Chem. Ch. 10 ~ THE MOLE

NOTE: Vocabulary terms are in **boldface and underlined**. Supporting details are in *italics.*

10.1 Notes

I. Measuring Matter

A. SI unit of chemical quantity = the **mole** (abbreviated *mol*)

1) 6.0221367×10^{23}
2) **6.02×10^{23} = Avogadro's number** (Amedeo Avogadro, 1776-1856)
   TO MAKE IT EASIER, WE ROUND TO 6.02×10^{23}
3) the mass of 12 g of pure C-12
4) the **mole** is a *counting unit used in science to count particles*
   - **representative particle (r.p.)** = *atom, ion, molecule, formula unit*

- **atom** = *one symbol, no charge: Br, U, Cs*
- **ion** = *one symbol with charge (monatomic) or more than one symbol with charge (polyatomic: Na^+, N^{3-}, $(C_2H_3O_2)^-$*
- **molecule** = *compound with all nonmetals (BM, TM): CO, BF_3, Cl_2*
- **formula unit ("fun")** = *compound with metal and nonmetal (BI, TI): KI, Na_2SO_4*

1 MOLE = 6.02×10^{23} representative particles

6.02×10^{23} atoms = 1 mol atoms	6.02×10^{23} molecules = 1 mol molecules
$6.02\ 10^{23}$ ions = 1 mol ions	6.02×10^{23} formula units = 1 mol f.un.

B. **stoichiometry**—using balanced chemical equations to obtain info.

C. mole-to-r.p. and r.p.-to-mole example problems:

E1) How many moles of Ca are in 9.00×10^{16} atoms of calcium?

$$9.00 \times 10^{16}\ \cancel{\text{atoms Ca}} \times \frac{1\text{ mol Ca}}{6.02 \times 10^{23}\ \cancel{\text{atoms Ca}}} = \boxed{1.50 \times 10^{-7}\text{ mol Ca}}$$

E2) How many moles of sulfide ions are in 4.14×10^{30} sulfide ions$^-$?

$$4.14 \times 10^{30}\ \cancel{\text{ions } S^{2-}} \times \frac{1\text{ mol S}}{6.02 \times 10^{23}\ \cancel{\text{ions } S^{2-}}} = \boxed{6.88 \times 10^{6}\text{ mol } S^{2-}}$$

E3) How many molecules are in 0.0221 mol oxygen gas?

$$0.0221\ \cancel{\text{mol } O_2} \times \frac{6.02 \times 10^{23}\text{ molecules } O_2}{1\ \cancel{\text{mol } O_2}} = \boxed{1.33 \times 10^{22}\text{ molecules } O_2}$$

D. Finding the number of atoms in a compound—look at the subscripts

E4) How many hydrogen atoms are in 0.89 mol water?

$$0.89\ \cancel{\text{mol } H_2O} \times \frac{6.02 \times 10^{23}\ \cancel{\text{molecules } H_2O}}{1\ \cancel{\text{mol } H_2O}} \times \frac{2\text{ H atoms}}{1\ \cancel{\text{molecule } H_2O}} = \boxed{1.1 \times 10^{24}\text{ H atoms}}$$

E5) How many sodium ions are found in 0.129 mol of sodium phosphate?

$$0.129\ \cancel{\text{mol } Na_3PO_4} \times \frac{6.02 \times 10^{23}\ \cancel{\text{f.un.}}}{1\ \cancel{\text{mol } Na_3PO_4}} \times \frac{3\ Na^+\text{ ions}}{1\ \cancel{\text{f.un. } Na_3PO_4}} = \boxed{2.33 \times 10^{23}\ Na^+\text{ ions}}$$

10.2 & 10.3 Notes

II. Mass and the Mole

[The atomic masses on the periodic table have a unit of **atomic mass unit** (amu, or u).]

A. **GAM** = **gram atomic mass**

1) *the atomic mass (listed on the periodic table) written in grams*

1 atom Xe = 131.30 u & 1 mol Xe = 131.30 g & GAM of Xe = 131.301 g

2) *these numbers are usually rounded to 0.1 (tenths)*

B. **GMM** = **gram molecular mass**

1) *the sum of all masses of atoms in a molecular compound*

1 molecule Cl_2 = 70.9 u & 1 mol Cl_2 = 70.9 g & GMM Cl_2 = 70.9 g

2) example:

E6) Find the GMM of methane, CH_4.

