

# Chemistry Midterm Notes

## Average Atomic Mass

**Atomic Mass** – weighted average of the respective atomic masses of the natural isotopes in given element

$$AM = M_1f_1 + M_2f_2 + M_3f_3 + \dots, \text{ where:}$$

M = mass of each isotope.

f = fractional abundance of that isotope.

Ex) Carbon has two naturally occurring isotopes:

	$^{12}_6\text{C}$	$^{13}_6\text{C}$
Mass	12 u	13.00335 u
Abundance	98.892%	1.108%

The average atomic mass of carbon is:

$$(0.98892)(12) + (0.01108)(13.00335) = 12.011 \text{ u}$$

## Molecular Mass

For H<sub>2</sub>O:

Molecular Mass: mass of one molecule 18.015 u

Molar Mass: mass of one mole 18.015 g/mol

## Molecular Mass vs Exact Mass

**Molecular Mass:** Use the average atomic mass of each element

$$6 \times 12.01 + 12 \times 1.01 + 6 \times 16.00 = 180.18 \text{ u}$$

**Exact Mass:** Use exact mass of the most abundant isotope of each element

$$6 \times 12.000000 + 12 \times 1.007825 + 6 \times 15.994915 = 180.06339 \text{ u}$$

## Polyatomic Ions

**TABLE 2.5** Common Polyatomic Ions\*

Formula	Name
<b>Cations</b>	
$\text{NH}_4^+$	Ammonium
$\text{H}_3\text{O}^+$	Hydronium
$\text{Hg}_2^{2+}$	Mercury(I)
<b>Anions</b>	
$\text{CH}_3\text{COO}^-$ (or $\text{C}_2\text{H}_3\text{O}_2^-$ )	Ethanoate (or Acetate)
$\text{CN}^-$	Cyanide
$\text{OH}^-$	Hydroxide
$\text{ClO}^-$	Hypochlorite
$\text{ClO}_2^-$	Chlorite
$\text{ClO}_3^-$	Chlorate
$\text{ClO}_4^-$	Perchlorate

$\text{NO}_2^-$	Nitrite
$\text{NO}_3^-$	<b>Nitrate</b>
$\text{MnO}_4^-$	<b>Permanganate</b>
$\text{CO}_3^{2-}$	<b>Carbonate</b>
$\text{HCO}_3^-$	<b>Hydrogen carbonate</b> (or <b>bicarbonate</b> )
$\text{CrO}_4^{2-}$	Chromate
$\text{Cr}_2\text{O}_7^{2-}$	<b>Dichromate</b>
$\text{O}_2^{2-}$	Peroxide
$\text{PO}_4^{3-}$	<b>Phosphate</b>
$\text{HPO}_4^{2-}$	Hydrogen phosphate
$\text{H}_2\text{PO}_4^-$	Dihydrogen phosphate
$\text{SO}_3^{2-}$	Sulfite
$\text{SO}_4^{2-}$	<b>Sulfate</b>
$\text{HSO}_4^-$	Hydrogen sulfate (or bisulfate)

## Naming Inorganic Compounds

Hydrates: An ionic compound containing a fixed number of molecules of water

Ex)  $\text{MgSO}_4 \cdot 2\text{H}_2\text{O} = \text{Magnesium sulfate dihydrate}$

## Naming Acids

Binary Acids/Hydroacids: Compound with the general formula  $H_nX$ .

Named using prefix *hydro-* and suffix *-ic*.

