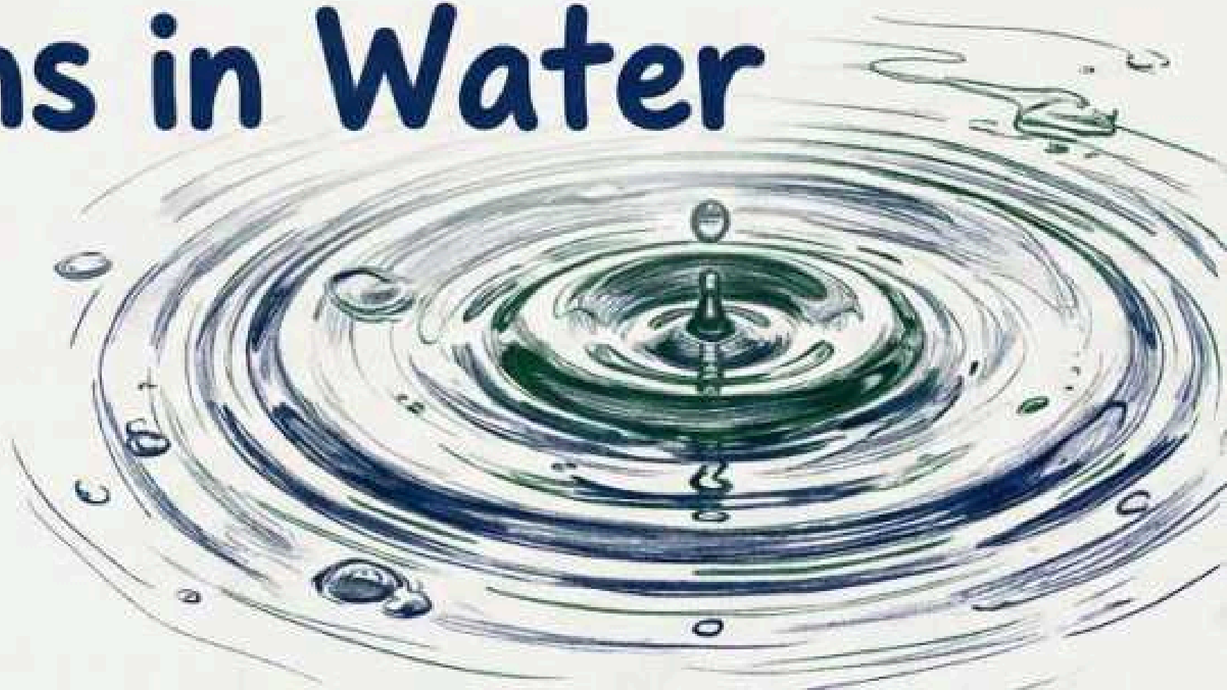
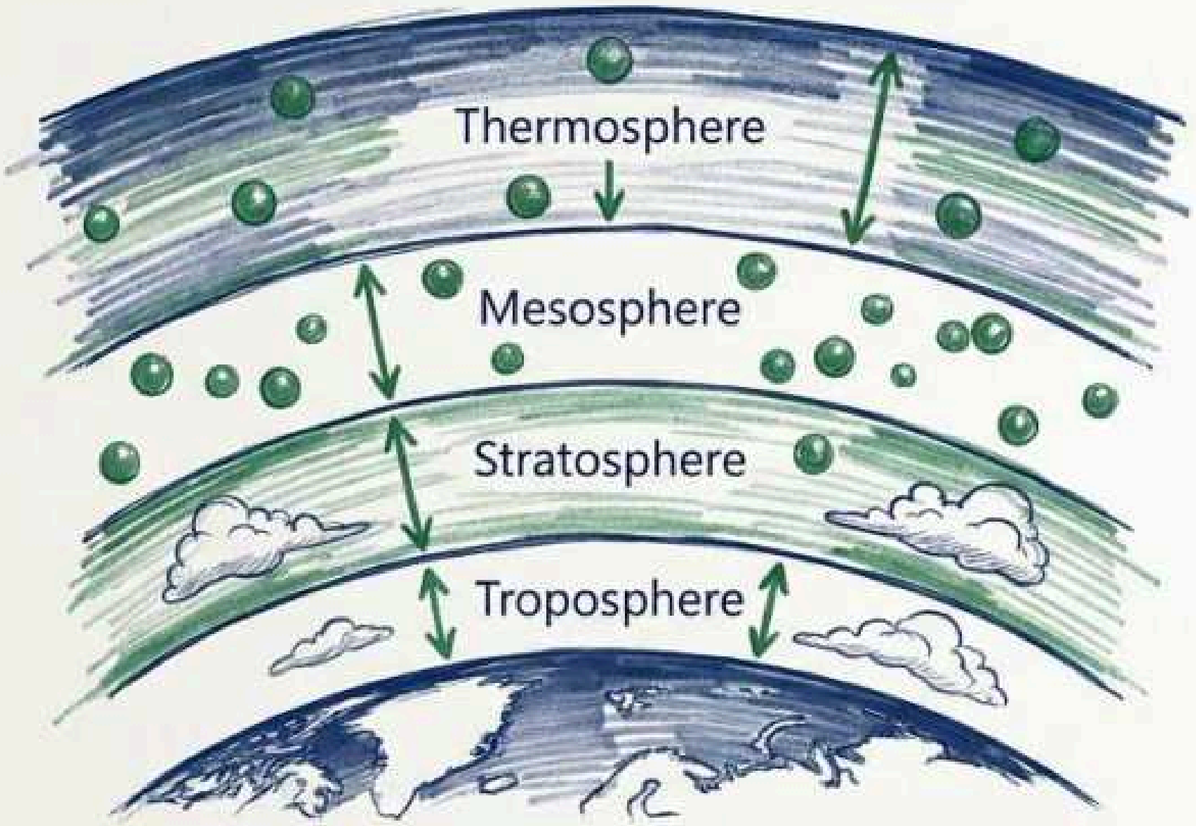
¶ 5.1

