Moles and Stoichiometry
Unit 5

Moles and Stoichiometry

Compounds

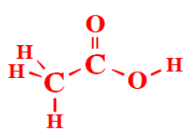
- chemical compounds are *substances* composed of two or more *elements* that are *chemically bonded* in a *definite proportion*
  - Examples:
    - H_2O and H_2O_2 are two different chemical compounds with different chemical and physical properties
    - Water is never $H_{3.5}O_{1.7}$
- chemical compounds can be broken down by *chemical means*
  - Examples:
    - $2\ H_2O_{(\ell)} \rightarrow 2\ H_{2\,(g)} + O_{2\,(g)}$
    - $CaCO_{3\,(s)} \rightarrow CaO_{(s)} + CO_{2\,(g)}$
- chemical compounds can be represented by a specific *chemical formula* and given a specific name based on the IUPAC system of nomenclature
  - Examples:
    - H_2O is water (accepted IUPAC name) or oxidane (recommended IUPAC systematic name)
    - $CaCO_3$ is calcium carbonate

Chemical names and formulas are:

- systematic (the whole world does it the same way)
- unambiguous (they tell the *exact* composition)

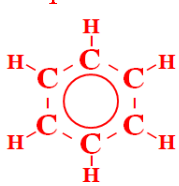
Types of chemical formulas:

- empirical (the smallest whole number ratio)
  - Example: the molecule $C_2H_4O_2$ has the empirical formula CH_2O
    - Ionic compounds all have empirical formulas, except mercury(I): Hg_2Cl_2 (See Table E)
- molecular (the *exact* composition of a molecule)
  - Example: the molecule glucose has the molecular formula $C_6H_{12}O_6$
- structural (shows position and bonding)
  - Example: the molecule C_2H_3OOH has the structural formula



Required by NYS Regents

- condensed (shows position and bonding as text)
  - Example: the molecule C_2H_3OOH has the structural formula
- skeletal (a chemical shorthand)
  - Example: benzene



structural

skeletal

Sometimes used by NYS Regents

Writing Chemical Formulas

Type-I Binary Compounds (ionic with monovalent metal)

1. Write the cation symbol first
2. Write the anion symbol last
3. If the oxidation number sum is = 0, use easy criss-cross
4. Reduce subscripts to least whole number ratio, if necessary

Notes: ammonium is treated as a monoatomic cation for naming purposes; peroxide, cyanide, and hydroxide are named as monoatomic anions (see Table E)

Example: sodium chloride

NaCl	Check ox. no.s: $+1 + -1 = 0$	formula complete = NaCl

Example: magnesium chloride

MgCl	Check ox. no.s: $+2 + -1 = +1$	
	use criss-cross	$Mg^{+2}Cl^{-1} \rightarrow Mg_1Cl_2$
$MgCl_2$	Check ox. no.s: $+2 + 2(-1) = 0$	formula complete = $MgCl_2$

Example: aluminum oxide

AlO	Check ox. no.s: $+3 + -2 = +1$	
	use criss-cross	$Al^{+3}O^{-2} \rightarrow Al_2O_3$
Al_2O_3	Check ox. no.s: $2(+3) + 3(-2) = 0$	formula complete = Al_2O_3

Example: hydrosulfuric acid (acids are a special case)

HS	Check ox. no.s: $+1 + -2 = -1$	
	use criss-cross	$H^{+1}S^{-2} \rightarrow H_2S_1$
H_2S	Check ox. no.s: $2(+1) + -2 = 0$	formula balanced = H_2S
H_2S	always add (aq) to acid formulas	formula complete = $H_2S_{(aq)}$

Type-II Binary Compounds (ionic with multivalent metal will have Roman numeral)

1. Write the cation symbol first (oxidation number already known)
2. Write the anion symbol last
3. If the oxidation number sum is = 0, use easy criss-cross
4. Reduce subscripts to least whole number ratio, if necessary

Notes: mercury(I) must occur in even numbers, e.g. Hg_2Cl_2
same notes for special polyatomic ions as for Type-I (see Table E)

Example: lead(IV) oxide

PbO	Check ox. no.s: $+4 + -2 = +2$	
	use criss-cross	$Pb^{+4}O^{-2} \rightarrow Pb_2O_4$
Pb_2O_4	Check ox. no.s: $2(+4) + 4(-2) = 0$	formula balanced = Pb_2O_4
PbO_2	Reduce subscripts, if needed	formula complete = PbO_2

