Spr 17 B3

Chemistry Unit 2

Primary reference: *CHEMISTRY*, Addison-Wesley

Topic	Essential Knowledge	Study and Practice
Scientific Investigation 1.2 **SOL 1b, 1g**	Understand and use **Material Safety Data Sheet** (MSDS) warnings including: handling chemicals, lethal dose (LD), disposal and chemical spill clean-up. The **percent by mass of an element in a compound** can be determined: % by mass of element = $\frac{\text{total mass of element in compound}}{\text{molar mass of the compound}}$ X 100	**Ch 7: Read pp. 188-191** on percent composition
Atomic Structure and Periodic Relationships **2.2** **SOL 2a,2c,2d, 2e,2i**	**Democritus**:Greek philosopher who suggested the idea of atoms @ 400 BC. **John Dalton** atomic theory of 4 postulates was based on experimentation—early 1800s **J.J. Thomson** and **Millikan** discovered the electron and it's charge respectively. **Ernest Rutherford's** gold foil experiment showed the atom was mostly empty space with a small, dense, positively charge nucleus. Atoms are made of **protons, neutrons** in the nucleus. A cloud of electrons surrounds the nucleus. An atom's **atomic number** = the number of **protons**. All atoms of the same element have the same number of **protons**. A **proton** has a **positive** charge and a relative mass of one. The number of **electrons** equals the number of protons in a neutral atom. An **electron** has a **negative** charge and a relative mass of zero. A **neutron** has **no charge** and a relative mass of one. Isotopes are atoms of the same element with a different number of neutrons (Example C-12 and C-13). $\boxed{\textbf{mass number = \#protons + \# neutrons}}$ The **atomic masses** on the periodic tables are a weighted average of the isotope masses. **Dmitri Mendeleev** created a Periodic Table based on the elements' masses and physical and chemical properties. **Moseley** reordered the table slightly based on atomic number. Rows are called **periods** and columns are called **groups** or **families**. Named families are **alkali metals, alkaline earth metals, halogens**, and **noble gases.**	**Chapter 5:** pp. 107-112 on early atomic models. pp. 113-117 about atomic numbers. pp. 123-126 about the perioidic table.
Nomenclature, Formulas, and Reactions **3.2** **SOL 3a, 3b, 3c**	**Subscripts** in a **chemical formula** represent the relative number of each type of atom. The subscript follows the element symbol. Example: a water molecule, H_2O, has 2 hydrogen atoms and one oxygen atom. **Parentheses** are used when a **subscript** affects a group of atoms. Example: $Mg(NO_3)_2$ has a ratio of one magnesium atom, 2 nitrogen atoms and 6 oxygen atoms in the compound. **Molecules** form from non-metals and **ionic compounds** form from a metal cation and a non-metal anion. Metals lose electrons to become **cations**. Non-metals gain electrons to form **anions**. For ionic compounds, the charges of the anions and cations must add to zero. In binary ionic compounds, we name the metal first followed by the anion ending with –ide. Roman numerals are used to show the charge/oxidation state of metals other than alkali or alkaline earth metals. In binary molecular compounds, we use prefixes in front of the element names and end with –ide. A **chemical equation** shows the formulas of all the **reactants** on the left hand side of the arrow, and the formulas for all the **products** on the right hand side. Chemical reactions follow the **Law of Conservation of Mass**—matter is neither created nor destroyed during a chemical reaction. We **balance** chemical equations using **coefficients in front of each substance** in the equation so that each side has the same number of atoms of each element.	**Chapter 6:** pp 133-137 on molecular and ionic compounds. pp 138-140 on chemical formulas. pp 149-151 and 158-159 **Chapter 8:** pp. 203-211 on chemical equations
Molar Relationships 4.2 **SOL 4a, 4b, 4d**	**Molar mass** is the sum of all the atomic masses in a compound. The **mole** can be used to convert between mass, particles and gas volume using unit cancelation. $\boxed{\text{1 mole} = 6.02 \times 10^{23} \text{ things = molar mass = 22.4 L(gas at 0°C \& 1atm only)}}$ Ionic compounds **dissociate** in water to form **electrolyte** solutions (conduct electricity) whereas molecular compounds do not.	**Ch 7** pp 176-190 on molar conversions and % composition. **Chapter 17** pp. 482-485 on electrolytes
Phases of Matter and Kinetic Molecular Theory 5.2 **SOL 5a, 5d**	**Kinetic Molecular Theory** describes the behavior of gases based on a model of an ideal gas. Ideal gases do not exist but help us understand how real gases behave. **Real gases** exist, have intermolecular forces, particle volume and can change states, whereas ideal gases do not. **Avogadro's hypothesis**: Equal volumes of gas at the same pressure and temperature will contain the same number of gas particles. $\boxed{\text{1 mole gas = 22.4 Liters at 0°C and 1 atm.}}$	**Chapter 10 and Chapter 12:** pp. 267-272 and pp. 327-328 and p. 347 on gases.