Lecture 5: Gases, Solutions, and Reactions in Water



¶ 5.2

The Air We Breathe



Bodies, Oceans, and Kitchens



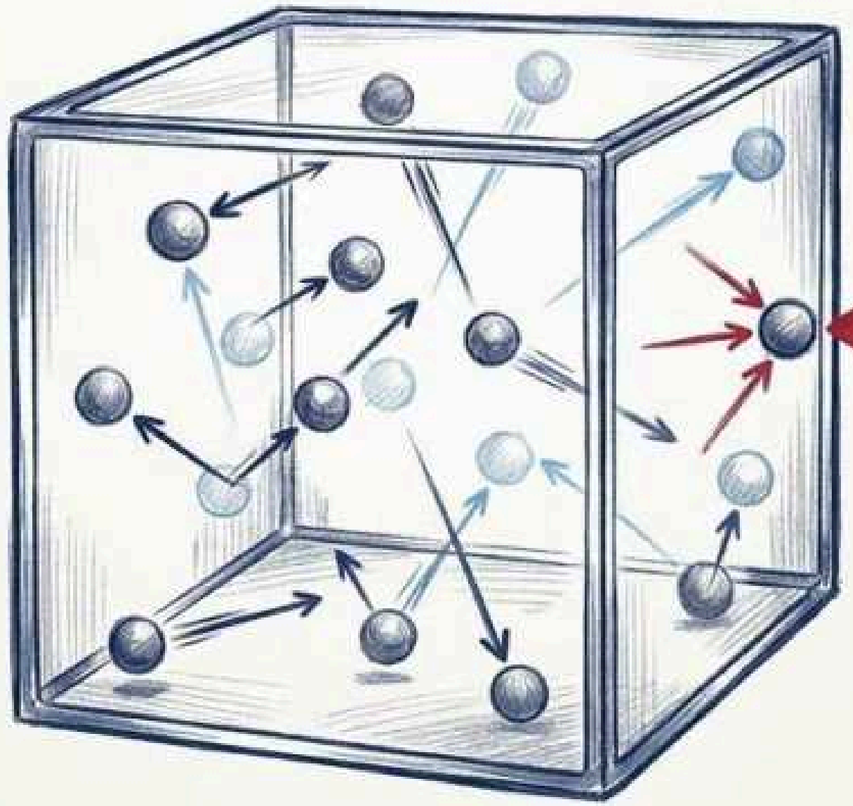
Our Path: Gas Laws → Solution Concentration → Reactions in Water.

¶ 5.3

The 4 Properties of a Gas:

Pressure (P), Volume (V), Temperature (T), and Amount (n)

Gas particles are tiny and far apart.



