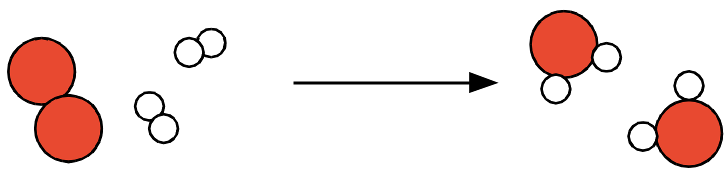
The Mole Concept

A. Atomic Masses and Avogadro's Hypothesis

1. We have learned that compounds are made up of two or more different elements and that elements are composed of atoms. Therefore, compounds must be composed of molecules made up of two more different kinds of atoms. During a chemical reaction, the atoms that make up the starting materials rearrange to form new and different molecules



The question that arises however, is how many atoms and molecules are involved in the reaction or **how much of one element will combine with another element?** In addition, **since atoms are so small, how can we count them?**

2. Early experimental work by English chemist **John Dalton (1766–1844)** was concerned with how much of one element could combine with a given amount of another element.

Dalton assumed that elements formed simple compounds in a 1:1 ratio and he assigned an **ARBITRARY MASS** to each element. Since hydrogen was the lightest element, it was assigned a mass of "1" and all other elements were compared to that of hydrogen.

Carbon was found to be 6 times heavier than hydrogen so it was assigned a mass of 6. Oxygen was found to be 16 times heavier than hydrogen so it was assigned a mass of 16.

In this way, Dalton was able to calculate the **"RELATIVE MASSES"** for several elements.

3. Dalton's atomic mass scale was partly in error because not all the molecules he studied actually contained only one atom of each element. During the time that Dalton's mass scale was just being introduced, the French chemist **Joseph Gay-Lussac** began to study how gases reacted. When Gay-Lussac reacted pairs of gases at the same temperature and pressure, he found that gases combined in simple whole number ratios.

1 L of hydrogen gas reacts with 1 L of chlorine gas to make 2 L of $HCl_{(g)}$

1 L of nitrogen reacts with 3 L of hydrogen gas to make 2 L of $NH_{3(g)}$

2 L of CO gas react with 1 L of oxygen gas to make 2 L of $CO_{2(g)}$

By itself, Gay-Lussac's findings did not seem to be related to atomic mass but then the Italian chemist **Amadeo Avogadro** proposed the following explanation for Gay-Lussac's data.

AVOGADRO'S HYPOTHESIS

Equal volumes of different gases, at the same temperature and pressure, contain the same number of particles.

In other words, if **1 L of gas A** reacts with **1 L of gas B**, then there exactly the same number of particles of A and B present. Therefore, the molecule formed by reacting A with B is **AB**.

Similarly, if **2 L of gas A** reacts with **1 L of gas B**, the molecules formed have the formula **A_2B**.

B. The Mole

1. Avogadro's hypothesis allows us to predict the formula of a compound by determining the ratio of the volumes of gases needed to make the compound.

   e.g. If 1 L of nitrogen reacts with 3 L of hydrogen to form ammonia, then its formula is NH_3.

   If 2 L of hydrogen reacts with 1 L of oxygen to form water, then its formula is H_2O.

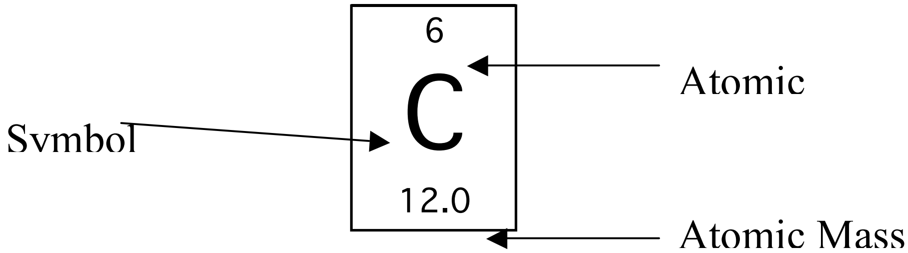
   So if we want to make a particular compound, all we need to do is react volumes of gases in the ratio given by their formulas **BUT** how do we determine how much of one element reacts with another element when they are not gases?

   e.g. How much iron is required to react with sulphur to produce iron (II) sulphide, FeS, so that neither element is left over?

   Since the easiest way to measure solids is to measure their mass, we need to relate mass to the number of atoms.

2. The periodic table shows us the **relative masses of the elements**. Its units are "u" which stands for "unified atomic mass units". Unlike Dalton's mass scale, the present day scale is not based on hydrogen. Instead, **1 u is defined as $^1/_{12}$ the mass of carbon–12** (carbon–12, ^{12}C, is a particular isotope of carbon).

