4.1 An Overview of the Physical States of Matter

- Most substances can *exist as a solid, a liquid or a gas* under appropriate conditions of pressure and temperature
- A gas adopts the shape of the container and fills the container, because its particles are far apart and move randomly.
- A liquid adopts the shape of the container to the extent of the containers volume, because its particles are close together but free to move around each other.
- A solid has a fixed shape regardless of the shape of the container, because its particles are close together and held rigidly in place.

Several aspects distinguish gases from liquids and solids:

1. *The volume of a gas changes significantly with pressure:*
    - Gases under pressure do a lot of work
    - When in a container, *increasing* the external force hitting the container will *decrease* the volume of the gas
2. *The volume of a gas changes significantly with temperature:*
    - When gas is heated, it expands; when it is cooled, it contracts/shrinks
3. *Gases flow very freely:*
    - This behavior allows gases to be transported more easily through pipes, but it also means that they leak more rapidly
4. *Gases have relatively low densities:*
    - The density of a gas is usually measured in units of grams per litre (g/L)
    - Solids and liquids are usually grams per millilitre (g/mL)
5. *A gas forms a solution in any proportions:*
    - Air is a solution of 18 gases

4.2 Gas Pressure and Its Measurement:

- **Pressure (p):** is defined as the force exerted per unit of surface area:
    - **Pressure = force / area**

Laboratory Devices for Measuring Gas Pressure:

- **Barometer** is used to measure atmospheric pressure
- (Pressure *decreases* with altitude)
- **Manometers** are devices that are used to measure the pressure of a gas in an experiment.

Units of Pressure:

- Pressure results from a force exerted on an area.
- Pressure is the **Pascal (Pa)** unit; equals a force of one Newton exerted on an area of one square metre:
    - 1 Pa = 1 N/m^2
- A larger unit is the **Standard Atmosphere (atm)**, the average atmospheric pressure measured at sea level at 0° C
    - 1 atm = 101.325 kPa (kilopascals) = $1.01325 * 10^5$ Pa
- Another unit is **millimetre of mercury (mmHg) –** renamed to the **torr (Torr)**
    - 1 Torr = 1 mmHg = $\frac{1}{760}$atm = $\frac{101.325}{760}$kPa = 133.322 Pa
- The **bar** is currently used for defining **standard pressure** and when describing thermodynamic standards:
    - 1 bar = $1 * 10^2$ kPa = $1 * 10^5$ Pa

4.3 The Gas Laws and Their Experimental Foundations:

- The individual gas laws are special cases of unifying relationship called the *ideal gas law*, which quantitatively describes the behavior of an **ideal gas**;
    - A gas that exhibits linear relationships among volume, pressure, temperature, and amount.
    - No *deal gas actually exists*, most simple gases such as Nitrogen Gas, Oxygen Gas, Hydrogen Gas and Noble gases, behave nearly ideally at ordinary temperatures and pressures

The Relationship Between Volume and Pressure: Boyle's Law:

- The product of corresponding p and V values is a constant
- V is *inversely* proportional to P
- V is *directly* proportional to P to 1/p, and plot of V versus 1/p is linear. This *linear relationship between two gas variables* is a hallmark of ideal gas behavior
- **Boyle's Law:** *at constant temperature, the volume occupied by a fixed amount of gas is* ***inversely*** *proportional to the applied (external) pressure*
- **pV = constant**
- If one *increases,* the other *decreases*
- **Charles's Law:** *at constant pressure, the volume occupied by a fixed amount of gas is* ***directly*** *proportional to its absolute (Kelvin) temperature*
- Absolute Zero = (0 K or -273.15 C)
- At *constant volume, the pressure exerted by a fixed amount of gas is directly proportional to the absolute temperature.* In other words, as temperature goes up, so does pressure.