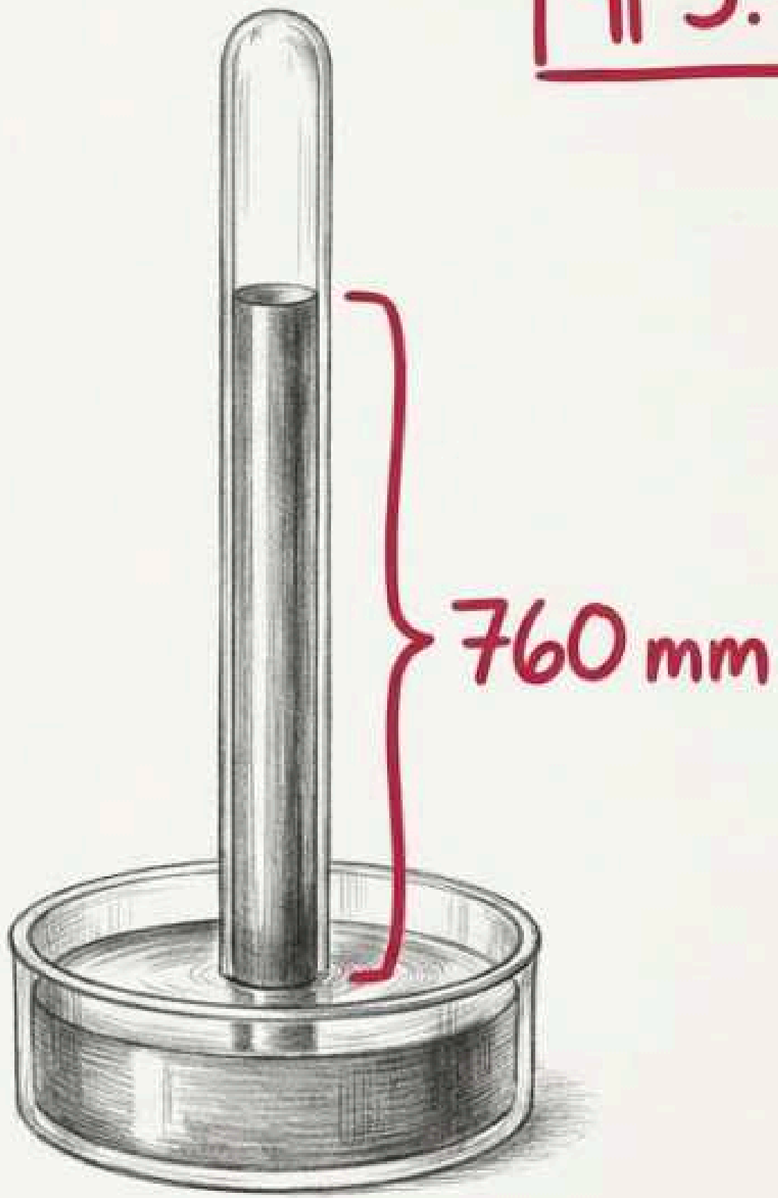
Pressure is the literal force of these collisions spread over the wall's area.

¶ 5.4

$$P = \frac{\text{Force}}{\text{Area}}$$

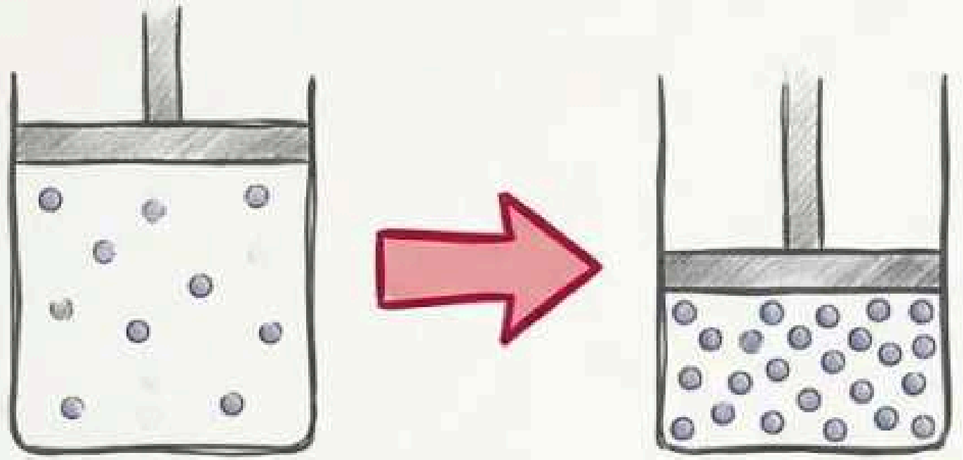
1 atm = 101.3 kPa = 760 mmHg

Weather forecasters track small changes to predict sorms. storms.