Ex) HF                  Hydrofluoric acid

## Naming Oxoacids

Compound with general formula  $H_mXO_n$  that ionizes in water

Oxoacid = oxoanion +  $\text{H}^+$  ions

*-ite* becomes *-ous*, *-ate* becomes *-ic*

Ex)  $\text{ClO}^-$  - Hypochlorite                   $\rightarrow$                    $\text{HClO}$  - Hypochlorous acid

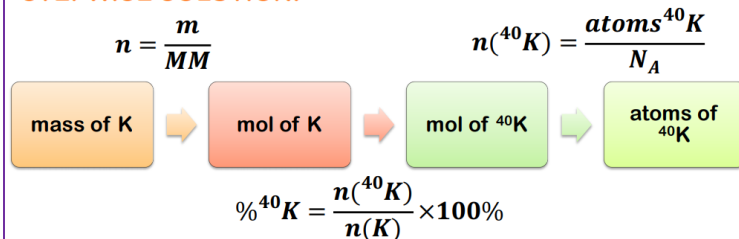
## Percent Composition

Number of parts of a component in 100 parts of the whole.

Ex) A rock contains 3.5% gold by mass means 3.5 g of gold per 100 g of rock

Potassium-40 is one of the few naturally occurring radioactive isotopes of elements of low atomic number. Its percent natural abundance among K isotopes is 0.012%. How many  $^{40}\text{K}$  atoms do you ingest by drinking one cup of whole milk containing 371 mg of K?

**STEPWISE SOLUTION:**



$$n = \frac{m}{M}$$

$$\%n_A = \frac{n_A}{n_T} \times 100$$

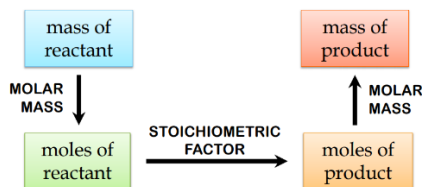
$$n = \frac{N}{N_A}$$

## Stoichiometry

Step 1: Identify all reactants and products

Step 2: Balance the chemical equation

Step 3: Follow the mole method



## Reaction Yield

The theoretical yield is the amount of product expected if the reactants react to completion (Calculated number).

$$\% \text{Yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

## Chemical Reactions in Solution

**Solution:** solute dissolved in solvent to form a homogeneous mixture, **Solute:** present in smallest amount, **Solvent:** present in largest amount. When the **solvent is water** = aqueous solutions.

**Concentration:** the measure of the amount of solute in a solution.

## Concentrations of Solutions

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{volume of solution in litres}} \rightarrow C = \frac{n}{V}$$

We want to prepare exactly 0.2500 L of an 0.250 M  $\text{K}_2\text{CrO}_4$  solution in water. What mass of  $\text{K}_2\text{CrO}_4$  should we use?

**WANT:** ? g  $\text{K}_2\text{CrO}_4$

**HAVE:**  $V = 0.2500 \text{ L}$   
 $C = 0.250 \text{ mol/L}$


**NEED:** MM of  $\text{K}_2\text{CrO}_4 = 194.02 \text{ g/mol}$

$$C = n/V$$

$$n = m/MM$$

$$? \text{ g } \text{K}_2\text{CrO}_4 = 0.2500 \text{ L solution} \times \frac{0.2500 \text{ mol } \text{K}_2\text{CrO}_4}{\text{L solution}} \times \frac{194.02 \text{ g } \text{K}_2\text{CrO}_4}{\text{mol } \text{K}_2\text{CrO}_4}$$

$$= 12.1 \text{ g}$$

Molality	$m$ or $b$	$\frac{n_{\text{solute}}}{m_{\text{solvent}}} = \frac{x \text{ mol solute}}{1 \text{ kg solvent}}$	
Mole Fraction	$\chi$	$\frac{n_A}{n_{\text{total}}} = \frac{x \text{ mol of "A"}}{1 \text{ mol solution}}$	
Parts per million	ppm	$\frac{1 \text{ part solute}}{10^6 \text{ parts solution}} = \frac{x \text{ g solute}}{10^6 \text{ g solution}}$ $\approx \frac{x \text{ mg solute}}{\text{L solution}}$	for dilute aqueous solutions ONLY!