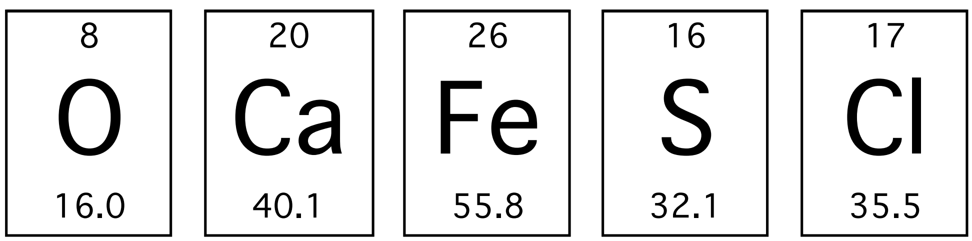
> A **MOLE** is the number of carbon atoms in exactly **12 g of carbon**.
>
> **MOLAR MASS** is the mass of one mole of particles.



one **"C" atom** has a mass of **12.0 u**

one **MOLE of "C" atoms** has a mass of **12.0 g**

The periodic table gives us the **molar mass of each of the elements expressed in grams**.



Finding the molar mass of a compound simply involves adding together the molar masses of each of the atoms that make up the compound (remember the units of molar mass are grams).

EXAMPLE V.1	**CALCULATING MOLAR MASSES**
Problem:	What is the molar mass of iron (III) sulphate?
Solution:	Iron (III) sulphate is $Fe_2(SO_4)_3$ To calculate the molar mass we need to add the masses of $2\ Fe + 3\ S + 12\ O$ $2\ (55.8\ g) + 3\ (32.1\ g) + 12\ (16.0\ g) = \mathbf{399.9\ g}$ Alternatively, $2\ Fe + 3\ (SO_4)$ $2\ (55.8\ g) + 3\ (32.1 + 3\ (16.0\ g))$ $2\ (55.8\ g) + 3\ (96.1\ g) = \mathbf{399.9\ g}$

SAMPLE PROBLEM V.1	**CALCULATING MOLAR MASSES**
Problem:	Calculate the molar masses of (a) $Na_2Cr_2O_7$ (b) Ag_2SO_4 (c) $Pb_3(PO_4)_4$
Solution:	(a) $2\ (23.0\ g) + 2\ (52.0\ g) + 7\ (16.0\ g) = \mathbf{262.0\ g}$ (b) $2\ (107.9\ g\ ) + 32.1\ g + 4\ (16.0\ g) = \mathbf{311.9\ g}$ (c) $3\ (207.2\ g) + 4\ (31.0\ g + 4\ (16.0\ g)) = \mathbf{1001.6\ g}$

C. Relating Moles to Mass, Volume of Gas, and Number of Particles

1. The molar mass of a compound allows us to calculate the mass of a given number of moles of a substance and the number of moles in a given mass of a substance.

   Since we know that

1 mol of "X" has a mass of **(molar mass of "X") g**

   we have two conversion factors:

$$\frac{1\text{ mol}}{(\text{molar mass of "X"})\text{ g}} \quad \text{or} \quad \frac{(\text{molar mass of "X"})\text{ g}}{1\text{ mol}}$$

EXAMPLE V.2	**RELATING MASS AND MOLES**
Problem:	(a) What is the mass of 3.25 mol of CO_2? (b) What is the mass of 1.36×10^{-3} mol of SO_3? (c) How many moles of N_2 are there in 50.0 g of N_2? (d) How many moles of CH_3OH are there in 0.250 g of CH_3OH?
Solution:	(a) 1 mol CO_2 = 44.0 g mass $CO_2 = 3.25\text{ mol} \times \frac{44.0\text{ g}}{1\text{ mol}} = \mathbf{143\ g}$ (b) 1 mol SO_3 = 80.1 g mass $SO_3 = 1.36 \times 10^{-3}\text{ mol} \times \frac{80.1\text{ g}}{1\text{ mol}} = \mathbf{0.109\ g}$

	(c) 1 mol N_2 = 28.0 g mol $N_2 = 50.0\text{ g} \times \frac{1\text{ mol}}{28.0\text{ g}}$ = **1.79 mol** (d) 1 mol CH_3OH = 32.0 g mol $N_2 = 0.250\text{ g} \times \frac{1\text{ mol}}{32.0\text{ g}}$ = **7.81 x 10^{-3} mol**

