EXAM 1
Review Session

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YALE UNIVERSITY

CHEMISTRY 161

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www.mioy.org/chem161

OUTLINE

1. Significant Figures
2. Dimensional Analysis
3. Elements and Atoms
4. Naming Compounds
5. The Mole: Atomic Mass, Molar Mass, and Avogadro's Number
6. Chemical Composition: Mass Percent, Empirical/Molecular Formulas
7. Balancing Chemical Equations
8. Stoichiometry and Mole-Mole Ratios
9. Limiting Reactants & Yields

Counting Significant Figures

- Nonzero numbers are always significant

What about zeroes?

- Leading zeroes = <u>not</u> significant
- Captive/trapped zeroes = significant
- Trailing zeroes (<u>after</u> decimal) = significant

0.0003040 meters

<u>not</u> significant   significant

Calculations with Significant Figures

ROUND AT THE END

Multiplication/Division:

number with *less* significant figures

Addition/Subtraction:

the less precise number

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$$7.23 + 70 + 3.7795 = 81.0795$$

$$\frac{4.771 + 2.3}{3.12} = 2.26634615$$

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Addition/Subtraction:
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Addition/Subtraction:
the less precise number

$$\frac{4.771 + 2.3}{3.12} = 2.26634615 = 2.3$$

Back to the Basics: Fractions

We can do these types of calculations already

$$\frac{4}{9} \times \frac{1}{6} = \frac{4}{54} = \frac{2}{27}$$

$$\frac{4}{5} \times \frac{3}{4} \times \frac{5}{2} = \frac{3}{2}$$

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Extend Fractions to Multiplying Units

Unit conversion is just multiplying fractions!

$$\frac{\cancel{4}}{\cancel{5}} \times \frac{3}{\cancel{4}} \times \frac{\cancel{5}}{2} = \frac{3}{2}$$

$$\frac{centimeter}{seconds} \times \frac{meter}{centimeter} \times \frac{kilometer}{meter} = \frac{kilometer}{second}$$

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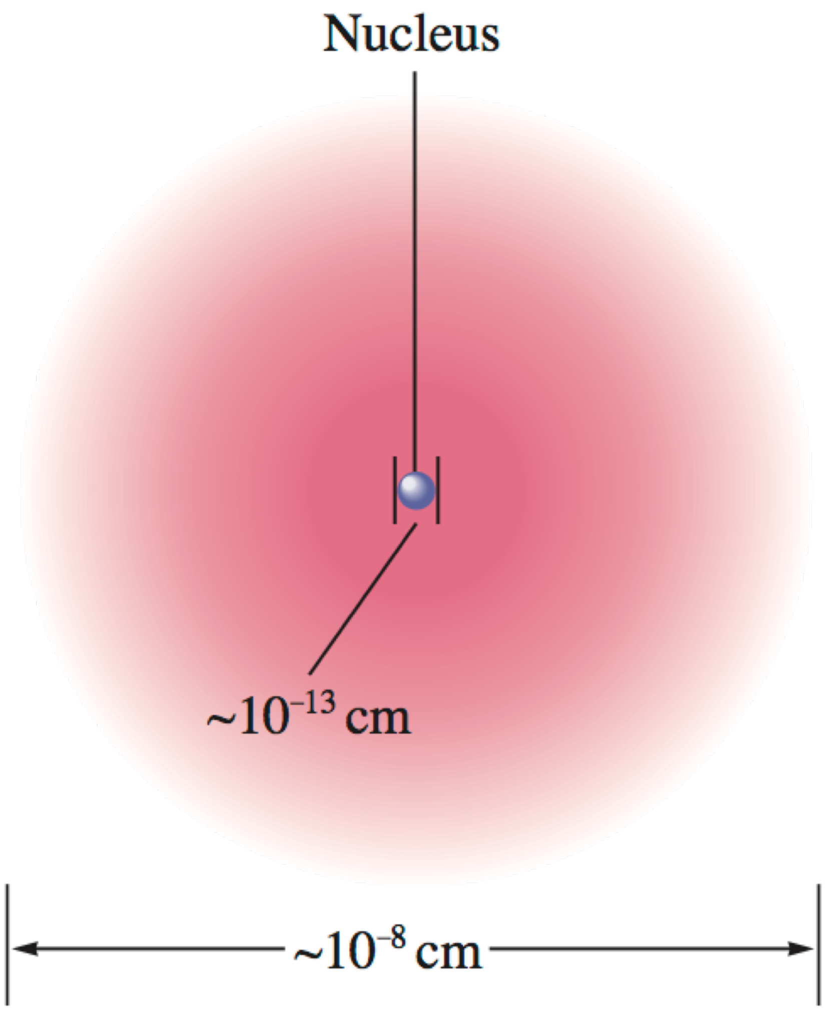
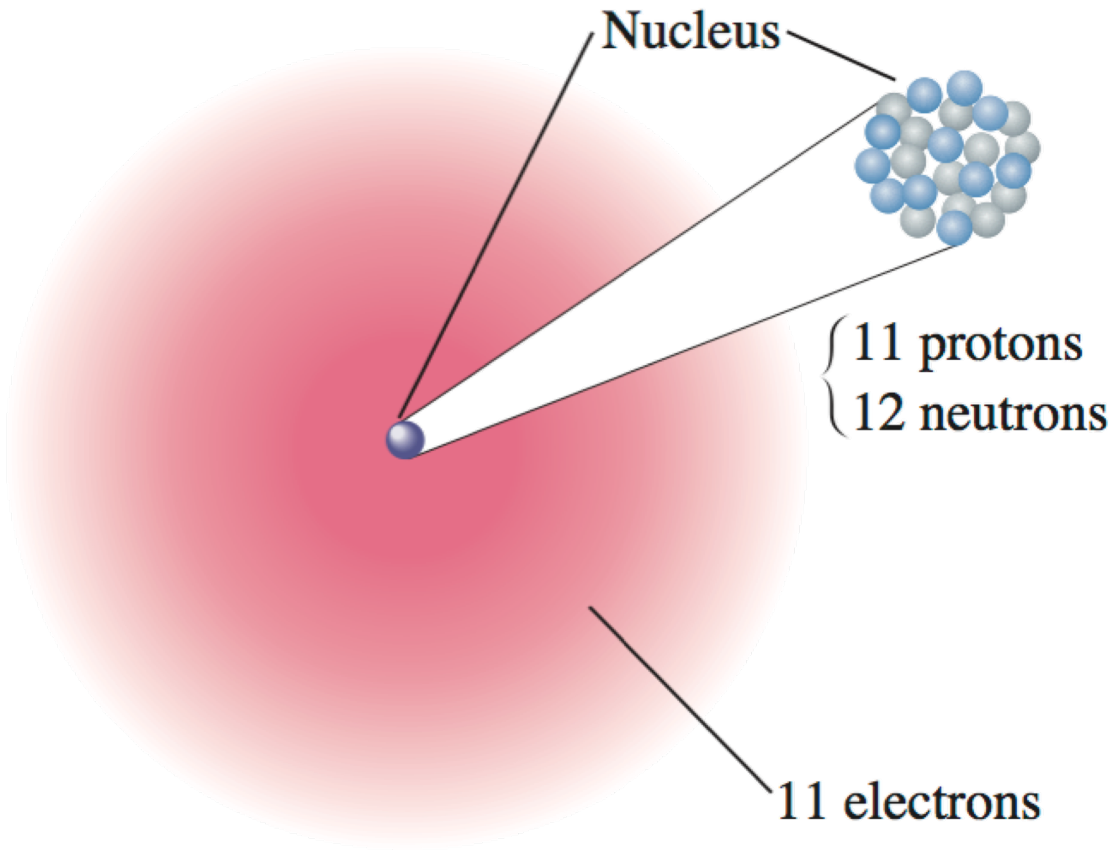
The Atom