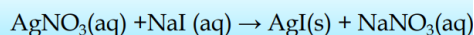
**Dilution:** Addition of extra solvent does not change the amount of dissolved solute, therefore:

$$C_{\text{concentrated}} V_{\text{concentrated}} = \text{moles} = C_{\text{dilute}} V_{\text{dilute}}$$

$$C_1 V_1 = C_2 V_2$$

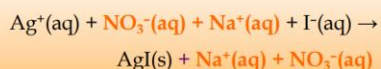
## Writing Aqueous Ionic Reactions

### A. Molecular equation



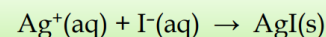
- species are written as intact compounds
- in reality: they are dissociated ions!

### B. Ionic equation



- more realistic: represents all species in their ionic forms
- in this example,  $\text{Na}^+$  and  $\text{NO}_3^-$  are known as **spectator ions**

### C. Net ionic equation



- this equation drops all of the spectator ions, and so includes only those ions participating in the reaction

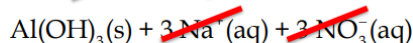
Write and balance the net ionic equation.



Molecular Equation:



Ionic Equation:  $\text{Al}^{3+}(\text{aq}) + 3 \text{NO}_3^-(\text{aq}) + 3 \text{Na}^+(\text{aq}) + 3 \text{OH}^-(\text{aq}) \rightarrow$



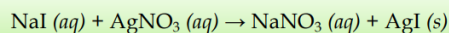
Net Ionic Equation:  $\text{Al}^{3+}(\text{aq}) + 3 \text{OH}^-(\text{aq}) \rightarrow \text{Al}(\text{OH})_3(\text{s})$

## Redox Reactions

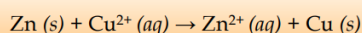
- transfer of a proton = **ACID-BASE reaction**



- formation of a solid = **PRECIPITATION reaction**



- transfer of electrons = **REDOX reaction**



An oxidation-reduction, or **REDOX**, reaction involves the transfer of electrons - causes changes to the **OXIDATION STATE**.

## Oxidation State vs Charge

- Charge = measurable parameter = #protons - #electrons
- Oxidation state = theoretical parameter
  - what if we took away the bonding electrons?

Na <sup>+</sup>	H-Cl
charge = +1	no charge
oxidation state = +1	oxidation state of H = +1 oxidation state of Cl = -1

## Assigning Oxidation States

**TABLE 19.1** Rules for Assigning an Oxidation Number

### General Rules

- For an atom in its elemental form (such as Na, O<sub>2</sub>, and Cl<sub>2</sub>), oxidation number = 0.
- For a monatomic ion, oxidation number = ion charge (with the sign *before* the numeral).
- The sum of the oxidation number values for the atoms in a molecule or formula unit of a compound equals zero. The sum of the oxidation number values for the atoms in a polyatomic ion equals the charge of the ion.

### Rules for Specific Atoms or Periodic Table Groups

- For group 1: Oxidation number = +1 in all compounds
- For group 2: Oxidation number = +2 in all compounds
- For hydrogen: Oxidation number = +1 in combination with nonmetals  
Oxidation number = -1 in combination with metals and boron
- For fluorine: Oxidation number = -1 in all compounds
- For oxygen: Oxidation number = -2 in most cases, unless coupled to a more electronegative centre (such as F) or a group 1 or group 2 metal (in which case, it might be +2 or -1)
- For group 17: Oxidation number = -1 in combination with metals, nonmetals (except O), and other halogens lower in the group



$$\sum \text{OS} = \text{charge}$$

1	+1
2	+2
H	+1
F	-1
O	-2
X	-1

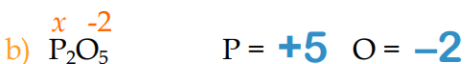
### Tips:

If two rules contradict each other, follow the rule that appears higher in the list.

For a multi-atom species, label the easy oxidation states, and solve for the unknown atoms.

## Deriving Oxidation States

What are the oxidation states of the elements in each of the following?