Unit 2 Objectives

Chemistry, Addison-Wesley, 2002

I. Basic Atomic Structure
   A. Early Atomic Models through Rutherford
   B. Atomic number, mass number, atomic mass and isotopes
II. Introduction to the Periodic Table
   A. Parts of the periodic table
III. Chemical Names and Formulas
   A. Differentiating between molecular and ionic compounds
   B. Ionic charges of Elements
   C. Names ↔ Formulas Binary Ionic Compounds
   D. Names ↔ Formulas Binary Molecular Compounds
   E. Diatomic Elements (Review)
IV. Mole Calculations
   A. Molar Mass
      1. Review of counting atoms in formulas (p198#460)
      2. Calculating molar mass(p179#7;p181#9,10;p198#50,51)
      3. Converting between moles and molar mass(p183#16-19;p186#24,27;p198#55,56)
   B. Molar Volume of Gases at STP
      1. Avogadro's hypothesis
      2. Converting between moles and molar volume at STP (1 mole gas = 22.4 L)(p184#20,21;p198#57)
   C. More Molar Conversions
      1. Conversions: mass ↔ volume, mass ↔ count, volume ↔ count (p186#25;p198#59)
V. Chemical Reactions
   A. Understanding chemical reaction symbols
   B. Balancing Chemical Reactions

<u>Objectives (SOL)</u> *<u>book problems</u>*

1. Identify the contributions of Democritus, Dalton, Thomson, Rutherford, and Millikan, to the development of the modern atomic model. (2i)
2. Describe the structure of an atom, including the location of protons, electrons and neutrons.(2c) *p129#37, 38, 39, 40, 41.*
3. Define the charges and relative masses of electrons, protons and neutrons.
4. Determine the number of protons, neutrons and electrons in elements and isotopes. (2a)*p115#7,8;p116#9, p121#23;p129#42*
5. Explain how isotopes differ, yet are still the same element.(2a)*p121#21*
6. Calculate the atomic mass for an element given the weighted averages of the isotopes.(2b)*p129#53*
7. Identify the contributions of Mendeleev and Mosely to the modern periodic table.(2i)
8. Identify the following areas on the periodic table: alkali metals, alkaline earth metals, halogens, noble or inert gases, representative elements, transition metals, non-metals, metals, and metalloids.(2d)
9. Distinguish between ionic and molecular compounds.(2g, 3a) *p167:#49*
10. Count the number of atoms present in compound formulas(3c)*p166#71*
11. Explain how anions and cations are formed.(2g) *p136#1,2;p145#17; p166#46,53*
12. Predict monatomic ion charges using the periodic table (2g) *p145#16;p148#20, 22*
13. Use the roman numeral Stock System to identify and name transition metal ions.(3a)*p148#23*
14. Predict the ionic compound formed from any two monatomic ions.(3c)*p151#24,25*
15. Write the formulas for binary ionic and molecular compounds given their names and *visa versa*.(3a&3c) *Ionic(p155:#29, p166#58,61, p167#68) molecular (p159#38, p160#4`, p167#64)* Ionic(*p153#26,27; p156#30,31;p167#67,69) Molecular(p159#37; p160#41; p167#64)*
16. Name the seven diatomic elements.(3a)
17. Explain Avogadro's Hypothesis.(4a)
18. Memorize molar volume = 22.4 Liter at 1 atmosphere and 0°C(4a)
19. Calculate the molar mass of a substance given the formula.(4a)
20. Calculate conversions between moles, molar masses, molar volumes, and particle counts.(4a)
21. Master reading and writing chemical equations using chemical formulas and symbols correctly. (3b)
22. Explain the Law of Conservation of Mass
23. Balance equations (3b)
24. Explain a catalyst's role in a chemical reaction. (3f)

