Unit 4: States of Matter, Solutions and Gases

Chapter 10: States of Matter

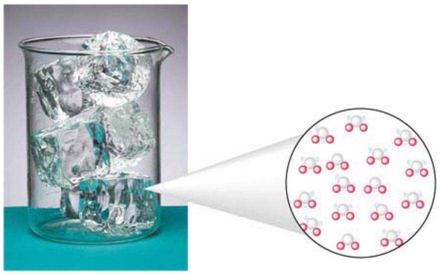
10.1: The Nature of Gases

Solids: - the state of matter where it has a definite volume with a constant shape.

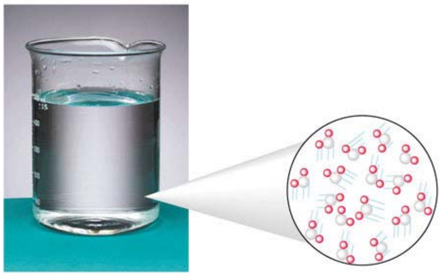
Liquid: - the state of matter where it has a definite volume but an indefinite shape.

Gas: - the state of matter where it has an indefinite volume and shape (compressible).

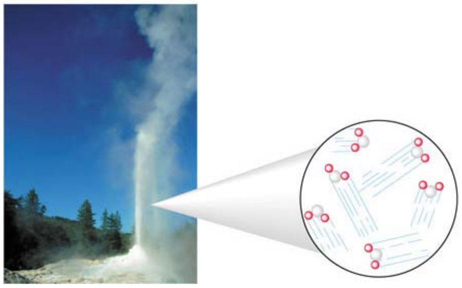
Kinetic Molecular Theory and States of Matter



Solids have particles are in fixed positions.

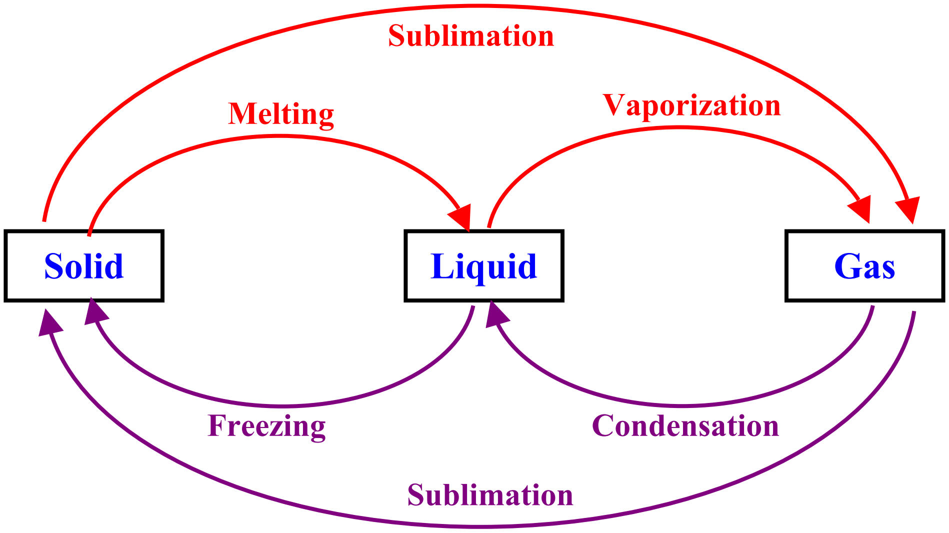


Liquids have particles that can "rolled" past one another.



Gases have particles that have truly random motion and have very weak interactions between molecules.

Phase Change



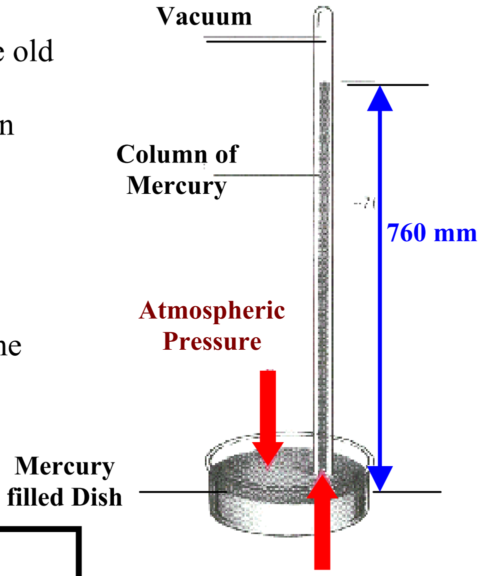
Sublimation: - the phase change from solid to gas directly or vice versa.

Example: Dry ice (Solid Carbon Dioxide) sublimates from solid to gas directly, skipping the liquid phase.

Pressure: - the amount of force per unit of area [1 N/m^2 = 1 Pascal (Pa)], measured in **kiloPascal (kPa), mm Hg**, **torr**, **atmospheric pressure (atm)**.
- in a pressurized container, a pressure can be felt as the particles pushed on the inside wall of the container.

Barometer: - an instrument (invented by Evangelista Torricelli, hence, the old unit of pressure *torr* was named after him) to measure atmospheric pressure using a dish filled with mercury and an inverted glass tube filled with mercury.

- when the glass tube is placed into the dish filled with mercury, the level of mercury in the tube will drop. However, the atmospheric pressure of the surface of mercury dish will keep the mercury column at a certain level.
- the higher the atmospheric pressure, the higher the mercury is in the column.
- at normal atmospheric pressure at sea level, the mercury column is at 760 mm.



Units of Pressure

1 standard atmosphere (atm) = 760 mm Hg = 760 torr = 101.325 kPa

Note: 1mm Hg = 1 torr

Example 1: Convert 525 mm Hg to atm and kPa.

$$525\,\cancel{\text{mm Hg}} \times \frac{1\text{ atm}}{760\,\cancel{\text{mm Hg}}} = \boxed{0.691\text{ atm}} \qquad 525\,\cancel{\text{mm Hg}} \times \frac{101.325\text{ kPa}}{760\,\cancel{\text{mm Hg}}} = \boxed{70.0\text{ kPa}}$$

Example 2: Convert 350 kPa to atm, and mm Hg.

$$350\,\cancel{\text{kPa}} \times \frac{1\text{ atm}}{101.325\,\cancel{\text{kPa}}} = \boxed{3.45\text{ atm}} \qquad 350\,\cancel{\text{kPa}} \times \frac{760\text{ mm Hg}}{101.325\,\cancel{\text{kPa}}} = 2625.2\text{ mm Hg} = \boxed{2.62 \times 10^3\text{ mm Hg}}$$

Gaseous Pressure and Temperature in Kinetic Molecular Theory

1. **The smaller the volume, the higher the pressure**. This is there is less space for the molecules to move about in the container. Hence, each molecule will exert more force on the wall of the container, causing a higher pressure.
2. **The higher the temperature, the higher the pressure**. When the temperature is high, it means molecules move faster and have a higher average kinetic energy. With more force they can exert on the wall of the container, the pressure increases.

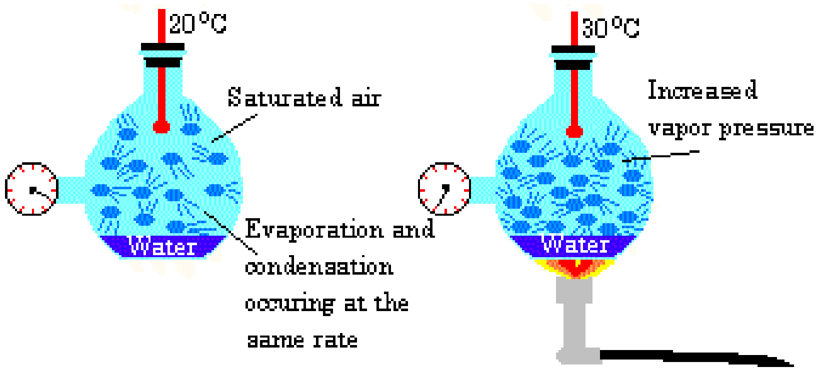
10.2: The Nature of Liquids

Vaporization: - the physical process of changing a pure substance from a liquid phase to a gaseous phase.

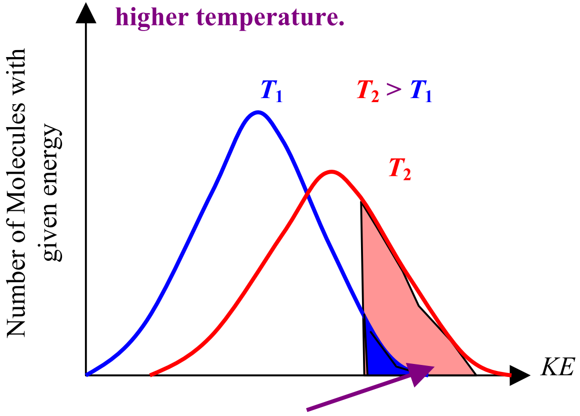
Evaporation: - vaporization below the boiling point.
- occurs because some molecules have attained sufficient energy to leave the liquid surface.
- as the temperature increases, more molecules will have the threshold energy to leave the surface of the liquid.
- when evaporation occurs, there is a cooling effect because it takes energy from the surrounding for molecules to move fast enough to escape from the liquid's surface.

Vapour Pressure: - the pressure existed above a liquid when its rate of evaporation is the same as the rate of its condensation.
- sometimes refers to as **equilibrium vapour pressure** because **equilibrium means the same rate of a two-way process**.
- in general, for a particular compound, **as temperature increases**, **the vapour pressure increases**.
- This is because the **higher the temperature**, **kinetic energy increases** and **more lquid molecules will overcome** the **intermolcular forces (force of attraction between molecules)** to become a gas. Therby, **increasing vapour presuure**.

Temperature ↑, Vapour Pressure ↑



Note that more molecules can overcome the intermolecular forces needed to be in gaseous state at a higher temperature.



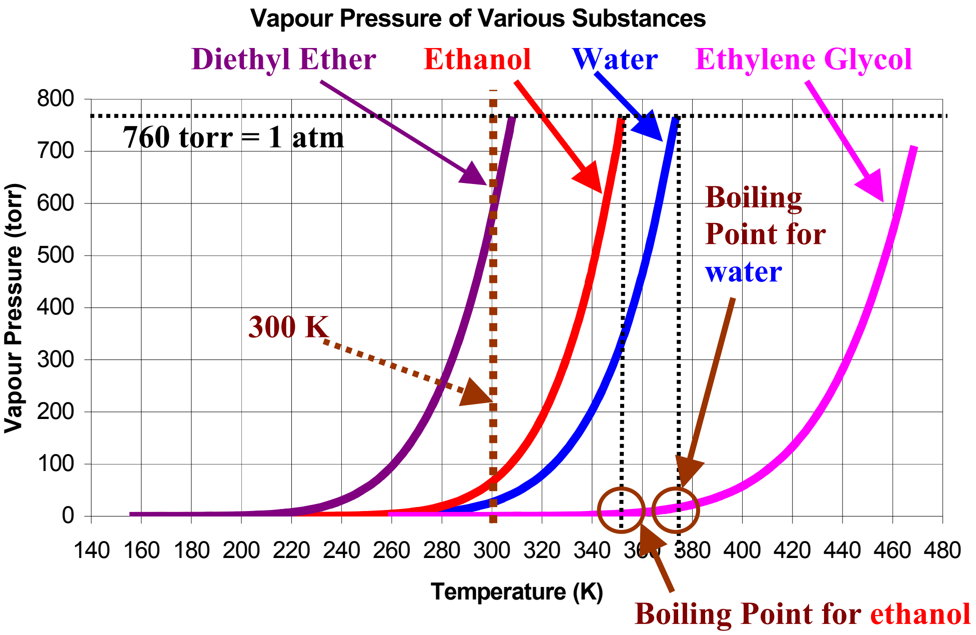
Energy required to overcome intermolecular forces.

Boiling: - vaporization at the boiling point.

Boiling Point (Bp): - the temperature at which the vapour pressure equals to the outside external pressure (like atmospheric pressure at sea level).
- **the lower the atmospheric pressure, the lower the boiling point**. This is because, there is less external air molecules pushing down on the surface of the liquid. Hence, the molecules under the liquid surface find it easier to attain that threshold energy to escape.

Normal Boiling Point: - the boiling point of a substance at 1 atm or 101.325 kPa (at sea level altitude) .

We can use this graph to find the boiling points of various substances by moving across from a desired vapour pressure (which is equal to external air pressure at boiling). The temperature corresponds to that particular pressure for a certain substance is its boiling point at the air pressure.



Note: For substances, when P_{vap} = 1 atm = 760 torr, the corresponding temperature represents the boiling point.

10.4: Changes of State

Phase Diagrm: - a digaram with axes of Presuure versus Temperature to illustrate all three phases of a substance.
- it is more useful than a heating curve because boiling and melting points are pressure dependent, and the user now can see where the melting and boiling point are at any pressure. The curve of the vapour pressure versus temperature of a liquid is only a line on the phase diagram.
- it allows the user to identify the phase of a substance at a certain pressure and temperature.

Solid-Liquid Line: - a line that outlines the temperatures and the corresponding pressures where solid and liquid phases coexist.
- used to find the **melting point** of a substance at any given pressure.