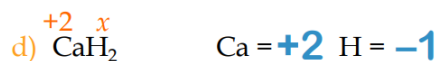


$$2x + 5(-2) = 0$$

$$x = +5$$

1
2
H
F
O
X

What are the oxidation states of the elements in each of the following?



$$(+2) + 2x = 0$$

$$x = -1$$

Ca is **+2**  
H is **+1**

1
2
H
F
O
X

Compounds in which hydrogen has a -1 ox. state are called **HYDRIDES**

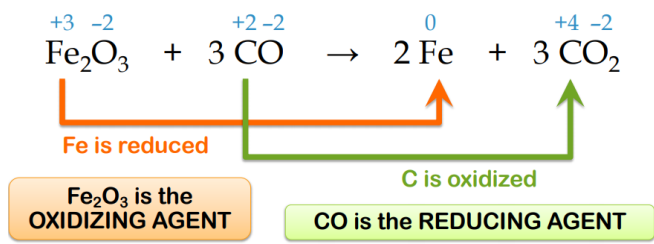
## Redox Reactions

**LEO GER** (Loss Electrons = Oxidation, Gain Electrons = Reduction)

**Oxidizing Agent (Oxidant):** Oxidation state decreases in a redox reaction (**reduced**).

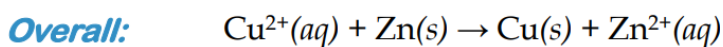
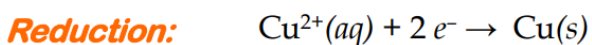
**Reducing Agent (Reductant):** Oxidation state increases in a redox reaction (**oxidized**).

Ex) • We can determine which species is oxidized and which species is reduced by looking at the oxidation states.



## Half Reactions

A redox reaction is represented by two half-reactions



Balancing redox reactions depends on whether it is an **acidic** or **basic** medium...

## Balancing Redox Reactions

### The Half-Equation Method:

The overall reaction is separated into two half-reactions – an oxidation and reduction.

Each half-reaction is balanced.

The balanced half-reactions are then added to give the overall reaction.

Step 1: Write out the net ionic form of the reaction.

Step 2: Separate into two half reactions.

Step 3: Balance all elements except O and H in the half reaction.

Step 4: Add enough molecules of H<sub>2</sub>O to balance the O atoms.

Step 5: Add enough H<sup>+</sup> ions to balance the H atoms.

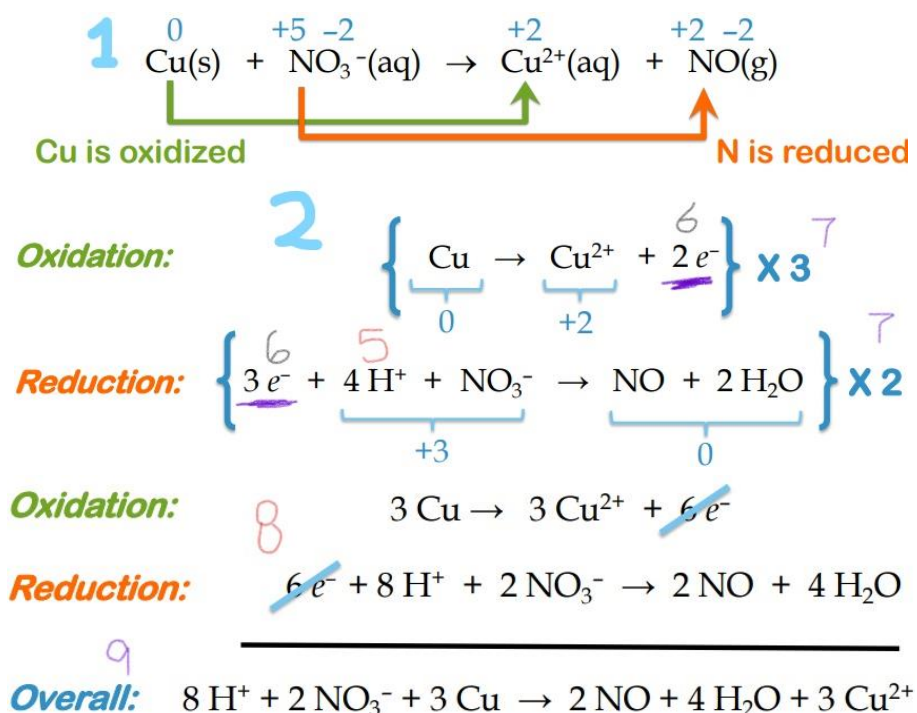
Step 6: Balance the net charge on each side by adding electrons.