n^0 e^-

Chapter 5 Atomic Structure Skeleton Notes

What is an atomic number of an element and where do we find it?	• The number of p^+ in the element • Always a whole-number integer • DEFINES (identifies) the element
What is a mass number?	• # of n^0 and p^+ (in nucleus) • Always a whole #
What is an isotope?	• An element w/ same # of p^+, BUT DIFF. # of NEUTRONS!!
How do you read isotope symbols? $^{6}_{3}Li$ or $^{6}_{3}Li$ or Li-6 $3p^+$ vs $3n^0$	top # = mass # (atomic mass) ★ n and p bottom # = atomic # ★ p
$^{7}_{3}Li$ or $^{7}_{3}Li$ or Li-7 $3p^+$ $4n^0$	← THIS IS AN ISOTOPE OF LITHIUM
How many protons, neutrons and electrons are in $^{35}_{17}Cl$? (neutral)	protons 17, neutrons 18, electrons 17
How many protons, neutrons and electrons are in $^{37}_{17}Cl$?	p^+ 17, n^0 20, e^- 17
How many protons, neutrons and electrons are in Calcium-42?	p^+ 20, n^0 22, e^- 20
What is the atomic mass of an element and where do we find it?	(aka: mass number) The weighted avg appears on PT

"neutral"
p=e

Calculating Atomic mass

Elements contain a mix of isotopes. If we are given the percent composition of each isotope, we can calculate the atomic mass using weighted averages.

Analogy: Weighted Grades

Type	% Weight	x Score	= Contribution
Tests	50%	75%	
Quizzes	25%	92%	
Homework	25%	95%	

We use mass number x % abundance (composition) to calculate the approximate atomic mass

Example: Find the atomic mass of chlorine using the data below. (Ans = 35.4846 amu)

Isotope	% Abundance	x mass number	= Contribution
Cl-35	75.77 %	35	
Cl-37	24.23	37	

$$(0.7577 \cdot 35) + (0.2423 \cdot 37) =$$

We can calculate a more accurate value by using % abundance and isotope mass in atomic mass units, amu, to calculate the value. (Ans = 35.4528)

Isotope	% Abundance	x amu	= Contribution
Cl-35	75.77	34.969	
Cl-37	24.23	36.966	

more exact

Practice: Naturally occurring oxygen contains 99.757 % Oxygen-16, 0.038% Oxygen-17 and 0.205% Oxygen-18. Calculate the approximate atomic mass.(Ans = 16.00448 amu)

Practice: Use the atomic mass unit data in the following table to calculate oxygen's atomic mass more accurately. (Ans = 15.999 amu)