Type-III Binary Compounds (molecular, all nonmetal atoms, and utilizes prefixes)

1. Use the prefixes to write the formula

Example: carbon dioxide

CO	adjust numbers with prefixes	CO_2
CO_2	formula complete = CO_2	

Example: dinitrogen monoxide

NO	adjust numbers with prefixes	N_2O
N_2O	formula complete = N_2O	

Polyatomic Ions (name ends with 'ate' or 'ite', use Table E)

1. Write the cation symbol first (oxidation number already known)
2. Write the anion symbol last
3. If the oxidation number sum is = 0, use easy criss-cross
4. Reduce subscripts to least whole number ratio, if necessary

Example: ammonium hydrogen sulfate

NH_4HSO_4 Check ox. no.s: $+1 + +1 + -2 = 0$ formula complete = NH_4HSO_4

Polyatomic Acids (name ends with 'acid' but does not start 'hydro', use Table E)

1. Write the H cation symbol first (oxidation number is always +1)
2. Write the polyatomic anion symbol last
   for ions ending in 'ic', search for 'ate' polyatomic ions in Table E
   for ions ending in 'ous', search for 'ite' polyatomic ions in Table E
3. If the oxidation number sum is = 0, use easy criss-cross
4. Reduce subscripts to least whole number ratio, if necessary

Example: sulfurous acid

HSO_3 Check ox. no.s: $+1 + -2 = -1$

use criss-cross $H^{+1}SO_3^{-2} \rightarrow H_2(SO_3)_1$

H_2SO_3 Check ox. no.s: $2(+1) + -2 = 0$ formula complete = H_2SO_3

Naming Compounds (Nomenclature)

Binary Compounds

1. Type-I ionic binary (metal + nonmental, the metal is monovalent)
   a. Write the metal name first
   b. Write the nonmetal name last changing the ending to 'ide'

Irregular anion names:

hydrogen	hydride	oxygen	oxide
nitrogen	nitride	phosphorus	phosphide
sulfur	sulfide		

Example: $MgCl_2$
magnesium + chlorine → chloride = magnesium chloride

Example: HCl should be a Type-III but hydrogen binaries are named like Type-I
hydrogen + chlorine → chloride = hydrogen chloride

Example: NaCN should be polyatomic but *cyanide ions* are named like Type-I
sodium + cyanide = sodium cyanide

Example: $Ca(OH)_2$ should be polyatomic but monovalent *bases* are Type-I
calcium + hydroxide = calcium hydroxide

Example: $(NH_4)_2S$ should be polyatomic but ammonium ion is named like Type-I
ammonium + sulfide = calcium hydroxide

2. Type-II ionic binary (metal + nonmetal, the metal has more than one oxidation state)
   a. Write the metal name first
   b. Write the metal charge (calculated from the nonmetal)
   c. Write the nonmetal name last changing the ending to 'ide'

Example: CuO (where Cu must be +2 because O is −2)
copper + (II) + oxygen → oxide = copper(II) oxide

Example: $Cu(OH)_2$ should be polyatomic but *divalent bases* are Type-II
copper(II) + hydroxide = copper(II) hydroxide

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3. Type-III binary (molecular – all elements are nonmetals, utilizes prefixes)
   a. Write the first element name with a prefix (never add mono to the first element)
   b. Write the second element name with a prefix

      List of prefixes:

1. mon(o)	2. di	3. tri
4. ter(a)	5. pent(a)	6. hex(a)
7. hept(a)	8. oct(a)	9. non(a)
	10. dec(a)	

   Example: N_2O
      di + nitrogen + mon + oxygen → oxide = dinitrogen monoxide

   Example: CO_2
      carbon + di + oxygen → oxide = carbon dioxide

4. Binary acids (molecular – H + nonmetal, must be aqueous, special naming system)
   a. Write the second element root name with a prefix 'hydro'
   b. Add the suffix 'ic'
   c. Add the name 'acid'

      Binary acid root names:

S	sulfur	Te	tellur	Se	selen
Cl	chlor	Br	brom	I	iod

   Example: $HCl_{(aq)}$
      hydro + chlorine → chlor+ic + acid = hydrochloric acid