Step 7: If necessary, multiply the half-reactions by the lowest common denominator.

Step 8: Add the two half-reactions together and simplify.

Step 9: Verify your final reaction.

### Balancing in Acidic Solution



### Basic Solution

When reaction is basic, **STEP 5** is modified as follows:

For every H<sup>+</sup>, add an OH<sup>-</sup> on each side of the half-reaction.

Where H<sup>+</sup> and OH<sup>-</sup> appear on the same side, combine these two ions to make H<sub>2</sub>O.

## Gas or Vapour?

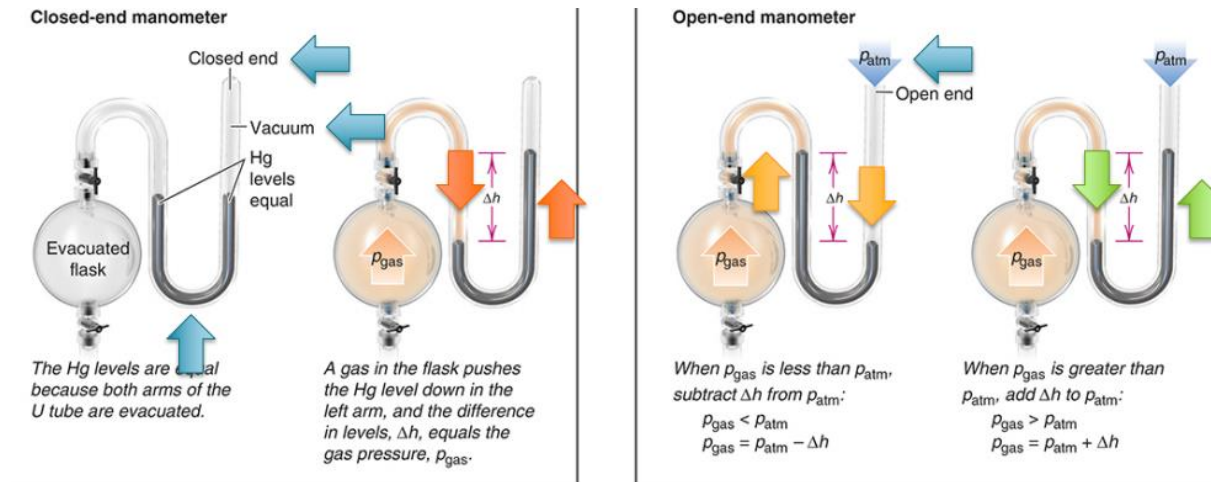
A gas is a substance which is **normally gaseous** under regular temperature and pressure conditions.

A vapour is the gaseous form of a substance which is **normally a liquid** under regular temperature and pressure conditions.

## Concept of Pressure

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

## Manometer



$$\Delta h = P_{\text{gas}} \text{ (in mmHg)}$$

$$P_{\text{gas}} = P_{\text{atm}} - \Delta h$$

$$P_{\text{gas}} = P_{\text{atm}} + \Delta h$$

## Ideal Gas Law

$$PV = nRT$$

## Boyle's Law

$$P_1V_1 = \text{constant} = P_2V_2$$

$$P_1V_1 = P_2V_2$$

## Charles's Laws

$$T \text{ (Kelvin)} = 273.15 + T \text{ (Celsius)}$$

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

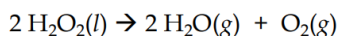
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

## Avogadro's Hypothesis

Equal volumes of gases at the same T and P have the same number of molecules.

$$\frac{V}{n} = \text{constant}$$

## The Gas Law and Stoichiometry



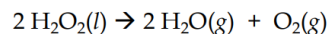
1.1 g of  $\text{H}_2\text{O}_2$  are decomposed in a flask with a volume of 2.50 L.  
What is the pressure of  $\text{O}_2$  at 25°C?