Isotope	% abundance	amu
O-16	99.757	15.995
O-17	0.038	16.999
O-18	0.205	17.999

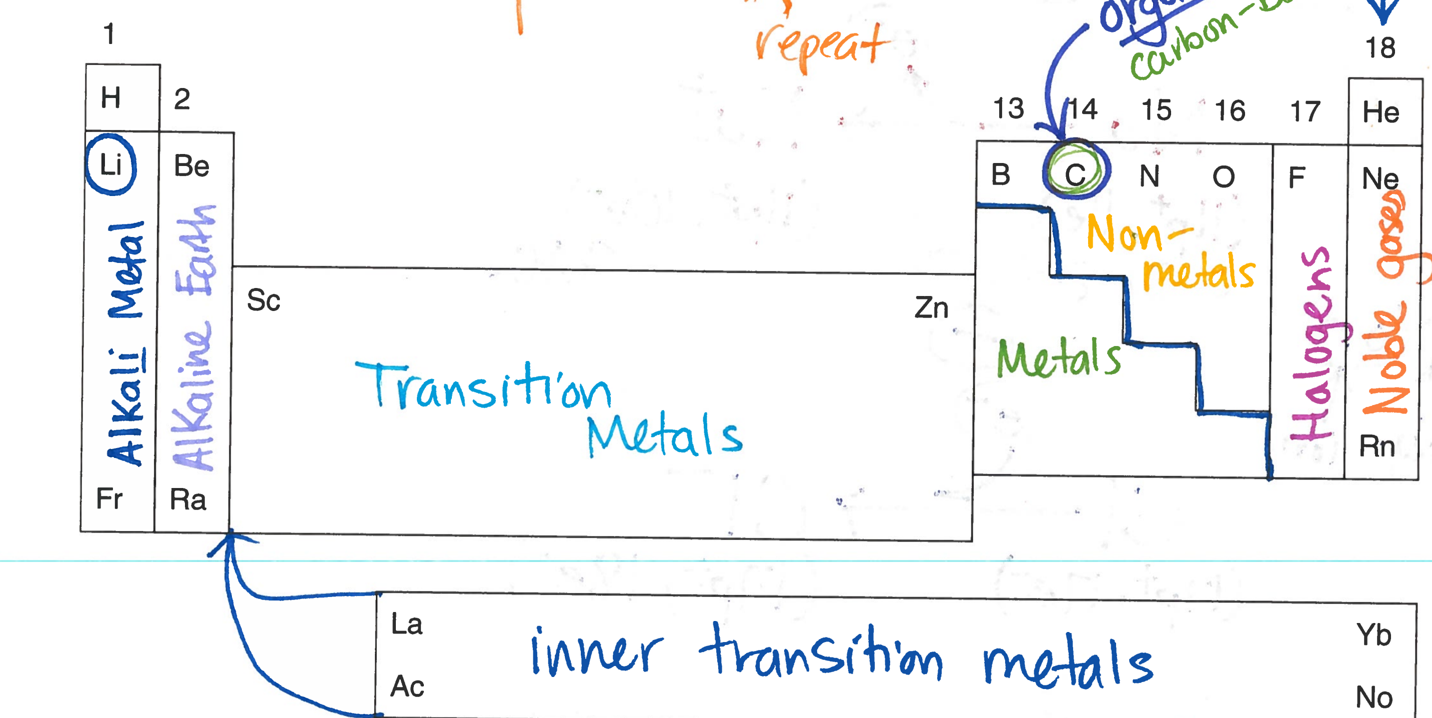
Animation for mass spec and isotopes at:
http://wps.prenhall.com/wps/media/objects/4974/5093961/emedia/ch02/MassSpectrometer/c2s4item20/MassSpectrometer.html

Chapter 6: Chemical Names and Formulas

Areas of the Periodic Table (Chapter 5)

Columns are called groups or families — Why? React similarly

Rows are called periods — repeat



Nonmetals are located in the above & to the right of

Metals are located in the below & to the left of

A. Types of Compounds

1) Compounds are combinations of ≥ 2 different elements

2) Elements may combine to form heterogeneous or homo. or compounds

ionic or molecular

binary (2), ternary

COMPOUND

IONIC
- A metal + a non-metal.

$NaCl$ KBr

$CaCl_2$ $Mg(OH)_2$

C_3H_8

C_4H_{10}

MOLECULAR / COVALENT
- 2 or more NON-metals!!!

H_2O, CO_2, CO,

NH_3, SiH_4,

C_8H_{18}, C_xH...