Polyatomic Ionic Compounds

5. Salts (Use Table E)
   a. Write the first element name (Type-I or Type-II)
   b. Write the polyatomic ion name (from Table E)

   Example: Na_2SO_4
      sodium + sulfate = sodium sulfate

   Example: $Cu(HSO_3)_2$
      copper(II) + hydrogen sulfite = copper(II) hydrogen sulfite

6. Acids (Begin with H and end with a polyatomic ion, use Table E + ate-ic ite-ous)
   a. Write the polyatomic ion name (from Table E) changing the ending
      ate → ic
      ite → ous
   b. Write the word 'acid'

   Example: H_2SO_4
      sulfate → sulfuric + acid = sulfuric acid

   Example: H_3PO_3
      phosphite → phosphorous + acid = phosphorous acid
      Note that only phosphate (PO_4^{3-}) is shown on Table E, so (PO_3^{3-}) is phosphite

   Example: HClO
      hypochlorite → hypochlorous + acid = hypochlorous acid

7. Bases (Start with a metal and end in -OH, name like Type-I or Type-II binary)
   See examples in Type-I and Type-II ionic binaries

Organic Compounds (carbon based with special rules that we will learn in a later unit)

Finding Gram Formula Mass (gfm)

Formula Mass – the sum of the atomic masses of the atoms in a substance

- based on atomic mass, the unit is the atomic mass unit (amu or u)
  - the amu is arbitrary
  - the amu is relative (1/12th the mass of a C-12)
  - 1 amu = 1.66054×10^{-24} grams or 6.02214×10^{23} u = 1 gram
  - 6.02214×10^{23} = Avogadro's Number ($\mathcal{N}_A$)
- the atomic mass of an element is given in the upper left corner on the Periodic Table

The Magic of Avogadro's Number – $\mathcal{N}_A$ turns amu (or u) into grams

Example: sodium

1 atom of sodium = 22.99 g

$22.99\ u \times 1\ \text{mole} = 22.99\ g$ 1 mole = 6.02214×10^{23}

$22.99\ u \times 6.02214 \times 10^{23} =$ 6.02214×10^{23} u = 1 gram

$22.99\ u \times 6.02214 \times 10^{23} \times \dfrac{1\ g}{6.02214 \times 10^{23}\ u} =$ cancel terms and units

$22.99\ \cancel{u} \times \cancel{6.02214 \times 10^{23}} \times \dfrac{1\ g}{\cancel{6.02214 \times 10^{23}}\ \cancel{u}} =$ 1 mole = 6.02214×10^{23}

$= 22.99\ g$

This will always work because the number in one mole and the number of atomic mass units in 1 gram are equal

(1 mole = 6.02214×10^{23} and 6.02214×10^{23} u = 1 g)

The gfm is always measured for a representative particle

Metals: 1 atom

Example: copper: the gfm is for Cu

Ionic compounds: 1 empirical formula

Example: sodium sulfate: the gfm is for Na_2SO_4

Network solids: 1 atom for elements or 1 empirical formula for compounds

Example: diamond: the gfm is for C

Example: quartz: the gfm is for SiO_2

Molecular solids: 1 molecule

Example: carbon dioxide: the gfm is for CO_2

Calculating the gfm

Example: $CaCO_3$

Ca:	40.08	=	22.99	u	
C:	12.011	=	12.011	u	
O:	15.9994 × 3	=	47.9982	u	
			100.09	u	(the formula mass)
			100.09	g	(the gfm)

Example: $Al_2(SO_4)_3$

Al:	26.98154	× 2	=	53.96308	u	
S:	32.065	× 3	=	96.195	u	
O:	15.9994	× 12	=	191.9928	u	
				342.151	u	(the formula mass)
				342.151	g	(the gfm)

Empirical and Molecular formulas

Empirical formula: the simplest whole-number ratio of atoms of the elements in a compound

Molecular formula: the actual ratio of atoms of the elements in a molecular compound