¶ 5.5

Boyle's Law (1662): Constant Temperature

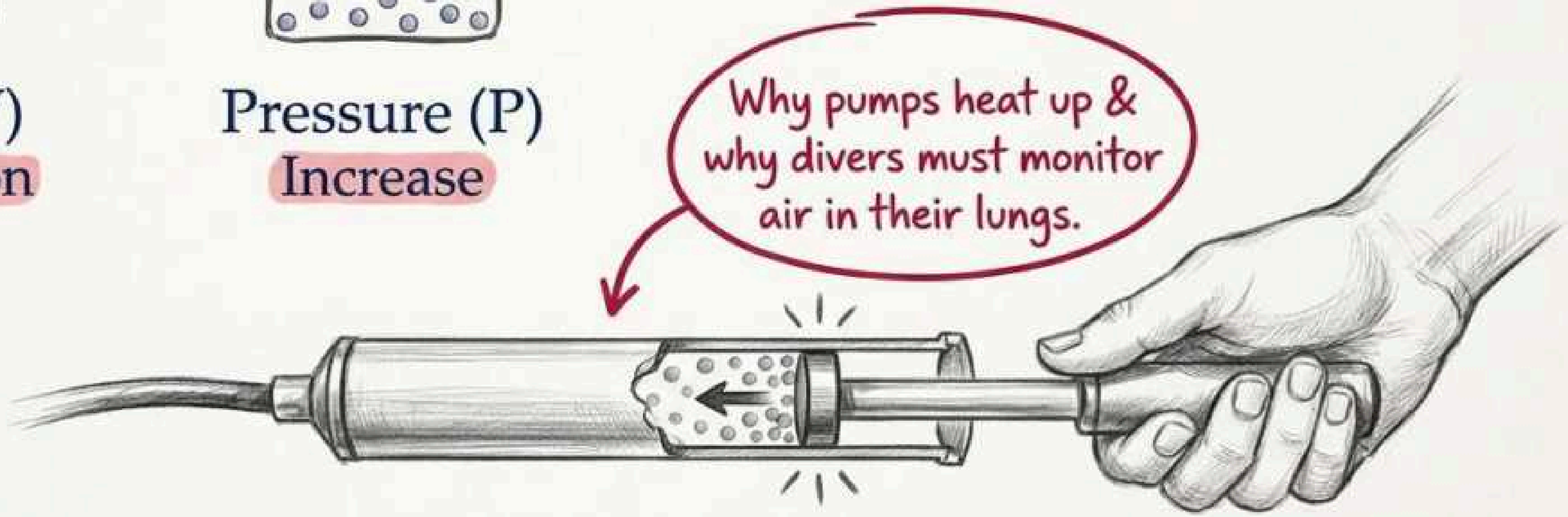


Volume (V)
Compression

Pressure (P)
Increase

$$P \times V = \text{Constant}$$

Example: 2L at 1 atm → 0.5L at 4 atm



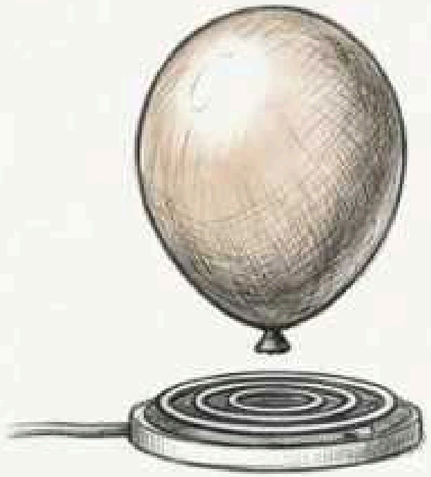
¶ 5.6

Charles's Law: Constant Pressure

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

Example: 3L at 300K → 3.5L at 350K

Before



After



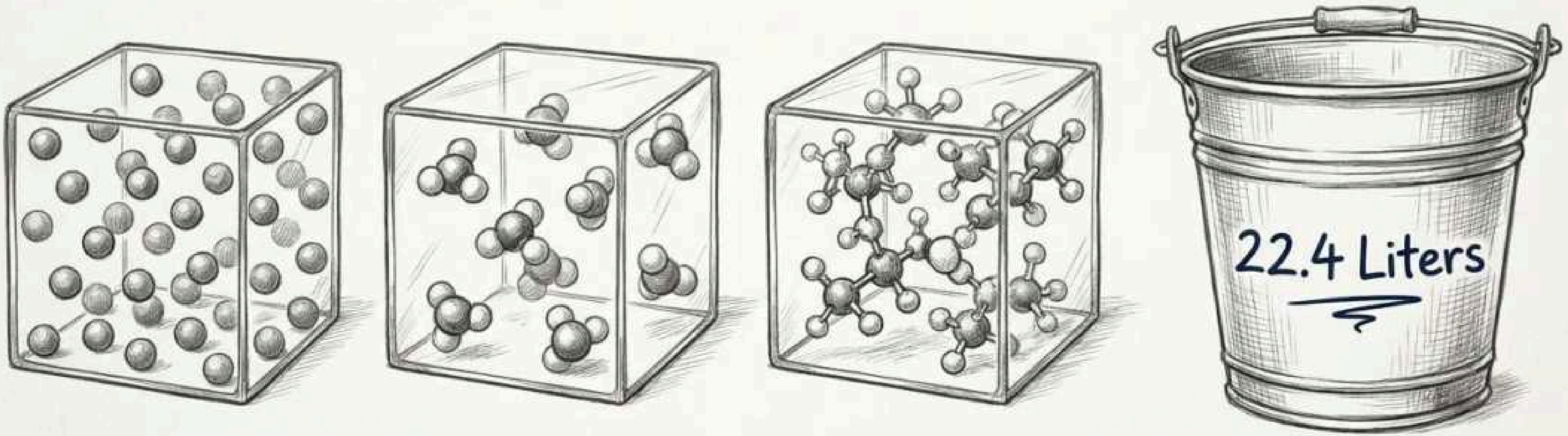
Must use Kelvin! Using Celsius is the most common mistake.

¶ 5.7

Avogadro's Law

Equal volume at same T and P = Equal number of particles.

1 mole of ANY gas = **22.4 Liters (at 0°C, 1 atm)**



¶ 5.8

$$PV = nRT$$

Pressure

Volume

Number of Moles

R (Constant) = $0.08206 \frac{\text{L-atm}}{\text{(mol-K)}}$

Temperature (Kelvin)

Substitute: $n = 1\ \text{mol}, T = 273\ \text{K}, P = 1\ \text{atm}$

$$1\ \text{atm} * V = 1\ \text{mol} * \left(0.08206 \frac{\text{L-atm}}{\text{mol-K}}\right) * 273\ \text{K} \rightarrow V = 22.4\ \text{L}$$

¶ 5.9

Case Study: How much space do 2.5 moles of gas take at 300K and 2 atmospheres?