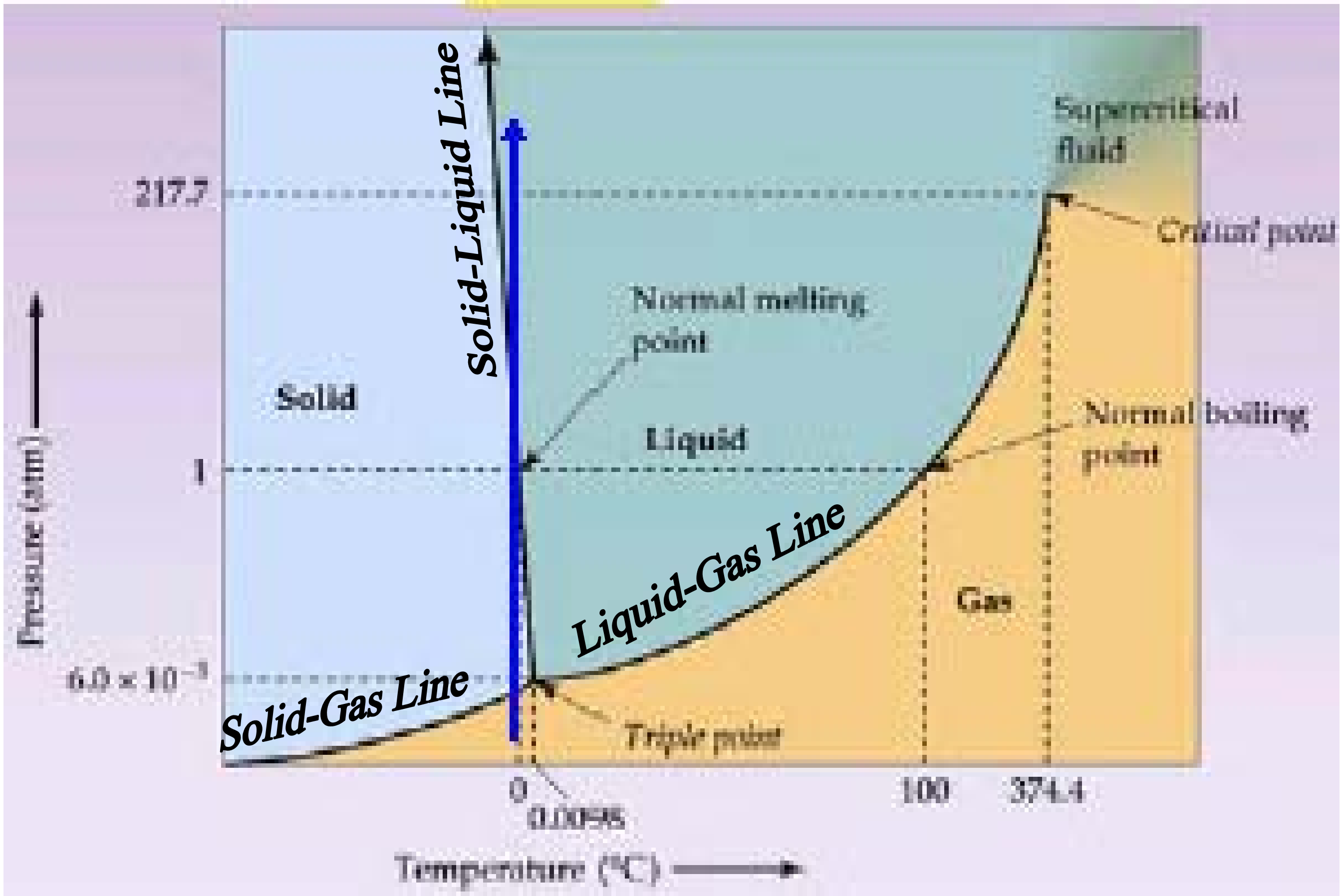
Liquid-Gas Line: - a line that outlines the temperatures and the corresponding pressures where liquid and gas phases coexist.
- used to find the **boiling point** of a substance at any given pressure.

Solid-Gas Line: - a line that outlines the temperatures and the corresponding pressures where solid and gas phases coexist.
- used to find the **sublimation point** of a substance at any given pressure.

Triple Point: - a point indicating the pressure and temperature where all solid-liquid, liquid solid, and solid-gas line meet.
- this is the **pressure and temepertaure conditions where all three phases can coexist**.

Critical Point: - a point at the end of the liquid-gas line which indicates an intermediate "fluid" region where liquid can be coverted to gas instantaneously without waiting for phase change .
- the temperature at critical point is called critical temperature and the pressure at critical point is called critical pressure.
- at this region (beyond the critical pressure and temperature), the liquid will become a gas withoutout going through a change of state.

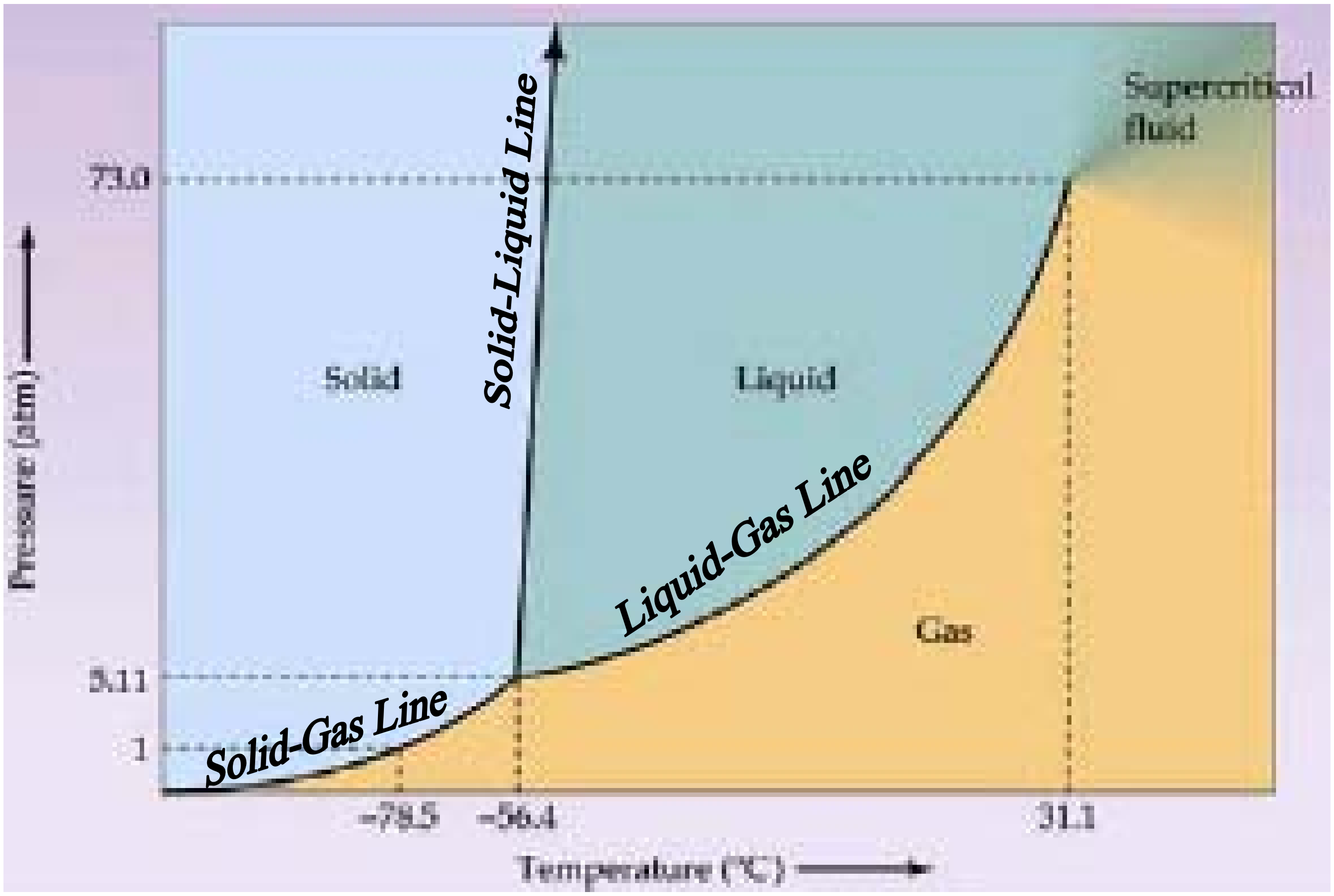
Phase Diagram of Water



From the phase diagram of water:

- at 1 atm, water has a normal melting point of 0°C and a normal boiling point at 100°C.
- there is no normal sublimation point because the solid-gas line ends well under 1 atm.
- the negative slope of the solid-liquid line indicates that as pressure increases, the melting point of water decreases.
- the positive slope of the liquid-gas line illustates that as pressure increases, the boiling point of water also increases.
- the positive slope of the solid-gas line shows that as pressure increases, the sublimation point of water increases as well.
- the triple point, where all three phases can coexist, is at 0.0098°C and 6.0×10^{-3} atm (4.56 torr).
- the critical point of water is at 374.4°C and 217.7 atm. Beyond this critical temperature and pressure, water will become gaseous instantenously.

Phase Diagram of Carbon Dioxide



From the Phase Diagram of Carbon Dioxide:

- at 1 atm, carbon dioxide has a normal sublimation point of −78.5°C.
- there is no normal melting and boiling points because carbon dioxide's solid-gas line is well above normal atmospheric condition of 1 atm.
- the positive slope of the solid-liquid line indicates that as pressure increases, the melting point of carbon dioxide increases.
- the positive slope of the liquid-gas line illustates that as pressure increases, the boiling point of carbon dioxide also increases.
- the positive slope of the solid-gas line shows that as pressure increases, the sublimation point of carbon dioxide increases as well.
- the triple point, where all three phases can coexist, is at −56.4°C and 5.11 atm (518 kPa or 3884 torr).
- the critical point of carbon dioxide is at 31.1°C and 73 atm. Beyond this critical temperature and pressure, liquid carbon dioxide will become gaseous instantaneously.

Assignment

10.1 pg. 271 #1, 2; pg. 272 #3 to 7
10.2 pg. 279 #8 to 13
10.4 pg. 286 #17 to 19
Ch 10 Review: pg. 289 #20 to 42, 46 and 47

Chapter 12: The Behaviour of Gases

12.1: The Properties of Gases

Properties of Gases:

1. **Compressibility**: - a main property of gas where the amount of volume can decrease under increase pressure.

2. **Lack of Particle Interaction**: - Unlike liquids and solids, gas particles have no attractive and repulsive forces between them as assumed in the kinetic theory.

3. **Rapid and Constant Motion**: - gas particles move in a straight line and independent of each other.
   - at collisions, gas particles are completely **elastic** (total kinetic energy remains constant before and after the collision).

Variables to Describe a Gas:

1. **Pressure** (**P**): - the amount of force per unit of area, measures in **kiloPascal (kPa)**.
   - in a pressurized container, a pressure can be felt as the particles pushed on the inside wall of the container.

2. **Volume** (**V**): - the amount of space the gas is occupied; measures in **Litre** (**L**).

3. **Temperature** (**T**): - the average of kinetic energy of the gas; measures in **Kelvin (K)**.

4. **Moles** (**n**): - the amount of gas particle in a closed system; measures in **moles (mol)**.

12.2: Factors Affecting Gas Pressure

1. **Amount of Gas**: as amount of gas particles increases, there are more particles pushing on the inside wall of the container, causing an increase in pressure.

   Amount of Gas ↑ **Pressure ↑**

2. **Volume**: - as the volume decreases, there is less room for the gas particles to move about. This causes the particles to push "harder" onto the inside wall of the container, causing an increase in pressure.

   Volume ↑ **Pressure ↓**

3. **Temperature**: - as the temperature increases, gas particles move faster and push harder on the inside wall of the container, causing an increase in pressure.

   Temperature ↑ **Pressure ↑**

Assignment
12.1 pg. 328 #1 to 4
12.2 pg. 332 #5 to 9

12.3: The Gas Laws

1. **Boyle's Law**: - at a constant temperature, **pressure is inversely proportional to the volume**.

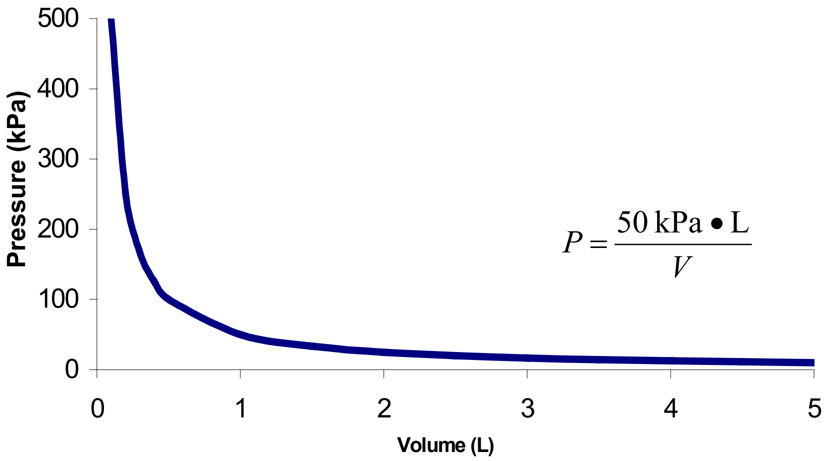
$$P \propto \frac{1}{V}$$

$$P = \frac{k}{V} \quad \text{or} \quad PV = k$$

k = constant of proportion

Volume ↑ Pressure ↓

Boyle's Law (Constant Temperature)



Boyle's Law

$$P_1V_1 = P_2V_2$$

P_1 = Pressure of Initial Condition
V_1 = Volume of Initial Condition
P_2 = Pressure of Final Condition
V_2 = Volume of Final Condition

Example 1: A gas cylinder changed its volume from 2.50 L to 6.25 L. If it were at 101.325 kPa initially, what would be its final pressure?

$P_1 = 101.325$ kPa
$P_2 = ?$
$V_1 = 2.50$ L
$V_2 = 6.25$ L

$$P_1V_1 = P_2V_2$$

$$\frac{P_1V_1}{V_2} = P_2$$

$$P_2 = \frac{(101.325\ \text{kPa})(2.50\ \cancel{\text{L}})}{(6.25\ \cancel{\text{L}})}$$

$$\boxed{P_2 = 40.5\ \text{kPa}}$$

As Volume ↑, Pressure ↓

2. **Charles's Law**: - at constant pressure, **volume is directly proportional to the temperature**.

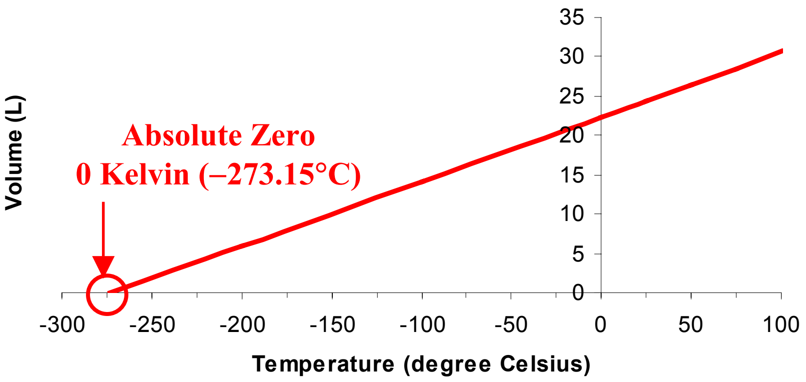
$$V \propto T$$

$$V = kT \quad \text{or} \quad \frac{V}{T} = k$$

k = constant of proportion

Temperature ↑ Volume ↑

Charles's Law (Constant Pressure)



Charles's Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

T_1 = Temperature of Initial Condition in Kelvin
T_2 = Temperature of Final Condition in Kelvin
V_1 = Volume of Initial Condition
V_2 = Volume of Final Condition

Example 2: A balloon is has a volume of 3.25 L at 25.0°C. Determine the volume of the same balloon when the temperature is dropped to 5.00°C.

