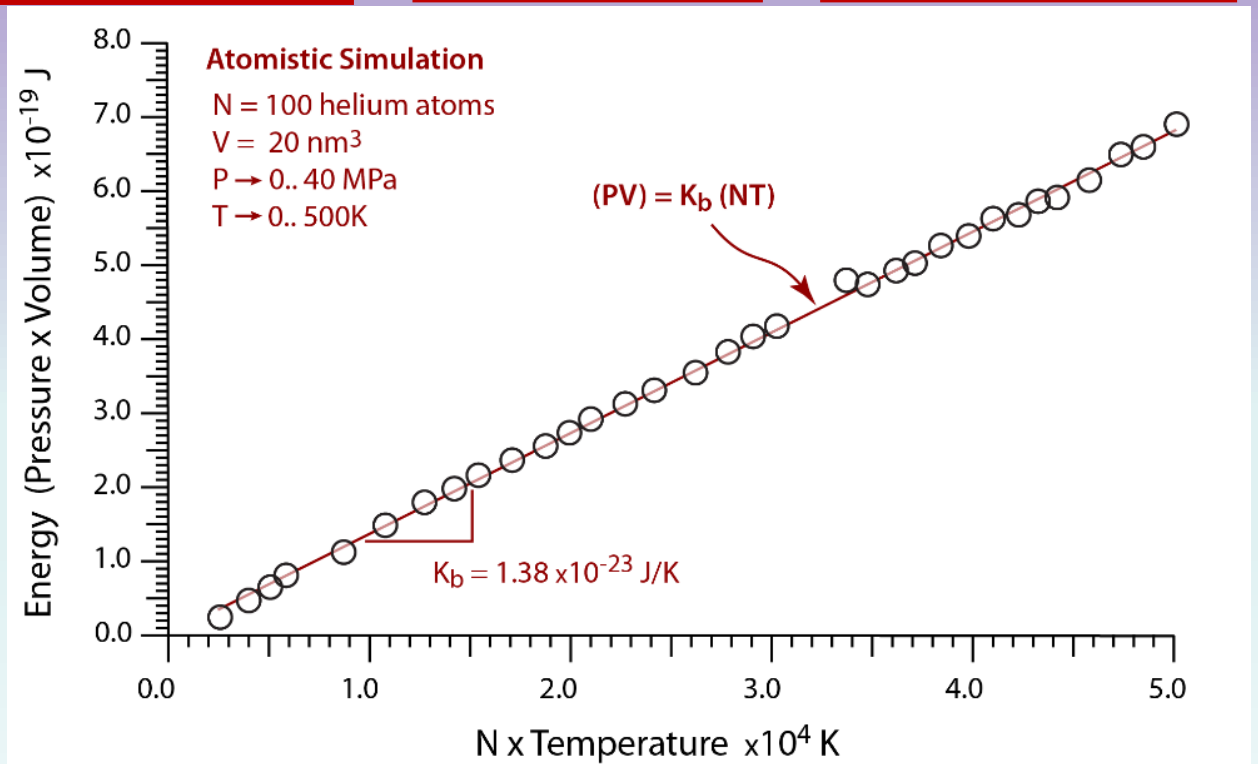
$$P \propto \frac{T}{V}$$

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

$$PV \propto nT$$

$$PV = k \times nT$$

$$PV = nRT$$



Calculate *R* at STP and SATP.

$$PV = nRT$$

$$\frac{P_1V_1}{n_1T_1} = R$$

At STP:

$$\frac{(101.3kPa)(22.4141L)}{(1.00mol)(273.15K)} = ?$$

$$8.31kPa \bullet L/(mol \bullet K)$$

At SATP:

$$\frac{(101.3kPa)(24.4714L)}{(1.00mol)(298.15K)} = ?$$

Write this!

- We can combine the gas laws to form the Ideal gas law.

$$\boxed{PV = nRT}$$

- Where *R* is the Ideal Gas Constant

$$R = 8.31 kPa \bullet L/(mol \bullet K)$$

Example

Write this!

You have 8.0 g of oxygen gas at 2.0×10^2 kPa & 15°C.

How many litres of oxygen are there?