PARTICLE	MASS	CHARGE
Electron	9.11×10^{-31} kg	1–
Proton	1.67×10^{-27} kg	1+
Neutron	1.67×10^{-27} kg	0

The nucleus is very dense:
A proton/neutron is ~2000 times heavier than an electron



Atomic Symbols



$^{23}_{11}\mathrm{Na}$

MASS NUMBER (A)
protons + neutrons

$^{A}_{Z}X$

ATOMIC NUMBER (Z)
# of protons
# of electrons

NEUTRONS
= A – Z

Atomic Symbols:
Practice Problems

SYMBOL	$^{64}_{30}Zn$	$^{32}_{16}S$	
# Protons			40
# Neutrons			50
# Electrons			40
Mass Number			90

MASS NUMBER (A)
protons + neutrons

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= A – Z

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Atomic Symbols:
Practice Problems

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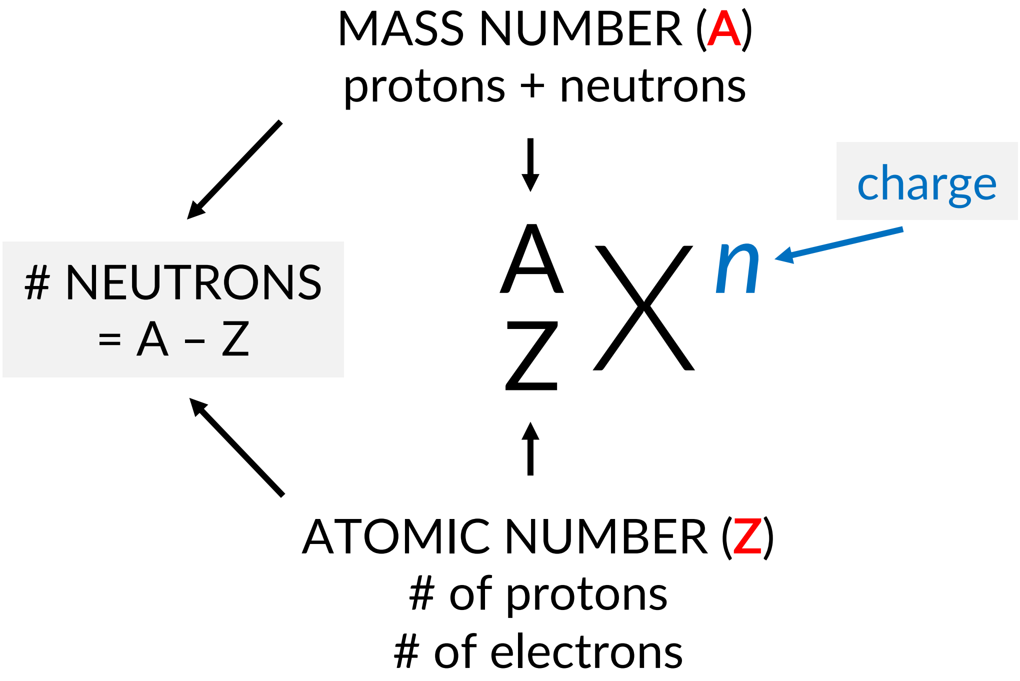
SYMBOL	${}^{64}_{30}\mathrm{Zn}^{2+}$	${}^{32}_{16}\mathrm{S}^{-}$	${}^{90}_{40}\mathrm{Zr}^{4+}$
# Protons			
# Neutrons			
# Electrons			
Mass Number			



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MASS NUMBER (A)
protons + neutrons

NEUTRONS = A – Z

$${}^{A}_{Z}X^{n}$$

charge

ATOMIC NUMBER (Z)
# of protons
# of electrons

Atomic Symbols: Practice Problems

SYMBOL	$^{64}_{30}\mathrm{Zn}$	$^{32}_{16}\mathrm{S}$	$^{90}_{40}\mathrm{Zr}$
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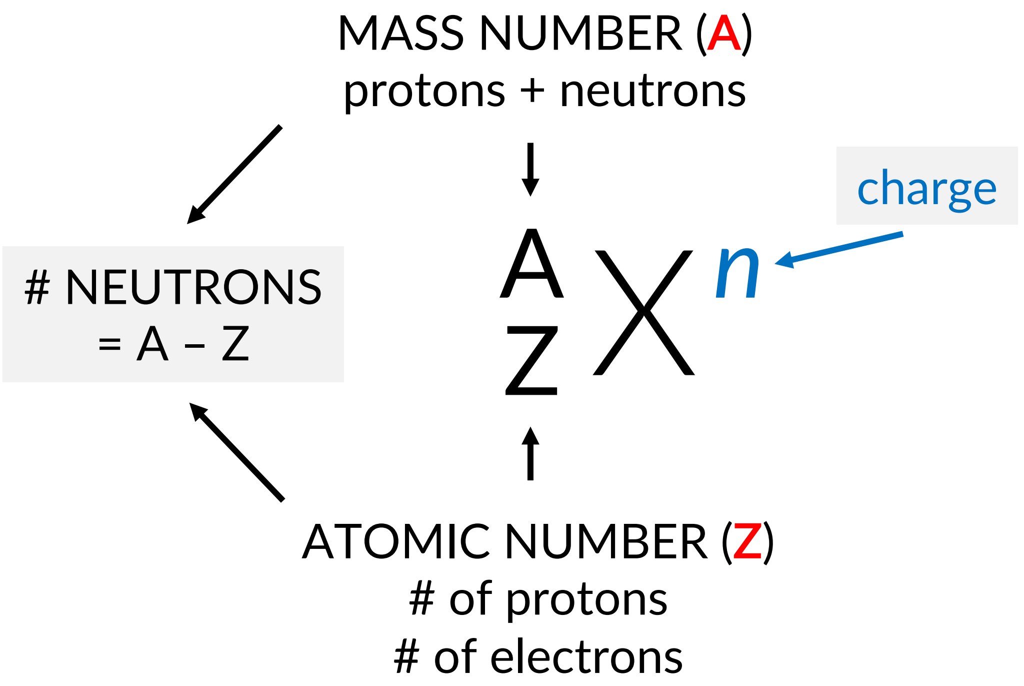
# Neutrons	34	16	50
# Electrons	28	18	36
Mass Number	64	32	90



General Rules for Naming Compounds

IONIC

- Metal + Nonmetal
- Cation + Anion
- Must be **neutral** overall!

Naming:

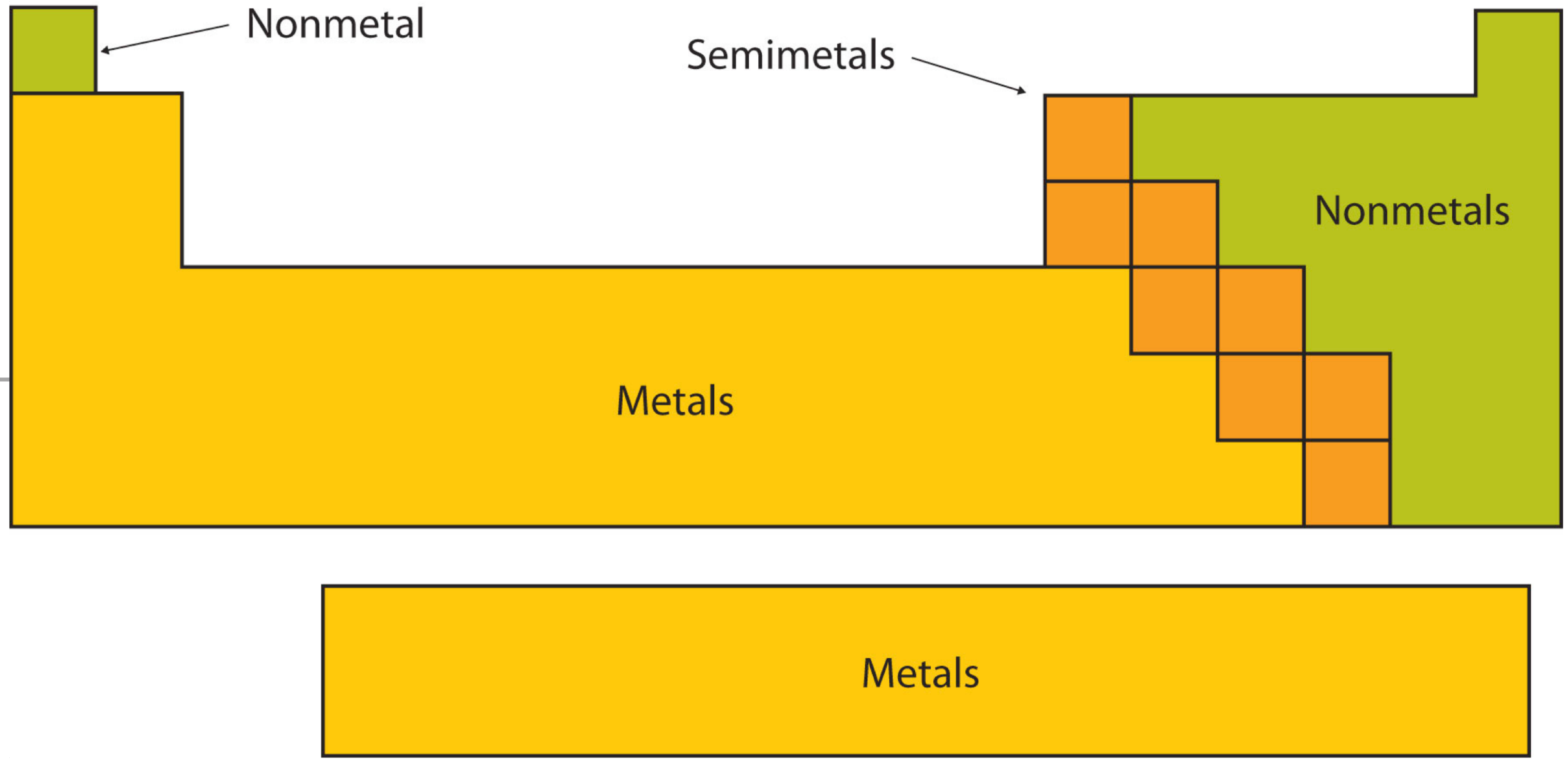
- Cation + Anion Root + “-ide”
  NaCl → Sodium Chloride

MOLECULAR

- Nonmetal + Nonmetal

Naming:

- 1^{st} element: full name
- 2^{nd} element: root + “-ide”
- Use prefixes (Table 2.2)
  BF_3 → Boron Trifluoride



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Naming:

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More exotic rules

- Transition metals require charge
  Hint: Find charge of anion first!
- Cation + Charge + Anion Root + “-ide”
  $FeCl_2$ → Iron (II) Chloride
  PbO_2 → Lead (IV) Oxide
- Polyatomics are “one ion” (Table 2.3)
  AgCN → Silver (I) Cyanide

MOLECULAR

- Nonmetal + Nonmetal

Naming:

- 1st element: full name
- 2nd element: root + “-ide”
- Use prefixes (Table 2.2)
  BF_3 → Boron Trifluoride

More exotic rules

- Don’t use “mono-” for first atom
  NO → Nitrogen Monoxide
- Drop “extra” vowels
  N_2O_5 → Dinitrogen Pentoxide
- Oxoanions: -ate has more O’s than -ite
  NO_3^- → Nitrate
  NO_2^- → Nitrite

How should I think about atomic mass?

- Atoms are very small and very light
- Periodic table reports the measured *average atomic mass*
  - 1 hydrogen atom = 1.008 amu (atomic mass unit)
  - Makes it really impractical to work with in a laboratory/life

Convince yourself of this:

The mass of one atom (even one molecule) is horribly small.

For instance: 1 H atom weighs 1.674×10^{-24} grams

Recognize this is not a useful unit to use in life!

How should I think about atomic mass?

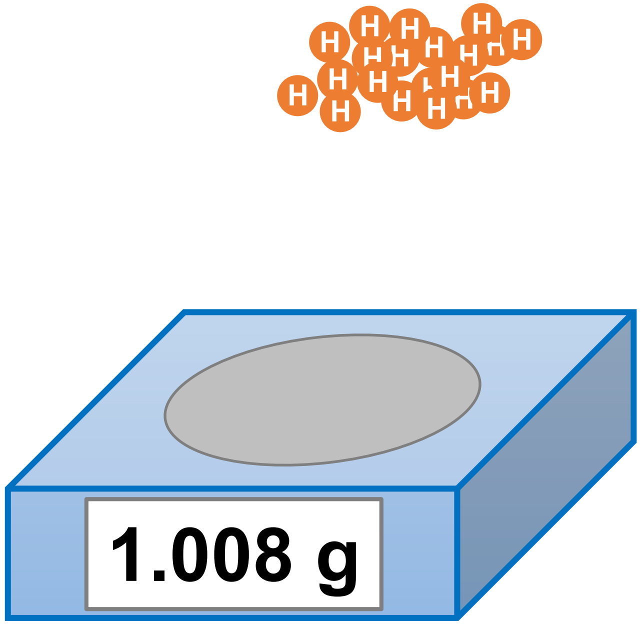
- Ask yourself: If an amu is useless, what units do we want to use for mass? [***Answer: grams***]

- So, we would like if: **H**: 1.008 amu means 1.008 g
  C: 12.01 amu means 12.01 g

- But how…?

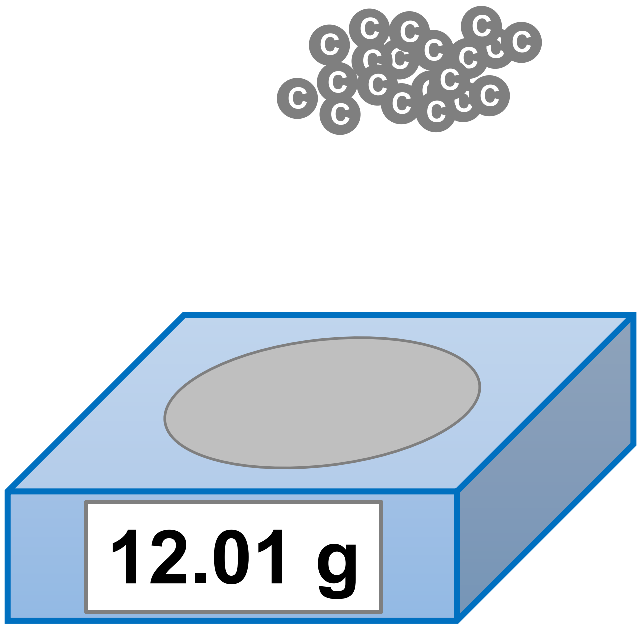
If you placed 2 hydrogen atoms onto a scale, it still wouldn’t register.

Now, imagine placing hydrogen atoms onto a scale until you get 1.008 g.



If you placed 2 carbon atoms onto a scale, it still wouldn’t register.

Now, imagine placing carbon atoms onto a scale until you get 12.01 g.



How should I think about atomic mass?

- Eventually you will put enough atoms to achieve a reading of 1.008 g for hydrogen and 12.01 g for carbon. In other words:

	1 atom	"x" atoms
H	1.008 amu	1.008 g
C	12.01 amu	12.01 g

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- But how many atoms ("x") did we need to get that mass in g?

