## Atomic mass:

- $p^+$ = protons; $e^-$ = electrons; $n^0$ = neutrons
- $p^+ + n^0$ = atomic mass
  - For carbon-12, $6p^+ + 6n^0$ = atomic mass of 12.0
  - For chlorine-35, $17p^+ + 18n^0$ = atomic mass of 35.0
- atomic mass - $p^+ = n^0$
  - Li-7 ? $n^0$ 7(atomic mass) – 3 $p^+$(atomic number) = $4n^0$
- atomic mass - $n^0 = p^+$
  - Al-27 ? $p^+$ 27 – 14 = 13*

    *Or, just find the atomic number for Al from the Periodic Table.

## Isotopes:

- Element of <u>*same atomic number*</u>, but ***different atomic mass***
  - Example: Hydrogen
    - H-1 ($1p^+$ + no $n^0$ = 1)
    - H-2 ($1p^+ + 1\ n^0 = 2$)
    - H-3 ($1p^+ + 2\ n^0 = 3$)
  - Carbon
    - C-12 ($6p^+ + 6\ n^0 = 12$)
    - C-14 ($6p^+ + 8\ n^0 = 14$)
  - The Periodic Table gives the average atomic mass of naturally occurring isotopes of each element, based on their abundance.
    - That is why the atomic masses have decimals.
    - In general, to find the number of neutrons, if we don't know the exact isotope, we have to round the mass.
    - From now on, to do many calculations, we will need to use masses to one decimal place.

**The mole**

- Molar mass
  - Of an atom= atomic mass + g (grams)
    - Examples:
      - Molar mass of Be= 9.0 g
      - Molar mass of Fe= 55.8 g
  - Of a molecule= sum of all atomic masses + g (grams)
    - Examples:
      - Molar mass of NaCl= 58.5 g (23.0 + 35.5)
      - Molar mass of $H_2O$= 18.0 g (2 x 1.0 + 16.0)
- Avogadro's number
  - The number of representative particles in 1 mole
    - The abbreviation for mole is mol.
    - Can be used to count atoms, molecules, ions, etc.
    - Experimentally determined to be about $6.02 \times 10^{23}$
    - A mole (mol) of any element or compound contains $6.02 \times 10^{23}$ atoms or molecules or ions, for our purposes.
    - A mol of any substance weighs the molar mass of that substance.
- Problem solving
  - Number of moles → Number of particles
    - Multiply by $6.02 \times 10^{23}$
    - Ex: 3.0 mol = how many atoms?
      - 3.0 mol x $6.02 \times 10^{23}$ atoms/mol = $1.8 \times 10^{24}$ atoms
      - NOTE: The number of moles is considered a measurement, so the significant digit rules apply to it. However, Avogadro's number ($6.02 \times 10^{23}$) is a conversion factor, to which those rules *do not apply*.

The answer must be rounded to the same number of significant digits as there are in the measurement *only*. So, the answer is $1.8 \times 10^{24}$ atoms, rounded to two significant digits, because 3.0 mol only has two.

- Number of particles → Number of moles
  - Divide by $6.02 \times 10^{23}$
  - Ex: $5.8 \times 10^{31}$ molecules of water = ? mol water
    - $5.8 \times 10^{31}$ molecules ÷ $6.02 \times 10^{23}$ molecules/mol = $9.6 \times 10^{23}$ mol
- Moles → Grams of an element or compound
  - Number of moles x molar mass
    - Ex: 2.50 mol $NH_3$ = ? grams $NH_3$
      - Molar mass $NH_3$ = 17.0 g (3 x 1.0 = 2.0 g for $H_3$, + 14 g for N)
      - 17.0 g/mol x 2.50 mol = 42.5 g $NH_3$
- Grams of an element or compound → Moles
  - Grams ÷ Molar mass
    - Ex: 267 g $H_2SO_4$ = ? mol $H_2SO_4$
      - Molar mass of $H_2SO_4$
        - 2 x 1.0 = 2.0 g for $H_2$
        - S = 32.1 g
        - 4 x 16.0 = 64.0 g for $O_4$
        - 2.0 + 32.1 + 64.0 = 98.1 g/mol $H_2SO_4$
      - 267 g ÷ 98.1 g/mol = 2.72 mol $H_2SO_4$
- Number of moles of each element in a compound
  - If no subscript, it's one mole
  - If there is a subscript, it's that many moles
  - Ex: How many moles of each element are there in one mole of $HC_2H_3O_2$?