X^+ $-(-1) = +1$ (M^+)

3) IONIC COMPOUNDS form when a metal combines with a non-metal (Nm^-)

CA⊕ION "paws-itive"

a) cations are metal atoms that have LOST electrons, so they acquire a ⊕ charge

$$Na \xrightarrow{e^-} Na^+$$

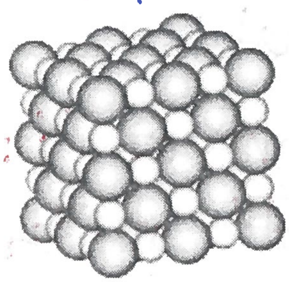
$(11p^+, 11e^-)$ $(11p^+, 10e^-)$

anion
a neg. ion

b) anions are nonmetal atoms that have GAIN electrons so they acquire a ⊖ charge

$$Cl \xrightarrow{e^-} Cl^-$$

$(17p^+, 17e^-)$ $(17p^+, 18e^-)$



Formula Unit: Lowest whole-number ration of cation:anion

$Cu_6N_4 \rightarrow Cu_3N_2$ ⋮ $Mg_8Cl_{16} \rightarrow MgCl_2$

M+Nm → Properties of Ionic Compounds

Melting points are VERY, VERY! high MP

Physical State at room temperature Solid

Ionic Compounds dissociate into ions in water when they dissolve.

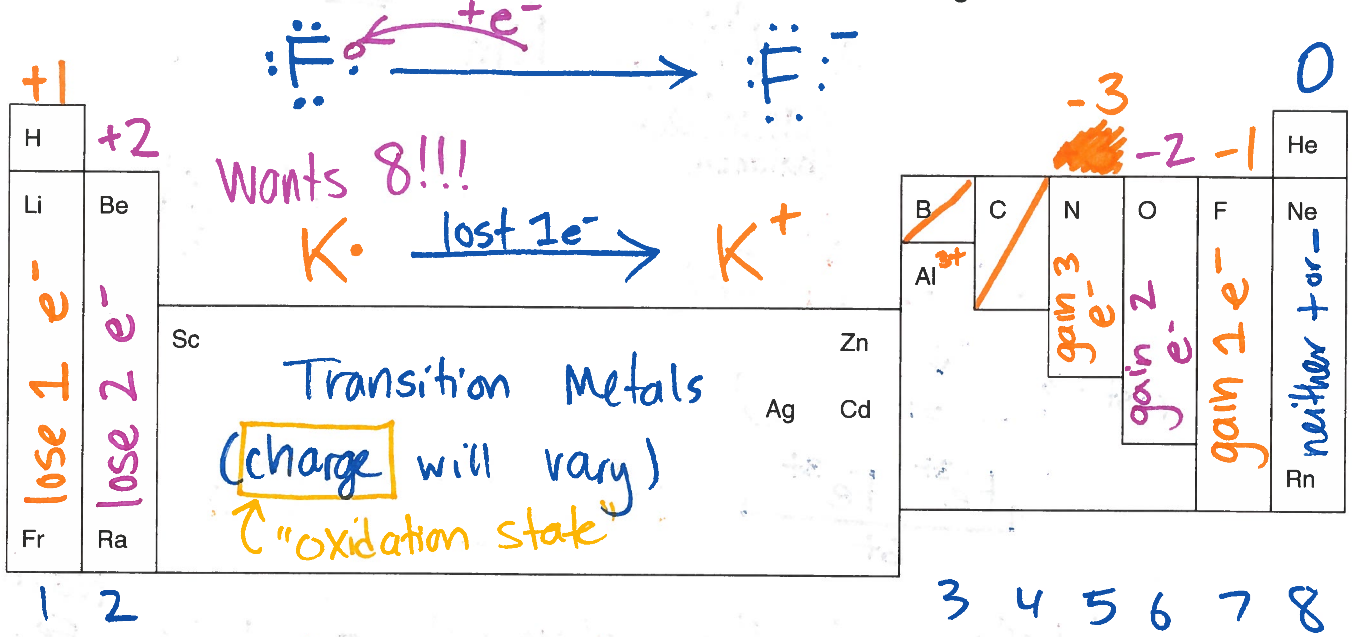
$NaCl(s) + H_2O(l) \longrightarrow NaCl(aq)$ aqueous ("dissolved in water")

$NaCl(s) + H_2O(l) \longrightarrow Na^+(aq) + Cl^-(aq)$

sodium cation chloride anion

B. Common Oxidation States (Charges) of Elements in Ionic Compounds

1. For alkali metals, alkaline earth metals, non-metals and noble gases.



2. For all other elements

- Transition metals may have more than one common oxidation state
- The oxidation state is indicated by a roman numerals in ( ) I, II, III, IV, V, VI, VII, VIII
- Example: iron(II) = Fe^{2+} iron(III) = Fe^{3+}
- Exceptions: silver Ag^{+}, Zinc Zn^{2+}. Cadmium Cd^{2+}, Aluminum Al^{3+}

Review: formulas tell us two things: $C_9H_{13}N$

1) which elements are present
2) ... and how many of each are there.