Step 1 (Rearrange): $V = \frac{nRT}{P}$

Step 2 (Plug In): $V = \frac{2.5 \times 0.08206 \times 300}{2}$

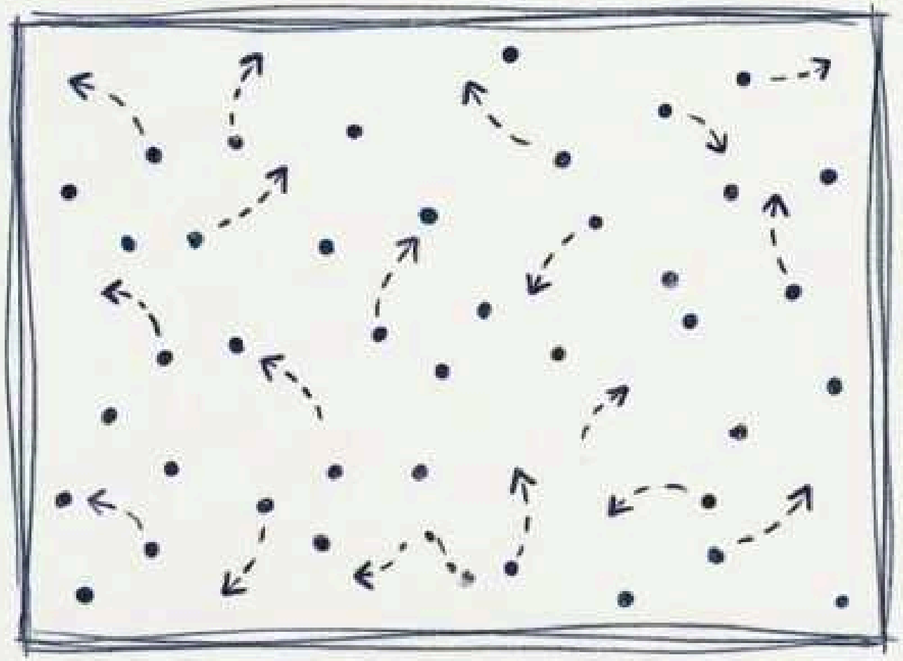
Step 3 (Solve): $V \approx 30.8 \text{ Liters}$

Always check your units before you trust a number.

¶ 5.10

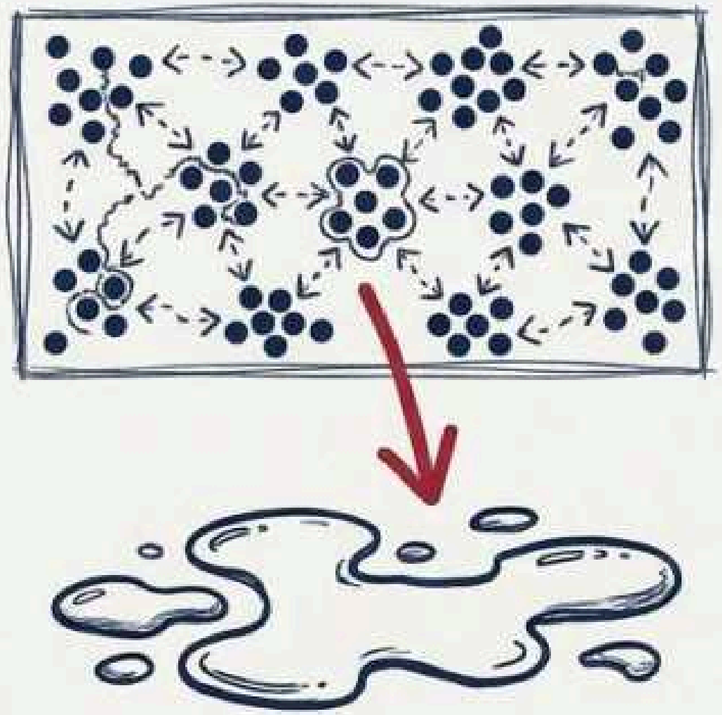
Ideal Gas

Assumes zero particle volume and zero attraction. Works well at ordinary conditions.



Real Gas

Particles possess actual volume and intermolecular attractions. Deviates from ideal at high pressure or low temperature.

