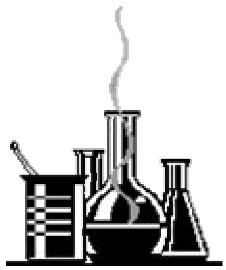

Chemical Amounts

2.8 The Mole

- Unit for amount of substance in terms of the number of entities (atoms, molecules, ...) in it.
- **1 mol of entities → # of atoms in 12 g of ^{12}C**
  - 1 atom of $^{12}C \rightarrow 1.99265\times10^{-23}$ g^{12}C (mass spectrometry)
  - 12 g^{12}C × [1 atom/1.99265×10^{-23} g^{12}C] = = 6.0221×10^{23} atoms

⇒1 mol of entities → 6.0221×10^{23} entities

- The Avogadro constant (N_A) - # of entities per mol

  6.0221×10^{23} /mol
- Conversion between moles and entities

  [1 mol entities/6.0221×10^{23} entities]

Example: How many atoms of hydrogen are present in 2.7 mol of water.

$$2.7\ \text{mol H}_2\text{O}\times\left(\frac{6.0221\times10^{23}\ \text{molec. H}_2\text{O}}{1\ \text{mol H}_2\text{O}}\right)\times\left(\frac{2\ \text{atoms H}}{1\ \text{molec. H}_2\text{O}}\right)=3.3\times10^{24}\ \text{atoms H}$$

2.9 Molar Mass (*M*)

- Mass of a substance per 1 mol of its particles
  - element → atoms
  - molecular compound → molecules
  - ionic compound → formula units
- $\boldsymbol{M = m_{particle} \times N_A}$

Example: What is the molar mass of ^{1}H, if the mass of 1 atom ^{1}H is 1.673×10^{-24} g?

$\boldsymbol{M = 1.673\times10^{-24}\,\text{g} \times 6.022\times10^{23}\,/\text{mol} = 1.007\ \text{g/mol}}$

- Naturally occurring elements are mixtures of isotopes
  - molar masses of isotopes ← mass spectrometry
  - **average molar masses** of elements

Example: Calculate the average atomic mass of Cu, given that it naturally occurs as 69.17% ^{63}Cu (M=62.94g/mol) and 30.83% ^{65}Cu (M=64.93g/mol).

$$0.6917\times62.94\ \text{g/mol} + 0.3083\times64.93\ \text{g/mol} = $$
$$= 63.55\ \text{g/mol}$$

- Conversion between moles (***n***) and mass (***m***)

$$\boldsymbol{m = n \times M} \quad \leftrightarrow \quad \boldsymbol{n = m/M}$$

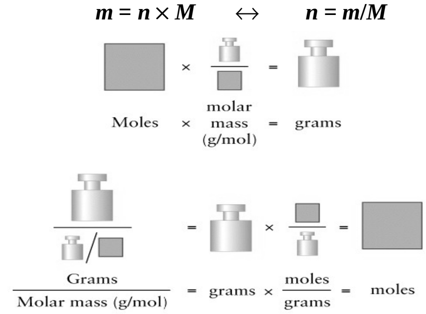

Fig. 2.13

Example: What is the mass of 1.221 mol Kr?

***m* = 1.221 mol × 83.80 g/mol = 102.3 g**

- ***M*** can be used as a conversion factor

  Example: How many moles of atoms are present in 1.23 g of Kr?

$$1.23 \text{ g Kr} \times \left(\frac{1 \text{ mol Kr}}{83.80 \text{ g Kr}} \right) = 1.47 \times 10^{-2} \text{ mol Kr}$$

Fig. 2.15

2.10 Measuring Out Compounds

- Molar mass of a compound - sum of the molar masses of its elements multiplied by their subscripts in the chemical formula

  Example: Determine the molar masses of CO_2 and K_2SO_4.

  (C→12.01; O→16.00; S→32.06; K→39.10 g/mol)

⇒1×12.01 g/mol + 2×16.00 g/mol = 44.01 g/mol

⇒2×39.10 + 1×32.06 + 4×16.00 = 174.26 g/mol

- Conversion between moles (***n***) and masses (***m***) of compounds (same as for elements)

  Example: Calculate the number of moles of urea, $(NH_2)_2CO$, in 2.3×10^5 kg of this compound.

  ***M* = 2×14.00 + 4×1.008 + 1×12.01 + 1×16.00 =**
  = 60.04 g/mol

$$2.3\times10^5 \text{ kg urea} \times \left(\frac{10^3 \text{ g urea}}{1 \text{ kg urea}} \right) \times \left(\frac{1 \text{ mol urea}}{60.04 \text{ g urea}} \right) = 3.8\times10^6 \text{ mol urea}$$