Dr Kepert, Chem 1901

Gases, Liquids and Solids II

Silberberg 2nd Ed, Chapter 5
Any questions or requests - email to c.kepert@chem.usyd.edu.au

Ideal Gas Equation

\[ PV = nRT \]

Ideal gas constant \( R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1} \)

Units must be consistent

Common Units of Pressure

<table>
<thead>
<tr>
<th>Unit</th>
<th>Atmospheric Pressure</th>
<th>Scientific Field</th>
</tr>
</thead>
<tbody>
<tr>
<td>pascal (Pa); kilopascal (kPa)</td>
<td>1.01325x10^5 Pa; 101.3 kPa</td>
<td>SI unit; physics, chemistry</td>
</tr>
<tr>
<td>atmosphere (atm)</td>
<td>1 atm</td>
<td>Chemistry</td>
</tr>
<tr>
<td>millimeters of mercury (mmHg)</td>
<td>760 mmHg*</td>
<td>Chemistry, medicine, biology</td>
</tr>
<tr>
<td>torr</td>
<td>760 torr</td>
<td>Chemistry</td>
</tr>
<tr>
<td>pounds per square inch (psi or lb/in^2)</td>
<td>14.7 lb/in^2</td>
<td>Engineering</td>
</tr>
<tr>
<td>bar</td>
<td>1.01325 bar</td>
<td>Meteorology, chemistry, physics</td>
</tr>
</tbody>
</table>

*This is an exact quantity; in calculations, we use as many significant figures as necessary.

R in Different Units

\[ R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1} \]

Energy = \( J = \text{ kg m}^2 \text{ s}^{-2} \) (eg \( E = \frac{1}{2} \text{ mv}^2 \))

Pressure = \( P = \text{ Pa} = \text{ force/area} = \text{ kg m}^{-1} \text{ s}^{-2} \) (force = \( \text{ kg m s}^{-2} \))

\[ P = \frac{\text{force}}{\text{area}} = \frac{\text{kg m s}^{-2}}{\text{m}^2} = \text{kg m}^{-1} \text{ s}^{-2} \]

\[ \Rightarrow P \text{ m}^3 \text{ has units kg m}^{-2} \text{ s}^{-2}, \text{ same as } J. \]

\[ R = 8.314 \text{ Pa m}^3 \text{ mol}^{-1} \text{ K}^{-1} \]

\[ = 8.314 \text{ kPa L mol}^{-1} \text{ K}^{-1} \]

Car Air Bags

Often made with sodium azide:

\[ 6\text{NaN}_3(s) + \text{Fe}_2\text{O}_3(s) \rightarrow 3\text{Na}_2\text{O}(s) + 2\text{Fe}(s) + 9\text{N}_2(g) \]

Example: How many grams of \( \text{NaN}_3 \) are required to produce 75 L of nitrogen at 25 °C and 1.31 atm?

\[ \frac{PV = nRT}{n} = \frac{PV}{RT} \]

\[ = \text{ moles of } \text{NaN}_3 = \text{ mass } \text{NaN}_3 \text{ (mol. wt. } = 65.0) = \]

Dalton’s Law of Partial Pressures

Total pressure of mixture of gases = sum of pressures that each would exert if it were alone.
Mixture of gases:

- Ideal gas: no. moles \( \propto \) partial pressure of \( i \), \( P_i \)

Mole fraction of gas \( i \):

\[
\text{mole fraction of } i = \frac{\text{no. moles}_i}{\text{total no. moles}}
\]

Dry Air:

- \( P_{N_2} = 0.7808 \text{ atm} \)
- \( P_{O_2} = 0.2095 \text{ atm} \)
- \( P_{Ar} = 0.0093 \text{ atm} \)
- \( P_{CO_2} = 0.00034 \text{ atm} \)

Find the volume of gas generated by dissolving 1.00 g Zn in excess HCl at 27 \(^\circ\)C, 1 atm pressure, with gas collected over water; vapour pressure of water at 27 \(^\circ\)C = 3.6 kPa. 1 atm = 101.3 kPa.

Example:

Find the volume of gas generated by dissolving 1.00 g Zn in excess HCl at 27 \(^\circ\)C, 1 atm pressure, with gas collected over water; vapour pressure of water at 27 \(^\circ\)C = 3.6 kPa. 1 atm = 101.3 kPa.

Example cont’d:

- Total pressure = 1 atm = (Dalton’s law)
- \( \Rightarrow \) partial pressure of \( H_2 = P = \)
- \( \Rightarrow \) no. mol \( H_2 = \)

<table>
<thead>
<tr>
<th>( P ) (kPa)</th>
<th>( V ) (L)</th>
<th>( n ) (mol)</th>
<th>( R ) (kPa L mol(^{-1}) K(^{-1}))</th>
<th>( T ) (K)</th>
</tr>
</thead>
<tbody>
<tr>
<td>104.7</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

⇒ \( \text{Volume} = \)

Implications of \( PV = nRT \)

- Boyle’s Law: \( PV = \) constant for constant temperature
- Charles’ Law: \( \frac{V}{T} = \) constant for constant pressure
- Lowest temperature: \(-273 \(^\circ\)C\) (ie \( V = 0 \) at this temperature)
- Gay-Lussac: (combining volumes)
- Avogadro’s hypothesis: (equal volumes of gas have the same number of molecules)
- Avogadro’s constant: \( N_A = \# \text{ molecules/mol} = 6.02 \times 10^{23} \text{ mol}^{-1} \)
- Dalton’s law of partial pressures

Diverse observations from single unifying theory: one of the goals of science!

Graham’s Law

- Rate of gas effusion/diffusion \( \propto \frac{1}{\sqrt{\text{molar mass}}} \)

Reason: diffusion is at constant energy = \( 1/2 \text{ mv}^2 \)
If \( 1/2 \text{ mv}^2 \) is constant, then velocity \( v \propto \frac{1}{\sqrt{m}} \)

Diffusion time \( \propto 1/\text{velocity} \) ⇒ Graham’s Law.

Distribution of Molecular Speeds

See Silberberg - Fig 5.14, page 200

Definition of Temperature

The Kinetic Theory of Gases assumes that temperature is proportional to the average kinetic energy of the gas

- Ideal gas thermometer: use gas which behaves ideally (e.g., He at low pressure) to measure \( T \) from observed volume and pressure using \( PV = nRT \).