Empirical formula from molecular formula

1. Find the greatest common factor (GCF) of the subscripts
2. Divide all subscripts by the GCF

Example: find the empirical formula of tetraphosphorus decaoxide

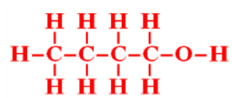
Formula from name → P_4O_{10}

GCF of 4 and 10 = 2

4 ÷ 2 = 2 and 10 ÷ 2 = 5

Empirical formula for P_4O_{10} is P_2O_5

Example: find the empirical formula of



Formula from structure → $C_4H_{10}O$

GCF of 4, 10, and 1 = 1

Empirical formula is $C_4H_{10}O$

Molecular formula from empirical formula and gfm

The molecular formula *must* be a multiple of the empirical formula

The molecular mass (gfm) *must* be a multiple of the empirical mass

1. The gfm of the empirical formula
2. Calculate the molecular mass ÷ empirical mass
3. Multiply the empirical formula by the quotient from Step 2

Example: find the molecular formula of a compound that has a molecular mass of 56 u and an empirical formula of CH_2

The empirical formula (CH_2) has a gfm = 12 + 2 (1) = 14 u

56 u ÷ 14 u = 4

4 (CH_2) = C_4H_8

Percent Composition

While not always explicitly stated, percent composition problems are implicitly by mass unless explicitly stated otherwise (the Regents does not do mole fraction problems)

From Table T:

$$\%\text{ composition by mass} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100$$

Example: find the percent composition by mass of sulfur in sodium thiosulfate

sulfur = S

sulfur = 32.065 u × 2

= 64.130 u

sodium thiosulfate = $Na_2S_2O_3$

$Na_2S_2O_3$ =

Na = 22.989 77 u × 2

S = 32.065 u × 2

O = 15.9994 u × 3

158.108 u

$$\%\text{ composition by mass} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100$$

$$= \frac{\text{mass of S}}{\text{mass of } Na_2S_2O_3} \times 100 = \frac{64.130\text{ u}}{158.108\text{ u}} \times 100$$

= 40.561 %

Example: find the percent composition of water in sodium thiosulfate pentahydrate

water = H_2O

H_2O =

H = 1.007 94 u

O = 15.9994 u

= 18.0153 u

sodium thiosulfate = $Na_2S_2O_3$

$Na_2S_2O_3$ =

Na = 22.989 77 u × 2

S = 32.065 u × 2

O = 15.9994 u × 3

158.108 u

$$\%\text{ composition by mass} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100$$

$$= \frac{\text{mass of } H_2O}{\text{mass of } Na_2S_2O_3{\cdot}5H_2O} \times 100 = \frac{(18.0153\text{ u})5}{158.108\text{ u} + (18.0153\text{ u})5} \times 100$$

= 36.2942 %

Example: determine the mass of iron in a 75.0 gram sample of iron(III) oxide.

iron(III) oxide = Fe_2O_3

Fe_2O_3 = 159.687 u

$$\text{mass Fe} = \text{actual mass of whole} \times \frac{\text{formula mass of part}}{\text{formula mass of whole}}$$

$$= \text{mass of ore} \times \frac{\text{mass of Fe}}{\text{mass of } Fe_2O_3} = 75.0\text{ g}_{Fe_2O_3} \times \frac{(55.845\text{ u})2}{159.687\text{ u}}$$

= 52.5 g

Mass Mole Problems

Using Table T

Formula for Mole Calculations:

$$\text{number of moles} = \frac{\text{given mass}}{\text{gram-formula mass}} \quad or \quad n = \frac{m}{gfm}$$

where:

value	symbol		unit
number of moles	(n)	=	mol
given mass	(m)	=	g
gram-formula mass	(gfm)	=	g/mol

Changing Mass to Moles

Example: how many moles of water are represented by 5.45 grams of water?

n = ?

m = 5.45 g

gfm = 18.0153 g/mol

$$n = \frac{m}{gfm} = \frac{5.45\text{ g}}{18.0\text{ g/mol}} = 0.303\text{ mol}$$

Changing Moles to Mass

Example: how many grams of sodium sulfate are represented by 0.731 mol of Na_2SO_4 (gfm = 142 g/mol)?

m = ?

n = 0.731 mol

gfm = 142 g/mol

$$n = \frac{m}{gfm} = 0.731\text{ mol} = \frac{m}{142\text{ g/mol}}$$

$$m = 0.731\text{ mol} \times 142\text{ g/mol} = 104\text{ g}$$

Mass Mole Problems (continued)

Using Conversion Factors

Formula for Mole Calculations:

where:

value	symbol		unit
number of moles	(n)	=	mol
given mass	(m)	=	g
gram-formula mass	(gfm)	=	g/mol

Changing Mass to Moles

Example: how many moles of water are represented by 5.45 grams of water?