Examples:

$NaCl$

#	Sym
1	Na
1	Cl

Al_2O_3

#	Sym
2	Al
3	O

Distribute!!!

$Al_2(SO_4)_3$

#	Sym
2	Al
3	S
12	O

MM = ____ g/mol

hydrocarbon

$CH_3(CH_2)_{12}CH_3$

#	Sym
14	C
30	H

metal & a non-metal

C. Binary Ionic Compounds

1) Binary compound: a compound w/ 2 diff elem.

2) Ionic compounds are a M^+ & Nm^-. The positive charges from the cation must be cancelled (balanced) by the negative charges from the anion.

3) The cation (metal) is always written first, and the anion ends in -ide.

Examples: sodium chloride is NaCl

rust → Iron(III)oxide is Fe_2O_3

$\underbrace{Fe^{3+} \quad Fe^{3+}}_{6+} \qquad \underbrace{O^{2-} \quad O^{2-} \quad O^{2-}}_{-6}$

4) Steps for writing Binary Ionic Formulas from names:

Na, Au, Fe, Cl

a) Write out the element symbols WITH THE CHARGE!!!!!!

b) Figure out how many cations & anions you need to balance.

Examples:

Calcium bromide

Ca^{2+} Br^{-}

$CaBr_2$

copper(II)chloride

1: Cu^{2+} 2 Cl^{-}

$CuCl_2$

Magnesium nitride

Mg^{2+} N^{3-}

Mg_3N_2

lead(IV)sulfide

Pb^{4+} S^{2-} → ~~Pb_2S_4~~

PbS_2

Zinc phosphide

Zn^{2+} P^{3-}

Zn_3P_2

Lithium oxide

Li^{+} O^{2-}

Li_2O

Formula → Name

$MgCl_2$ magnesium chloride.

5) Steps for naming binary ionic compounds from formulas with transition metals

a) Write out the metal (cation)'s name. DONE
b) If it's "shady" (trans. metal), open ( )
c) Write anion (non-metal) -ide
d) Fill in ( ) with oxidation state

K_3N
potassium nitride
K^+ K^+ K^+ $\xrightarrow{e^-}$ N^{3-}

SnF_2
tin(II) fluoride
$Sn^{\boxed{2+}}$ F^- F^-
$\underbrace{\quad}_{+2}$ $\underbrace{\quad}_{-2}$

Fe_2O_3
→ iron(III) oxide
$Fe^{\boxed{3+}}$ O^{2-}
$Fe^{\boxed{3+}}$ O^{2-}
O^{2-}
$\underbrace{\quad}_{+6}$ $\underbrace{\quad}_{6-}$

TiO_2
titanium(IV) oxide
$Ti^{\boxed{4+}}$ O^{2-}
O^{2-}
$\underbrace{\quad}_{+4}$ $\underbrace{\quad}_{-4}$

CaO

D. Binary Molecular Compounds

all NON-metals

↳ 2 NM

1) Composed of two non-metals
2) Because they are molecules, no parentheses or ions or charge are involved.
3) Prefixes are used to show the number of each type of atom present:

1 =	mono-	6 =	hexa-
2 =	di-	7 =	hepta-
3 =	tri-	8 =	octa-
4 =	tetra-	9 =	nona-
5 =	penta-	10 =	deca-

4) Exceptions: Do not use mono for the first element!!!

CO_2
carbon dioxide

Drop the prefixes ending ~~–a~~ or ~~–o~~ with oxygen compounds.

