12 • The Gas Laws
Boyle's Law (P and V)
(1 of 8)

General: When P↑, V↓ (inversely proportional)

Formula: $P \cdot V$ = constant or $P_1V_1 = P_2V_2$

Restrictions: P_1 and P_2 must be in the same units
V_1 and V_2 must be in the same units

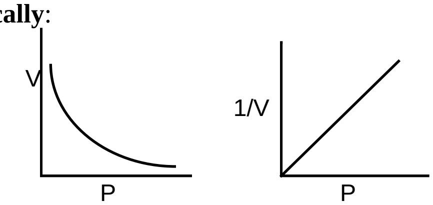
Convert pressures using conversion factors using the fact that 1 atm = 760 mmHg = 760 torr = 101.3 kPa = 14.7 psi

$\text{psi} = \frac{\text{lb}}{\text{in}^2}$

Example: $730 \text{ mmHg} \times \frac{101.3 \text{ kPa}}{760 \text{ mmHg}} = 97.3 \text{ kPa}$

12 • The Gas Laws
Boyle's Law Lab
(2 of 8)

Graphically:



In our lab, we had to **add** the **atmospheric pressure** to our measurements because tire **gauges** only measure the pressure **ABOVE** atmospheric pressure.

Consistent (" good") data form a **straight** line (P vs. $\frac{1}{V}$).

12 • The Gas Laws
Kelvin Temperature Scale
(3 of 8)

K = °C + 273 **°C = K – 273**

Examples: 0 °C + 273 = 237 K
25 °C + 273 = 298 K
100 °C + 273 = 373 K

300 K – 273 = –27 °C

The **Kelvin** scale is used in gas law problems because the pressure and volume of a gas depend on the **kinetic energy** or **motion** of the particles.

The **Kelvin** scale is **proportional** to the **KE** of the particles… that is, **0 K** (**absolute zero**) means **0 kinetic energy**. **0 °C** is simply the **freezing point** of water.

12 • The Gas Laws
Charles' Law (V and T)
Gay-Lussac's Law (P and T)
(4 of 8)

<u>**Charles' Law**</u>

General: When T↑, V↑ (directly proportional)

Formula: $\frac{V}{T}$ = constant or $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Restrictions: T must be in Kelvins
V_1 and V_2 must be in the same units

<u>**Gay-Lussac's Law**</u>

General: When T↑, P↑ (directly proportional)

Formula: $\frac{P}{T}$ = constant or $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Restrictions: T must be in Kelvins
P_1 and P_2 must be in the same units

12 • The Gas Laws
The Combined Gas Law
(5 of 8)

Formula: $\frac{P \cdot V}{T}$ = constant or $\frac{P_1 \cdot V_1}{T_1} = \frac{P_2 \cdot V_2}{T_2}$

Restrictions: T must be in Kelvins
V_1 and V_2 must be in the same units
P_1 and P_2 must be in the same units

STP ("standard temperature and pressure") is often used as one of the two conditions
T = 0 °C = 273 K **P** = 1 atm = 760 mmHg = 101.3 kPa

Each of the **three gas laws** is really a **special case** of this law.
Example: If $T_1 = T_2$, the law becomes $P_1V_1 = P_2V_2$

12 • The Gas Laws
The Ideal Gas Law
(6 of 8)

Formula: $P \cdot V = n \cdot R \cdot T$ or $PV = nRT$
where
P = pressure
V = volume
n = number of moles
R = the ideal gas constant
T = temperature (in Kelvins)

The value of R depends on the P and V units used.

$R = \frac{PV}{nT}$ so you can use the molar volume info to calculate R

$$R = \frac{(101.3 \text{ kPa})(22.4 \text{ L})}{(1 \text{ mole})(273 \text{ K})} = 8.31 \frac{\text{L} \cdot \text{kPa}}{\text{mol} \cdot \text{K}}$$

$$R = 62.4 \frac{\text{L} \cdot \text{mmHg}}{\text{mol} \cdot \text{K}} = 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

12 • The Gas Laws
Dalton's Law of Partial Pressure
(7 of 8)

When you have a **mixture** of gases, you can determine the pressure exerted by each gas separately. This is called the **partial pressure** of each gas.
Since each gas has the same power to cause pressure (see card #8) the partial pressure of a gas depends on how much of the mixture is composed of each gas (in moles)
Example: Consider air, a mixture of mostly O_2 and N_2

$$\frac{\text{moles } O_2}{\text{moles total}} = \frac{P_{O_2}}{P_{total}} \qquad \frac{\text{moles } N_2}{\text{moles total}} = \frac{P_{N_2}}{P_{total}}$$

Also: $P_{total} = P_{O_2} + P_{N_2}$
This idea is used when a **gas is collected over water**
$P_{atm} = P_{gas} + P_{H_2O}$ P_{H_2O} is found on a **chart**

12 • The Gas Laws
Why Do All Gases Cause the Same Pressure?
Graham's Law
(8 of 8)

The gas laws work (to 3 significant digits) for **all** gases… that is, all gases have the same **power** to cause **pressure**.

At the **same temperature**, the **KE** of each gas is the **same**.
KE = $^1/_2$ mass·velocity^2… if two particles have different **masses**, their **velocities** are also different. So…

SMALL particles move FAST $\quad mV^2$

LARGE particles move SLOWLY $\quad Mv^2$

We can use this idea with numbers as well: (Graham's Law)
$KE_A = KE_B \qquad m_Av_A^2 = m_Bv_B^2$
[another version of this formula is on pg 323 of the text]