$V_1 = 3.25\text{ L}$
$T_1 = 25.0°\text{C} = 298.15\text{ K}$
$V_2 = ?$
$T_2 = 5.00°\text{C} = 278.15\text{ K}$

(Change °C to K)

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{V_1 T_2}{T_1} = V_2$$

$$V_2 = \frac{(3.25\text{ L})(278.15\,\cancel{\text{K}})}{(298.15\,\cancel{\text{K}})}$$

As Temp ↓, Volume ↓

$$\boxed{V_2 = 3.03\text{ L}}$$

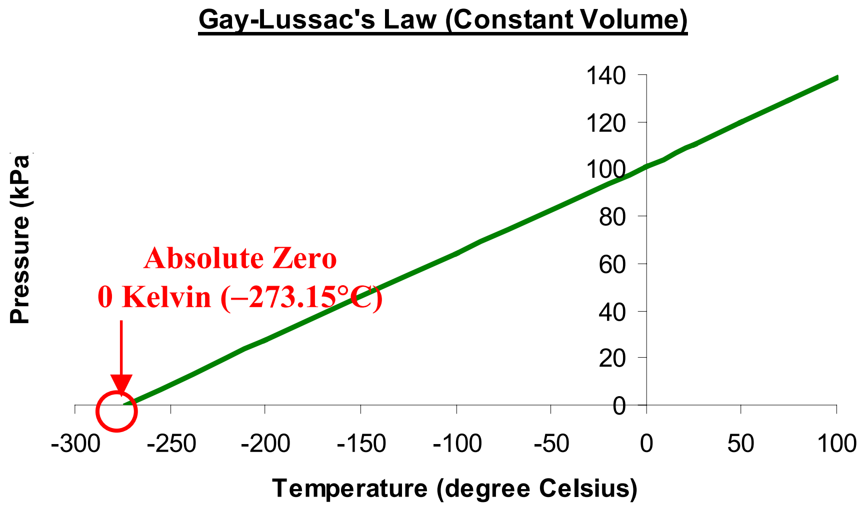
3. **Gay-Lussac's Law**: - at constant volume, **pressure is directly proportional to the temperature**.

$$P \propto T$$

$$P = kT \quad \text{or} \quad \frac{P}{T} = k$$

k = constant of proportion

Temperature ↑ Pressure ↑



Gay-Lussac's Law

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

T_1 = Temperature of Initial Condition in Kelvin
T_2 = Temperature of Final Condition in Kelvin
P_1 = Pressure of Initial Condition
P_2 = Pressure of Final Condition

Example 3: A canister is has a pressure of 8.00 atm at 15.0°C. Calculate its pressure if the temperature was to increase to 100.0°C. **(Don't do this at home!)**

$P_1 = 8.00$ atm
$T_1 = 15.0°\text{C} = 288.15$ K
$P_2 = ?$
$T_2 = 100.0°\text{C} = 373.15$ K

(Change °C to K)
(*P* can be in atm)

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{P_1 T_2}{T_1} = P_2$$

$$P_2 = \frac{(8.00\text{ atm})(373.15\,\cancel{\text{K}})}{(288.15\,\cancel{\text{K}})}$$

As Temp ↑, Pressure ↑

$$\boxed{P_2 = 10.4\text{ atm}}$$

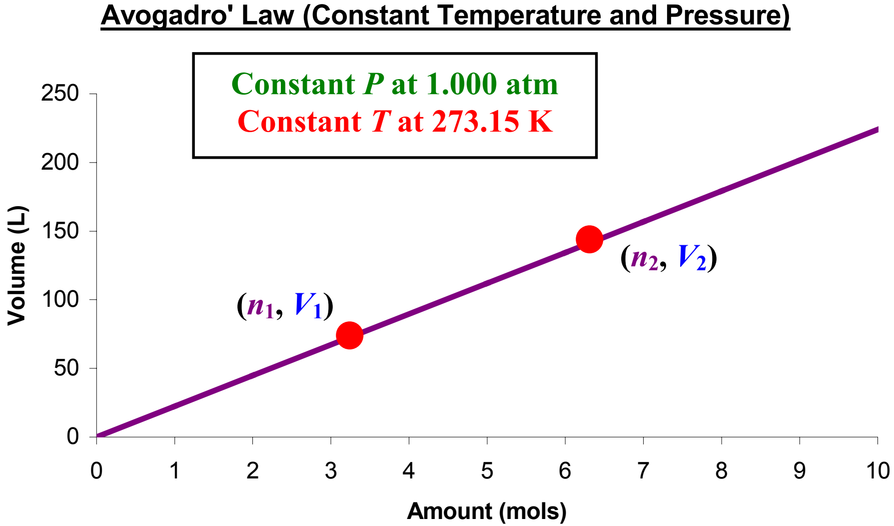
4. **Avogadro's Law**: - at constant pressure and temperature, **volume is directly proportional to amount of moles of gas present**.

$V \propto n$

$V = kn$ or $\frac{V}{n} = k$

k = constant of proportion

Moles ↑ Volume ↑



Avogadro's Law

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

n_1 = Amount of Moles at Initial Condition
n_2 = Amount of Moles at Final Condition
V_1 = Volume at Initial Condition
V_2 = Volume at Final Condition

Example 4: An excess amount of solid phosphorus is reacted with 9.00 mol of chlorine gas at 223.2 L to produce phosphorus trichloride gas under constant temperature and pressure.

a. Write a balanced equation for this reaction.
b. Assuming a complete reaction, what is the volume of phosphorus trichloride produced?

a. $P_{4\,(s)} + 6\,Cl_{2\,(g)} \rightarrow 4\,PCl_{3\,(g)}$

b.
$n_1 = 9.00$ mol (Cl_2)
$V_1 = 223.2$ L

$$n_2 = 9.00\,\cancel{\text{mol Cl}_2} \times \frac{4\text{ mol PCl}_3}{6\,\cancel{\text{mol Cl}_2}}$$

$n_2 = 6.00$ mol (PCl_3)
$V_2 = ?$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

$$\frac{n_2 V_1}{n_1} = V_2$$

$$V_2 = \frac{(6.00\,\cancel{\text{mol}})(223.2\text{ L})}{(9.00\,\cancel{\text{mol}})}$$

As Moles ↓, Volume ↓

$$\boxed{V_2 = 149\text{ L}}$$

5. **Combined Gas Law**: - a formula that summarizes Boyle's Charles's and Guy-Lussac's Gas Laws.
- allow the user of the formula to determine the change in conditions of the same amount of gas.

Combined Gas Law

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

P_1 = Pressure of Initial Condition
V_1 = Volume of Initial Condition
T_1 = Temperature of Initial Condition in Kelvin

P_2 = Pressure of Final Condition
V_2 = Volume of Final Condition
T_2 = Temperature of Final Condition in Kelvin

Example 5: A high altitude weather balloon has a volume of 57.2 L at SATP. Determine its volume at its maximum height where the pressure is 37.0 kPa and the temperature is –45.0°C.

V_1 = 57.2 L

At SATP:
P_1 = 100 kPa
T_1 = 25.0°C = 298.15 K

V_2 = ?
P_2 = 37.0 kPa
T_2 = –45.0°C = 228.15 K

(Change °C to K)

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{P_1V_1T_2}{T_1P_2} = V_2$$

$$V_2 = \frac{(100\ \cancel{\text{kPa}})(57.2\text{ L})(228.15\ \cancel{\text{K}})}{(298.15\ \cancel{\text{K}})(37.0\ \cancel{\text{kPa}})} = 118.2986978\text{ L}$$

$$\boxed{V_2 = 118\text{ L}}$$

Because *P* ↓ much more than *T* ↓, the final Volume ↑.

12.4: Ideal Gases

Ideal Gas Law: - a formula that relates pressure, volume, amount, and temperature of an **ideal gas** (gaseous volume does not account for total particles volumes) at one specific condition.

Ideal Gas Law

$$PV = nRT$$

P = Pressure (kPa)
V = Volume (L)
n = Amount of Gas (mol)
T = Temperature (K)
R = Gas Constant = 8.314 (L • kPa)/(K • mol)

Example 1: Determine the mass of propane if it is in a 200 L container at 15.0°C and at 32.0 atm.

$V = 200\ \text{L}$

$T = 15.0°\text{C} = 288.15\ \text{K}$

$$P = 32.0\ \cancel{\text{atm}} \times \frac{101.325\ \text{kPa}}{1\ \cancel{\text{atm}}}$$

$P = 3242.4\ \text{kPa}$

[need to change atm to kPa because R is in (L • **kPa**)/(K • mol)]

$R = 8.314$ (L • kPa)/(K • mol)

$\boldsymbol{m = ?}$

n = ? (need to find n first)

$$\boldsymbol{PV = nRT}$$

$$\frac{PV}{RT} = n$$

$$n = \frac{(3242.4\ \cancel{\text{kPa}})(200\ \cancel{\text{L}})}{\left(8.314\ \frac{\cancel{\text{L}} \bullet \cancel{\text{kPa}}}{\cancel{\text{K}} \bullet \text{mol}}\right)(288.15\ \cancel{\text{K}})}$$

$\boldsymbol{n = 270.670642\ \text{mol}}$

For propane, C_3H_8, $M = 44.11$ g/mol

$\boldsymbol{m = nM}$

$m = (270.670642\ \cancel{\text{mol}})(44.11\ \text{g}/\cancel{\text{mol}})$

$m = 11939.28202\ \text{g}$

$$\boxed{\boldsymbol{m = 1.19 \times 10^4\ \text{g} = 11.9\ \text{kg}}}$$

Example 2: It is known that air is less dense at higher altitude. Calculate the density of pure oxygen gas near the top of Mount Everest where the temperature is –30.0°C and 31.0 kPa. (By comparison, pure oxygen has a density of 1.29 g/L at SATP.)

$T = -30.0°\text{C} = 243.15\ \text{K}$

$P = 31.0\ \text{kPa}$

$R = 8.314$ (L • kPa)/(K • mol)

$M = 32.00$ g/mol

$$\boldsymbol{D} = \frac{m}{V} = ?$$

Substitute $\frac{m}{M}$ for n.

Then, solve for $\frac{m}{V}$.

$$\boldsymbol{PV = nRT}$$

$$PV = \frac{m}{M}RT$$

$$\frac{PM}{RT} = \frac{m}{V}$$

$$\boldsymbol{D} = \frac{PM}{RT} = \frac{(31.0\ \cancel{\text{kPa}})(32.0\ \text{g}/\cancel{\text{mol}})}{\left(8.314\ \frac{\text{L} \bullet \cancel{\text{kPa}}}{\cancel{\text{K}} \bullet \cancel{\text{mol}}}\right)(243.15\ \cancel{\text{K}})} = 0.4907127905\ \text{g/L}$$

$$\boxed{\boldsymbol{D = 0.491\ \text{g/L}}}$$

Gaseous Stoichiometry Procedure (Ideal Gas)

1. Predict the products and balance the chemical equation.
2. Put all the information given under the appropriate chemicals.
3. Find the moles of the given chemical $\left(\boldsymbol{n = \frac{PV}{RT}}\right)$.
4. Find the mole of the required chemical using mole ratio.
   $$\left(\textbf{mol of require} = \textbf{mol of given} \times \frac{\textbf{require coefficient}}{\textbf{given coefficient}}\right)$$
5. Convert mole of the required chemical to its volume equivalence **($PV = nRT$)**

Example 3: Ammonia is reacted with oxygen to form nitrogen monoxide and water vapour.

$$4\ NH_{3\ (g)} + 5\ O_{2\ (g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$$

If 50.0 L of ammonia at 90.0 kPa at 25.0°C were allowed to react with excess oxygen, what would be the pressure of nitrogen monoxide in a collector vessel measuring 30.0 L at a temperature of 10.0°C?

$4\ NH_{3\ (g)}$	+	$5\ O_{2\ (g)}$	→	$4\ NO_{(g)}$	+	$6\ H_2O_{(g)}$
50.0 L				30.0 L		
25.0°C = 298.15 K				10.0°C = 283.15 K		
90.0 kPa				? kPa		

① $n_{NH_3} = \dfrac{PV}{RT} = \dfrac{(90.0\ \cancel{kPa})(50.0\ \cancel{L})}{\left(8.314\ \frac{\cancel{kPa \bullet L}}{mol \bullet \cancel{K}}\right)(298.15\ \cancel{K})} = 1.815380558\ mol$

② $n_{NO} = 1.815380558\ \cancel{mol\ NH_3} \times \dfrac{4\ mol\ NO}{4\ \cancel{mol\ NH_3}} = 1.815380558\ mol\ NO$

③ $P_{NO} = \dfrac{nRT}{V} = \dfrac{(1.815380558\ \cancel{mol})\left(8.314\ \frac{kPa \bullet \cancel{L}}{\cancel{mol \bullet K}}\right)(283.15\ \cancel{K})}{30.0\ \cancel{L}} = 142.453463\ kPa$ $\boxed{P_{NO} = 142\ kPa}$

Example 4: If 15.25 L of hydrogen at STP is reacted with 7.85 L of nitrogen at 98.0 kPa and at −18.6°C, determine the volume of ammonia formed at SATP (24.8 L/mol).

$3\ H_{2\ (g)}$	+	$N_{2\ (g)}$	→	$2\ NH_{3\ (g)}$
15.25 L		7.85 L; 98.0 kPa		? L
STP = 22.4 L/mol		−18.6°C = 254.55 K		SATP = 24.8 L/mol
		R = 8.31 (L • kPa) / (K • mol)		

Since there is enough information to determine the moles of two reactants, we need to determine which one is the limiting reagent.

① $n_{H_2} = 15.25\ \cancel{L} \times \dfrac{1\ mol}{22.4\ \cancel{L}} = 0.6808035714\ mol\ H_2$

② $n_{N_2} = \dfrac{PV}{RT} = \dfrac{(98.0\ \cancel{kPa})(7.85\ \cancel{L})}{\left(8.31\ \frac{\cancel{L} \bullet \cancel{kPa}}{\cancel{K} \bullet mol}\right)(254.55\ \cancel{K})} = 0.363681833\ mol\ N_2$

Let's assume N_2 is the limiting reagent. Calculate the mol H_2 actually needed.

③ $n_{H_2} = 0.363681833\ \cancel{mol\ N_2} \times \dfrac{3\ mol\ H_2}{1\ \cancel{mol\ N_2}} = 1.091045499\ mol\ H_2$ needed

But we don't have 1.091045499 mol of H_2, we only have 0.6808035714 mol of H_2. Therefore, H_2 is the limiting reagent. (*Note: the limiting reagent is NOT always the chemical with the smaller number of moles. You have to always compare like we did above.*)

Now, we calculate the moles of NH_3 formed by using moles of limiting reagent, H_2.

④ $n_{NH_3} = 0.6808035714\ \cancel{mol\ H_2} \times \dfrac{2\ mol\ NH_3}{3\ \cancel{mol\ H_2}} = 0.4538690476\ mol\ NH_3$

Finally, we determine the volume of NH_3 produced.

⑤ $V_{NH_3} = (0.4538690476\ \cancel{mol}\ NH_3)(24.8\ L/\cancel{mol}) = 11.25595\ldots\ L$ $\boxed{V_{NH_3} = 11.3\ L}$

Real Gas: - a gas where its properties deviate from an ideal gas.

- **when we account for the fact that gaseous particles have volume, and they actually have attraction between particles (intermolecular bonds), the ideal gas law become less accurate**.
- This ***happens mainly when pressure is extremely high, volume is small and temperature is low***. At such conditions, the small volume will make the particle size an important matter. This is the same for attraction forces between particles because they are moving a lot slower.