SAMPLE PROBLE MS V.2	RELATING MASS AND MOLES
Problem:	(a) What is the mass of 0.834 mol of $FeSO_4$? (b) What is the mass of 2.84 x 10^{-2} mol of Na_3N? (c) How many moles of CH_4 are there in 27.5 g of CH_4? (d) How many moles of $Ca(NO_3)_2$ are there in 35.0 g of $Ca(NO_3)_2$?
Solution:	(a) 1 mol $FeSO_4$ = 151.9 g mass $FeSO_4 = 0.834\text{ mol} \times \frac{151.9\text{ g}}{1\text{mol}}$ = **127 g** (b) 1 mol Na_3N = 83.0 g mass $Na_3N = 2.84 \times 10^{-2}\text{ mol} \times \frac{83.0\text{ g}}{1\text{ mol}}$ = **2.36 g** (c) 1 mol CH_4 = 16.0 g mol $CH_4 = 27.5\text{ g} \times \frac{1\text{ mol}}{16.0\text{ g}}$ = **1.72 mol**

	(d) 1 mol $Ca(NO_3)_2$ = 164.1 g mol $Ca(NO_3)_2 = 35.0\text{ g} \times \frac{1\text{ mol}}{164.1\text{ g}}$ = **0.213 mol**

2. Calculations involving gas volumes are simplified by Avogadro's hypothesis. Recall that

> **AVOGADRO'S HYPOTHESIS**
>
> Equal volumes of different gases, at the same temperature and pressure, contain the same number of particles.

> The **MOLAR VOLUME** of a gas is the volume occupied by one mole of the gas.

Since the volume of a gas is drastically affected by the **temperature** and **pressure** we need to define **STANDARD** conditions.

> **STANDARD TEMPERATURE AND PRESSURE (STP)** = **0 °C** and **101.3 kPa**

Avogadro's hypothesis can be interpreted to mean that all gas samples with the same temperature, pressure, and numbers of particles occupy the same volume. This can be re–stated as **equal numbers of moles of any gas at STP occupy the same volume**.

Experimentally, it is determined that

> **1 mol** of any gas at STP has a volume of **22.4 L**

In other words, the **MOLAR VOLUME of any gas at STP is 22.4 L**. We can obtain two conversion factors:

$$\frac{1\ \text{mol}}{22.4\ \text{L}} \quad \text{or} \quad \frac{22.4\ \text{L}}{1\ \text{mol}}$$

EXAMPLE V.3	**RELATING VOLUME OF A GAS AND MOLES**
Problem:	(a) How many moles of gas are contained in a balloon with a volume of 10.0 L at STP? (b) What volume will 0.250 mol of CO_2 occupy at STP?
Solution:	(a) mol of gas = 10.0 L x $\frac{1\ \text{mol}}{22.4\ \text{L}}$ = **0.446 mol** (b) volume of CO_2 = 0.250 mol x $\frac{22.4\ \text{L}}{1\ \text{mol}}$ = **5.60 L**

3. The mole is the fundamental unit in chemistry for measuring the amount of substance or the "number of particles of a substance". In a sense, the mole is simply a counting number. Just as a

dozen = 12

experimentally,

$$\boxed{1\ \text{mol} = 6.02 \times 10^{23}}$$

This value, 6.02×10^{23}, is called Avogadro's number. Notice that there are no units in the same way a "dozen" stands for "12".

Conversion factors:

$$\frac{1\ \text{mol particles}}{6.02 \times 10^{23}\ \text{particles}} \quad \text{or} \quad \frac{6.02 \times 10^{23}\ \text{particles}}{1\ \text{mol particles}}$$