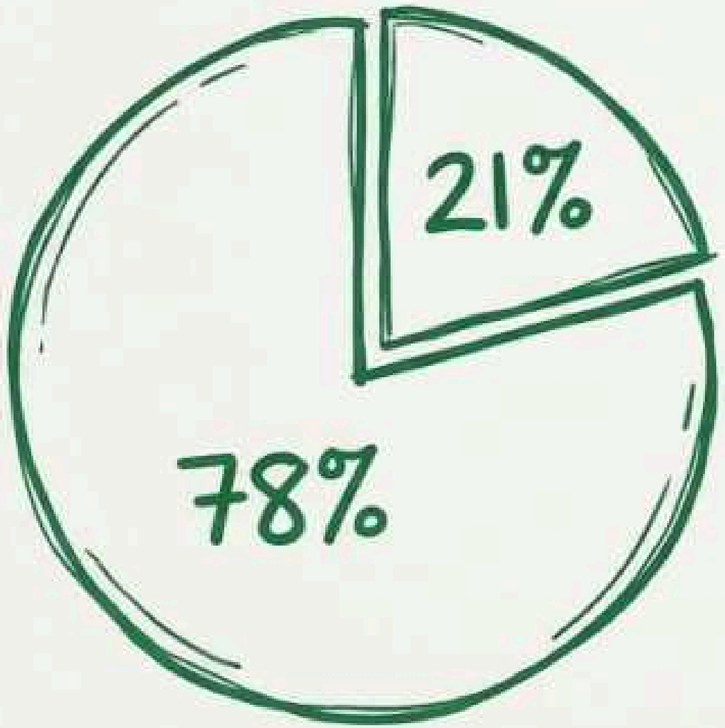


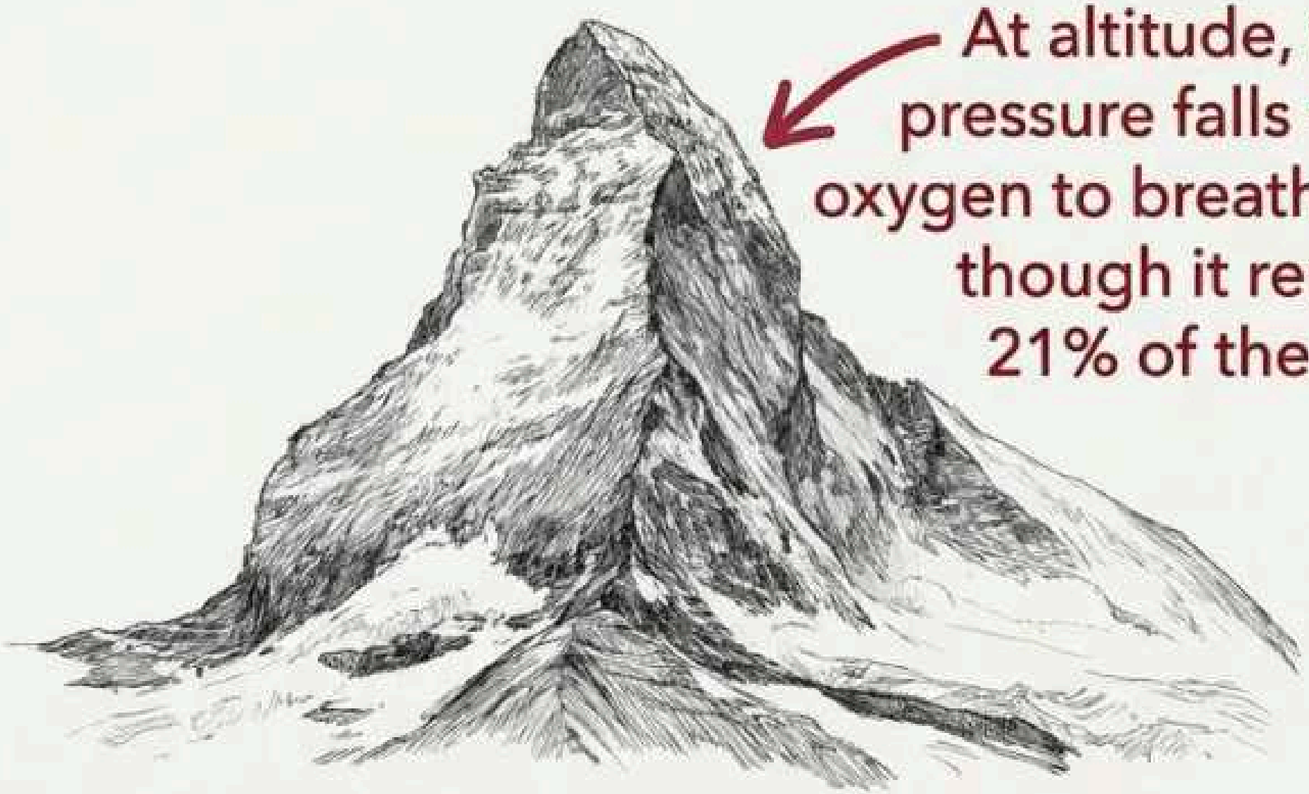
Dalton's Law: Total $P = P_1 + P_2 + P_3...$

¶ 5.11

Air mixture = 78% Nitrogen, 21% Oxygen.

At 1 atm, partial pressure of O_2 = 0.21 atm.