Assignment

12.3 pg. 335 #10, 11; pg. 337 #12, 13; pg. 338 #14; pg. 339 #15; pg. 340 #16 to 21
12.4 pg. 342 #22, 23; pg. 343 #24, 25; pg. 346 #26 to 30
Ch 12 Review: pg. 356 #45 to 63 and Gas Stoichiometry Worksheet

Gas Stoichiometry Worksheet

1. A 25.0 L propane tank at 500.0 kPa and 30.0°C is completely used up during a barbecue cookout. Calculate the volume of carbon dioxide when it reached the upper atmosphere where the pressure is at 50.0 kPa and the temperature is at −20.0°C.

2. 60.0 kL of hydrogen at SATP (24.8 L/mol) is reacted with excess amount of nitrogen. The resulting ammonia gas needs to be stored in a gas tank with a volume of 6.25 kL and under 2000 kPa of pressure. What should the temperature regulator of the gas tank be set at?

3. Nitrogen and oxygen in the air regularly react within an automobile engine to form nitrogen monoxide. If 40.0 L of oxygen at 300 kPa is reacted inside a 200°C engine, calculate the volume of smog produced at STP in the exhaust.

4. When nitrogen monoxide reacts with oxygen in the atmosphere under sunlight, it forms photochemical smog (nitrogen dioxide – a brown colour gas). If on an average day, 400,000 ML of nitrogen monoxide at SATP – 24.8 L/mol is produced in a large city during rush hour; determine the volume of smog produced when it reached a higher altitude with the pressure at 80.0 kPa and the temperature at −15.0°C.

5. A 15.0 g piece of zinc metal is completely reacted with excess amount of hydrochloric acid at 10.0°C and under 90.0 kPa. Determine the volume of gas produced in this reaction.

Answers:

1. 626 L **2.** 932 K **3.** 137 L **4.** 4.33×10^5 ML **5.** 6.00 L

12.5: Gas Molecules: Mixtures and Movements

Partial Pressure: - the pressure exerted by one component of a gas mixture.

Dalton's Law of Partial Pressures: - the total pressure in a gas mixture is equal to the sum of the pressures exerted by individual gaseous components.

Dalton's Law of Partial Pressures

$$P_{Total} = P_1 + P_2 + P_3 + \ldots$$

Example 1: Write the equations for partial pressures for air, which is composed of N_2, O_2, CO_2, Ar, and other trace gases (78.084%, 20.947%, 0.033%, 0.934%, 0.002% by volume respectively). Determine the partial pressures of these gases for 1 atm of air in kPa.

$$P_{total} = P_{N_2} + P_{O_2} + P_{CO_2} + P_{Ar} + P_{\text{trace gases}} = 101.325\text{ kPa}$$

$P_{N_2} = (0.788\,04)(101.325\text{ kPa})$ $P_{N_2} = 79.848\text{ kPa}$

$P_{O_2} = (0.209\,47)(101.325\text{ kPa})$ $P_{O_2} = 21.225\text{ kPa}$

$P_{CO_2} = (0.000\,33)(101.325\text{ kPa})$ $P_{CO_2} = 0.033\text{ kPa}$

$P_{Ar} = (0.009\,34)(101.325\text{ kPa})$ $P_{Ar} = 0.946\text{ kPa}$

$P_{\text{trace gases}} = (0.00\,002)(101.325\text{ kPa})$ $P_{\text{trace gases}} = 0.002\text{ kPa}$

Example 2: The total pressure of the gas collected over a closed ended column partially with water is 135.20 atm. If the vapour pressure of the water at that temperature is 3.15 atm, what is the actual pressure of a gas?

$$P_{Total} = P_{\text{water vapour}} + P_{\text{actual gas}}$$

$$P_{\text{actual gas}} = P_{Total} - P_{\text{water vapour}} = 135.20\text{ kPa} - 3.15\text{ kPa}$$

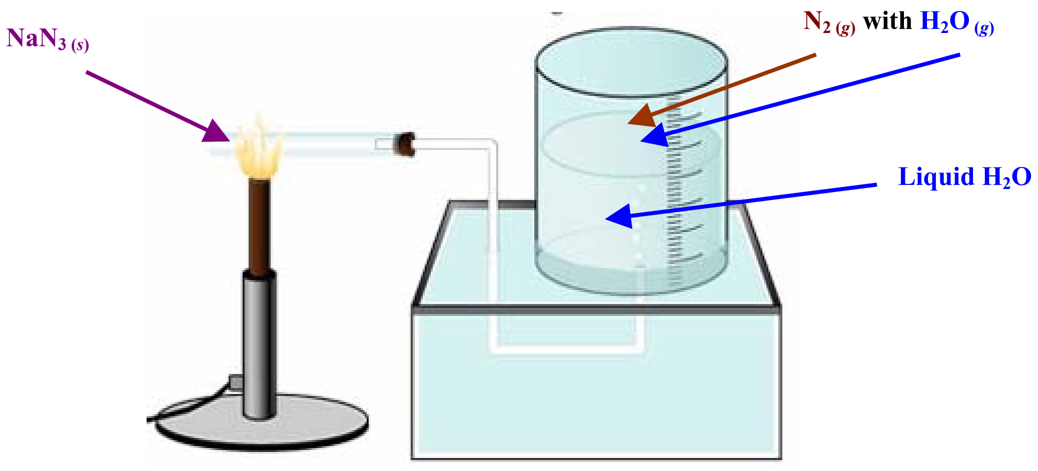
$P_{\text{actual gas}} = 132.05\text{ kPa}$

Example 3: Sodium azide, NaN_3, can be electronically ignited to produced nitrogen used in automobile airbags. The reaction is as follows.

$$2\,NaN_{3\,(s)} \xrightarrow{heat} 2\,Na_{\,(s)} + 3\,N_{2\,(g)}$$

An experiment is set up where the nitrogen produced is bubbled through water for collection and has become saturated with water vapour. Suppose 800 mL of nitrogen gas is collected at 20.0°C and has a total pressure of 105 kPa. (Vapour Pressure of water at 20.0°C is 17.535 torr – the concept of vapour pressure will be covered in section 10.8.)

a. Determine the partial pressure of the nitrogen gas in this experiment.
b. Calculate the mass of sodium azide needed to react to produce this amount of nitrogen.



$$P_{\text{water vapour}} = 17.535\ \cancel{\text{torr}} \times \frac{101.325\ \text{kPa}}{760\ \cancel{\text{torr}}}$$

$$P_{\text{water vapour}} = 2.33780773\ \text{kPa}$$

$\boldsymbol{P_{\text{nitrogen}} = ?}$

$V_{\text{Total}} = 800\ \text{mL} = 0.800\ \text{L}$

$P_{\text{Total}} = 105\ \text{kPa}$

$T = 20.0°\text{C} = 293.15\ \text{K}$

$\boldsymbol{m\ NaN_3 = ?}$

a. $P_{\text{Total}} = P_{\text{water vapour}} + P_{\text{nitrogen}} \quad \Rightarrow \quad P_{\text{nitrogen}} = P_{\text{Total}} - P_{\text{water vapour}}$

$$P_{\text{nitrogen}} = 105\ \text{kPa} - 2.33780773\ \text{kPa}$$

$$\boxed{\boldsymbol{P_{\text{nitrogen}} = 102.6621923\ \text{kPa}}}$$

b. Solving for $\boldsymbol{n_{N_2}}$ allows us to use Stoichiometry to find $\boldsymbol{m}$ of $\boldsymbol{NaN_3}$

$$P_{\text{nitrogen}}V = n_{\text{nitrogen}}RT$$

$$n_{\text{nitrogen}} = \frac{P_{N_2}V}{RT} = \frac{(102.6621923\ \cancel{\text{kPa}})(0.800\ \cancel{\text{L}})}{\left(8.31\,\frac{\cancel{\text{L}}\bullet\cancel{\text{kPa}}}{\cancel{\text{K}}\bullet\text{mol}}\right)(293.15\ \cancel{\text{K}})} = 0.0337139469\ \text{mol}$$

$$2\ NaN_{3\,(s)} \xrightarrow{heat} 2\ Na_{(s)} + 3\ N_{2\,(g)}$$

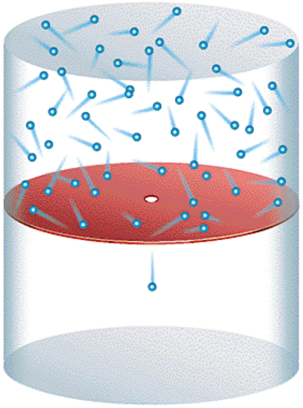
NaN_3: 65.02 g/mol, ? g

N_2: 0.0337139469 mol

① $$n_{NaN_3} = 0.0337139469\ \cancel{\text{mol } N_2} \times \frac{2\ \text{mol } NaN_3}{3\ \cancel{\text{mol } N_2}} = 0.0224759646\ \text{mol } NaN_3$$

② $$m_{NaN_3} = nM = (0.0224759646\ \text{mol } NaN_3)(65.02\ \text{g/mol})$$

$$\boxed{\boldsymbol{m_{NaN_3} = 1.46\ \text{g}}}$$



Effusion: - when gas passes through from one container to the next through a small hole.

Rate of Effusion: - the speed of which a gas pass through a small hole between two containers.
- it is the same as the root mean square velocity, u_{rms}. This is because the small hole on the separting wall of the two containers allow one gas particle to pass through at a time. Therefore, the rate of effusion = u_{rms}
- **rate of effusion is generally measures in mL/min**

Graham's Law of Effusion: - the ratio between the effusion rate of two gases.

Graham's Law of Effusion

$$\frac{\text{Rate of Effusion for Gas 1}}{\text{Rate of Effusion for Gas 2}} = \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

M = Molar Mass in g/mol or kg/mol

$M\downarrow$ (g/mol), Effusion Rate $\uparrow$ (mL/min)

Example 1: Calculate the ratio of effusion rates between $NH_{3\,(g)}$ and $HCl_{\,(g)}$ under the same temperature.

M_{HCl} = 36.46 g/mol

M_{NH_3} = 17.04 g/mol

$\frac{r_{NH_3}}{r_{HCl}} = ?$

$$\frac{r_{NH_3}}{r_{HCl}} = \sqrt{\frac{M_{HCl}}{M_{NH_3}}} = \sqrt{\frac{36.46\ \cancel{g/mol}}{17.04\ \cancel{g/mol}}}$$

$$\boxed{\frac{r_{NH_3}}{r_{HCl}} = 1.463}$$

NH_3 effuse about 1.5 times faster than HCl.

Example 2: Determine the effusion rate F_2 through a porous barrier if the effusion rate of N_2 through the same barrier is 274 mL/min.

r_{N_2} = 274 mL/min

M_{F_2} = 38.00 g/mol

M_{N_2} = 28.02 g/mol

$r_{F_2} = ?$

$$\frac{r_{F_2}}{r_{N_2}} = \frac{\sqrt{M_{N_2}}}{\sqrt{M_{F_2}}} \qquad r_{F_2} = \frac{(r_{N_2})\sqrt{M_{N_2}}}{\sqrt{M_{F_2}}} = \frac{(274\ \text{L/min})\sqrt{28.02\ \cancel{g/mol}}}{\sqrt{38.00\ \cancel{g/mol}}}$$

Since F_2 is bigger than N_2, r_{F_2} is less than r_{N_2}

$$\boxed{r_{F_2} = 235\ \text{mL/min}}$$

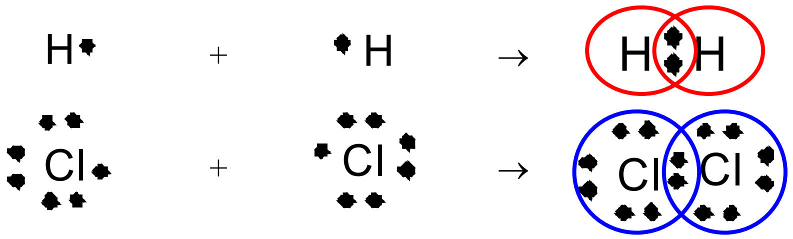
Assignment

12.5 pg. 348 #31 to 33; pg. 349 #35, 36; pg. 351 #37, 38; pg. 353 #39 to 44; pg. 356 #64 to 76

Chapter 16: Covalent Bonding

16.1: The Nature of Covalent Bonding

Covalent Bond: - an attraction force between two non-metal elements where the sharing of valence electrons is taken place so each atom can achieve the stable duet (for hydrogen) or octet (second and third row elements) condition.

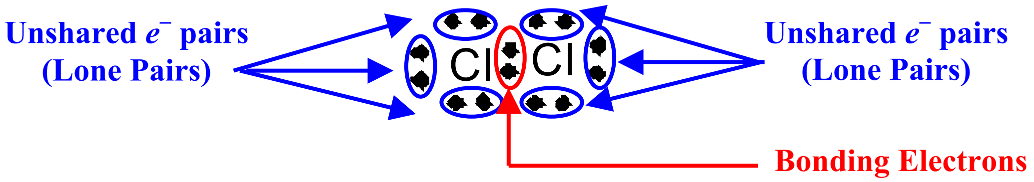


2 valence e^- around each H atom (duet rule)

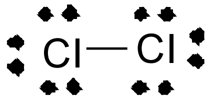
8 valence e^- around each Cl atom (octet rule)

Single Covalent Bond**: - covalent bond where there are only** two valence electrons **being shared between two atoms (also refer to as** bonding electrons**).**

Unshared Electron Pairs (Lone Pairs): - pairs of electrons **not** involving in the **actual covalent bonds**.

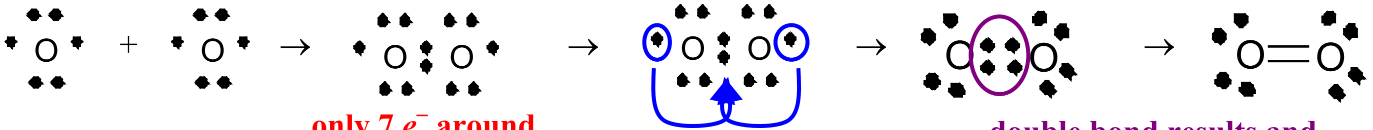


Structural Formula: - a Lewis dot diagram of a molecule where the bonding electrons are replaced by a short line.



Double Covalent Bond**: - covalent bond where there are only** four valence electrons **being shared between two atoms.**

- denotes by a double line in the structural formula.