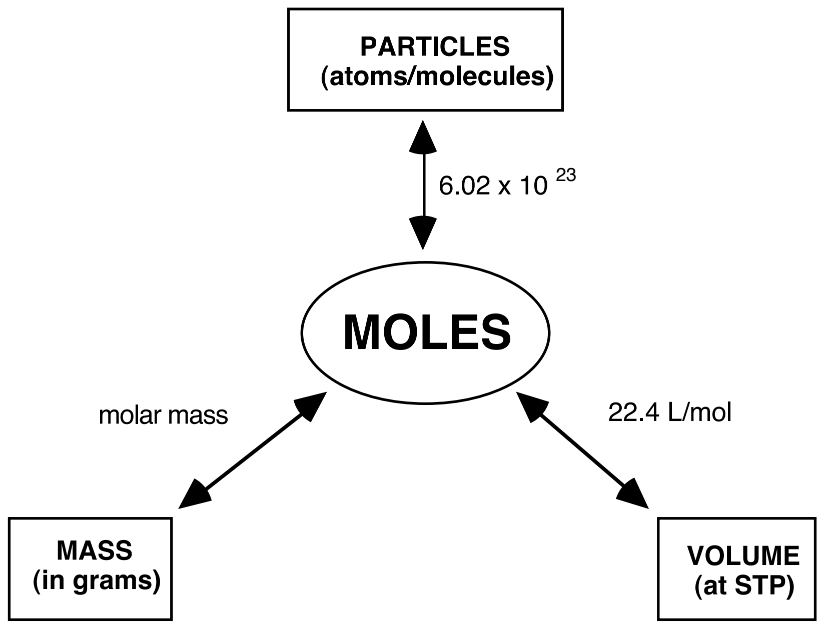
EXAMPLE V.4	**RELATING NUMBER OF PARTICLES AND MOLES**
Problem:	(a) How many molecules are there in 0.125 mol of molecules? (b) How many moles of N are there in 5.00×10^{17} N atoms? (c) How many atoms are in 5 molecules of $CuSO_4 \cdot 5H_2O$?
Solution:	(a) molecules = $0.125 \text{ mol} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol molecules}}$ $= \mathbf{7.53 \times 10^{22}}$ **molecules** (b) moles $N_2 = 5.00 \times 10^{17} \text{ atoms} \times \frac{1 \text{ mol atoms}}{6.02 \times 10^{23} \text{ atoms}}$ $= \mathbf{8.31 \times 10^{-7}}$ **mol** (c) atoms = 5 molecules $\times\ 21 \frac{\text{atoms}}{\text{molecule}}$ = **105 atoms**

SAMPLE PROBLEMS V.3	**RELATING VOLUME OF GAS / NUMBER OF PARTICLES AND MOLES**
Problem:	(a) How many moles of gas are contained in a balloon with a volume of 17.5 L at STP? (b) What volume of gas will 0.074 mol of gas occupy at STP? (c) How many atoms are there in 0.0185 mol of atoms? (d) How many moles of Fe_2O_3 are there in 8.75×10^{20} Fe_2O_3 molecules? (e) How many atoms of H are in 30 molecules of $Ca(H_2PO_4)_2$?
Solution:	(a) mol of gas = $17.5\ \text{L} \times \frac{1\ \text{mol}}{22.4\ \text{L}}$ = **0.781 mol** (b) volume of gas = $0.074\ \text{mol} \times \frac{22.4\ \text{L}}{1\ \text{mol}}$ = **1.7 L** (c) atoms = $0.0185\ \text{mol} \times \frac{6.02 \times 10^{23}\ \text{atoms}}{1\ \text{mol atoms}}$ = **1.11×10^{22} atoms** (d) moles Fe_2O_3 = $8.75 \times 10^{20} \times \frac{1\ \text{mol}}{6.02 \times 10^{23}}$ = **1.45×10^{-3} mol** (e) H atoms = $30\ \text{molecules} \times 4\ \frac{\text{atoms}}{\text{molecule}}$ = **120 H's**

4. All of the previous problems have involved single–step conversions between moles, mass, volume, or number of particles. The following is a summary of the conversion factors needed:

CONVERSION	CONVERSION FACTOR
MOLES ←→ NUMBER OF PARTICLES	$\frac{6.02 \times 10^{23} \text{ particles}}{1 \text{ mol}}$ or $\frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ particles}}$
MOLES ←→ MASS	$\frac{\textit{(molar mass)} \text{ g}}{1 \text{ mol}}$ or $\frac{1 \text{ mol}}{\textit{(molar mass)}}$
MOLES ←→ VOLUME (gases @ STP)	$\frac{22.4 \text{ L}}{1 \text{ mol}}$ or $\frac{1 \text{ mol}}{22.4 \text{ L}}$
MOLECULES ←→ ATOMS	$\frac{\textit{(atom count)} \text{ atoms}}{1 \text{ mol}}$ or $\frac{1 \text{ mol}}{\textit{(atom count)} \text{ atoms}}$

The following flow chart will help to simply calculations that involve multiple conversions.



EXAMPLE V.5	**MOLE CALCULATIONS INVOLVING MULTIPLE CONVERSIONS**
Problem:	(a) What is the volume occupied by 50.0 g of $NH_{3(g)}$ at STP? (b) What is the mass of 1.00 x 10^{12} atoms of Cl? (c) How many oxygen atoms are contained in 75.0 L of $SO_{3(g)}$ at STP?
Solution:	(a) **MASS → MOLES → VOLUME** $? \text{ L} = 50.0 \text{ g} \times \frac{1 \text{ mol}}{17.0 \text{ g}} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = \mathbf{65.9 \text{ L}}$