WANT: P ( $\text{O}_2$ ) = ? bar

HAVE: m ( $\text{H}_2\text{O}_2$ ) = 1.1 g  
V = 2.50 L  
T = 25°C = 298 K

NEED: MM of  $\text{H}_2\text{O}_2$  = 34.0 g/mol

$$\begin{aligned} ? \text{ mol O}_2 &= 1.1 \text{ g H}_2\text{O}_2 \times \frac{\cancel{\text{mol H}_2\text{O}_2}}{34.0 \text{ g H}_2\text{O}_2} \times \frac{1 \text{ mol O}_2}{2 \cancel{\text{mol H}_2\text{O}_2}} \\ &= 0.0162 \text{ mol O}_2 \end{aligned}$$



1.1 g of  $\text{H}_2\text{O}_2$  are decomposed in a flask with a volume of 2.50 L.  
What is the pressure of  $\text{O}_2$  at 25°C?

WANT: P ( $\text{O}_2$ ) = ? bar

HAVE: m ( $\text{H}_2\text{O}_2$ ) = 1.1 g  
V = 2.50 L  
T = 25°C = 298 K

NEED: MM of  $\text{H}_2\text{O}_2$  = 34.0 g/mol

$$\begin{aligned} P &= \frac{nRT}{V} = \frac{(0.0162 \text{ mol})(0.083145 \text{ L} \cdot \text{bar} / \text{mol} \cdot \text{K})(298 \text{ K})}{2.50 \text{ L}} \\ &= 0.16 \text{ bar O}_2 \end{aligned}$$

## The Ideal Gas Law

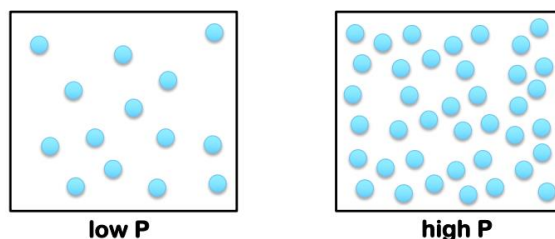
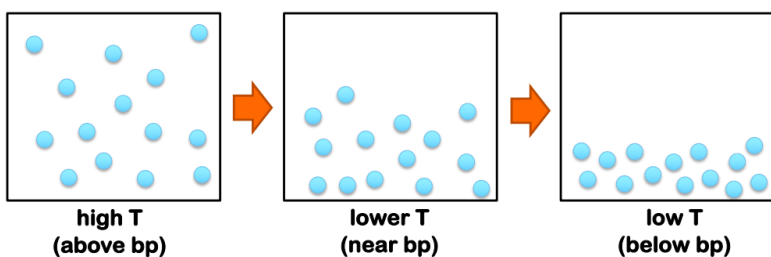
An ideal gas is a gas whose P, V, and T obey the ideal Gas Law

1. No attraction or repulsion between the molecules of an ideal gas [**Breaks down at low T**]
2. The volume occupied by the gas molecules themselves is negligible with respect to the gas container (i.e., the space in the container is empty) [**Breaks down at high P**]

This approximation of an ideal gas works best at **high temperature and low pressure**

1. no attraction or repulsion between the molecules of an ideal gas

2. the volume occupied by the gas molecules themselves is negligible with respect to the gas container (i.e., the space in the container is empty)



## Standard Temperature and Pressure

**STP:** Reference points for normal temperature and pressure conditions

$$P = 1 \text{ bar} = 0.987 \text{ atm} \quad T = 0^\circ\text{C} = 273.15 \text{ K}$$

Experiments revealed that, at STP, the space occupied by one mole of gas is:

$$1 \text{ mol gas} = 22.7 \text{ L} \leftarrow \text{Molar Volume}$$

## Another Form of PV=nRT

In cases where *the number of moles of gas is constant*:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

