Class #9
Introduction to Gases

CHEM 107
L.S. Brown
Texas A&M University

Gases
• All gases have several properties in common:
  – Relatively low density
  – Compressible
  – Expand to fill any container

Ideal Gases
• Gas Law:
  \[ PV = nRT \]
• Originally based on empirical observations
• Can also be derived mathematically from the kinetic theory of gases
Pressure

- Dimensions of force/area
- SI unit is N/m², or Pascal
- Other more common units:
  - atmospheres:
    - 1 atm = 101,325 Pa
  - torr (mm Hg):
    - 760 torr = 1 atm
  - lbs/in²:
    - 14.7 psi = 1 atm

Gas Constant - R

- Proportionality constant, found by experiment
- In SI units:
  - R = 8.314 J mol⁻¹ K⁻¹
- Other units:
  - R = 0.08206 L atm mol⁻¹ K⁻¹
  - = 62.36 L torr mol⁻¹ K⁻¹

Units

- Consistent units are ALWAYS required!
- Use any convenient units for P, V
- Then choose R so that units all match
Estimate number of gas molecules in a typical balloon.

Temperature

\[ V = \frac{nRT}{P} \]

- If \( T \to 0 \), then \( V \to 0 \)
- MUST use absolute \( T \) (Kelvin, NOT °C)
  - \( 0°C = 273.15 \, K \)
  - \( 0 \, K = -459.7°F \)

What makes a gas “ideal?”

- In most simple terms, an ideal gas obeys the ideal gas law.
- But why? Can we predict when a gas might NOT obey the law?

Molecular viewpoint
Properties of an Ideal Gas

- Huge # of molecules
- Point masses, no volume
- No forces between molecules
- Constant, chaotic motion
- Elastic collisions

Gas Law = Limiting Case

- Any gas will behave ideally in the limits of
  - LOW PRESSURE
  - HIGH TEMPERATURE
- Explain this in terms of previous properties?

Applicability of Ideal Gas Law

- OK for most gases at ordinary T, P
- MUST fail at some point, since solids and liquids exist.
- Deviations at high P, low T.
Using Gas Law

• Change in conditions
  ➢ “Before” and “after” eq’s, collect constant terms.

Example

• A bulb is filled with 760 torr of CH₄ at 25°C
• The bulb may burst if the pressure exceeds 2 atm
• If the bulb is heated, at what temperature would the pressure reach 2 atm?