n = ?

m = 5.45 g

gfm = 18.0153 g/mol

$$n = 5.45\ \text{g} \times \frac{1\ \text{mol}}{18.0\ \text{g}} = 0.303\ \text{mol}$$

Changing Moles to Mass

Example: how many grams of sodium sulfate are represented by 0.731 mol of Na_2SO_4 (gfm = 142 g/mol)?

m = ?

n = 0.731 mol

gfm = 142 g/mol

$$m = 0.731\ \text{mol} \times \frac{142\ \text{g}}{1\ \text{mol}} = 104\ \text{g}$$

Chemical Reactions

For languages we use: letters, words, and sentences

For chemical reactions we use: symbols (representing elements), formulas (representing compounds), and equations (representing reactions)

All chemical reactions exhibit:

- conservation of mass (matter is neither created nor destroyed)

Example 1: modeling conservation of mass

$Ca + 2\ HCl \rightarrow CaCl_2 + H_2$ Let Ca = ○, H = ⊘, and Cl = ●

○ ⊘● ⊘● ●○● ⊘⊘ When # atoms is =, the mass is =

Example 2: calculating conservation of mass

$Ca + 2\ HCl \rightarrow CaCl_2 + H_2$

40. g + 73 g = 111 g + 2 g The mass is = on both sides

Example 3: atom tracking

$Ca + 2\ HCl \rightarrow CaCl_2 + H_2$

1 Ca 1

2 H 2

2 Cl 2 The atom count is = on both sides

- conservation of energy (energy is neither created nor destroyed)

Example: $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(\ell)}$ + energy

This example is an endothermic reaction (bond formation) but how we do not know if the PE of the ions on the left is equal to the heat released on the right

Table I shows that the amount of heat energy released should be 55.8 kJ/mol

$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(\ell)} + 55.8\ \text{kJ}$ (an energy balanced equation)

- conservation of charge (the total charge must be the same on both sides of an equation)

Example: Which equations represents the conservation of charge?

(1) $Cu + Ag^{+} \rightarrow Cu^{2+} + Ag$	$1+ \neq 2+$	
(2) $Mg + Zn^{2+} \rightarrow 2\ Mg^{2+} + Zn$	$2+ \neq 4+$	
(3) $2\ F_2 + Br^{-} \rightarrow 2\ F^{-} + Br_2$	$1- \neq 2-$	
(4) $2\ I^{-} + Cl_2 \rightarrow I_2 + 2\ Cl^{-}$	$2- = 2-$	(this is the correct answer)

Balanced equations show:

- changes in bonding of atoms
- reactants (starting substances, R)
- products (substances that form, P)
- + signs to separate substances
- → (yields sign) separates R and P
- coefficients (numbers) to balance atoms
- phase symbols (s), (ℓ), (g), or (aq)

Sample balanced equations:

$Ca_{(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2\,(aq)} + H_{2\,(g)}$

$CaCO_{3\,(s)} + \text{energy} \rightarrow CaO_{(s)} + CO_{2\,(g)}$

$Cu_{(s)} + 4\ HNO_{3\,(aq)} \rightarrow Cu(NO_3)_{2\,(aq)} + 2\ NO_{2\,(g)} + H_2O_{(\ell)} + \text{energy}$

Hints for balancing equations:

- the 'gens' are diatomic as elements
  H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2 (How Nice O F Cleo to Bring Ice)
- start with the most complicated particle, or start with a singular occurring element, or work symmetrically left – right – left
- treat polyatomic ions as *one* thing
- if H_2O is present, balance H and O last
- check your work with atom tracking

Examples of how to balance equations:

sodium and chlorine form table salt

write correct symbols	$Na + Cl \rightarrow NaCl$		
correct all formulas	$Na + Cl_2 \rightarrow NaCl$		
use coefficients L → R to fix # Cl atoms	$Na + Cl_2 \rightarrow 2\ NaCl$		
use coefficients R → L to fix # Na atoms	$2\ Na + Cl_2 \rightarrow 2\ NaCl$		
atom track to check work	2 Na 2		
	2 Cl 2		work checks

iron(III) oxide and carbon monoxide forms iron and carbon monoxide

write correct symbols	$FeO + CO \rightarrow Fe + CO$		
correct all formulas	$Fe_2O_3 + CO \rightarrow Fe + CO_2$		
use coefficients L → R to fix # Fe atoms	$Fe_2O_3 + CO \rightarrow 2\ Fe + CO_2$		
back to L, must remove 3 O w/ 3 CO	$Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$		
back to R to fix # C atoms	2 Fe 2		
	6 O 6		
	3 C 3		work checks