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C	12.01 amu	12.01 g

- But how many atoms ("x") did we need to get that mass in g?

$$1.008\ \mathrm{g\ H} \times \frac{1\ \mathrm{atom\ H}}{1.674\times10^{-24}\ \mathrm{g}} = 6.022\times10^{23}\ \mathrm{atoms\ H}$$

Look familiar?!

Avogadro's Number = 6.022×10^{23} and the MOLE

- Avogadro's number is what connects the atomic masses (in amu) to masses in grams! So…

1.008 g H	=	6.022×10^{23} atoms H	=	1 mol H
12.01 g C	=	6.022×10^{23} atoms C	=	1 mol C
16.00 g O	=	6.022×10^{23} atoms O	=	1 mol O
14.01 g N	=	6.022×10^{23} atoms N	=	1 mol N

- Take-home message: The units on the periodic table that are useful are g/mol (molar mass). Often, we are dealing with large numbers of atoms/molecules, so g/mol is easier to deal with.

How many atoms of Mg are in 10.0 g Mg?

How many atoms of Mg are in 10.0 g Mg?

The molar mass of Mg is 24.31 g/mol.

Let's do it in steps just for clarity.

First, convert mass to moles using molar mass of Mg:

$$10.0\ \text{g}\ \cancel{\text{Mg}} \times \frac{1\ \text{mol Mg}}{24.31\ \text{g}\ \cancel{\text{Mg}}} = 0.411_4\ \text{mol Mg}$$

Second, convert moles to atoms using Avogadro's number:

$$0.411_4\ \cancel{\text{mol}}\ \text{Mg} \times \frac{6.022 \times 10^{23}\ \text{atoms}}{1\ \cancel{\text{mol}}} = 2.48 \times 10^{23}\ \text{atoms Mg}$$

Everything in one giant conversion:

$$10.0\ \text{g}\ \cancel{\text{Mg}} \times \frac{1\ \cancel{\text{mol}}\ \text{Mg}}{24.31\ \text{g}\ \cancel{\text{Mg}}} \times \frac{6.022 \times 10^{23}\ \text{atoms}}{1\ \cancel{\text{mol}}} = 2.48 \times 10^{23}\ \text{atoms Mg}$$

SUMMARY: MOLE IS CENTRAL

use: $\dfrac{6.022\times10^{23}\text{ atoms}}{1\text{ mol}}$ or $\dfrac{1\text{ mol}}{6.022\times10^{23}\text{ atoms}}$

 # ATOMS
– or –
 # MOLECULES

↔

 # MOLES

↔

MASS (in grams)

use molar mass from periodic table: $\dfrac{12.01\text{ g C}}{1\text{ mol C}}$ or $\dfrac{1\text{ mol C}}{12.01\text{ g C}}$

Which has more atoms?

2 kg of Cu or 2 kg Mg

Which has more atoms?

2 kg of Cu or 2 kg Mg

The atomic mass of Cu is 63.55 amu.

The atomic mass of Mg is 24.31 amu.

→ Since each atom of Mg weighs less (1 atom Mg = 24.31 amu), it will take more atoms of Mg to reach 2 kg of Mg.

Avogadro's number also applies to molecules/compounds

- Similar to how we don't really encounter the mass of one atom, we rarely encounter the mass of one molecule in real life.
- Again, these masses would be very small numbers.
- For compounds or molecules, we also work in moles.
- The mass of 1 mole of a compound:

  Ex.) The molar mass of H_2O is 18.02 g/mol.

$$\begin{array}{lcccccc}
1\text{ molecule }H_2O & \rightarrow & 2\text{ atoms H} & + & 1\text{ atom O} & & \\
1\text{ mol }H_2O & \rightarrow & 2\text{ mol H} & + & 1\text{ mol O} & & \\
\hline
1\text{ mol }H_2O & = & 2\,(1.008\text{ g}) & + & 1\,(16.00\text{ g}) & = & 18.02\text{ g }H_2O
\end{array}$$

Find the molar mass of ammonium carbonate.

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First, write the formula: $(NH_4)_2CO_3$

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$$2\text{ mol N} \rightarrow 2\text{ mol N} \times \frac{14.01\text{ g N}}{1\text{ mol N}} = 28.02\text{ g N}$$

$$8\text{ mol H} \rightarrow 8\text{ mol H} \times \frac{1.008\text{ g H}}{1\text{ mol H}} = 8.064\text{ g H}$$

$$1\text{ mol C} \rightarrow 1\text{ mol C} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = 12.01\text{ g C}$$

$$3\text{ mol O} \rightarrow 3\text{ mol O} \times \frac{16.00\text{ g O}}{1\text{ mol O}} = 48.00\text{ g O}$$

96.09 g

1 mol $(NH_4)_2CO_3$ = 2 mol N + 8 mol H + 1 mol C + 3 mol O
= 2 (14.01 g) + 8 (1.008 g) + 1 (12.01 g) + 3 (16.00 g)
= 96.09 g

SUMMARY: MOLE IS CENTRAL

use: $\dfrac{6.022\times10^{23}\ \text{molecules}}{1\ \text{mol}}$ or $\dfrac{1\ \text{mol}}{6.022\times10^{23}\ \text{molecules}}$

 # ATOMS
– or –
 # MOLECULES

⟷

 # MOLES

⟷

MASS (in grams)

use molar mass of compound: $\dfrac{18.02\ \text{g}\ H_2O}{1\ \text{mol}\ H_2O}$ or $\dfrac{1\ \text{mol}\ H_2O}{18.02\ \text{g}\ H_2O}$

How many molecules are in a 50.0 g sample of ammonium carbonate?

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The molar mass of $(NH_4)_2CO_3$ is 96.09 g/mol.

The procedure is similar to what we did for atoms, but now use *molecules*.

First, convert from mass to number of moles using the molar mass:

$$50.0\text{ g }(NH_4)_2CO_3 \times \frac{1\text{ mol }(NH_4)_2CO_3}{96.09\text{ g }(NH_4)_2CO_3} = 0.520_3\text{ mol }(NH_4)_2CO_3$$

Second, convert moles to number of *molecules* using Avogadro's number:

$$0.520_3\text{ mol }(NH_4)_2CO_3 \times \frac{6.022 \times 10^{23}\text{ molecules}}{1\text{ mol}} = 3.13 \times 10^{23}\text{ molecules }(NH_4)_2CO_3$$

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Everything in one giant conversion:

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How many hydrogen atoms are in a 50.0 g sample of ammonium carbonate?

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Remember that:

- 1 molecule of $(NH_4)_2CO_3$ = 2 atoms N + 8 atoms H + 1 atom C + 3 atoms O
- 1 mole $(NH_4)_2CO_3$ = 2 mole N + 8 mole H + 1 mole C + 3 mole O

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- 1 mole $(NH_4)_2CO_3$ = 2 mole N + 8 mole H + 1 mole C + 3 mole O

So:

$$3.13_4 \times 10^{23}\ \mathrm{molecules}\ (\mathrm{NH_4})_2\mathrm{CO_3} \times \frac{8\ \mathrm{atoms\ H}}{1\ \mathrm{molecule}\ (\mathrm{NH_4})_2\mathrm{CO_3}} = 2.51 \times 10^{24}\ \mathrm{atoms\ H}$$

If you have equal mole samples of each compound, which contains the greatest number of oxygen atoms?

Magnesium nitrate

Dinitrogen pentoxide

Iron(III) phosphate

Barium oxide

Potassium acetate

If you have equal mole samples of each compound, which contains the greatest number of oxygen atoms?