Work:

#1 Find # of moles of O_2

$$\frac{1 mol O_2}{x} = \frac{32.00 g}{8.0 g}$$

$$x = 0.25 mol O_2$$

#2 $PV = nRT$

$$200 kPa \bullet V = 0.25 mol \bullet 8.31 \frac{L \cdot kPa}{K \cdot mol} \bullet 288 K$$

$$V = \frac{0.25 mol \bullet 8.31 \frac{L \cdot kPa}{K \cdot mol} \bullet 288 K}{200 kPa}$$

$$V = 2.99 L$$

Ans: There are 3.0 L of oxygen (2 sig.figs.)

- Start worksheet:
  - due next class
  - counts for term 2

Write this!

- One can test a gas to check if it is an "ideal gas" for certain P, V, T & n conditions. By checking if the calculated constant R is in fact

$$8.31 kPa \bullet L/(mol \bullet K)$$

Ex. A sample of gas contains 1 mole, occupies 25L, at 100 kPa & 27°C. Is the gas ideal?

- Convert to kelvins: 27°C+273=300K

  PV=n*R*T

  ***R* = PV/nT**

  R=(100*kPa)(25L)*÷(300*K)(*1*mol*)

  R=8.33 $^{kPa \bullet L}/_{K \bullet mol}$ we expected 8.31 $^{kPa \bullet L}/_{K \bullet mol}$

- So the gas is <u>not</u> perfectly ideal, but it is very close to an ideal gas,
  - It varies from ideal by only 0.24%

Write this!

- We can rearrange PV=nRT to give:

Before: $$\frac{P_1V_1}{n_1T_1} = R$$

After: $$\frac{P_2V_2}{n_2T_2} = R$$

- R is always 8.31 for ideal gases so…..

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

- This is the Combined Gas Law

The Ideal Gas Law & the Combined Gas Law are given on the formula sheet!

The neat thing about the Combined gas law is that it can replace the three original gas laws.

Just cross out or cover the parts that don't change, and you have the other laws:

- If the temperature is constant, then you have Boyle's law.
- If, instead, pressure remains constant, you have Charles' Law
- And finally, if the volume stays constant, then you have Gay-Lussac's Law

Most of the time, the number of moles stays the same, so you can remove moles from the equation.

$$\frac{P_1V_1}{\square} = \frac{P_2V_2}{\square}$$

John Dalton

1766-1844 England

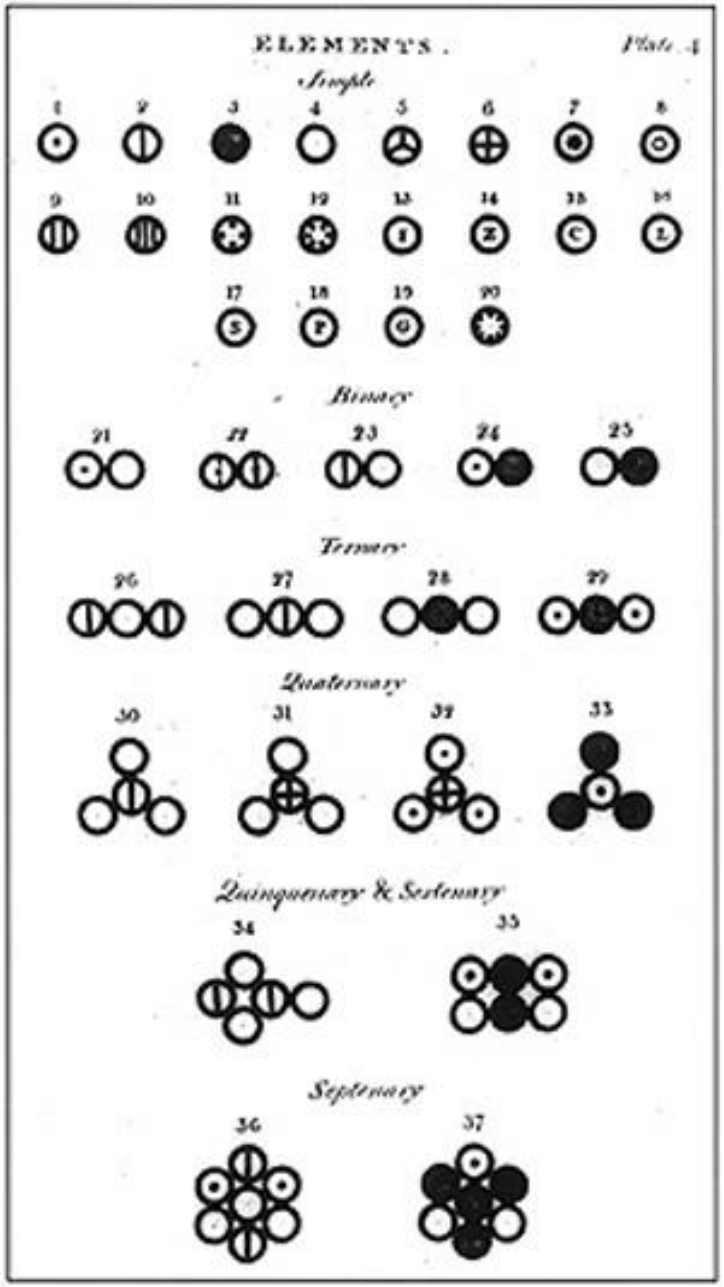
Known for:

modern atomic theory.

Studying colorblindness

experimentation on gases

first to estimate the composition of the atmosphere:



Kinetic Theory Connection

- Hypothesis 3 of the kinetic theory states that **gas particles do not attract or repel each other.**
- Dalton established that each type of gas in a mixture **behaved independently** of the other gases.
- The pressure of **each gas contributes towards the total pressure** of the mixture.
- This is called **Partial Pressure**

Write this!

Dalton's Law of partial pressures of gases

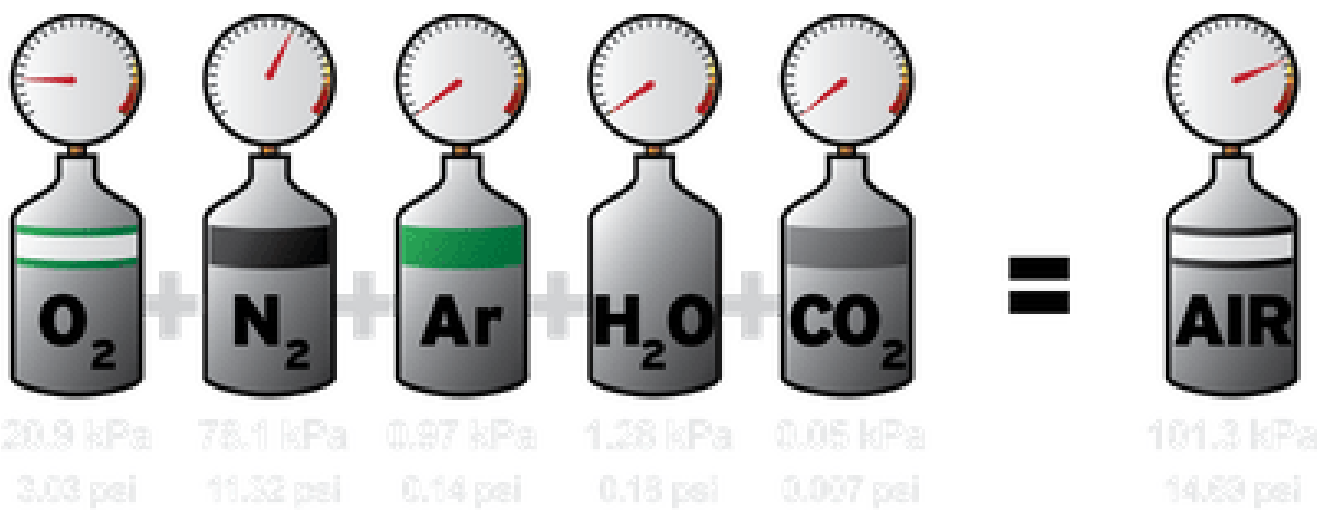
$$P_T = P_1 + P_2 + \ldots$$

Where: P_T = total pressure of mixed gases

P_1 = pressure 1st gas

P_2 = pressure 2nd gas

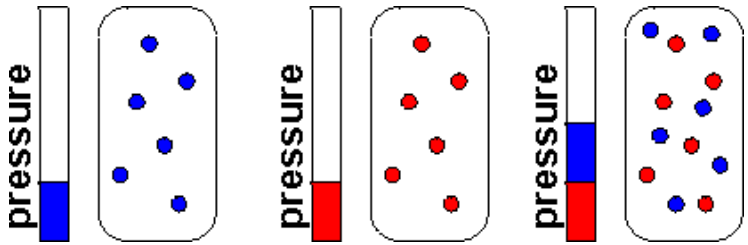
etc...