- 4 mol H (H + $H_3$ = 4 total)
- 2 mol C
- 2 mol O
- (In two mol $HC_2H_3O_2$ there would be twice as many mol of each element: 8 mol H, 4 mol C, & 4 mol O)

## Percent composition

- *Percent composition* is the relative weight of each element in a compound expressed as a percentage.
  - Example: Find the percent composition of $NH_3$
    - Atomic mass of N = 14.0
    - Atomic mass of H = 1.0
      - 1.0 x 3 = 3.0 total
    - Total of all atomic masses = 14.0 + 3.0 = 17.0
    - Divide each mass by the total
      - N= 14.0 ÷ 17.0 = 82.4% (To 0.1%)
      - $H_3$ = 3.0 ÷ 17.0 = 17.6% (To 0.1%)
      - NOTE: The individual % should add up to ≈ 100%
        - 82.4% + 17.6% = 100.0%
        - *You can be +/- 1% from 100%; this is close enough*
  - You can also find the formula for a compound from its percent composition, given the molar mass.
    - Divide each percent's number by the atomic mass of the element; round to $100^{ths}$.
    - Divide those results by the smallest result and round to the nearest whole number.
    - These become the subscripts; however, if it's 1, do not write a subscript.

- Example: A compound is 17.6% hydrogen and 82.4% nitrogen, and its molar mass is 17.0g. What is the compound's formula?
  - 17.6 ÷ 1.0 = 17.60 for H
  - 82.4 ÷ 14.0 = **5.89** for N
  - 17.60 ÷ **5.89** rounds to 3 (subscript for H)
  - 5.89 ÷ **5.89** is 1 (for N, but we don't write it if it's 1)
  - The formula = $NH_3$
- Empirical and Molecular formulas
  - An empirical formula is the reduced ratio of all elements in a compound, and is what you get by looking at just the percent composition.
    - Ex: $CH_4$ and $C_2H_8$ have the same empirical formula $CH_4$
      - $CH_4$ is 75.0% carbon and 25.0% hydrogen
      - $C_2H_8$ is also 75.0% carbon and 25.0% hydrogen
  - You must know the molecular weight in order to find the molecular formula
    - Ex: Let's work backward using $C_2H_8$
      - Percent composition = 75.0% carbon and 25.0% hydrogen; the molecular weight is 32.0
      - $1^{st}$, find the *empirical* formula
        - 75.0 ÷ 12.0 = **6.25** (for carbon)
        - 25.0 ÷ 1.0 = 25.0 (for hydrogen)
        - 6.25 ÷ **6.25** = 1
        - 25.0 ÷ **6.25** = 4
        - The *empirical* formula is therefore $CH_4$
      - $2^{nd}$, use the molecular mass to adjust the subscripts
        - 1 C weighing 12.0 + 4 H weighing 4.0 = 16.0

        - 32.0 (molecular mass) ÷ 16.0 = **2**
        - So, the *molecular* formula is $C_2H_8$, *twice* the empirical one.
    - Ex: Molecular mass = 30.0; 80.0% C and 20.0% H
        - 80.0 ÷ 12.0 = **6.67** for C
        - 20.0 ÷ 1.0 = 20.0 for H
        - 6.67 ÷ **6.67** = 1 for C
        - 20.0 ÷ **6.67** = 3 for H
        - $CH_3$ is the empirical formula, with a mass = 15.0
        - 30.0 ÷ 15.0 = 2
        - $C_2H_6$ is the actual *molecular* formula

## Solutions

- A homogeneous mixture, often of a solid dissolved in a liquid
    - Example: Salt dissolved in water
    - The substance that is dissolved is the *solute.*
    - The part that dissolves the solute is the *solvent.*
    - Solutions do not separate under normal circumstances; however, the solvent may evaporate over time.
- Concentration of solutions
    - Refers to the relative strength of the mixture
        - If the solution is made stronger by adding solute, the more concentrated the solution is.
        - When the solution is diluted by adding more solvent, the less concentrated (more dilute) it is.
    - May be described using
        - Percent by weight

        - Weight of solute ÷ total weight of solution, converted to a percent
    - Percent by volume
        - Volume of solute ÷ total volume of solution, converted to a percent
    - Molarity
        - Moles of solute dissolved in 1L of solution
        - Unit for molarity is ***M***
        - Calculated as Moles of solute ÷ Liters of solution
        - Ex: 2 mol sugar dissolved in 4 L of solution
            - 2 ÷ 4 = 0.5 *M*
        - If the mass of a solute is given, the mass must 1$^{st}$ be converted to moles
            - Ex: 71 g $Na_2SO_4$ in 6.0 L
                - 71 g ÷ 142.1 g/mol = 0.50 mol
                - 0.50 mol ÷ 6.0 L = 0.083 *M*

- Dilution formula
    - $C_1V_1 = C_2V_2$
        - $C_1V_1$ are the concentration (C) and volume (V) of the known solution
        - $C_2V_2$ are the concentration and volume of the other solution, where either C or V is unknown
        - Ex: What volume of a stock solution of 4.0 M HCl is needed to make 500.0 L of 0.6 M HCl?
            - $C_1$ = 0.6 M HCl
            - $V_1$ = 500.0 L HCl
            - $C_2$ = 4.0 M HCl
            - $V_2$ = ? L HCl

- 0.6 M x 500.0 L = 4.0 M x ? L
  - 300 M-L = 4.0 M x ? L
  - 300 M-L ÷ 4.0 M =75 L
  - NOTE: If 75 L of the 500.0 L is HCL, how much is water?
    - 500.0 L solution – 75 L HCl = 425 L $H_2O$