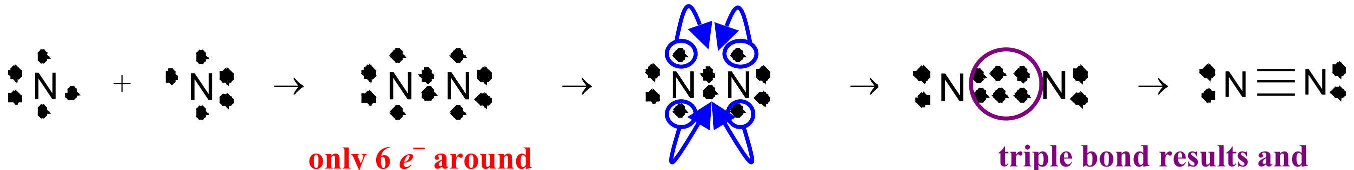
only 7 e^- around each O atom

2 unshared e^- become another bonding pair

double bond results and $8e^-$ around each O atom

Triple Covalent Bond**: - covalent bond where there are only** six valence electrons **being shared between two atoms.**

- denotes by a triple line in the structural formula.



only 6 e^- around each N atom

4 unshared e^- become 2 other bonding pairs

triple bond results and $8e^-$ around each N atom

Coordinate Covalent Bond: - where an atom contributes an unshared electron pair (lone pair) to form double or triple bonds to achieve a stable octet around both atoms.
- sometimes refer to as **multiple bond**.

Procedure to draw Lewis Dot Diagram for Covalent Molecules

1. **Add up all the valence electrons available from all atoms and extra charges from complex ions (ex: charge of 1– means one more electron; charge of 1+ means one less electron).**

2. **Decide on the central atom. Usually, it is the atom with the most available bonding sites (carbon and silicon has the most bonding sites – 4 valence e^-) or it is element with the least number of atoms. Place a pair of electrons to form a single bond between each atom.**

3. **Arrange any remaining electrons around all atoms such that the octet rule is followed for the second and third periods atoms and duet rule is satisfied for the first period atoms. This may sometimes mean setting up multiple bonds (moving lone pairs to the central atom so that it satisfies the octet rule).**

4. **Final Lewis structure of Polyatomic Ion must include a big square bracket and the indication of net charge.**

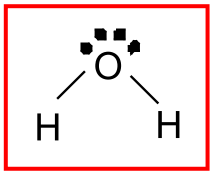
Example 1: Draw the Lewis structure for the following molecules and polyatomic ions.

a. H_2O

$2e^-$ (2 H atoms)
+ $6e^-$ (1 O atom)
$8e^-$ Total
– $4e^-$ (2 single bonds)
$4e^-$ Left
– $4e^-$ (O needs $4e^-$ to achieve octet)
$0e^-$

H O H

H – O – H



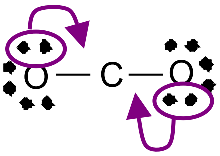
Reason for why the H_2O structure is bent will be explained in section 16.2

b. CO_2

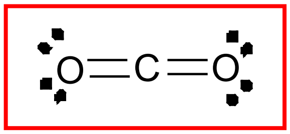
$4e^-$ (1 C atom)
+ $12e^-$ (2 O atoms)
$16e^-$ Total
– $4e^-$ (2 single bonds)
$12e^-$ Left
– $12e^-$ (each O need $6e^-$ to achieve octet)
$0e^-$

O C O

O – C – O

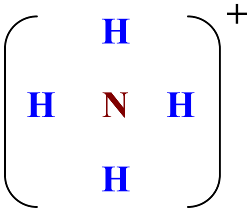


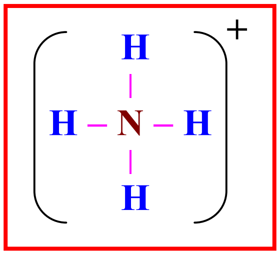
There are only have 4 e^- around C. Oxygen atoms need to contribute unshared e^- pairs (2 lone pairs) to form two double bonds.



c. NH_4^+

   5e^- (1 N atom)
+ 4e^- (4 H atoms)
− 1e^- (+1 charge of NH_4^+ ion)
   8e^- Total
− 8e^- (4 single bonds)
   0e^-



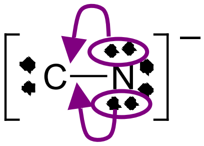


d. CN^-

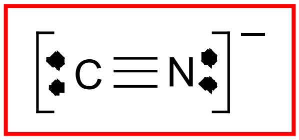
   5e^- (1 N atom)
+ 4e^- (1 C atom)
+ 1e^- (−1 charge of CN^- ion)
   10e^- Total
− 2e^- (1 single bond)
   8e^- Left
− 6e^- (N need 6e^- to achieve octet)
   2e^- (the rest will go to the C atom)

$[C \quad N]^-$

$[C – N]^-$



There are only have 4 e^- around C. Nitrogen atoms need to contribute unshared e^- pairs (2 lone pairs) to form a triple bond.



Bond Energy: - the energy needed to break the covalent bond.
- commonly known as **bond dissociation energy**.

Bond Length: - the distance between centres of the two atoms involved in the covalent bond.

***Note*: In general, bond energy increases and bond length decreases with multiple bonds.**

Bond and Bond Type	Bond Length (pm)	Bond Energy (kJ/mol)
C – C (single bond)	154	247
C = C (double bond)	134	614
C ≡ C (triple bond)	120	839

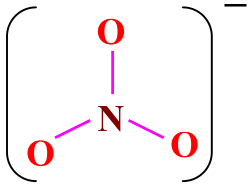
Resonance: - occurs when more than one Lewis structure exists for a molecule.

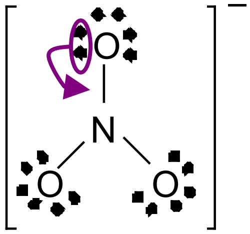
Resonance Structure: - different variations of a resonance molecule can be drawn using a double headed arrow, ↔, between the diagrams.

Example 2: Draw the Lewis structure for nitrate.

Nitrate = NO_3^-

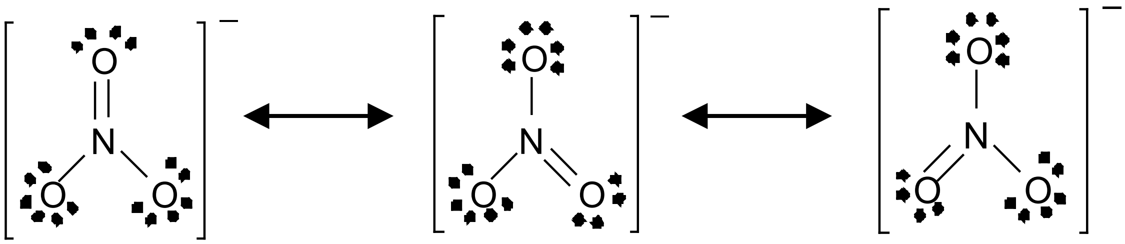
$5e^-$ (1 N atom)
+ $18e^-$ (3 O atoms)
+ $1e^-$ (−1 charge of NO_3^- ion)
$24e^-$ Total
− $6e^-$ (3 single bond)
$18e^-$ Left
−$18e^-$ (each O atom need $6e^-$ to achieve octet)
$0e^-$





There are only have $6e^-$ around N. One oxygen atom needs to contribute unshared e^- pair (1 lone pair) to form a double bond.

There will be three resonance structures as the double bond can "rotate" around the central N atom.

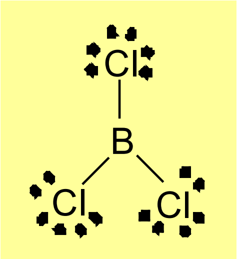


Exceptions to the Octet Rule

1. **In the second row, due to the 2*s* and the 2*p* orbitals, C, N, O, and F should always follow the octet rule. Overall, the second period elements will never exceed the octet rule.**
2. **Boron (B) follows the $6e^-$ rule.**
3. **In the third period, heavier atoms towards the right of the Table often satisfy the octet rule. This being said, there are times that they can use the empty 3*d* orbitals to exceed the 8 electrons limit. (Examples: P can have $10e^-$ rule; S and Cl can have $10e^-$ or $12e^-$ rule)**
4. **Elements in higher row can definitely exceed the octet rule. This sometimes applies to larger noble gases as well. (Examples: Br, I, Xe, Rn can have $10e^-$ or $12e^-$)**

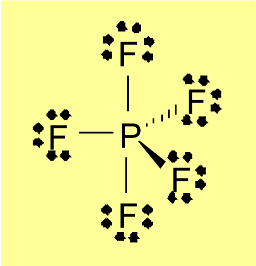
Example 3: Draw the Lewis structure for the following molecules.

a. BCl_3



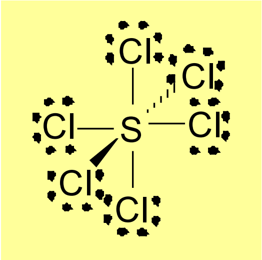
$6e^-$ around B

b. PF_5



$10e^-$ around P

c. SCl_6



$12e^-$ around S

16.2: Bonding Theories

Valence Shell Electron-Pair Repulsion (VSEPR) Model:

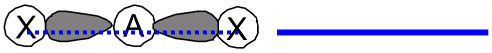
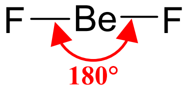
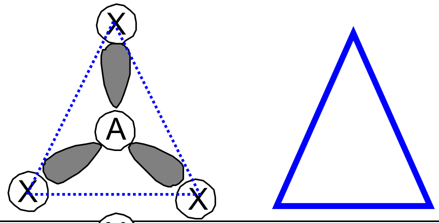
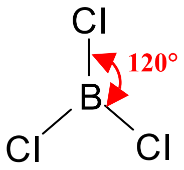
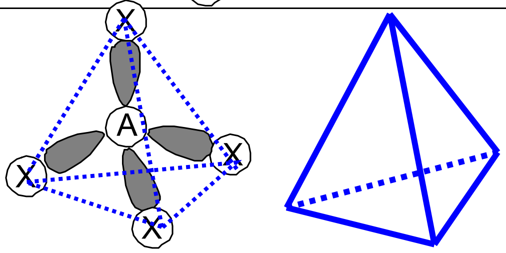
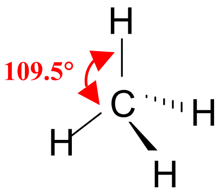
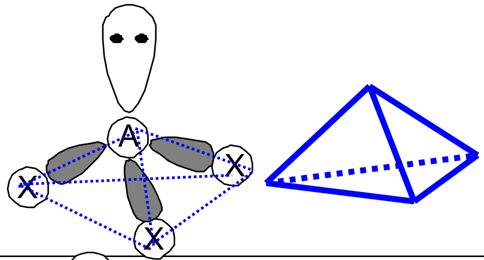
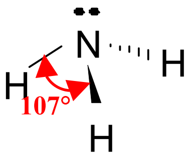
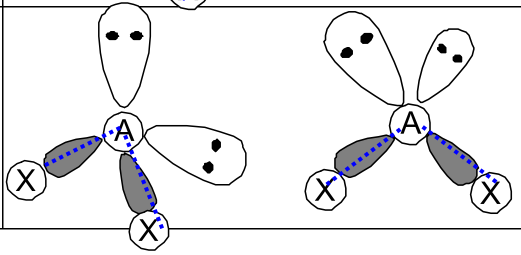
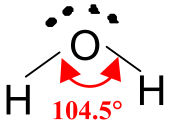
- ***the best structure for a molecule is one that minimizes electrons lone pairs repulsion***.
- most often used to predict molecular structures involving non-metals.

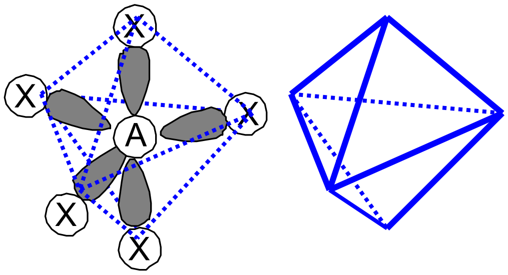
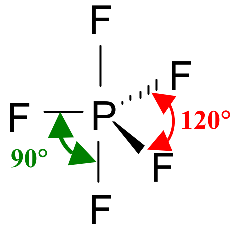
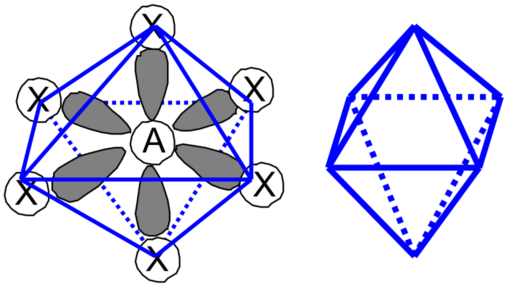
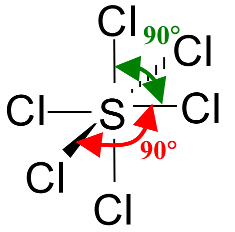
Example: For molecules with a total of 4 e^- pairs, the bond angles decreases from 109.5° as more lone pairs added. (Repulsion of Lone Pair(s) with bond electrons pushed the angle down.)

Effective Electron Pairs:- sometimes refer to as **substituents**

- the number of lone pairs on the central atom of a molecule and the number of connections between the central atom with the outer atom(s). Each set of Multiple Bond (like double and triple bonds) count as one connection or one effective electron pair.

Summary of Geometrical Shape of Single Bond Molecules

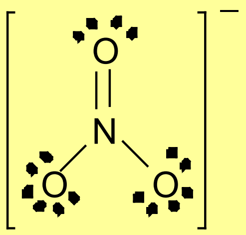
Around Central Atom		**Molecular Structures**	**Geometrical Shape**	**Example and Bond Angles**
Total # of Eff. e^- Pairs	**# of Lone Pairs**			
2	0		Linear	BeF_2 
3	0		Trigonal planar	BCl_3 
4	0		Tetrahedral	CH_4 
4	1		Trigonal pyramid	NH_3 
4	2		V-Shape (Bent)	H_2O 

Around Central Atom		Molecular Structures	Geometrical Shape	Example and Bond Angles
Total # of Eff. e^- Pairs	**# of Lone Pairs**			
5	0		Trigonal bipyramid	PF$_5$ 
6	0		Octahedral	SCl$_6$ 

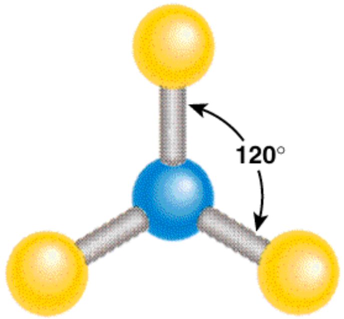
Using VSEPR Model with Multiple Bonds Molecule.

1. **Look at multiple bonds as single bonds (only to determine molecule geometry).**
2. **Using the number of bonds and line pairs around the central atom, determine the geometry of the molecule from the charts above.**

Example 1: From Example 2 of Section 16.1, determine the geometrical shape of NO_3^-. What is the possible bond angle in this complex ion?



Since there are 3 bonds around nitrogen (we look at the double bond as a single bond when deciding on molecular geometry), and there are no lone pairs around the central atom, the NO_3^- ion is in the shape of <u>trigonal planar</u> with <u>bond angles at 120°</u>.