$CH_4 = 1(12.0) + 4(1.0) = 12.0 + 4.0 = \boxed{16.0\ g}$

C. **GFM** = **gram formula mass**

1) *the mass of one mole of ionic compound*

1 f.unit NaCl = 58.4 amu & 1 mol NaCl = 58.4 g & GFM NaCl = 58.4 g

2) example:

E7) Find the GFM of calcium hydroxide.

$Ca(OH)_2 = 1(40.1) + 2(16.0) + 2(1.0) = \boxed{74.1\ g}$

D. **molar mass**—*the mass, in g, of 1 mole of a substance*

- molar mass is a general term for doing GAM, GMM, or GFM

WE OUR ROUND MOLAR MASSES TO TENTHS (0.1 g), ONE DECIMAL PLACE

III. Mole-to-Mass and Mass-to-Mole Conversions (dimensional analysis)

$\cancel{\text{grams A}} \times \dfrac{\text{1 mol A}}{\cancel{\text{MOLAR MASS (g) A}}} = \boxed{\text{mol A}}$ $\qquad$ $\cancel{\text{mol B}} \times \dfrac{\text{MOLAR MASS (g) B}}{\cancel{\text{1 mol B}}} = \boxed{\text{grams B}}$

examples:

E8) How many grams are in 0.70 mol of carbon dioxide?

$$0.70\ \cancel{\text{mol } CO_2} \times \frac{\overset{44.0}{[1(12.0) + 2(16.0)]}\ \text{g } CO_2}{1\ \cancel{\text{mol } CO_2}} = \boxed{31\ \text{g } CO_2}$$

E9) How many moles are in 362 g of sodium bromide?

$$362\ \cancel{\text{g NaBr}} \times \frac{\text{1 mol NaBr}}{\underset{102.9}{(23.0 + 79.9)}\ \cancel{\text{g NaBr}}} = \boxed{3.52\ \text{mol NaBr}}$$

10.4 – 10.5 Notes

Preview of gas laws…

IV. Molar Volume: volume-to-mole and mole-to-volume conversions

A. **STP** = **standard temperature and pressure**

1) *standard temperature = 0 °C, 273 K*

2) *standard pressure = 101.3 kPa, 1.00 atm, 760 mm Hg, 14.7 psi*

B. At STP, all gases occupy the same amount of space:

MOLAR VOLUME of any gas at STP: 22.4 L = 1 mol

C. examples

E10) What is the volume, in L, of 0.495 mol of nitrogen dioxide gas at STP?

$$0.495\ \cancel{\text{mol } NO_2} \times \frac{22.4\text{ L } NO_2}{1\ \cancel{\text{mol } NO_2}} = \boxed{11.1\text{ L } NO_2}$$

E11) How many moles are found in 84 L of neon gas at STP?

$$84\ \cancel{\text{L Ne}} \times \frac{1\text{ mol Ne}}{22.4\ \cancel{\text{L Ne}}} = \boxed{3.8\text{ mol Ne}}$$

MOLE CONVERSION FACTOR SUMMARY:

$$\frac{6.02 \times 10^{23}\text{ representative particles}}{1\text{ MOLE}} \qquad \frac{1\text{ MOLE}}{6.02 \times 10^{23}\text{ representative particles}}$$

$$\frac{\text{MOLAR MASS (g)}}{1\text{ MOLE}} \qquad \frac{1\text{ MOLE}}{\text{MOLAR MASS (g)}}$$

$$\frac{22.4\text{ L (for a gas at STP)}}{1\text{ MOLE}} \qquad \frac{1\text{ MOLE}}{22.4\text{ L (for a gas at STP)}}$$

V. Mass-to-r.p. and r.p.-to-mass conversions

E12) How many grams of barium sulfide are in 3.39×10^{25} r.p. of barium sulfide?