## Dalton's Law of Partial Pressures

$$P_{\text{total in gas mixture}} = P_A + P_B + \dots$$

**Dalton's Law:** Total P is sum of **PARTIAL** pressures.

The Partial Pressure of a gas, A,  $P_A$ , in a gaseous mixture is:

$$P_A = \frac{n_A RT}{V}$$

The law of partial pressures says that the total pressure of the mixture is given by:

$$P_T = (n_A + n_B + n_C + \dots) \frac{RT}{V} = n_T \frac{RT}{V}$$

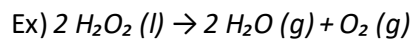
## Molar Fractions

The mole fraction of gas A in a mixture,  $x_A$ , is:

$$x_A = \frac{n_A}{n_T} = \frac{n_A}{n_A + n_B + \dots}$$

Therefore, according to the law of partial pressures:

$$P_A = x_A P_T$$



$$1.1\text{g} \quad 0.32 \text{ bar} \quad 0.16 \text{ bar}$$

$$\text{Therefore, } P_{\text{total}} = 0.16 \text{ bar} + 0.32 \text{ bar} = 0.48 \text{ bar}$$

When Dr. Fox goes scuba diving, she uses NITROX, a special blend of enriched air. The local scuba shop prepares 6.50 L tanks of NITROX by mixing 26.0 g of O<sub>2</sub> with 44.2 g of N<sub>2</sub> at a temperature of 25.0°C. What is the mole fraction and partial pressure of each gas in the mixture?

$$? \text{ mol O}_2 = 26.0 \text{ g O}_2 \times \frac{\text{mol O}_2}{31.998 \text{ g O}_2} = 0.812 \text{ mol O}_2$$

$$? \text{ mol N}_2 = 44.2 \text{ g N}_2 \times \frac{\text{mol N}_2}{28.014 \text{ g N}_2} = 1.578 \text{ mol N}_2$$

$$\text{Total mol } n_T = 0.812 + 1.578 = 2.39 \text{ mol}$$

$$P_T = \frac{n_T RT}{V} = \frac{(2.39 \text{ mol})(0.083145 \text{ L} \cdot \text{bar/mol} \cdot \text{K})(298 \text{ K})}{6.50 \text{ L}} = 9.11 \text{ bar}$$

$$\therefore \chi_{\text{N}_2} = \frac{1.578 \text{ mol}}{2.39 \text{ mol}} = 0.660$$

$$\therefore \chi_{\text{O}_2} = \frac{0.812 \text{ mol}}{2.39 \text{ mol}} = 0.340$$

$$P_{\text{N}_2} = \chi_{\text{N}_2} \times P_T = (0.660)(9.11 \text{ atm}) = 6.0 \text{ bar}$$

$$P_{\text{O}_2} = \chi_{\text{O}_2} \times P_T = (0.340)(9.11 \text{ atm}) = 3.1 \text{ bar}$$

## Kinetic Molecular Theory

In an ideal gas, there are no attractions or repulsions between the gas molecules, therefore the energy of the gas must entirely come from the **kinetic energy** of the individual gas molecules.

The kinetic energy of a molecule depends only on its mass and velocity.

### The Assumptions of the Theory:

1. The volume occupied by the molecules is negligible.
2. There are no attractive or repulsive forces between molecules; all collisions are perfectly elastic; thus the average kinetic energy is constant.
3. Gaseous molecules are in constant, random, straight-line motion in all directions; they collide fleetingly and frequently with each other and the walls of the container.