Assignment

16.1 pg. 440 #1 and 2; pg. 446 #3, 4; pg. 447 #5, 6; pg. 451 #7 to 11
16.2 pg. 459 #14

16.3: Polar Bonds and Molecules

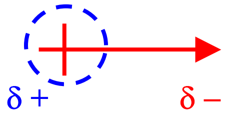
1. **Intramolecular Forces**: - attraction forces between atoms **WITHIN** a molecule.
   - **Ionic Bonding is the intramolecular force of an ionic compound** (metal and non-metal).
   - **Covalent Bonding is the intramolecular force of a molecular compound** (non-metal and non-metal)
   - **the strengths of the intramolecular forces explain the amount of energy involved in a chemical change.**

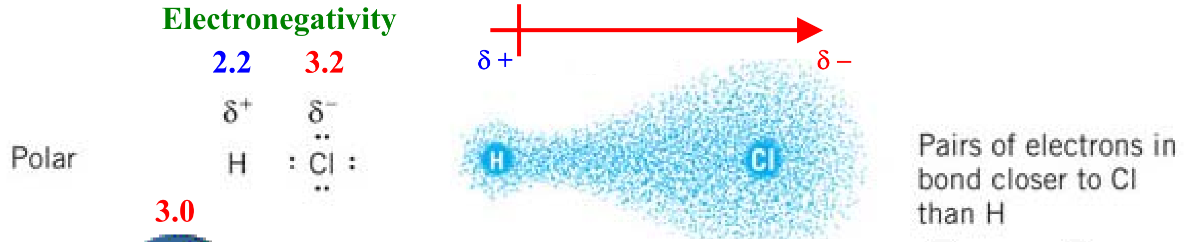
Bond Polarity: - the distribution of electrons within a molecule based on the electronegativities of the individual atoms.

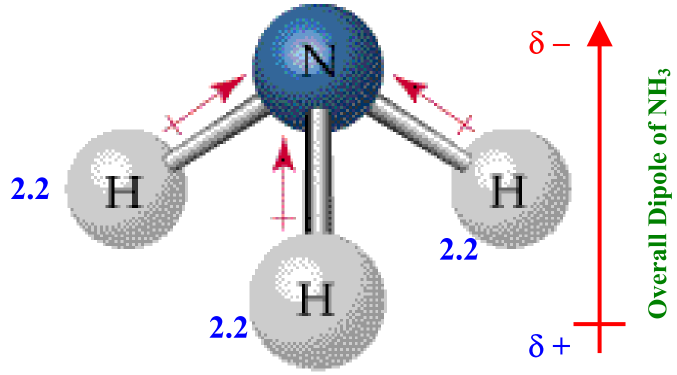
a. **Polar Covalent Bonds**: - the bonds in a covalent compound that exhibit an uneven bond polarity.

Polar Molecule: - when the entire molecule has a net resultant polarity.

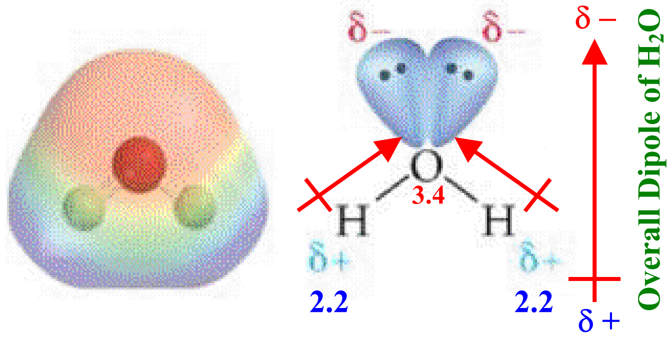
Dipole Moment: - the arrow which indicates the **two poles** (**partial negative, δ –** and **partial positive, δ +**) of a molecule or an individual bond. (δ - Small case Greek letter "delta").
- is a vector that has magnitude (difference in electronegativities) and a direction (based on the geometry of the molecule as dictated by VSERP).







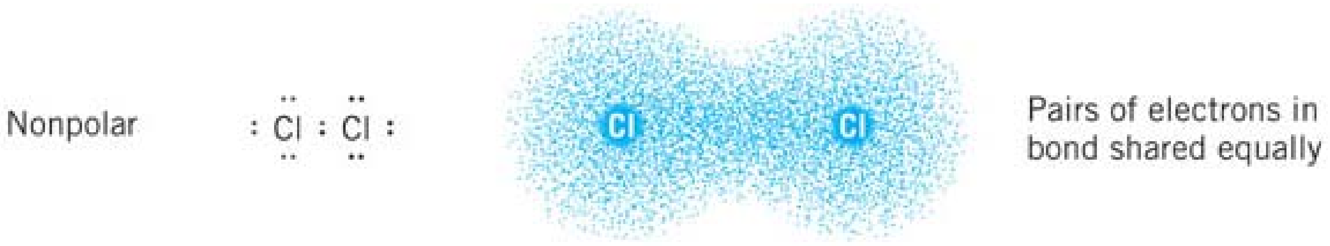
Trigonal pyramid geometry with all N–H bond polarities pointing upward causes an overall dipole for NH_3, making it a polar molecule



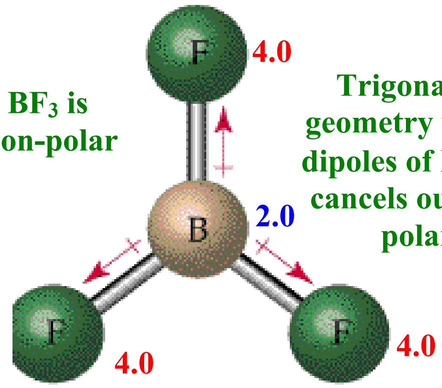
V-shape or Bent geometry with both O–H bond polarities pointing upward causes an overall dipole for H_2O, making it a polar molecule

b. **Non-polar Covalent Bonds**: - when the bond in a covalent compound has an evenly distributed bond polarity.
- occurs when two atoms of the same electronegativity are sharing electrons.

Non-polar Molecule: - when the entire molecule has a net resultant of zero polarity.
- occurs when all dipole moments of all bonds cancel one another (in directions and magnitudes).

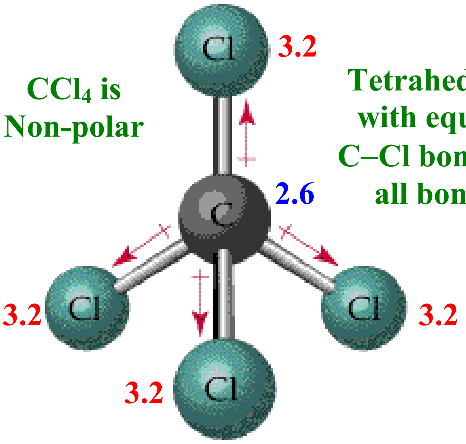


BF_3 is Non-polar



Trigonal planar geometry with equal dipoles of B–F bonds cancels out all bond polarities

CCl_4 is Non-polar



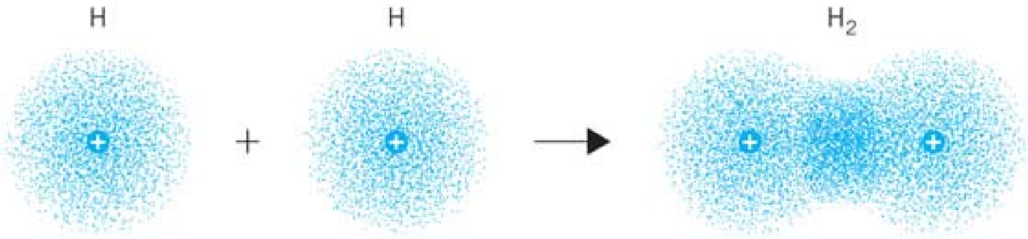
Tetrahedral geometry with equal dipoles of C–Cl bonds cancels out all bond polarities

2. **Intermolecular Forces**: - attraction forces between molecules in a compound
- **the strengths of the intermolecular forces explain the physical properties of compounds (solubility, boiling and freezing points).**

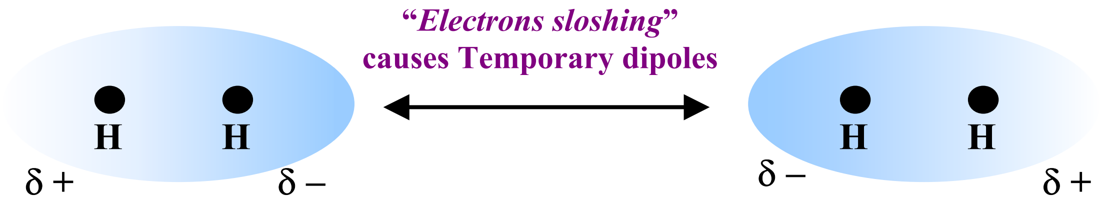
a. **van der Waals Forces**: - named after Johannes van der Waals who studied real gases and interaction between molecules.
- there are two kinds of van der Waals forces.
- they are **Dispersion Forces** and **Dipole-Dipole Interactions**.

i. **Dispersion Forces**: - also known as **London Dispersion Forces** (named after Fritz London who first proposed how this force works).

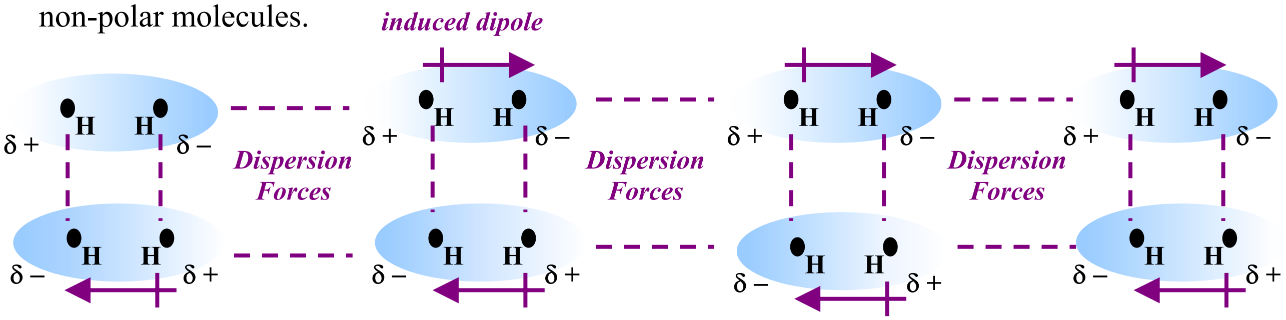
- on average, the non-polar molecules do not have any permanent dipoles like polar molecules



- the "*dispersion*" is the **temporary dipole** that forms within the molecules even in non-polar molecules due the constant motions of electrons. In one instance, they can move to one side of the molecule making it temporary polar. In another instance, electrons will move and the direction of this temporary dipole will switch.



- **This constant "sloshing around" of electrons causes non-polar molecules to have these temporary dipoles**. These temporary "***induced***" dipoles are what cause the attractions between non-polar molecules.



- even monoatomic element like Helium has London Forces. **(Check out animation at http://www.super-tech.ro/teoretic.html)**

-in general, ***the higher the molar mass or the more electrons there are in a molecule*, *the stronger the London Dispersion Force* (attraction between molecules – intermolecular force). *This causes an increase in melting and boiling points of the chemical*.**

-***Note:* All molecules have electrons. Hence, ALL molecules have London Dispersion Force.**

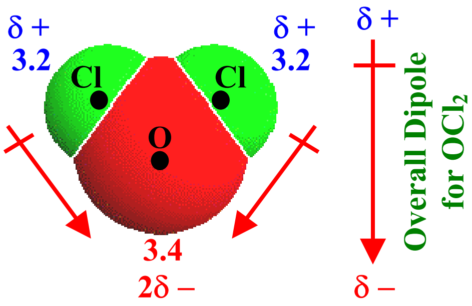
# of e^- or molar mass in atom or molecule ↑, London Dispersion Force ↑, Melting and Boiling Point ↑

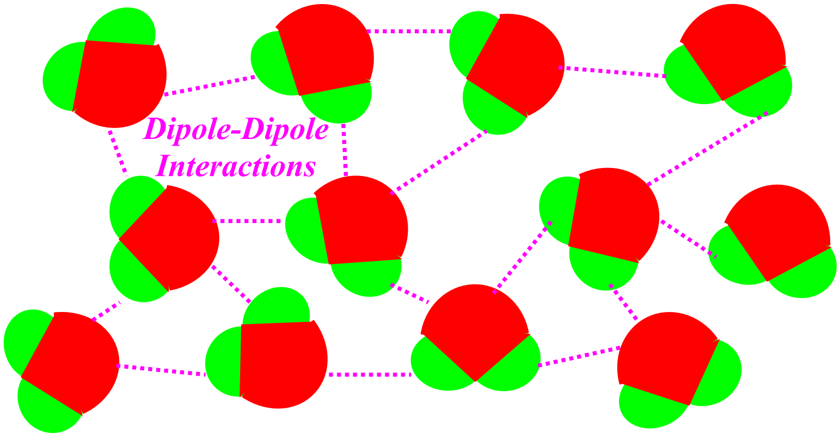
Example 1: Explain the boiling points and the melting points of the noble gases.

Noble Gases	# of e^-	Molar Mass (g/mol)	Melting Point	Boiling Point
He	2	4.00	−272°C (1 K)	−269°C (4 K)
Ne	10	20.18	−249°C (24 K)	−246°C (27 K)
Ar	18	39.95	−189°C (84 K)	−186°C (87 K)
Kr	36	83.80	−157°C (116 K)	−153°C (120 K)
Xe	54	131.29	−112°C (161 K)	−108°C (165 K)
Rn	86	222.00	−71°C (202 K)	−62°C (211 K)

All atoms of noble gases are monoatomic non-polar. The only intermolecular force that governs the melting and boiling points is the London Dispersion Force. As the number of electrons in the noble gases increase, London dispersion force makes the attraction between the atoms greater. This in turn has an effect of increasing the boiling and melting point of the noble gas as one goes down the column.

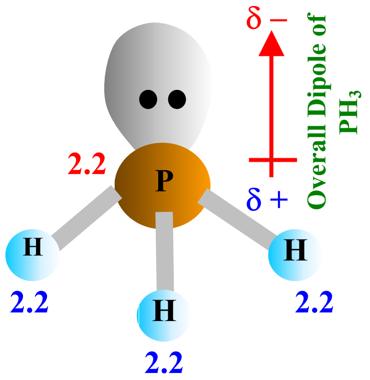
- intermolecular forces resulted from polar molecules.
- dipole interaction is **much stronger than Dispersion Force**.