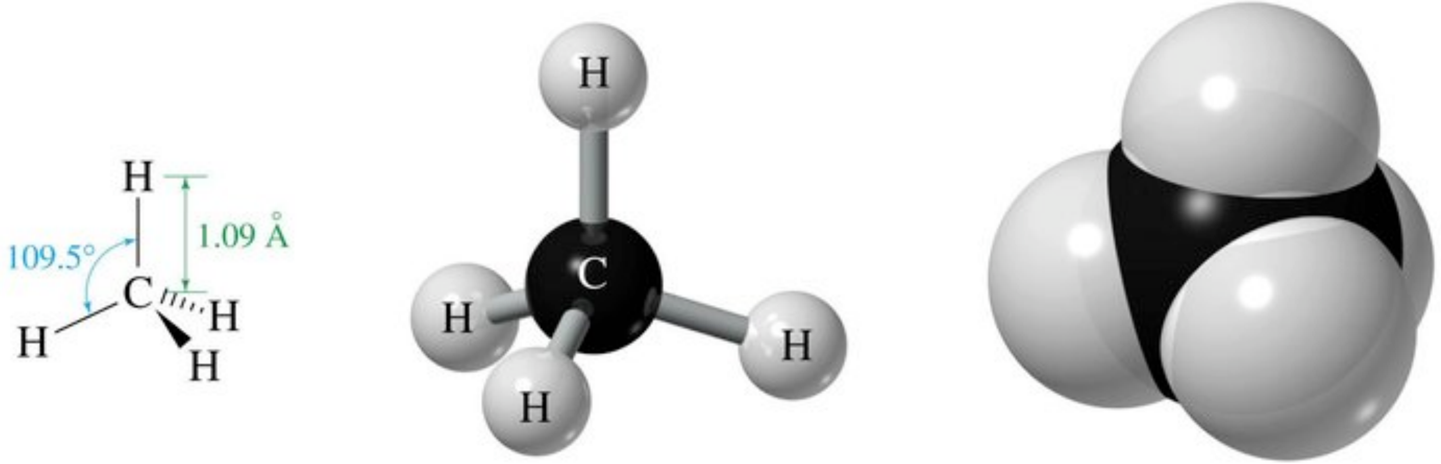
	Formula	Moles of O
Magnesium nitrate	$Mg(NO_3)_2$	
Dinitrogen pentoxide	N_2O_5	
Iron(III) phosphate	$FePO_4$	
Barium oxide	BaO	
Potassium acetate	KCH_3CO_2	

If you have equal mole samples of each compound, which contains the greatest number of oxygen atoms?

	Formula	Moles of O
Magnesium nitrate	$Mg(NO_3)_2$	6 mol O
Dinitrogen pentoxide	N_2O_5	5 mol O
Iron(III) phosphate	$FePO_4$	4 mol O
Barium oxide	BaO	1 mol O
Potassium acetate	KCH_3CO_2	2 mol O

Introduction to mass percent

CHEMISTS CARE ABOUT MASS PERCENT!



$$\% \text{ Mass} = \frac{\text{mass part}}{\text{mass whole}} \times 100\%$$

How do I calculate the mass percentages for 1 mol CH_4?

- Remember that the molar mass of CH_4 is 16.04 g/mol:

$$
\begin{aligned}
1 \text{ mol } CH_4 &= 1 \text{ mol C} + 4 \text{ mol H} \\
&= 1\,(12.01 \text{ g}) + 4\,(1.008 \text{ g}) = 16.04 \text{ g}
\end{aligned}
$$

$$\%\ \text{Mass} = \frac{\text{mass part}}{\text{mass whole}} \times 100\%$$

$$\%\ \text{C} \rightarrow \frac{1(12.01)\ \text{g}}{16.04\ \text{g}} \times 100\% = 74.90\%\ \text{C}$$

$$\%\ \text{H} \rightarrow \frac{4(1.008)\ \text{g}}{16.04\ \text{g}} \times 100\% = 25.10\%\ \text{H}$$

$$100.0\%\ \text{total}$$

How do I calculate the mass percentages for 2 mol CH_4?

- The molar mass of CH_4 is 16.04 g/mol, but now:

$$2 \text{ mol } CH_4 = 2 \text{ mol C} + 8 \text{ mol H}$$

$$= 2\,(12.01 \text{ g}) + 8\,(1.008 \text{ g}) = 32.08 \text{ g}$$

$$\%\,\text{Mass} = \frac{\text{mass part}}{\text{mass whole}} \times 100\%$$

$$\%\,\text{C} \rightarrow \frac{2(12.01)\text{ g}}{32.08\text{ g}} \times 100\% = 74.90\%\text{ C}$$

$$\%\,\text{H} \rightarrow \frac{8(1.008)\text{ g}}{32.08\text{ g}} \times 100\% = 25.10\%\text{ H}$$

$$100.0\%\text{ total}$$

Percent composition is <u>independent</u> of the starting amount!

This is why we usually *assume* we have 100 g or 1 mol.
These are just super easy numbers to work with.

<u>Note</u>: If you wanted to use a strange amount, like 0.27 mol or 74.5 g of substance, your answers would be the same but the math isn't as convenient. BUT you'll still be right. ☺

What is the mass percent of nitrogen in barium nitrate?

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First, write the formula: $Ba(NO_3)_2$

Second, determine the molar mass = 261.35 g/mol

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Second, determine the molar mass = 261.35 g/mol

We know that there are 2 mol N for 1 mol of $Ba(NO_3)_2$, so:

$$\frac{2(14.01)\ \mathrm{g}}{261.35\ \mathrm{g}} \times 100\% = 10.72\%\ \mathrm{N}$$

If you have equal mass samples of each compound, which contains the greatest number of oxygen atoms?

	Formula
Magnesium nitrate	$Mg(NO_3)_2$
Dinitrogen pentoxide	N_2O_5
Iron(III) phosphate	$FePO_4$
Barium oxide	BaO
Potassium acetate	KCH_3CO_2

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One way to solve this is to assume you have 100 g of each compound, then find the total number of O atoms in each compound.

This is a lot of work though!

Consider mass percentages!

If you have equal mass samples of each compound, which contains the greatest number of oxygen atoms?

	Formula	
Magnesium nitrate	$Mg(NO_3)_2$	$\frac{6(16.00\text{ g})}{148.33\text{ g}} \times 100\% = 64.72\%\text{ O}$
Dinitrogen pentoxide	N_2O_5	$\frac{5(16.00\text{ g})}{108.02\text{ g}} \times 100\% = 74.06\%\text{ O}$
Iron(III) phosphate	$FePO_4$	$\frac{4(16.00\text{ g})}{150.82\text{ g}} \times 100\% = 42.43\%\text{ O}$
Barium oxide	BaO	$\frac{1(16.00\text{ g})}{153.3\text{ g}} \times 100\% = 10.4\%\text{ O}$
Potassium acetate	KCH_3CO_2	$\frac{2(16.00\text{ g})}{98.14\text{ g}} \times 100\% = 32.61\%\text{ O}$

Most often, we use mass percentages to help us figure out what compound we have.

These are called EMPIRICAL FORMULAS.

You have some “nitrogen oxide” compound and you want to figure out what it is (both formula and name). You know that it’s 46.7% nitrogen by mass.

- We want to know: N_xO_y = formula? name?

- Remember: the amount doesn’t matter for percent composition!
- Let’s assume we have 100 g of our N_xO_y.
- This means that for every 100 g of N_xO_y, we have:
  - 46.7 g of N
  - 53.3 g of O

You have some "nitrogen oxide" compound and you want to figure out what it is (both formula and name). You know that it's 46.7% nitrogen by mass.

- We want to know: N_xO_y = formula? name?

We must convert the masses to *moles*.

$$\mathrm{N} \quad \rightarrow \quad 46.7\ \mathrm{g\ N} \times \frac{1\ \mathrm{mol\ N}}{14.01\ \mathrm{g\ N}} \quad = \quad 3.33\ \mathrm{mol\ N}$$

$$\mathrm{O} \quad \rightarrow \quad 53.3\ \mathrm{g\ O} \times \frac{1\ \mathrm{mol\ O}}{16.00\ \mathrm{g\ O}} \quad = \quad 3.33\ \mathrm{mol\ O}$$

You have some "nitrogen oxide" compound and you want to figure out what it is (both formula and name). You know that it's 46.7% nitrogen by mass.