Using reaction types to help balance equations:

Types of chemical reactions:

- Synthesis — $A + X \rightarrow AX$
  $2\ H_{2\,(g)} + CO_{2\,(g)} \rightarrow 2\ H_2O_{(g)}$
- Decomposition — $AX \rightarrow A + X$
  $H_2CO_{3\,(aq)} \rightarrow CO_{2\,(g)} + H_2O_{(\ell)}$
- Single Replacement — $X + AY \rightarrow AX + Y$
  $Zn_{(s)} + CuSO_{4\,(aq)} \rightarrow ZnSO_{4\,(aq)} + Cu_{(s)}$
- Double Replacement — $AX + BY \rightarrow AY + BX$
  $NaCl_{(aq)} + AgNO_{3\,(aq)} \rightarrow NaNO_{3\,(aq)} + AgCl_{(s)}$
- Combustion — $CH + O_2 \rightarrow CO_2 + H_2O$
  $CH_{4\,(g)} + 2\ O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\ H_2O_{(g)}$

Examples of using reaction types to balance equations:

Example 1: finish then balance the following equation:

$Mg_{(s)} + AgNO_{3\,(aq)} \rightarrow$

Match to generalized equation

$A + BX \rightarrow AX + B$ single-replacement

Swap A and B to write the products

$Mg_{(s)} + AgNO_{3\,(aq)} \rightarrow MgNO_{3\,(aq)} + Ag_{(s)}$

Check new formulas, use criss-cross if needed

$Mg^{2+}NO_3^{1-} \rightarrow Mg(NO_3)_{2\,(aq)}$

Work L→R→L, start with $(NO_3^{1-})_2$ on the right → add coefficient to left

$Mg_{(s)} + 2\ AgNO_{3\,(aq)} \rightarrow Mg(NO_3)_{2\,(aq)} + Ag_{(s)}$

Work L→R→L, now have 2 Ag on left, add coefficient to make 2 Ag on the right

$Mg_{(s)} + 2\ AgNO_{3\,(aq)} \rightarrow Mg(NO_3)_{2\,(aq)} + 2\ Ag_{(s)}$

Example 1: finish then balance the following equation:

$NH_{3\,(g)} \rightarrow$

Match to generalized equation

$AX \rightarrow A + X$ decomposition

Decompose to A and X to write the products

$NH_{3\,(g)} \rightarrow N_{(g)} + H_{(g)}$

Check new formulas, the '*gens*' are diatomic molecules

$NH_{3\,(g)} \rightarrow N_{2\,(g)} + H_{2\,(g)}$

Work L→R→L, start with N_2 on the right → add coefficient to left

$2\ NH_{3\,(g)} \rightarrow N_{2\,(g)} + H_{2\,(g)}$

Work L→R→L, now have 6 H on left, add coefficient to make 6 H on the right

$2\ NH_{3\,(g)} \rightarrow N_{2\,(g)} + 3\ H_{2\,(g)}$

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Mole and Mass Stoichiometry Problems

Conservation of Mass

Example: How many grams of NaCl can be formed by the complete reaction of 46 grams of sodium with 71 grams of chlorine?

$2\ Na + Cl_2 \rightarrow 2\ NaCl$

$46\ g + 71\ g = ?\ g$ The total mass is the same on both sides, $46\ g + 71\ g = 117\ g$

$= 117\ g$

Mole to Mole

Coefficients in balanced equations represent mole to mole ratios of reactants and products

The coefficients can be used as conversion factors to calculate the correct relative amounts of reactants and products to satisfy the requirements of the conservation of mass

Example: Given the following reaction, how many moles of oxygen gas are required to completely react with 47.5 moles of aluminum metal?

$2\ Al_{(s)} + 3\ O_{2\,(g)} \rightarrow 2\ Al_2O_{3\,(s)}$

$$47.5\ \text{mol}_{Al} \times \frac{3\ \text{mol}_{O_2}}{4\ \text{mol}_{Al}} = 35.6\ \text{mol}_{O_2}$$