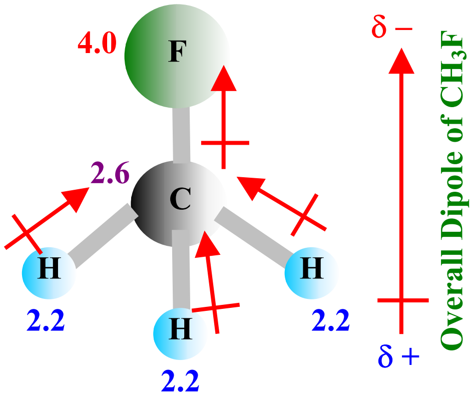




Example 2: Order the boiling points from the least to greatest for the following compounds with similar molar mass. PH_3 (34.00 g/mol), CH_3F (34.04 g/mol), and SiH_4 (32.13 g/mol)

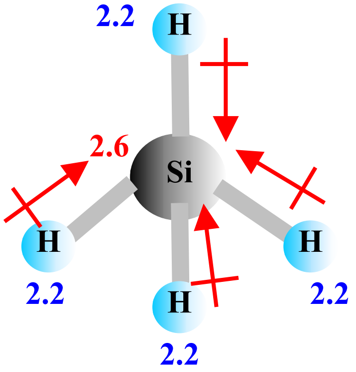
Since PH_3, CH_3F and SiH_4 have similar molar mass, any differences in boiling points cannot be due to London Dispersion forces. Since dipole-dipole interactions exist in polar molecules, we have to examine the molecular geometry and structure of each compound.





PH_3 has a trigonal pyramid geometry (VSEPR) and is Polar. Even though the P−H bonds have no polarity (electronegativities of P and H are the same), the lone pair on one end of the P atom.

CH_3F has a tetrahedral geometry and is very polar. The C−F bond along with the C−H bonds have strong polarity. The overall dipole moment for the molecule has electrons around the F atom.



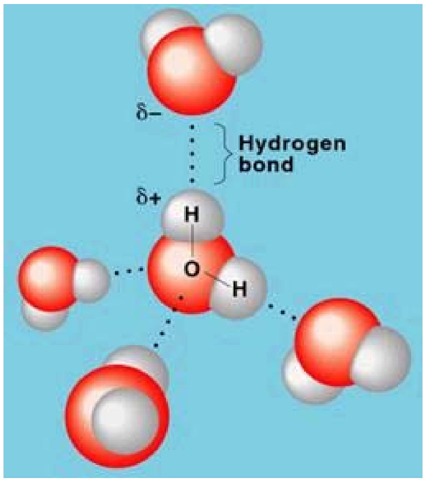
SiH_4 has a tetrahedral geometry with equal dipoles of Si−H bonds cancels out all bond polarities. Hence, SiH_4 is Non-Polar.

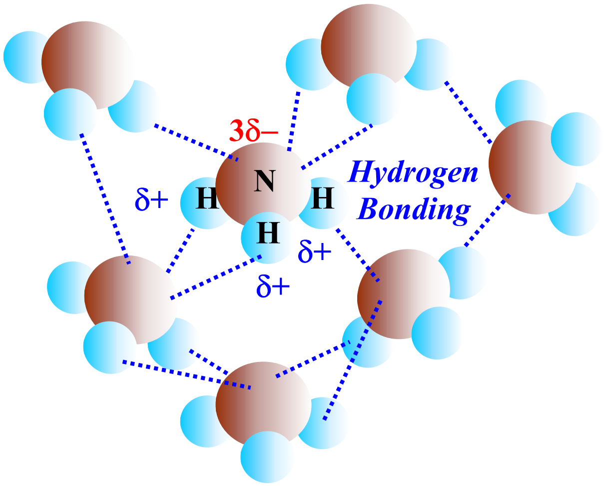
Since non-polar molecule have no dipole interactions, SiH_4 should have the lowest boiling point. PH_3 is less polar than CH_3F due to the difference in electronegativities between P−H bond and C−F with C−H bonds. Therefore, CH_3F must have the highest boiling point.

Boiling Point: $SiH_4 < PH_3 < CH_3F$

b. **Hydrogen Bonds**: - are intermolecular bonds that involve **hydrogen atom with very electronegative atom that also consists of lone pairs**.

- these **include O–H, N–H, and H–X (halogen atoms) bonds**.
- the resulting molecule is always polar. Therefore, **all hydrogen bonding molecules also have dipole interactions**.
- hydrogen bond is the **STRONGEST** of the intermolecular bonds.



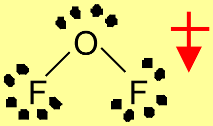


(Check out the Hydrogen Bond Animation at http://www.northland.cc.mn.us/biology/Biology1111/animations/hydrogenbonds.html)

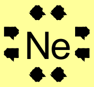
Example 3: Account for the differences in the boiling points of the compounds listed below.

Molecule	Molar Mass (g/mol)	London Dispersion Forces	Dipole Interactions	Hydrogen Bonds	Boiling Point
OF_2	54.00	✓	✓	✗	−145°C (128 K)
Ne	20.18	✓	✗	✗	−246°C (27 K)
HF	20.01	✓	✓	✓	19°C (292 K)
H_2O	18.02	✓	✓	✓	100°C (373 K)
NH_3	17.04	✓	✓	✓	−33°C (240 K)
CH_4	16.05	✓	✗	✗	−161°C (112 K)

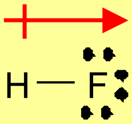
Again, we need to draw the structural formulas of these molecules and compare their polarities.



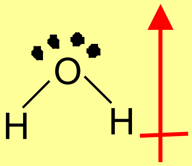
OF_2 is polar with dipole interactions



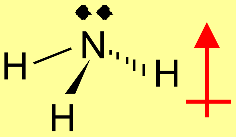
Ne is non-polar with dispersion forces only



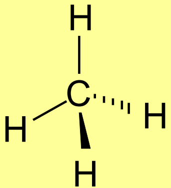
HF is polar with hydrogen bonds



H_2O is polar with hydrogen bonds



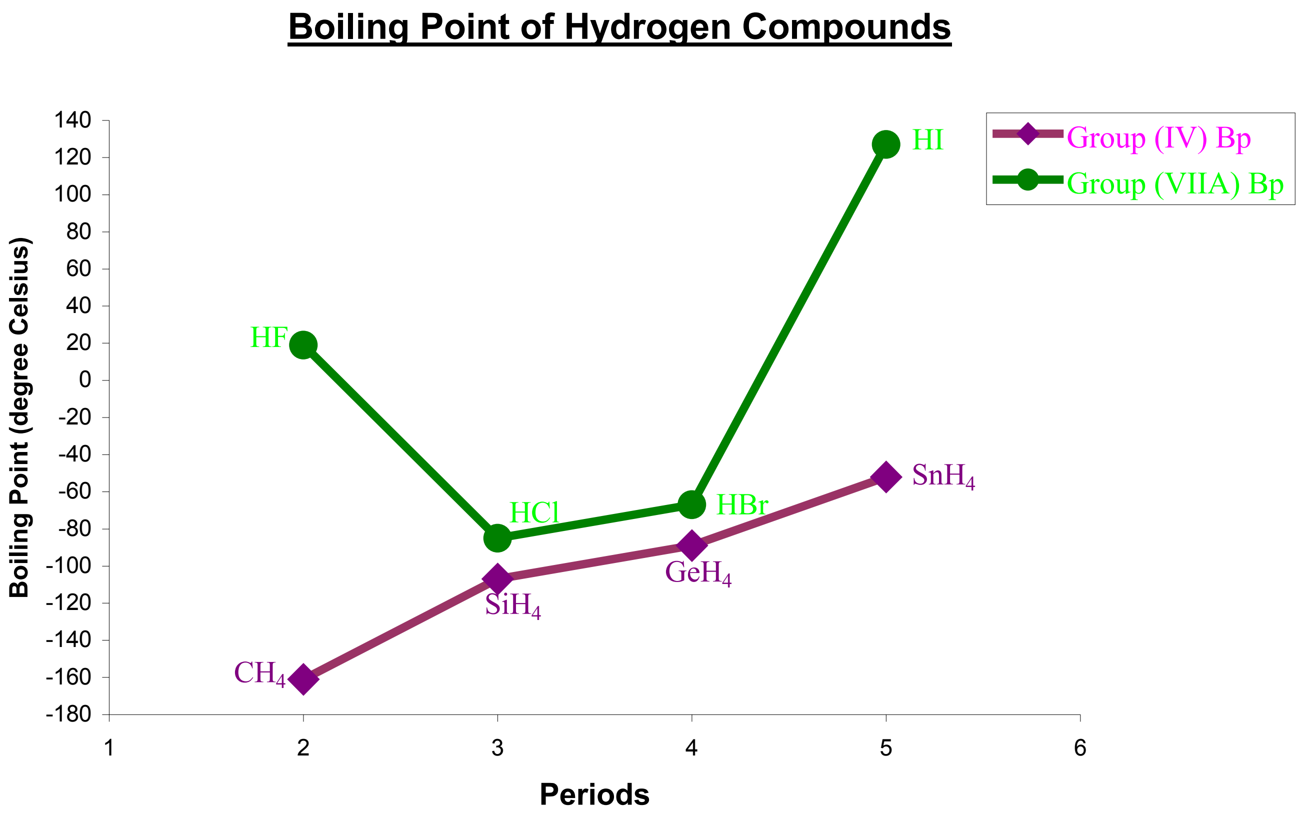
NH_3 is polar with hydrogen bonds



CH_4 is non-polar with dispersion forces only

Notice that the hydrogen bond molecules (HF, H_2O and NH_3) have boiling points much higher than molecule with just dipole interactions (OF_2) and the ones with only London Dispersion Forces (Ne and CH_4).

Example 4: Given the graph below on the boiling points of hydrogen compounds with different group series, explain the following using the concepts of chemical bonding.



a. The hydrogen compounds in the Group (VIIA) series have higher boiling points than hydrogen compounds in the Group (IVA) series.

b. HF has a higher boiling point than HCl and HBr in the Group (VIIA) series while CH_4 has a lowest boiling point in its own Group (IVA) series.

a. All hydrogen compounds in the Group (VIIA) series are polar and have hydrogen bonds. On the other hand, all hydrogen compounds in the Group (IVA) series are non-polar and only have London dispersion forces. Since hydrogen bonds are stronger intermolecular forces than London dispersion forces, the hydrogen compounds in the Group (VIIA) series have generally higher boiling points than the counterparts in the Group (IVA) series.

b. HF has a stronger hydrogen bond than HCl and HBr. The difference between the electronegativities in HF is 2.2–4.0 while for HCl and HBr are 2.2–3.2 and 2.2–3.0 respectively. This huge difference in electronegativities in HF is what causes its boiling point to *buckle the trend*. After HF, the rest of the hydrogen compounds in the Group (VIIA) series follow the effect of London dispersion forces, the higher the molar mass, the stronger the dispersion forces, and the increase in boiling points is the result.

CH_4 in the Group (IVA) series do not buckle the trend because the entire series are non-polar. The only intermolecular force at work is the London dispersion force. Hence, CH_4 has a lower boiling point than SiH_4 and HF has a higher boiling point than HCl and HBr.

Properties of Covalent (Molecular) Compounds

1. **Covalent Compounds tend to have much Lower Boiling and Melting Points than ionic compounds.** This is because solid covalent compounds' **crystalline structures** are made up of mostly **weak intermolecular forces**. Their boiling points are lower than ionic compounds because there are **no ion interactions in liquid state, only intermolecular forces**.
2. **Solid Covalent Compounds are Soft.** Again, covalent compounds' **weak crystalline structure** makes them soft.
3. **Covalent Compounds tend to be More Flammable than ionic compound.** This is due to the some non-metals like carbon and sulfur, which **combine readily with oxygen in combustion reactions**.
4. **Most Covalent Compounds are Insoluble in Water.** Because **water is very polar** and has lots of hydrogen bonds, it can only dissolve covalent compounds that are polar as well "***Like Dissolves Like***". Since **most covalent compounds are fairly non-polar**, they do not dissolve in polar water well.
5. **Soluble Covalent Compounds do NOT Conduct Electricity in Water.** This is simply due to the fact that covalent compounds **do not dissociate into ions or electrolytes** like soluble ionic compounds do.

Summary of Chemical Bonding

1. **Intramolecular bonds such as covalent and ionic bonds deal with bond strength between atoms within an atom. They explain amount of energy in a chemical change.**
2. **Intermolecular Bonds involve in a compound explain its physical properties such as solubility ("*like dissolves like*" – section 17.3, boiling and melting points – energy involved in physical phase change).**
3. **van der Waals Forces consist of London Dispersion forces (apply to all molecules) and Dipole Interactions (apply to polar molecules).**
4. **Hydrogen Bonding is the *strongest* of all intermolecular bonds.**

Strength of Chemical Bonds

$$\underbrace{\text{Ionic Bond} >> \text{Covalent Bond}}_{\text{Intramolecular Bonds}} \underset{\text{(much stronger than)}}{>>} \underbrace{\text{Hydrogen Bond} > \overbrace{\text{Dipole Interaction} >> \text{London Dispersion Force}}^{\text{van der Waals Forces}}}_{\text{Intermolecular Bonds}}$$

Assignment

16.3 pg. 466 #21 to 24, 26

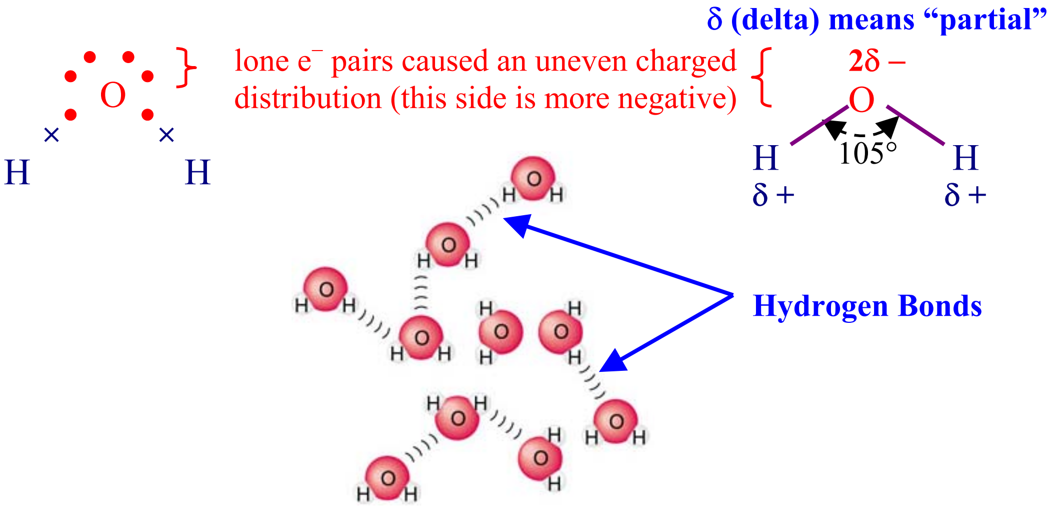
Chapter 16 Review: pg. 470–471 #27 to 33, 37, 38, 41, 43, 48, 51, 52, 54, 56 to 58

Chapter 17: Water as Aqueous Systems

17.1: Liquid Water and its Properties

Structure of Water:

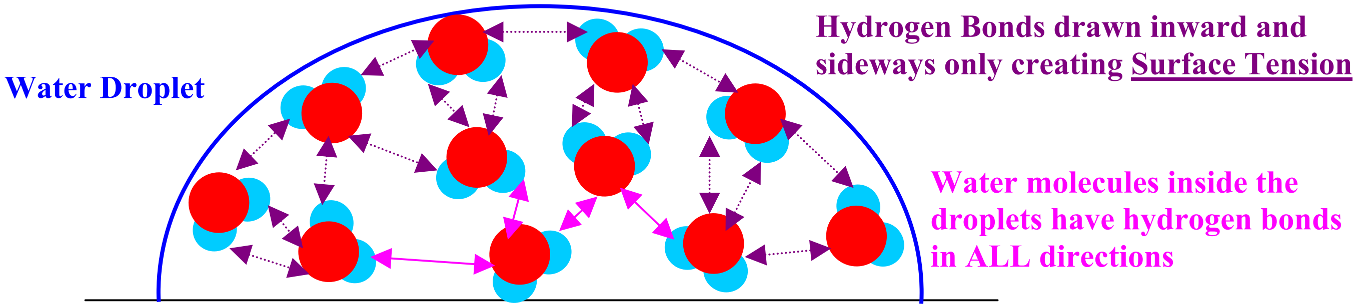
1. **V-Shaped**: - the two O–H bonds form 105° from each other, which leads to its polarity.
2. **Polar Molecule**: **-** unequal charged distribution due to the electron pairs around the oxygen atom.
3. **Strong O–H Hydrogen Bond**: - a type of hydrogen bond that is fairly strong compared to other types of intermolecular bonds (bonds between molecules).