- We want to know: N_xO_y = formula? name?

We need the simplest whole number ratio!

$$N \rightarrow 46.7 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 3.33 \text{ mol N} \rightarrow \frac{3.33 \text{ mol N}}{3.33} = 1 \text{ N}$$

$$O \rightarrow 53.3 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.33 \text{ mol O} \rightarrow \frac{3.33 \text{ mol O}}{3.33} = 1 \text{ O}$$

Divide the number of moles by the SMALLEST value!

You have some "nitrogen oxide" compound and you want to figure out what it is (both formula and name). You know that it's 46.7% nitrogen by mass.

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Empirical Formula:
NO | Nitrogen Monoxide

Aluminum oxide (Al_xO_y) is 41.51% Al and 36.92% O. Determine the empirical formula.

General procedure:

1. Assume a 100 g sample.

$$\mathrm{Al} \rightarrow 41.51\ \mathrm{g\ Al}$$

$$\mathrm{O} \rightarrow 36.92\ \mathrm{g\ O}$$

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General procedure:

1. Assume a 100 g sample.
2. Convert masses to moles.

$$\mathrm{Al} \quad \rightarrow \quad 41.51\ \mathrm{g\ Al} \times \frac{1\ \mathrm{mol\ Al}}{26.98\ \mathrm{g\ Al}} = 1.54\ \mathrm{mol\ Al}$$

$$\mathrm{O} \quad \rightarrow \quad 36.92\ \mathrm{g\ O} \times \frac{1\ \mathrm{mol\ O}}{16.00\ \mathrm{g\ O}} = 2.31\ \mathrm{mol\ O}$$

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General procedure:

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Uh... these aren't whole numbers!

Aluminum oxide (Al_xO_y) is 41.51% Al and 36.92% O. Determine the empirical formula.

General procedure:

1. Assume a 100 g sample.
2. Convert masses to moles.
3. Divide the mole amounts by the smallest mole value.
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$$\text{Al} \rightarrow 41.51\text{ g Al} \times \frac{1\text{ mol Al}}{26.98\text{ g Al}} = 1.54\text{ mol Al} \rightarrow \frac{1.54\text{ mol Al}}{1.54} = 1\text{ Al} \quad \times 2 \rightarrow 2\text{ Al}$$

$$\text{O} \rightarrow 36.92\text{ g O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 2.31\text{ mol O} \rightarrow \frac{2.31\text{ mol O}}{1.54} = 1.5\text{ O} \quad \times 2 \rightarrow 3\text{ O}$$

Multiply to get whole numbers!

Aluminum oxide (Al_xO_y) is 41.51% Al and 36.92% O. Determine the empirical formula.

General procedure:

1. Assume a 100 g sample.
2. Convert masses to moles.
3. Divide the mole amounts by the smallest mole value.
4. Write empirical formula from simplest whole number ratio.

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Al_2O_3

Calculate the empirical formula for a halohydrocarbon if it is 71.65% Cl, 24.27% C, and 4.07% H by mass.

Assuming a 100 g sample of the halohydrocarbon ($Cl_aC_bH_c$):

$$\mathrm{Cl} \rightarrow 71.65\ \mathrm{g\ Cl} \times \frac{1\ \mathrm{mol\ Cl}}{35.45\ \mathrm{g\ Cl}} = 2.021\ \mathrm{mol\ Cl} \rightarrow \frac{2.021\ \mathrm{mol\ Cl}}{2.021} = 1\ \mathrm{Cl}$$

$$\mathrm{C} \rightarrow 24.27\ \mathrm{g\ C} \times \frac{1\ \mathrm{mol\ C}}{12.01\ \mathrm{g\ C}} = 2.021\ \mathrm{mol\ C} \rightarrow \frac{2.021\ \mathrm{mol\ C}}{2.021} = 1\ \mathrm{C}$$

$$\mathrm{H} \rightarrow 4.07\ \mathrm{g\ H} \times \frac{1\ \mathrm{mol\ H}}{1.008\ \mathrm{g\ H}} = 4.04\ \mathrm{mol\ H} \rightarrow \frac{4.04\ \mathrm{mol\ H}}{2.021} = 2\ \mathrm{H}$$

The empirical formula is $ClCH_2$.

Determine the molecular formula for the same halohydrocarbon if it has a molar mass of 98.96 g/mol.

We determined previously that the empirical formula is $ClCH_2$.

The empirical formula mass is 49.48 g/mol.

The molecular formula is <u>always</u> a multiple of the empirical formula. So: $(ClCH_2)_n$

We can determine the multiple (n) by taking the ratio between the molecular formula's molar mass and the empirical formula mass:

$$n = \frac{Molar\ mass}{Empirical\ formula\ mass} = \frac{98.96\ \text{g}}{49.48\ \text{g}} = 2$$

The molecular formula is $(ClCH_2)_2$ or $Cl_2C_2H_4$.

A compound containing only sulfur and nitrogen is 69.6% S by mass. If its molar mass is 184 g/mol, what is the correct name for it?

First, determine the empirical formula and empirical formula mas:

$$\mathrm{S} \quad \rightarrow \quad 69.6\ \mathrm{g\ S} \times \frac{1\ \mathrm{mol\ S}}{32.06\ \mathrm{g\ S}} = 2.17\ \mathrm{mol\ S} \quad \rightarrow \quad \frac{2.17\ \mathrm{mol\ S}}{2.17} = 1\ \mathrm{S}$$

$$\mathrm{N} \quad \rightarrow \quad 30.4\ \mathrm{g\ N} \times \frac{1\ \mathrm{mol\ N}}{14.01\ \mathrm{g\ N}} = 2.17\ \mathrm{mol\ N} \quad \rightarrow \quad \frac{2.17\ \mathrm{mol\ N}}{2.17} = 1\ \mathrm{N}$$

So the empirical formula is SN and the empirical mass is 46.07 g.

We can determine the multiple (*n*) by taking the ratio between the molecular formula's molar mass and the empirical formula mass:

$$n = \frac{Molar\ mass}{Empirical\ formula\ mass} = \frac{184\ \mathrm{g}}{46.07\ \mathrm{g}} = 4$$

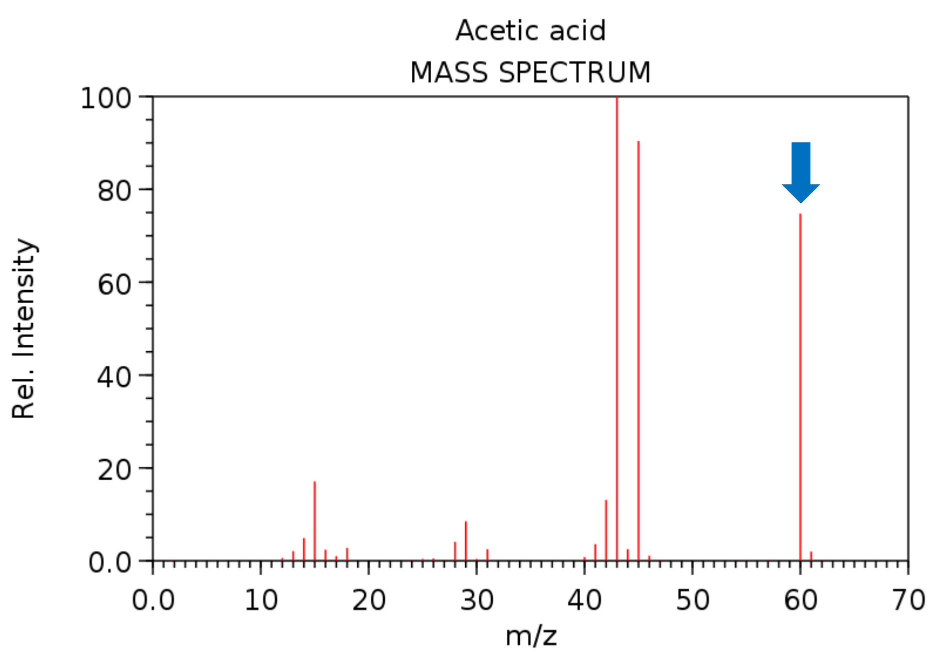
The molecular formula is $(SN)_4$ or S_4N_4 or tetrasulfur tetranitride.