The pressure of each gas in a mixture is determined by the # moles.

Write this!

It is calculated by:



$$P_A = \frac{n_A}{n_T} \cdot P_T$$

Where: P_A=Pressure of gas A

n_A = moles of gas A in the mixture

n_T = total moles of all gases in the mixture

P_T = Total Pressure of all gases

Write this!

Ex.1

A sample of air has a total pressure of 101kPa.

What is the partial pressure of oxygen if,

P_{N_2}=79.1kPa, P_{CO_2}= 0.04kPa & P_{others}=0.947kPa?

$$P_{Total} = P_{O_2} + P_{N_2} + P_{CO_2} + P_{others}$$

$$101 = ? + 79.1 + 0.04 + 0.947$$

$$P_{O_2} = 21 \text{ kPa}$$

Write this!

Ex. 2

A gas mixture contains 0.25 moles of H_2 gas and 1.20 moles O_2 gas.

What is the partial pressure of O_2 gas if the total pressure is 200.0 kPa?

$n_T = 0.25 + 1.20$

$= 1.45$

$$P_{O_2} = \frac{n_{O_2}}{n_T} \cdot P_T$$

$$P_{O_2} = \frac{1.20}{1.45} \cdot 200$$

$$P_{O_2} = 166\ kPa$$

Uses of Dalton's Law

In the 1960s NASA used the law of partial pressures to reduce the launch weight of their spacecraft. Instead of using air at 101 kPa, they used pure oxygen at 20kPa.

Breathing low-pressure pure oxygen gave the astronauts just as much "partial pressure" of oxygen as in normal air.

Lower pressure spacecraft reduced the chances of explosive decompression, and it also meant their spacecraft didn't have to be as strong or heavy as those of the Russians (who used normal air).. This is one of the main reasons the Americans beat the Russians to the moon.

Carelessness with pure oxygen, however, lead to the first major tragedy of the American space program…

At 20 kPa, pure oxygen is very safe to handle, but at 101 kPa pure oxygen makes everything around it extremely flammable, and capable of burning five times faster than normal.

On January 27, 1967, during a pre-launch training exercise, the spacecraft Apollo-1 caught fire. The fire spread instantly, and the crew died before they could open the hatch.

Crew of Apollo 1

Gus Grissom, Ed White, Roger Chaffee

Exercises
:

- Page 113 in new textbook, # 1 to 8

Extra practice (if you haven’t already started):

- Study guide: pp 2.12 to 2.17 # 1 to 22
  - There is an answer key in the back for these
  - Do these on your own as review

Summary:

- Dalton's Law: The total pressure of a gas mixture is the sum of the partial pressures of each gas.

$$P_T = P_1 + P_2 + \ldots$$

- Graham's Law: light molecules diffuse faster than heavy ones

$$\frac{Rate_1}{Rate_2} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

- Avogadro's hypothesis
  - A mole of gas occupies 22.4L at STP and contains 6.02×10^{23} particles

Summary of Kinetic Theory

- Hypotheses (re. Behaviour of gas molecules):
  1. Gases are made of molecules moving randomly
  2. Gas molecules are tiny with lots of space between.
  3. They have elastic collisions (no lost energy).
  4. Molecules don't attract or repel each other (much)
- Results:
  - The kinetic energy of molecules is related to their temperature (hot molecules have more kinetic energy because they move faster)
    - Kinetic theory is based on averages of many molecules (graphed on the Maxwell distribution "bell" curve)
    - Pressure is caused by the collision of molecules with the sides of their containers.
    - Hotter gases and compressed gases have more collisions, therefore greater pressure.

- The end of module 2
- The teacher will set a test date.