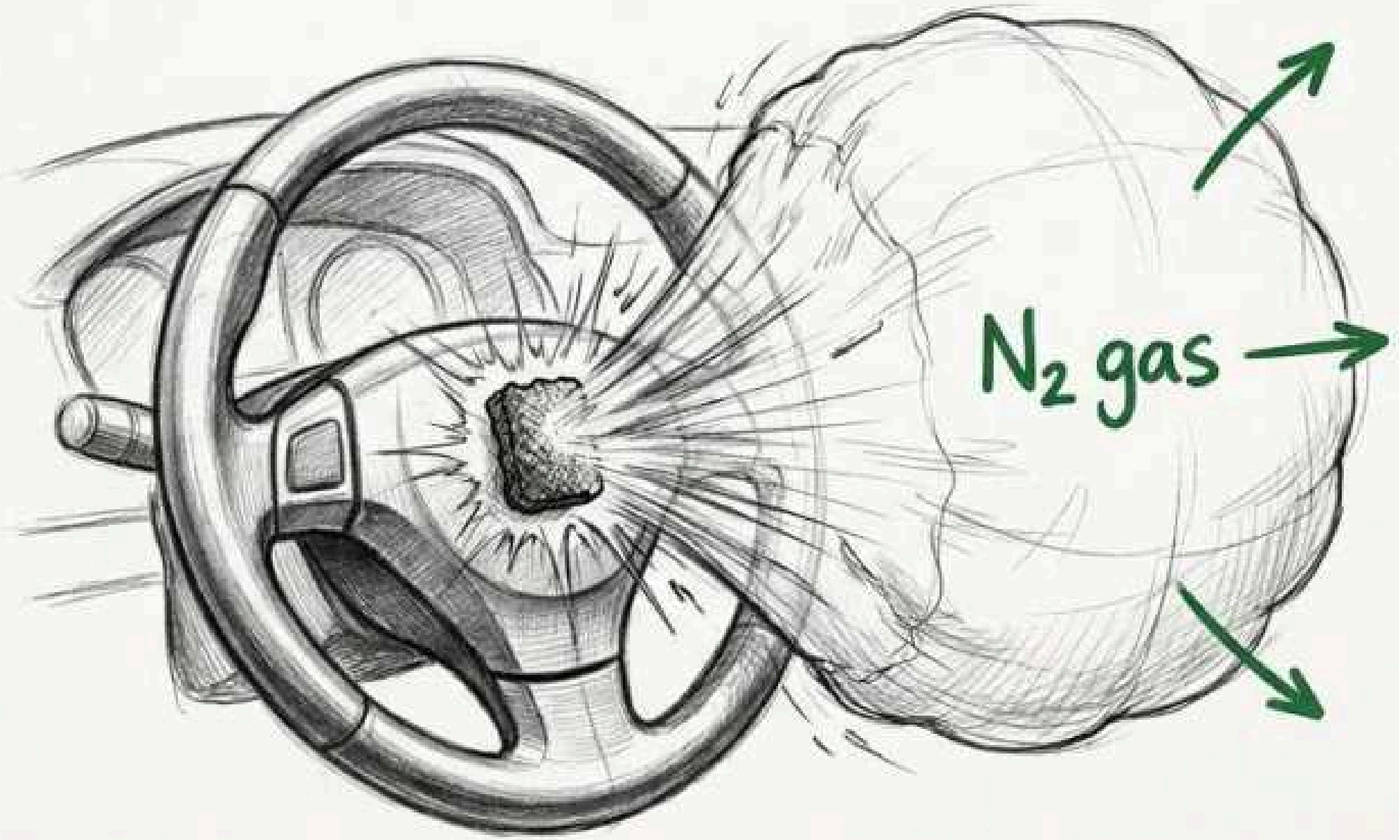


At altitude, total pressure falls → less oxygen to breathe, even though it remains 21% of the mix.

Stoichiometry meets Gas Laws:

¶ 5.12

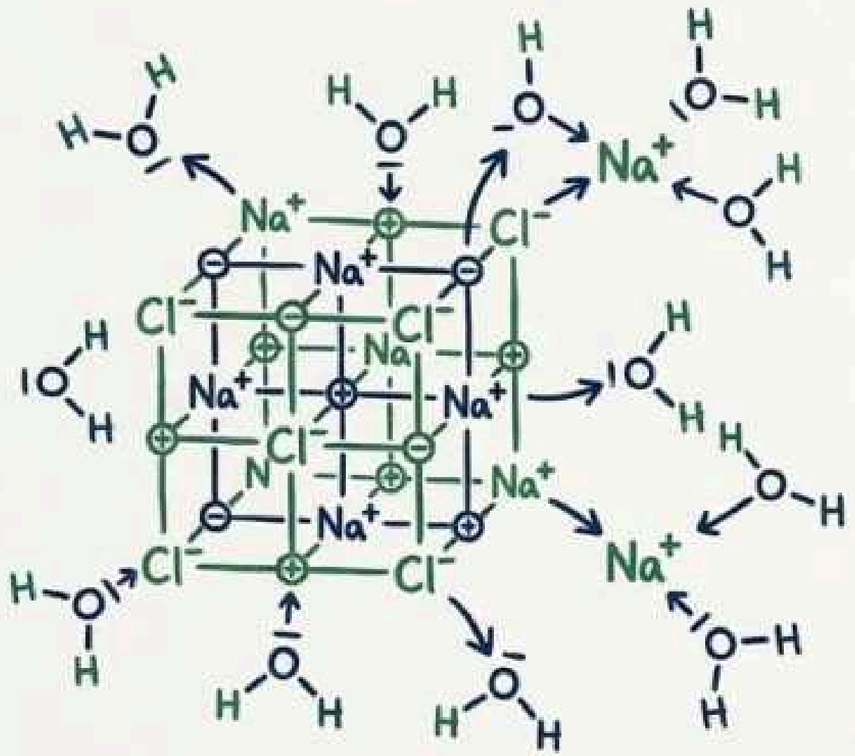
Solid Reaction → Moles of Gas formed (n) → Use PV=nRT → Predict Exact Volume