Gas behavior at Standard Conditions:

- To better understand the factors that influence gas behavior, chemists have assigned a baseline of *standard conditions* called **standard temperature and pressure (STP)**
    - STP – 0 C (273.15 K) and 1 bar (10^5 Pa)
- Under these conditions, the volume of 1 mol of an ideal gas is called the **standard molar volume:**
    - Standard molar volume = 22.710953(21) L or 22.7 K (3 sf)

Ideal Gas Law:

- Boyle's Law focuses on pressure (V is prop. to 1/p)
- Charles's Law focuses on temperature (V is prop. to T)
- Avogadro's Law focuses on the amount (mol) of gas (V is prop to n)
- Combining these, you get the **Ideal Gas Law**
    - pV = nRT
    - R is the universal gas constant
        - 8.312 J/(mol*k)

Approaching Ideal Gas Law Problems:

- Summarize the changing gas variables – known and unknown – and the variables held constant
- Convert units, if necessary
- Rearrange the ideal gas law to obtain the needed relationship of variables, and solve for the unknown

4.4 Rearrangements of the Ideal Gas Law:

The Density of a Gas:

- m/V = d = (M * p)/RT

The Molar Mass of a Gas:

- We can determine the molar mass of an unknown gas of a volatile liquid (liquid easily vaporized)
    - n= m/M; M = mRT/pV

Dalton's Law of Partial Pressures:

- When water vapor is added to dry air, the total air pressure increases by the pressure of the water vapor:
- He concluded that each gas is in the mixture exerts a **partial pressure** equal to the pressure it would exert by *itself*. This can be stated as **Dalton's Law of Partial Pressures:** *in a mixture of unreacting gases, th*e total pressure is the sum of the partial pressures of the individual gases:
    - Ptotal = P1 + P2 + P3 + …
- Each component in mixture contributes a fraction of the total number of moles in the mixture; this portion is the **mole fraction (X)** of that component.

4.5 The Kinetic-Molecular Theory: A Model for Gas Behavior:

Postulates of the Kinetic-Molecular Theory:

- *Particle Volume:* the concept of ideality relies largely on the necessity that the particles move about independently without exerting any forces upon one another.
- *Particle Motion:* The particles are in constant, random, straight-line motion, expect when they collide with the container walls or with each other.
- *Particle Collisions:* the collision are *elastic,* the colliding molecules exchange energy but do not lose energy but do not lose any energy through friction. Thus *their total kinetic energy (E_k) is constant*. Between collisions, the molecules do not influence each other by attractive or repulsive forces.
- $(E_k) = (mv^2) / 2$

The Central Importance of Kinetic Energy:

- *Implications of Avogadro's Law*
- *The meaning of Temperature:*
  - Using the definitions of velocity, momentum, force and pressure, we can also express this relationship by the following equation:
    - $(E_k) = 3[(R / N_A)(T)] / 2$
    - R is the gas constant and Na is Avogadro's number
  - This equation is an essential point that *temperature is a* ***measure*** *of the average kinetic energy of the particles*

Root-Mean-Square Speed:

- Deriving an expression for the speed of a gas particle that has the average kinetic energy of the particles in the sample
- **rms speed:** *a particle moving at this speed has the average kinetic energy*
  - $\mathbf{u_{rms}} = \sqrt{\frac{3RM}{\mathcal{M}}}$

Effusion and Diffusion:

- The movement of a gas into a vacuum and the movement of gases through one another an phenomena with some vital applications
- **Graham's Law of effusion:** *the rate of effusion of a gas is inversely proportional to the square root of its molar mass.*
- $\sqrt{\frac{1}{\mathcal{M}}}$

The process of Diffusion:

- **Diffusion:** is the movement of one gas through another.
- **Graham's Law** is the same here.
- Lighter molecules have higher average speeds than heavier molecules, so they move farther in a given time.

4.6 Non-ideal Gases: Deviations form Ideal Behavior

- *Gas particles are* ***not*** *points of mass,* but have volumes determined by the sizes of their atoms and the lengths and directions of their bonds.
- *Attractive and repulsive forces* ***do*** *exist among gas particles* because atoms contain charged subatomic particles and many bonds are polar

The van der Waals Equation: Adjusting the Ideal Gas Law

- The **van der Waals equation** for *n* moles of a non-ideal gas is
    - $[p + (n^2a/V^2)] \times (V - nb) = nRT$
- *a, b* are **van der Waals** constants