$$3.39 \times 10^{25}\ \cancel{\text{f.un.BaS}} \times \frac{1\text{ mol BaS}}{6.02 \times 10^{23}\ \cancel{\text{f.un. BaS}}} \times \frac{\overset{169.4}{(137.3 + 32.1)}\text{ g BaS}}{1\ \cancel{\text{mol BaS}}} = \boxed{9540\text{ g BaS}}$$

E13) How many particles of rubidium nitrate are in 45.00 g of rubidium nitrate?

$$45.00\ \cancel{\text{g } RbNO_3} \times \frac{1\ \cancel{\text{mol } RbNO_3}}{\underset{147.5}{[85.5 + 14.0 + 3(16.0)]}\ \cancel{\text{g } RbNO_3}} \times \frac{6.02 \times 10^{23}\ \cancel{\text{f.un } RbNO_3}}{1\ \cancel{\text{mol } RbNO_3}} = \boxed{1.84 \times 10^{23}\text{ f.un } RbNO_3}$$

There are many types of mole problems:

1 step:	*r.p. → mol*	*&*	*mol → r.p.*	
	mass → mol	*&*	*mol → mass*	
	volume → mol	*&*	*mol → volume*	*(of gas at STP)*
2 step:	*mass → r.p.*	*&*	*r.p. → mass*	
	mass → volume	*&*	*volume → mass*	*(of gas at STP)*
	r.p. → volume	*&*	*volume → r.p.*	*(of gas at STP)*

Preview of gas laws…

VI. Gas Density and Molar Mass

A. Density *D = M / V*

B. *gas density usually measured in g/L*

C. *use 22.4 L = 1 mol to calculate molar masses (g/mol, the mass of 1 mole)*

D. examples

E14) The density of a gas is 3.64 g/L in STP conditions. What is its molar mass?

$$\frac{3.64\ g}{\cancel{L}} \times \frac{22.4\ \cancel{L}}{1\ mol} = \boxed{81.5\ g/mol}$$

E15) At STP, 6.00 L of a gas has a mass of 25.10 g.
Calculate the density of the gas and its molar mass.

$$D = \frac{M}{V} = \frac{25.10\ g}{6.00\ L} = \boxed{4.18\ g/L} \qquad \text{MOLAR MASS} = \frac{4.18\ g}{\cancel{L}} \times \frac{22.4\ \cancel{L}}{1\ mol} = \boxed{93.6\ g/mol}$$

VII. Percent Composition

A. **<u>Percent composition</u>**—*% by mass of each individual element in a compound*

B. remember to list all percentages

C. double-check that the % total is 100% (or very close if rounding)

D. formulas

$$\textbf{\%} = \frac{\textbf{\# g element}}{\textbf{\# g cmpd.}} \times \textbf{100} \qquad \textbf{\%} = \frac{\textbf{MOLAR MASS of element}}{\textbf{MOLAR MASS of cmpd.}} \times \textbf{100}$$

E. examples

E16) Find the % by mass of hydrogen and oxygen in water.

MOLAR MASS H_2O = 2(1.0) + 1(16.0) = 18.0 g
2 H's + 1 O = total mass H_2O

$$\% H = \frac{2.0\ g}{18.0\ g} \times 100 = \boxed{11\%\ H} \qquad \% O = \frac{16.0\ g}{18.0\ g} \times 100 = \boxed{88.9\ \%\ O}$$

E17) Calculate the % composition of sulfuric acid.

MOLAR MASS H_2SO_4 = 2(1.0) + 32.1 + 4(16.0) = 98.1 g
2 H's + 1 S + 4 O's = total mass H_2SO_4

$\% H = \frac{2.0\ g}{98.1\ g} \times 100 = \boxed{2.0\ \%\ H}$ (2.04)

$\% S = \frac{32.1\ g}{98.1\ g} \times 100 = \boxed{32.7\%\ S}$ (32.72)

$\% O = \frac{4(16.0)\ g}{98.1\ g} \times 100 = \boxed{65.2\ \%\ O}$ (65.24)

VIII. Empirical Formula

A. **Empirical formula**—*the simplest whole-number ratio of elements in a cmpd.*

B. it is a non-reducible ratio of moles

C. problem procedure

1. *convert % to grams directly*
2. *find numbers of moles*
3. *make mole ratios using the smallest mole number as the denominator*
4. *use these whole number ratios as the subscripts of the formula*

(If a number with 0.5 is observed, multiply everything by 2. Don't round up!)