## Kinetic Energy of Gas Molecules

Kinetic energy of an **object**:  $E_k = \frac{1}{2}mv^2$

**Average kinetic energy** of the ensemble of molecules:  $\overline{E_k} = \frac{1}{2}m\overline{u^2}$

Where  $\overline{u^2}$  is the mean-square speed:

$$\overline{u^2} = \frac{u_1^2 + u_2^2 + \dots + u_N^2}{N}$$

More Equations:

$$\overline{E_k} = \frac{3}{2} \left( \frac{R}{N_A} \right) T$$

$$\overline{E_k} = \frac{1}{2} m \overline{u^2}$$

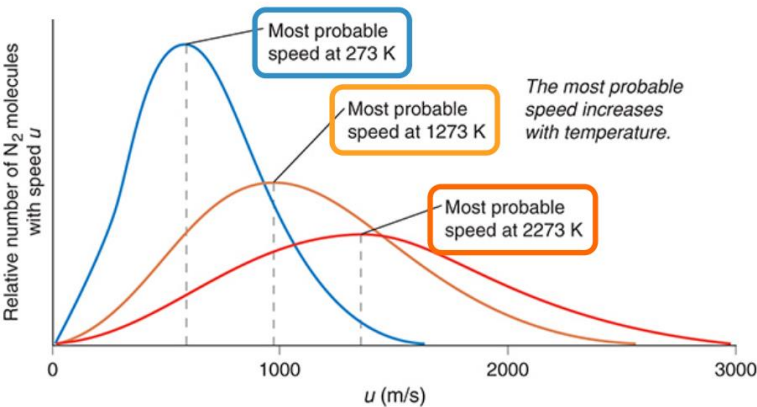
At constant m:

$$T \uparrow \overline{E_k} \uparrow u \uparrow$$

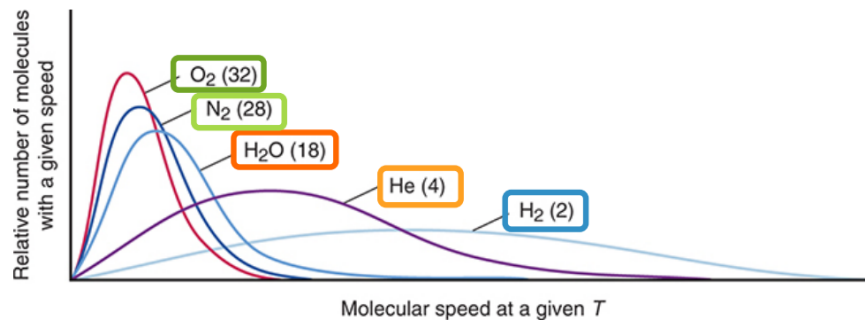
At constant T:

$$m \uparrow u \downarrow$$

# Effect of Temperature



# Effect of Mass



At constant  $m$ :  $T \uparrow \bar{E}_K \uparrow u \uparrow$

At constant  $T$ :  $m \uparrow u \downarrow$

**Root-mean-square Speed,  $u_{rms}$**

$$\sqrt{u^2} = u_{rms} = \sqrt{\frac{3RT}{MM}}$$

The root-mean-square,  $u_{rms}$ , **increases when  $T$  increases or  $MM$  decreases.**

Usually,  $u_{rms}$  is in m/s, so  **$MM$  must be in kg/mol** (SI) and  **$R$  is 8.3145** Pa m<sup>3</sup> / mol•K (SI).

## Gas Diffusion and Effusion

**Effusion:** Movement of molecules through a small hole into an empty container.

**Diffusion:** Gradual mixing of molecules of different gases.

## Gas Effusion

Graham's Law govern effusion of gas molecules:

$$\frac{\text{effusion of A}}{\text{effusion of B}} = \frac{(u_{rms})_A}{(u_{rms})_B} = \sqrt{\frac{MM_B}{MM_A}}$$

Rate of effusion is inversely proportional to the square root of the gas particle's molar mass.

## Deviations from Ideal Gas Law

We account for volume of molecules and intermolecular forces with **VAN DER WAAL'S EQUATION**:

$$\left(P + \frac{n^2 a}{V^2}\right)(V - nb) = nRT$$

$a$  &  $b$  must be determined experimentally.

Correction for intermolecular forces

Correction for volume