	(b) **ATOMS → MOLES → MASS** $? \text{ g} = 1.00 \times 10^{12} \text{ atoms g} \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ atoms}} \times \frac{35.5 \text{ g}}{1 \text{ mol}}$ $= \mathbf{5.90 \times 10^{-11} \text{ g}}$ (c) **VOLUME → MOLES → MOLECULES → ATOMS of O** $? \text{ O's} = 75.0 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} \times \frac{6.02 \times 10^{23} \text{ SO}_3}{1 \text{ mol}} \times \frac{3 \text{ O's}}{\text{SO}_3}$ $= \mathbf{6.05 \times 10^{24} \text{ O's}}$

4. So far, all the volumes have been of gases at STP. If **DENSITY** is mentioned at any point in a problem, recall that $\mathbf{d = {}^{m}/_{V}}$

- **If the volume of a solid or liquid is the unknown**, calculate the volume from $\left(V = {}^{m}/_{d}\right)$. (Note you cannot use the molar volume of a gas, 22.4 L, when calculating the volume of a solid or liquid.)

- **If the density is unknown**, you will need both mass and volume to calculate density from $\left(d = {}^{m}/_{V}\right)$.

- **If the number of moles is unknown**, use the density and volume to calculate mass from (m = d • V) and then convert mass to moles.

- **If the density of gas is unknown**, it can be calculated by the equation:

  $$\left(d = \frac{\textit{(molar mass)}}{\textit{(molar volume)}}\right)$$

EXAMPLE V.6	**DENSITY AND MOLE CALCULATIONS**
Problem:	(a) What is the volume occupied by 3.00 mol of ethanol, C_2H_5OH? (d = 0.790 g/mL) (b) How many moles of $Hg_{(l)}$ are contained in 100 mL of $Hg_{(l)}$? (d = 13.6 g/mL) (c) What is the density of $O_{2(g)}$ at STP?
Solution:	(a) We want volume of liquid so V = $^{m}/_{d}$, and mass can be related to moles, so $? \text{ mL} = 3.00 \text{ mol} \times \frac{46.0 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ mL}}{0.790 \text{ g}} = \mathbf{175 \text{ mL}}$ (b) To get moles, we need to calculate mass first and then convert to moles. $? \text{ mol} = 100 \text{ mL} \times \frac{13.6 \text{ g}}{\text{mL}} \times \frac{1 \text{ mol}}{200.6 \text{ g}} = \mathbf{6.78 \text{ mol}}$ (c) There are no numbers given but the term STP suggests a volume of 22.4 L for one mole of gas and the mass of one mole of $O_{2(g)}$ is 32.0 g. $\text{density} = \frac{m}{V} = \frac{\text{mass of 1 mol } \textit{(molar mass)}}{\text{volume of 1 mol } \textit{(molar volume)}}$ $= \frac{32.0 \text{ g}}{22.4 \text{ L}} = \mathbf{1.43 \text{ g/L}}$

EXAMPLE V.7	**MORE DENSITY CALCULATIONS**
Problem:	(a) A 2.50 L bulb contains 4.91 g of a gas at STP. What is the molar mass of the gas? (b) $Al_2O_{3(s)}$ has a density of 3.97 g/mL. How many atoms of Al are in 100 mL of Al_2O_3?
Solution:	(a) Density can be found from d = $^{m}/_{V}$, so $d = \frac{4.91\ g}{2.50\ L} = \mathbf{1.96\ g/L}$ (b) Change volume to mass, then to moles, then to molecules, and finally atoms. $? \text{ Al} = 100 \text{ mL} \times \frac{3.96 \text{ g}}{\text{mL}} \times \frac{1 \text{ mol}}{102.0 \text{ g}} \times \frac{6.02 \times 10^{23}\ Al_2O_3}{1 \text{ mol}} \times \frac{2 \text{ atoms Al}}{Al_2O_3}$ $= \mathbf{4.69 \times 10^{24}\ atoms\ Al}$

C. Percentage Composition

1. The **PERCENTAGE COMPOSITION** is the percentage (by mass) of each element or species in a chemical formula.

EXAMPLE V.8	PERCENTAGE COMPOSITION
Problem:	(a) What is the percentage composition of H_2SO_4? (b) What is the percentage composition of water in $CuSO_4 \cdot 5H_2O$?
Solution:	(a) Assume that there is 1 mole of the compound. Molar mass = 98.1 g Total mass of H = 2 x 1.0 g = 2.0 g Total mass of S = 1 x 32.1 g = 32.1 g Total mass of O = 4 x 16.0 g = 64.0 g $\% \text{ H} = \frac{2.0 \text{ g}}{98.1 \text{ g}} \times 100\% = \mathbf{2.0\%}$ $\% \text{ S} = \frac{32.1 \text{ g}}{98.1 \text{ g}} \times 100\% = \mathbf{32.7\%}$ $\% \text{ O} = \frac{64.0 \text{ g}}{98.1 \text{ g}} \times 100\% = \mathbf{65.2\%}$ (b) Assume that there is 1 mole of the compound. Molar mass = 249.6 g Total mass of H_2O = 5 x 18.0 g = 90.0 g $\% \text{ H}_2\text{O} = \frac{90.0 \text{ g}}{249.6 \text{ g}} \times 100\% = \mathbf{36.1\%}$

D. Empirical and Molecular Formulas

1. The empirical formula is called the simplest formula and is the smallest whole-number ratio of atoms which represents the molecular composition of a species.