Some Properties of Water:

1. **Strong Surface Tension**

Surface Tension: - the inward force of a liquid to minimize its surface area.
- water's hydrogen bond cannot attract with molecules in the air. Therefore, the higher net force is pushed inward, leaving a spherical surface with a high surface tension.



Surfactant: - a surface-active agent that decreases the surface tension (Example: detergent)

2. High Specific Heat Capacity

Specific Heat Capacity: - the amount of heat it takes to raise the temperature of 1 g of substance by 1°C.
- water has a high specific heat capacity, making it a good storage of solar energy

Specific Heat Capacity of Some Common Substances

Substance	Specific Heat Capacity	Substance	Specific Heat Capacity
Water	4.19 J/(g • °C)	Copper	0.385 J/(g • °C)
Ice	2.00 J/(g • °C)	Iron	0.449 J/(g • °C)
Ammonia	2.06 J/(g • °C)	Oxygen	0.918 J/(g • °C)
Ethanol	2.44 J/(g • °C)	Nitrogen	1.040 J/(g • °C)

17.2: Water Vapour and Ice

Other Properties of Water:

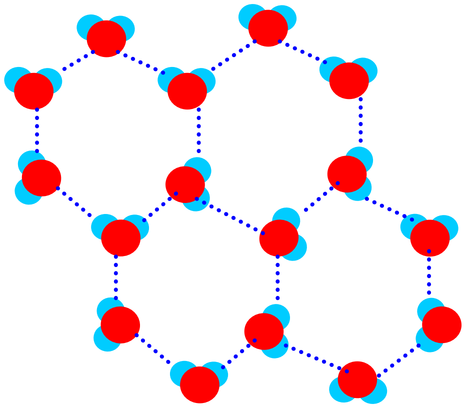
3. High Boiling Point and Heat of Vaporization

Boiling Point: - the temperature where the liquid form of a substance evaporates into its gaseous form (or a gaseous form of a substance condenses into its liquid form).
- water has a high boiling point (100°C) due to the strong hydrogen bonds. It takes more energy to break these intermolecular bonds.

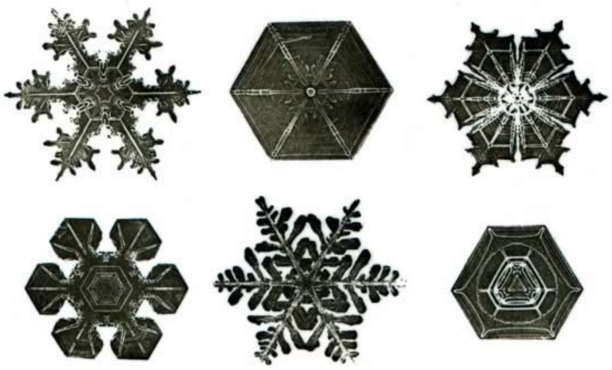
Heat of Vaporization: - the amount of heat to turn 1 mole or 1 gram of liquid substance into its gaseous form.
- water has a molar heat of vaporization of 40.65 kJ/mol or 2.256 kJ/g.
- because it takes an enormous amount of energy to evaporate water, temperatures are regulated on Earth. (Tropical regions are cooled as water absorbs heat. Polar regions are warmed as water vapour condenses to liquid giving off heat).

4. Low Density in Ice compared to Liquid Water

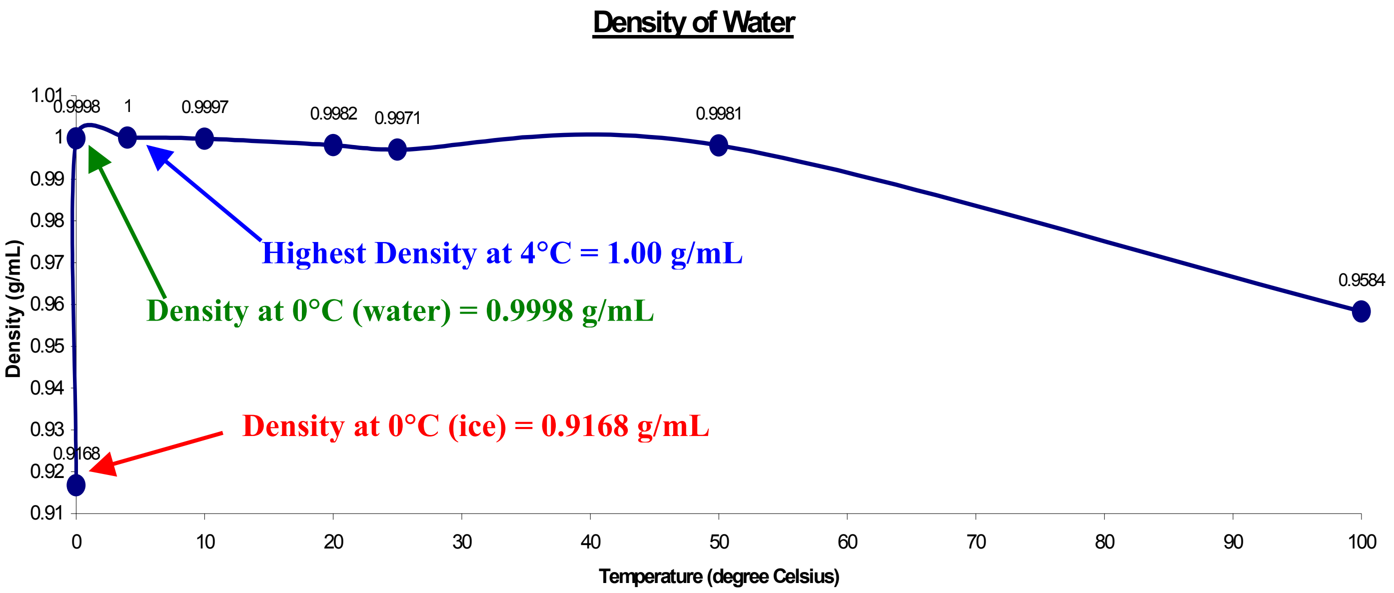
Hexagonal Arrangement of Ice Crystals: - due to the hydrogen bonds in water, it forms a honeycomb shape and expands in volume when it crystalizes into ice.



Due to the hydrogen bonds in water, it forms a *honeycomb shape* and expands in volume when it crystallizes into ice



Even though no two snowflakes are alike, all of them have a basic hexagonal shape as dictated by the bent shape of water molecule and its hydrogen bonds



17.3: Aqueous Solutions

Solute: - the matter that is being dissolved. **Solvent**: - the matter that is doing the dissolving.

Solvation: - the process where a solute is dissolving into a solvent.

Aqueous Solution: - solution where water is the solvent (*aq*).

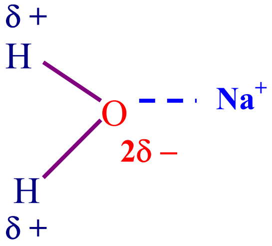
Reason that Water is a Common Solvent:

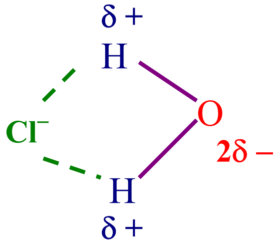
1. **Polar Molecule**: - dissolves many ionic compounds due to its ability to attract cations and anions (electrolytes).
   - *Note*: Some ionic compounds have low solubility (ability to dissolve) in water.

Hydration: - when ionic compound dissolves in water.

Example 1: Sodium Chloride (Salt)

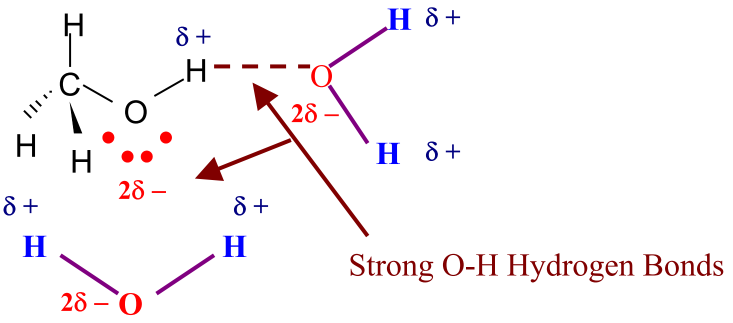
$NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$





2. **Strong O–H Hydrogen Bond**: - dissolves many molecular compounds that have O–H hydrogen bonds.

Example 2: CH_3OH (Methanol)



"**Like-Dissolves-Like**": - polar solvents tend to dissolve polar solutes; non-polar solvents tend to dissolve non-polar solutes.

Solubility Table: - a chart that shows the ability of various ion combinations to dissolve in water.

Solubility of Some Common Ionic Compounds in Water at 298.15 K

Ion	**Group 1 NH_4^+ H_3O^+ (H^+)**	**ClO_3^- NO_3^- ClO_4^-**	**CH_3COO^-**	**Cl^- Br^- I^-**	**SO_4^{2-}**	**S^{2-}**	**OH^-**	**PO_4^{3-} SO_3^{2-} CO_3^{2-}**
Solubility greater than or equal to 0.1 mol/L (very soluble)	all	all	most	most	most	Group 1 Group 2 NH_4^+	Group 1 NH_4^+ Sr^{2+} Ba^{2+} Tl^+	Group 1 NH_4^+
Solubility less than 0.1 mol/L (slightly soluble)	none	none	Ag^+ Hg^+	Ag^+ Pb^{2+} Hg^+ Cu^+ Tl^+	Ca^{2+} Sr^{2+} Ba^{2+} Ra^{2+} Pb^{2+} Ag^+	most	most	most

Example 3: Determine if the following ionic compounds are soluble in water.

a. silver chloride

Insoluble

Under the Cl^- column, Ag^+ is in the "slightly soluble" row.

b. ammonium hydroxide

Soluble

Under the NH_4^+ column, all cations combinations are soluble.

or

Under the OH^- column, NH_4^+ is listed in the "very soluble" row.

c. lead (IV) sulfate

Soluble

Under the SO_4^{2-} column, Pb^{4+} is NOT in the "slightly soluble" row.

(Pb^{4+} is NOT the same as Pb^{2+})

d. chromium (III) nitrate

Soluble

Under the NO_3^- column, all cations are combinations are soluble.

e. aluminum hydroxide

Insoluble

Under the OH^- column, Al^{3+} is NOT in the "very soluble" row.

f. copper (I) bromide

Insoluble

Under the Br^- column, Cu^+ is in the "slightly soluble" row.

(Cu^+ is NOT the same as Cu^{2+})

Electrolytes: - when soluble ionic compound dissolves, the ions has to ability to conduct electricity.

1. **Strong Electrolytes**: - ionic compounds that dissociates completely into their ions and conduct electricity very effectively.

Example 4: a. All ionic compounds containing NH_4^+, NO_3^-, ClO_3^-, or ClO_4^-.
b. Strong Acids: $HClO_{4\,(aq)}$, $HI_{(aq)}$, $HBr_{(aq)}$, $HCl_{(aq)}$, $H_2SO_{4\,(aq)}$ and $HNO_{3\,(aq)}$
c. Strong Bases: $NH_4OH_{(aq)}$, $Sr(OH)_{2\,(aq)}$, $Ba(OH)_{2\,(aq)}$, $TlOH_{(aq)}$, and Group 1 with OH^- ($LiOH_{(aq)}$, $NaOH_{(aq)}$, $KOH_{(aq)}$, $RbOH_{(aq)}$, $CsOH_{(aq)}$, and $FrOH_{(aq)}$)

2. **Weak Electrolytes**: - ionic compounds that **dissociate partially** into their ions and **conduct electricity poorly**.

Example 5: a. $AgCl_{(s)}$, $PbCl_{2(s)}$, $HgCl_{(s)}$, $CuCl_{(s)}$, and $TlCl_{(s)}$
b. Weak Acids: $HF_{(aq)}$, $HCH_3COO_{(aq)}$, $H_2SO_{3\,(aq)}$, and other weak acids.
c. Weak Bases: $Ca(OH)_{2\,(aq)}$, $Al(OH)_{3\,(aq)}$, $NH_{3\,(aq)}$, and other weak bases.

Non-Electrolytes: - ionic compounds that **do not dissociate** (generally molecular compounds) in the solvent and **do not conduct electricity at all**.

Example 6: $C_{12}H_{22}O_{11\,(aq)}$ and $C_6H_{12}O_{6\,(aq)}$

Hydrates: - sometimes called water of hydration or water crystallization.
- when water particles are locked into the crystal structure of an ionic compound.
- naming contains the name of the ionic compound with a prefix followed by the word "*hydrate*".

Prefixes for Hydrates

1 – mono
2 – di
3 – tri
4 – tetra
5 – penta
6 – hexa
7 – hepta
8 – octa
9 – nona
10 – deca

Example 7: Write the equation of the formation of iron (II) sulfate heptahydrate and determine the percent mass of water in this hydrate.

iron (II) sulfate + water → iron (II) sulfate heptahydrate

$$FeSO_{4\,(s)} + 7\,H_2O_{(l)} \rightarrow FeSO_4 \bullet 7\,H_2O$$

$M = 18.02$ g/mol

$M = 55.85 + 32.06 + 4(16.00) + 7(18.02)$

$M = 277.91$ g/mol

$$\%\text{ water by mass} = \frac{7(18.02\ \cancel{\text{g/mol}})}{277.91\ \cancel{\text{g/mol}}} \times 100\% \qquad \boxed{\%\ H_2O = 45.4\%}$$

Assignment

17.1 pg. 478 #1 to 4
17.2 pg. 481 #5 and 6
17.3 pg. 488 #8 to 10 and 12
Ch 17 Review: pg. 496 – 497 #19, 20, 22, 23, 24, 29, 31 to 35, 38, 40 and 41