MASS SPECTROMETRY

The peak farthest to the right (the one with the largest m/z value) tells you the *estimated* molecular mass of the compound.

Ex) Acetic Acid, C_2H_4O,

Molecular Mass = 60.05 g/mol



__Notes you don't need to know:__

- *The other peaks are broken fragments of the molecule.*
- *The height tells you the abundance of a fragment.*
- *We use mass spec. to identify elements with near-even (50%:50%) ratio of isotopes, like ^{79}Br and ^{81}Br.*

What do chemical equations tell us?

- Formulas for the reactants (left side)
- Formulas for the products (right side)
- Phases, most of the time
- **<u>Relative</u>** amounts of reactants of reactants and products

$$\text{REACTANT} \rightarrow \text{PRODUCTS}$$

What does it mean to be "balanced"?

- Same number of each type of atom on the left (reactants) and right (products) side.
- **Law of Conservation of Mass**

REACTANT $\rightarrow$ PRODUCTS

How do we balance chemical equations?

- Mainly trial-and-error (some general strategies though).
- Make sure you have the same number of each type of atom on both sides of the equation.
- Do <u>NOT</u> balance by changing subscripts! Seriously, don't.
- Balance the most complicated molecule *first*.

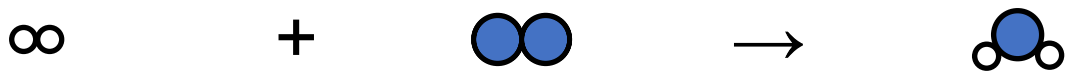
REACTANT → PRODUCTS

What makes a chemical reaction?

Hydrogen gas and oxygen gas react to form water vapor.

(words) hydrogen gas + oxygen gas → water vapor

(drawings)



Reactants	Products
2 H atoms	2 H atoms
2 O atoms	1 O atoms

Reactants	Products
4 H atoms	4 H atoms
4 O atoms	4 O atoms

(equation) $2H_2\ (g) + O_2\ (g) \rightarrow 2H_2O\ (g)$

Pictures aren't always convenient though…

How do I read a chemical equation?

COEFFICIENTS
Tell us the *relative* amounts of reactants/products

SUBSCRIPTS
Tell us the # of atoms in a molecule

PHASES
Tell us the nature of reactants/products

$$2H_2\ (g) \quad + \quad O_2\ (g) \quad \rightarrow \quad 2H_2O\ (g)$$

- Subscripts are not conserved!
- Coefficients have no real meaning by themselves…
- RATIO of coefficient is what's important.
- Read it like a recipe:

  "For every 2 H_2 molecules, we need 1 O_2 molecule to produce 2 H_2O molecules."

Write a <u>balanced chemical equation</u> for ammonia synthesis from nitrogen and hydrogen gases.

Write a balanced chemical equation for ammonia synthesis from nitrogen and hydrogen gases.

Write out the core of the equation from the description:

$$\underline{\ 1\ }\ N_2\ (g) + \underline{\ 3\ }\ H_2\ (g) \rightarrow \underline{\ 2\ }\ NH_3\ (g)$$

"To make 2 moles NH_3, we need 1 mole N_2 and 3 moles H_2."

If 5.00 g of CH_4 (methane) is burned, what mass of water can be produced?

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Write out the core of the equation from the description:

$$\_\_\_\ CH_4\,(g) + \_\_\_\ O_2\,(g) \rightarrow \_\_\_\ CO_2\,(g) + \_\_\_\ H_2O\,(g)$$

WHAT IS A COMBUSTION REACTION?	When we "burn" a hydrocarbon (a compound with C, H, and/or O atoms), it ***always*** reacts with O_2 gas in the air to form CO_2 and H_2O gases as products.

If 5.00 g of CH_4 (methane) is burned, what mass of water can be produced?

Write out the core of the equation from the description:

$$\underline{\ 1\ }\ CH_4\ (g) + \underline{\ 2\ }\ O_2\ (g) \rightarrow \underline{\ 1\ }\ CO_2\ (g) + \underline{\ 2\ }\ H_2O\ (g)$$

"For every 1 mol CH_4, we need to react with 2 mol O_2 to produce 1 mol CO_2 and 2 mol H_2O."

If 5.00 g of CH_4 (methane) is burned, what mass of water can be produced?

Write out the core of the equation from the description:

$$\underline{\ 1\ }\ CH_4\ (g)\ +\ \underline{\ 2\ }\ O_2\ (g)\ \rightarrow\ \underline{\ 1\ }\ CO_2\ (g)\ +\ \underline{\ 2\ }\ H_2O\ (g)$$

"For every 1 mol CH_4, we need to react with 2 mol O_2 to produce 1 mol CO_2 and 2 mol H_2O."

REMEMBER MY TIP: If you don't know how to start a problem, convert whatever they give you into **moles** first.

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REMEMBER MY TIP: If you don't know how to start a problem, convert whatever they give you into **moles** first.

1. Use molar mass of CH_4 to convert from mass to moles.

$$5.00\ \mathrm{g\ CH_4} \times \frac{1\ \mathrm{mol\ CH_4}}{16.04\ \mathrm{g\ CH_4}} = 0.311_7\ \mathrm{mol\ CH_4}$$

If 5.00 g of CH_4 (methane) is burned, what mass of water can be produced?

Write out the core of the equation from the description:

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"For every 1 mol CH_4, we need to react with 2 mol O_2 to produce 1 mol CO_2 and 2 mol H_2O."

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2. *Use 2:1 H_2O:CH_4 mole-mole ratio to find moles of H_2O.* $0.311_7\ \mathrm{mol\ CH_4} \times \frac{2\ \mathrm{mol\ H_2O}}{1\ \mathrm{mol\ CH_4}} = 0.623_4\ \mathrm{mol\ H_2O}$

If 5.00 g of CH_4 (methane) is burned, what mass of water can be produced?

Write out the core of the equation from the description:

$$\underline{\ 1\ }\ CH_4\ (g)\ +\ \underline{\ 2\ }\ O_2\ (g) \rightarrow \underline{\ 1\ }\ CO_2\ (g)\ +\ \underline{\ 2\ }\ H_2O\ (g)$$

"For every 1 mol CH_4, we need to react with 2 mol O_2 to produce 1 mol CO_2 and 2 mol H_2O."

REMEMBER MY TIP: If you don't know how to start a problem, convert whatever they give you into **moles** first.

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3. *Use molar mass of H_2O to convert from moles to mass.* $0.623_4\ \mathrm{mol\ H_2O} \times \frac{18.02\ \mathrm{g\ H_2O}}{1\ \mathrm{mol\ H_2O}} = 11.2\ \mathrm{g\ H_2O}$

Pouring an aqueous solution of hydrochloric acid onto a solid block of magnesium metal produces an aqueous solution of magnesium chloride and hydrogen gas.

$$\_\_\_ + \_\_\_ \rightarrow \_\_\_ + \_\_\_$$

A) Given 3.00 g Mg, how many moles of hydrochloric acid do we need?

B) If we produce 5.00 g H_2 gas, what mass of $MgCl_2$ solution is produced?

C) If we produce 4.00 g H_2 gas, what mass of HCl did we need?