   CH_2, C_2H_4, C_3H_6, C_4H_8, and C_5H_{10} all contain twice as many H's as there are C's. The empirical formula (simplest formula) for all of these molecules is $\mathbf{CH_2}$.

   Finding empirical formula is essentially the opposite of determining percentage composition.

EXAMPLE V.9	EMPIRICAL FORMULAS
Problem:	(a) What is the empirical formula of a compound consisting of 80.0% C and 20.0% H? (b) A compound contains 58.5 % C, 7.3% H, and 34.1 % N. What is the empirical formula of the compound?
Solution:	(a) Assume 100.0 g of the compound mass of C = 80.0% x 100.0 g = 80.0 g mass of H = 20.0% x 100.0 g = 20.0 g Use the mass to determine the moles of each element $\text{mole C} = 80.0\text{ g} \times \frac{1\text{ mol}}{12.0\text{ g}} = 6.67\text{ mol}$ $\text{mole H} = 20.0\text{ g} \times \frac{1\text{ mol}}{1.0\text{ g}} = 20\text{ mol}$ Determine smallest ratio by dividing by smallest

	number of moles (÷ 6.67) $C = \frac{6.67\ mol}{6.67\ mol} = 1\ C$ $H = \frac{20\ mol}{6.67\ mol} = 3\ H$ Empirical Formula is $\mathbf{CH_3}$ (b) Assume 100.0 g of the compound mass of C = 58.5 g, mass of H = 7.3 g, mass of N = 34.1 g Use the mass to determine the moles of each element $\text{mole C} = 58.5\ g \times \frac{1\ mol}{12.0\ g} = 4.88\ mol$ $\text{mole H} = 7.3\ g \times \frac{1\ mol}{1.0\ g} = 7.3\ mol$ $\text{mole N} = 34.1\ g \times \frac{1\ mol}{14.0\ g} = 2.44\ mol$ Determine smallest ratio by dividing by smallest number of moles (÷ 2.44) $C = \frac{4.88\ mol}{2.44\ mol} = 2\ C$ $H = \frac{7.3\ mol}{2.44\ mol} = 2.99 \approx 3\ H$ $N = \frac{2.44\ mol}{2.44\ mol} = 1\ N$ Empirical Formula is $\mathbf{C_2H_3N}$

EXAMPLE V.10	**MORE EMPIRICAL FORMULAS**
Problem:	What is the empirical formula of a compound consisting of 81.8% C and 18.2% H?
Solution:	Assume 100.0 g of the compound mass of C = 81.8 g, mass of H = 18.2 g Use the mass to determine the moles of each element $\text{mole C} = 81.8\text{ g} \times \frac{1\text{ mol}}{12.0\text{ g}} = 6.82\text{ mol}$ $\text{mole H} = 18.2\text{ g} \times \frac{1\text{ mol}}{1.0\text{ g}} = 18.2\text{ mol}$ Determine smallest ratio by dividing by smallest number of moles (÷ 6.82) $C = \frac{6.82\text{ mol}}{6.82\text{ mol}} = 1\text{ C}$ $H = \frac{18.2\text{ mol}}{6.82\text{ mol}} = 2.67\text{ H}$ **DO NOT** just round off ratio, you must multiply both number by 2, 3, 4, or 5 until both are whole numbers Multiplying both numbers by 3 gives whole numbers 3C : 7 H Empirical Formula is $\mathbf{C_3H_7}$

ALWAYS carry out calculations to **3 or 4 digits** and **NEVER** round off intermediate values. Improper round-off calculations will cause you to multiply by the wrong number when trying to obtain whole numbers.

2. The molecular formula can be found be using the molar mass of the empirical formula; that is, the **EMPIRICAL MASS**.

   The molecular formula is made up of whole number multiples of the empirical formula.

   CH_2, C_2H_4, C_3H_6, C_4H_8, and C_5H_{10} all have the same empirical formula CH_2.