D. examples

E18) Calculate the empirical formula of a compound composed of 67.6% mercury, 10.8% sulfur, and 21.6% oxygen. *(check the PROBLEM PROCEDURE steps)*

STEP 1… mercury: 67.6% Hg = 67.6 g Hg out of 100 g cmpd.
sulfur: 10.8% S = 10.8 g S out of 100 g cmpd.
oxygen: 21.6% O = 21.6 g O out of 100 g cmpd.

STEP 2... Hg: $67.6\ g\ Hg \times \frac{1\ mol\ Hg}{200.6\ g\ Hg} = \mathit{0.337\ mol\ Hg}$ S: $10.8\ g\ S \times \frac{1\ mol\ S}{32.1\ g\ S} = \mathit{0.336\ mol\ S}$

O: $21.6\ g\ O \times \frac{1\ mol\ O}{16.0\ g} = \mathit{1.35\ mol\ O}$

STEP 3… $Hg = \frac{0.337}{0.336} = 1$ $S = \frac{0.336}{0.336} = 1$ $O = \frac{1.35\ mol}{0.336\ mol} = 4$

STEP 4… $Hg_1S_1O_4 = \boxed{HgSO_4}$

E19) What is the empirical formula of a compound with 25.9% nitrogen and 74.1% oxygen?

25.9% N = 25.9 g N out of 100 g cmpd. 74.1% O = 74.1 g O out of 100 g cmpd.

N: $25.9\ g\ N \times \frac{1\ mol\ N}{14.0\ g\ N} = \mathit{1.85\ mol\ N}$ O: $74.1\ g\ O \times \frac{1\ mol\ O}{16.0\ g\ O} = \mathit{4.63\ mol\ O}$

N: $\frac{1.85}{1.85} = 1$ O: $\frac{4.63}{1.85} = 2.5$ $N_1O_{2.5}$ – can't have .5 subscripts x 2 = $\boxed{N_2O_5}$

IX. Molecular Formula

F. **Molecular formula**—*a multiple of the empirical formula*

G. still whole number ratios

H. examples

E20) A compound with an empirical formula of CH has a molecular weight of 78.0 g/mol. What is the molecular formula?

molar mass of CH = 12.0 + 1.0 = 13.0 g 78 / 13 = 6 molecular formula = $\boxed{C_6H_6}$

E21) A compound is 75.46% carbon, 4.43% hydrogen, and 20.10% oxygen by mass. It has a molecular mass of 318.31 g/mol. What is the molecular formula for this compound?

75.46% C = 75.46 g C out of 100 g cmpd. 4.43% H = 4.43 g H out of 100 g cmpd.
20.10% O = 20.10 g O out of 100 g cmpd.

(75.46 g C) (1 mol/ 12.0 g C) = *6.29 mol C* (4.43 g H) (1 mol/ 1.0 g H) = *4.4 mol H*
(20.10 g O) (1 mol/ 16.0 g O) = *1.26 mol O*

(6.29 mol C)/ (1.26) = 4.99 = 5 mol C
(4.4 mol H)/ (1.26) = 3.49 = 3.5 mol H
(1.26 mol O)/ (1.26) = 1 mol O

.5 value means multiply subscripts by 2: empirical fmla. = $C_{10}H_7O_2$

Now that you have the emp.fmla., you can find the molecular fmla like in problem E18.
emp. fmla. mass = 10(12.0) + 7(1.0) + 2(16.0) = 159.1 g/mol
The problem says the molecular mass is 318.31 g per mole.

$\frac{318.31 \text{ g/mol}}{159.0 \text{ g/mol}}$ = 2.001 = 2 ratio

Since there are two empirical units in a molecular unit, the molecular formula = $\boxed{C_{20}H_{14}O_4}$