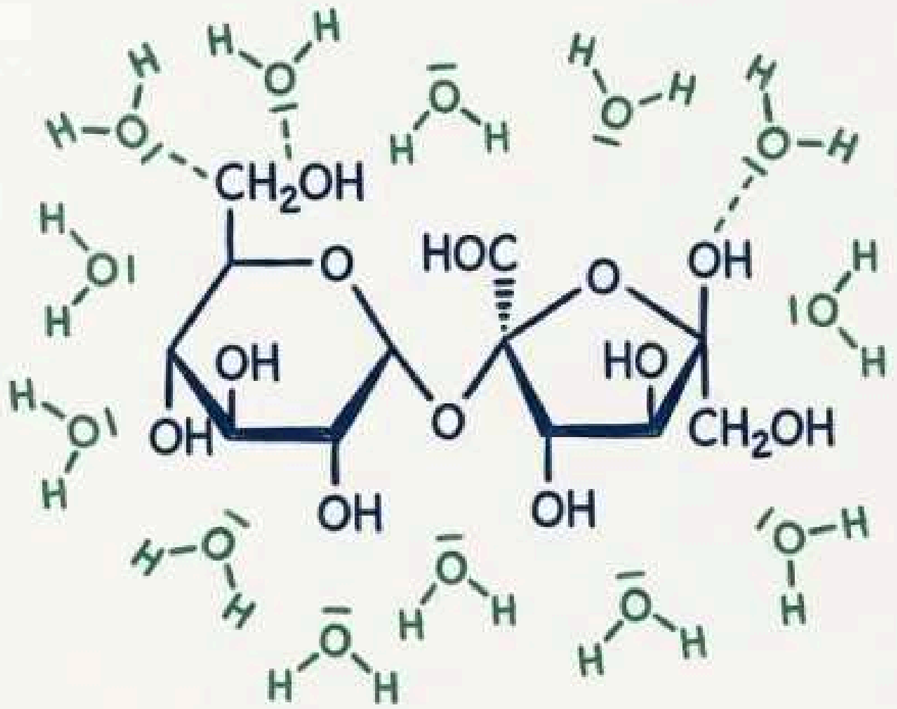
¶ 5.13

Solutions: Solute (dissolved) + Solvent (dissolver).
Water is the ultimate solvent.

Electrolyte: Conducts electricity (e.g., Salt)



Non-electrolyte: Does not conduct (e.g., Sugar)



Molarity (M)

¶ 5.14

$$\text{Molarity (M)} = \frac{\text{Moles of Solute}}{\text{Liters of Solution}}$$

The Math: 11.7g table salt (NaCl, molar mass 58.5) = 0.200 moles.

$$\frac{0.200 \text{ mol}}{0.5 \text{ L}} = 0.400 \text{ M}$$



¶ 5.15

The Math of Dilution: $M_1 \times V_1 = M_2 \times V_2$

(Moles of solute stay constant)

Goal: Make 250mL of 1M solution from a 6M stock.

Execution: Need 1/6 of the final volume = 41.7mL of stock.



Always add acid to water, never water to acid.

9.5.16

Precipitation: Soluble Solution + Soluble Solution -> Insoluble Solid.

Silver Nitrate + Sodium Chloride → Silver Chloride (solid).

Solubility Rules: Nitrates are always soluble; most silver salts are not.



¶ 5.17

Full Equation:

$$Ag^+ + \cancel{NO_3^-} + \cancel{Na^+} + Cl^- \longrightarrow AgCl_{(s)} + \cancel{Na^+} + \cancel{NO_3^-}$$

Net Ionic Equation:

$$Ag^+ + Cl^- \longrightarrow AgCl_{(s)}$$

The net equation strips away the spectators to show the chemistry that really happens.

¶ 5.18

Acid-Base

- Transfer of a hydrogen ion (H^+)

Redox

- Transfer of electrons (e^-)
- Oxidation = Losing electrons.
- Reduction = Gaining electrons.



¶ 5.19

Challenge 1:

A sealed can of air is left in a hot car. What happens to the pressure inside, and which law explains it?



Challenge 2:

Making a weaker solution from a stronger one. What do you measure, and what stays strictly constant?

¶ 5.20

Coming up in Lecture 6 (The Finale):

- Energy in reactions
- Reaction kinetics (speed)
- Chemical equilibrium
- Deep dive into Acids and Bases