REMEMBER : We only care about the **ratio** of coefficients, so we can use mole-mole ratios to go between reactants-to-reactants, reactants-products, products-to-reactants, or products-to-products.

Pouring an aqueous solution of hydrochloric acid onto a solid block of magnesium metal produces an aqueous solution of magnesium chloride and hydrogen gas.

$$\underline{\ 1\ }\ \mathrm{Mg}\ (s)\ +\ \underline{\ 2\ }\ \mathrm{HCl}\ (aq)\ \rightarrow\ \underline{\ 1\ }\ \mathrm{MgCl_2}\ (aq)\ +\ \underline{\ 1\ }\ \mathrm{H_2}\ (g)$$

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A) Given 3.00 g Mg, how many moles of hydrochloric acid do we need?

$$3.00\ \text{mol Mg} \times \frac{2\ \text{mol HCl}}{1\ \text{mol Mg}} = 6.00\ \text{mol HCl}$$

B) If we produce 5.00 g H_2 gas, what mass of $MgCl_2$ solution is produced?

C) If we produce 4.00 g H_2 gas, what mass of HCl did we need?

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Pouring an aqueous solution of hydrochloric acid onto a solid block of magnesium metal produces an aqueous solution of magnesium chloride and hydrogen gas.

__1__ Mg (s) + __2__ HCl (*aq*) → __1__ $MgCl_2$ (*aq*) + __1__ H_2 (*g*)

A) Given 3.00 g Mg, how many moles of hydrochloric acid do we need?

$$3.00\ \text{mol Mg} \times \frac{2\ \text{mol HCl}}{1\ \text{mol Mg}} = 6.00\ \text{mol HCl}$$

B) If we produce 5.00 g H_2 gas, what mass of $MgCl_2$ solution is produced?

$$5.00\ \text{g}\ \text{H}_2 \times \frac{1\ \text{mol}\ \text{H}_2}{2.016\ \text{g}\ \text{H}_2} \times \frac{1\ \text{mol}\ \text{MgCl}_2}{1\ \text{mol}\ \text{H}_2} \times \frac{95.21\ \text{g}\ \text{MgCl}_2}{1\ \text{mol}\ \text{MgCl}_2} = 236\ \text{g}\ \text{MgCl}_2$$

C) If we produce 4.00 g H_2 gas, what mass of HCl did we need?

REMEMBER : We only care about the **ratio** of coefficients, so we can use mole-mole ratios to go between reactants-to-reactants, reactants-to-products, products-to-reactants, or products-to-products.

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A) Given 3.00 g Mg, how many moles of hydrochloric acid do we need?

$$3.00\ \text{mol Mg} \times \frac{2\ \text{mol HCl}}{1\ \text{mol Mg}} = 6.00\ \text{mol HCl}$$

B) If we produce 5.00 g H_2 gas, what mass of $MgCl_2$ solution is produced?

$$5.00\ \text{g H}_2 \times \frac{1\ \text{mol H}_2}{2.016\ \text{g H}_2} \times \frac{1\ \text{mol MgCl}_2}{1\ \text{mol H}_2} \times \frac{95.21\ \text{g MgCl}_2}{1\ \text{mol MgCl}_2} = 236\ \text{g MgCl}_2$$

C) If we produce 4.00 g H_2 gas, what mass of HCl did we need?

$$4.00\ \text{g H}_2 \times \frac{1\ \text{mol H}_2}{2.016\ \text{g H}_2} \times \frac{2\ \text{mol HCl}}{1\ \text{mol H}_2} \times \frac{36.46\ \text{g HCl}}{1\ \text{mol HCl}} = 145\ \text{g HCl}$$

REMEMBER : We only care about the **ratio** of coefficients, so we can use mole-mole ratios to go between reactants-to-reactants, reactants-to-products, products-to-reactants, or products-to-products.

SUMMARIZING STOICHIOMETRY RELATIONSHIPS

THE MOLE IS *STILL* CENTRAL

MASS
VOLUME
MOLES

BALANCED CHEMICAL EQUATION

$mole_1 \longrightarrow mole_2$

MASS
VOLUME
MOLES

GIVEN
reactant or product

LOOKING FOR
reactant or product

*I hope now you understand why I say to convert to moles before you do anything else. It's because a balanced chemical equation gives us **mole-to-mole ratios** that we can use to convert between one reactant/product to another reactant/product.*

$$2H_2\,(g) \quad + \quad 1O_2\,(g) \quad \rightarrow \quad 2H_2O\,(g)$$

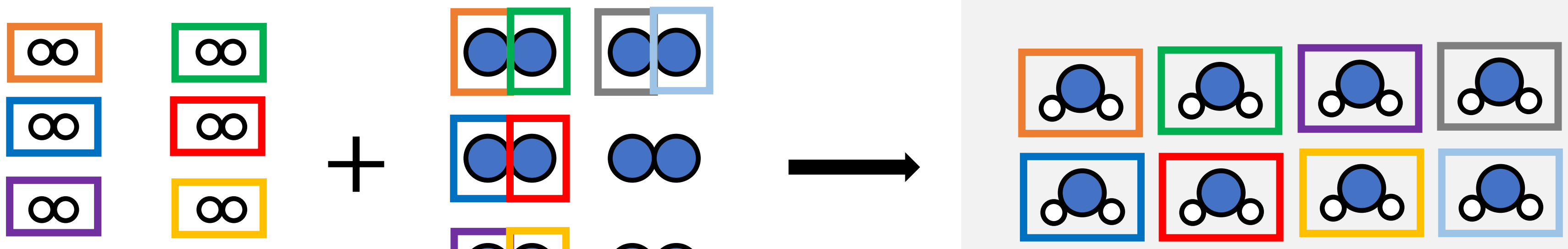
A) If you have 8 moles of hydrogen and all the oxygen you need, how many moles of water can you make?

$$8 \text{ mol } H_2 \times \frac{2 \text{ mol } H_2O}{2 \text{ mol } H_2} = 8 \text{ mol } H_2O$$

B) If you have 6 moles of oxygen and all the hydrogen you need, how many moles of water can you make?

$$6 \text{ mol } O_2 \times \frac{2 \text{ mol } H_2O}{1 \text{ mol } O_2} = 12 \text{ mol } H_2O$$

C) If you have 8 moles of hydrogen <u>and</u> 6 moles of oxygen, how many moles of water can you make?



Q: Which limits our reaction?

A: H_2 is limiting; 2 mol O_2 leftover

8 mol H_2O

When do I know if I need to figure out the limiting reactant?

When they give you the amounts of ALL reactants.

CHOOSE THE METHOD THAT WORKS BEST FOR YOU!

$$\underline{\ 2\ }\ H_2\,(g) + \underline{\ 1\ }\ O_2\,(g) \rightarrow \underline{\ 2\ }\ H_2O\,(g)$$

• METHOD 1

1. Assume one reactant is limiting and then determine amount of product you can form.
$$8\text{ mol }H_2 \times \frac{2\text{ mol }H_2O}{2\text{ mol }H_2} = 8\text{ mol }H_2O$$
2. Assume *other* reactant is limiting and then determine amount of product you can form.
$$6\text{ mol }O_2 \times \frac{2\text{ mol }H_2O}{1\text{ mol }O_2} = 12\text{ mol }H_2O$$
3. Reactant that *limits* amount of products formed is **limiting reactant**.

H_2 produces less H_2O so it is limiting.

• METHOD 2

1. Start with one reactant and determine how much of the other reactant you *need*.
$$8\text{ mol }H_2 \times \frac{1\text{ mol }O_2}{2\text{ mol }H_2} = 4\text{ mol }O_2$$
2. Compare what you *have* vs. what you *need*:
   Have: 6 mol O_2
   Need: 4 mol O_2
3. We *have* more O_2 than we *need*
   → O_2 excess
   → **H_2 is limiting.**