   The whole number multiple (N) is given by the formula

$$\text{Multiple} = N = \frac{\text{molar mass}}{\text{empirical mass}}$$

molecular formula = N x (empirical formula)

3. It may be necessary to calculate the molar mass from information that is given in the question.

   (a) ***Finding molar mass from density of a gas at STP***

   If: density of gas "X" = 1.43 $^{g}/_{L}$ (at STP)

   Then: molar mass of "X" = 1.43 $^{g}/_{L}$ x 22.4 $^{L}/_{mol}$ = **32.0 $^{g}/_{mol}$**

   (b) ***Finding molar mass from mass and volume of a gas at STP***

   If you are told: "0.0425 L of gas 'X' at STP has a mass of 0.135 g"

   Then: density of gas "X" = $\frac{0.135\ g}{0.0425\ L}$ = 3.176 $^{g}/_{L}$

   And: molar mass of "X" = 3.176 $^{g}/_{L}$ x 22.4 $^{L}/_{mol}$ = **71.2 $^{g}/_{mol}$**

(c) ***Finding molar mass from mass and a given number of moles***

If you are told: "0.0250 mol of 'X' has a mass of 1.775 g"

Then: molar mass = $\frac{1.775\text{ g}}{0.0250\text{ mol}}$ = **71.0 $^{g}/_{mol}$**

(d) ***Finding molar mass from of the molar mass if given as a multiple of a known molar mass***

If you are told: "X" has a molar mass which is 1.64 times that of CO_2

Then: molar mass of CO_2 = 44.0 $^{g}/_{mol}$

And: molar mass of "X" = 1.64 x 44.0 $^{g}/_{mol}$ = 72.2 $^{g}/_{mol}$

EXAMPLE V.11	**MOLECULAR FORMULA**
Problem:	The empirical formula of a compound is SiH_3. If 0.0275 mol of the compound has a mass of 1.71 g, what is the compound's molecular formula?
Solution:	Empirical mass Molar mass of SiH_3 = 28.1 g + 3 (1.0 g) = 31.1 g Molar mass of compound Molar mass = $\frac{1.71\text{ g}}{0.0275\text{ mol}}$ = 62.2 $^{g}/_{mol}$ Molecular formula $N = \frac{62.2\text{ g}}{31.1\text{ g}} = 2$ 2 x (SiH_3) = Si_2H_6

E. Molar Concentrations

1. **SOLUTIONS** are homogeneous mixtures in which the substances are so thoroughly mixed that they cannot be distinguished from one another. Most solutions contain a solid **(solute)** dissolved in a liquid **(solvent)**; however, there are solutions of gases as well.

   The **CONCENTRATION** of a substance in solution provides a way to find how much of the substance exists in a given volume of the solution. Chemists use the "mole" to describe the amount of substance in a solution.

   MOLAR CONCENTRATION or **MOLARITY** of a substance is the number of moles of the substance contained in **1 L of solution**.

   e.g. If 2.0 L of solution contain 5.0 mol of NaCl, what is the molarity of the NaCl?

$$\text{molar concentration} = \frac{5.0\ \text{mol}}{2.0\ \text{L}} = 2.5\ \frac{\text{mol}}{\text{L}}$$

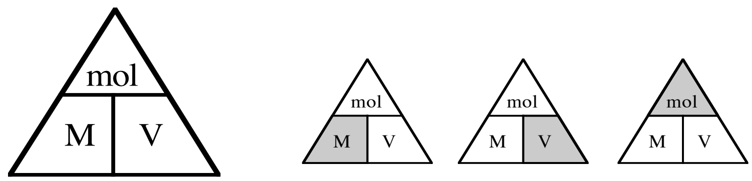
$$\text{The unit symbol for } \frac{\text{mol}}{\text{L}} \text{ is “M”}$$

   When expressed in words, the unit symbol **"M"** is written as **"molar"**.

   The short-hand symbol for **"molar concentration of …"** is a set of brackets: **[…]**, so **[NaCl]** means the **"molar concentration of NaCl."**

2. The definition of molar concentration leads to the following equation:

$$M = \frac{mol}{V}$$



M = molar concentration, in mol/L

mol = number of moles

V = volume, in litres

EXAMPLE V.12	**CALCULATING CONCENTRATION**
Problem:	What is the [NaCl] in a solution containing 5.12 g of NaCl in 250.0 mL of solution?
Solution:	In order to find molarity (M), the moles, and volume are needed. The volume is given and the mass must be converted to moles. moles of NaCl = 5.12 g x $\frac{1 \text{ mol}}{58.5 \text{ g}}$ = 0.0875 mol [NaCl] = M = $\frac{\text{mol}}{\text{V}}$ = $\frac{0.0875 \text{ mol}}{0.2500 \text{ L}}$ = **0.350 M**

EXAMPLE V.13	**CALCULATING MASS CONTAINED IN SOLUTIONS**
Problem:	What mass of NaOH is contained in 3.50 L of 0.200 M NaOH?
Solution:	The molarity (M) and volume (V) are given so moles can be found. Moles can then be converted to mass. solving $M = \frac{mol}{V}$ for n gives mol = M x V moles of NaOH = $0.200\ \frac{mol}{L} \times 3.50\ L = 0.700\ mol$ mass of NaOH = $0.700\ mol \times \frac{40.0\ g}{mol} = \mathbf{28.0\ g}$

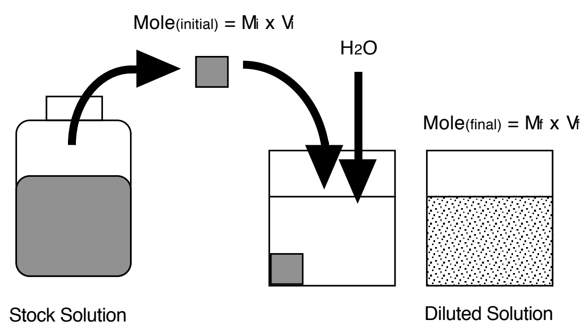
EXAMPLE V.14	**CALCULATING CONCENTRATION FROM DENSITY**
Problem:	What is the molarity of pure sulphuric acid, H_2SO_4, having a density of 1.839 g/mL?
Solution:	Since both density and molarity both have units of amount/volume, $\text{density} = \frac{\text{amount (as mass)}}{\text{volume}}$ and $\text{molarity} = \frac{\text{amount (as moles)}}{\text{volume}}$ therefore all we need to do is convert from mass to moles using molar mass. $[H_2SO_4] = \frac{1.839\ g}{0.001\ L} \times \frac{1\ mol}{98.1\ g} = \mathbf{18.7\ M}$

F. Dilution Calculations

1. When two solutions are mixed, the resulting mixture has a volume and total number of moles equal to the sum of the individual volumes and individual number of moles of chemical found in the separate solutions.

$$\text{molarity of mixture} = \frac{(\text{total moles of chemical})}{(\text{total volume of mixture})}$$

Consider the following dilution,



since the number of moles of solute in the initial solution are equal to the number of moles of solute in the diluted solution,

$$M_1 \text{ x } V_1 = M_2 \text{ x } V_2$$

this equation can be rearranged to give the following dilution equation

$$M_2 = M_1 \text{ x } \frac{V_1}{V_2}$$

EXAMPLE V.15	**SIMPLE DILUTION CALCULATIONS**
Problem:	If 200.0 mL of 0.500 M NaCl is added to 300.0 mL of water, what is the resulting [NaCl] in the mixture.
Solution:	$$M_2 = M_1 \times \frac{V_1}{V_2}$$ $$[NaCl] = 0.500\ M \times \frac{200.0\ mL}{(200.0 + 300.0)\ mL} = \mathbf{0.200\ M}$$

EXAMPLE V.16	**MAKING DILUTE SOLUTIONS FROM CONCENTRATED SOLUTIONS**
Problem:	What volume of 6.00 M HCl is needed to make 2.00 L of 0.125 M HCl?
Solution:	Since $$M_1 \times V_1 = M_2 \times V_2$$ then, $$V_1 = \frac{M_2 \times V_2}{M_1}$$ $$V_{HCl} = \frac{0.125\ M \times 2.00\ L}{6.00\ M} = \mathbf{0.0417\ L}$$

2. When two solutions having different concentrations of the same chemical are mixed, find the moles of solute in each solution and then add the results. Then divide by the final volume to get the overall concentration of the mixture.

EXAMPLE V.17	**MIXING SOLUTIONS OF DIFFERENT CONCENTRATION**
Problem:	If 300.0 mL of 0.250 M NaCl is added to 500.0 mL of 0.100 M NaCl, what is the resulting [NaCl] in the mixture?
Solution:	Calculate the moles of solute from each solution: $$M = \frac{\text{moles}}{\text{volume}}$$ $$\text{moles} = M \times V$$ $$\text{moles}_1 = 0.250\ M \times 0.300\ L = 0.075\ \text{moles}$$ $$\text{moles}_2 = 0.100\ M \times 0.500\ L = 0.050\ \text{moles}$$ $$\text{Total moles} = 0.075\ \text{mol} + 0.050\ \text{mol} = 0.125\ \text{mol}$$ $$M = \frac{0.125\ \text{mol}}{0.800\ L} = 0.156\ M$$