CO SO_3

N_2O
dinitrogen monoxide

Examples:

Phosphorus trichloride PCl_3

Nitrogen dioxide NO_2

Dinitrogen pentoxide N_2O_5

Carbon monoxide CO

CF_4 carbon tetrafluoride

P_2O_5 ______

SiO_2 (glass or sand)

$AsCl_3$ ______

Mixed Ionic and Molecular Naming

Molecular	Ionic
2 NM	1 M^+, 1 Nm^-
- use prefixes (di, tri, ...) - 2nd element: - ide No "mono"	- NO PREFIXES - Tran. Metals: use (X) - NM^- ends in -ide or -ate, -ite

Circle the molecular compounds, then name all compounds correctly.

CaF_2
calcium fluoride

CaO

NF_3
nitrogen trifluoride

P_2O_5

SiO_2

Fe_2N_3

E. DIATOMIC ELEMENTS: Professor Br_2 I_2 N_2 Cl_2 H_2 O_2 F_2

Pure

7 diatomic elements.

The Mole & Avogadro's Number

1 mole of ANYTHING (cars, people, atoms, molecules, books, protons…) is equal to ________________________. You definitely need to memorize this.

1 mole is equal to ↑, just like *1 dozen is equal to 12 things.*

Moles are usually referring to atoms and compounds (**1 mole of Na atoms** = **6.022 × 10^{23} Na atoms!**) because atoms and compounds are very, very, very small.

________________________ is Avogadro's number. It's equal to 1 mole. (you only need to know it to "6.02"…). GET OUT YOUR PERIODIC TABLE.

Video Guide: "How Big is a Mole?" (TED-Ed)

1) Who was the first guy to propose numerical "counting" of particles like atoms and molecules?

2) If you have 6.02 × 10^{23} (that's about 602,000,000,000,000,000,000,000) molecules of water (H_2O)…
   a. …how much will it weigh in grams? ____________g
   b. …and since the density of H_2O is 1 g/mL, 18.01 g of H_2O should also have a volume of ____________mL.

KNOW THIS:

1 mole of He atoms = 6.02 × 10^{23} He atoms.
1 mole of P atoms = 6.02 × 10^{23} P atoms
1 mole of Cu atoms = 6.02 × 10^{23} Cu atoms.
1 mole of Na atoms = 6.02 × 10^{23} atoms of Na.
1 mole of H_2 molecules = 6.02 × 10^{23} molecules of H_2.
1 mole of CO_2 molecules = 6.02 × 10^{23} molecules of CO_2. EASY!

Check out the next pattern. Have your PT ready:

1 mole of He atoms = 6.02 × 10^{23} He atoms = **4.003 g of He**
1 mole of P atoms = 6.02 × 10^{23} P atoms = **30.97 g of P**
1 mole of Cu atoms = 6.02 × 10^{23} Cu atoms = ______________ **g of Cu**
1 mole of Na atoms = 6.02 × 10^{23} atoms of Na = ______________ **g of Na**
1 mole of H_2 molecules = 6.02 × 10^{23} molecules of H_2 = **2.02 g of H_2**

The first tricky thing: How many H atoms are in each molecule of H_2? _______

So… you can easily find the **molar mass** of H on the PT. Why did we have to multiply that PT number (for a *single* H) by **2** in order to get the **molar mass** of H_2?

1 mole of CO_2 molecules = 6.02 × 10^{23} molecules of CO_2 = ________________ g of CO_2.
(1 carbon is ____________ **g/mol**. 2 oxygens = 2×______________ **g/mol**)

KNOW THIS before you practice: When you look on the PT and see that the **molar mass** for magnesium (**Mg**) is ____________, you've GOT to have units!

The unit for a regular pack of eggs is 12 **eggs/dozen**.
The unit for an average pack of paper is 500 **sheets/ream**.
The unit for a molar mass is **grams/mole**. (seen usually as **g/mol** for short.)

Worked Example for Calculating the Molar Mass of H_2O:

Worked Example for Calculating the Molar Mass of **nitrogen trifluoride**:

Worked Example for Calculating the Molar Mass of **ammonium chloride**:

Practice. Find the molar mass of each compound, using your PT. Must use appropriate units!

1) How much molar mass does 1 mole of zirconium have?
2) Determine the molar mass of 1 mole of sodium chloride.
3) Find the molar mass of 1 mole of CF_4.
4) What is the molar mass of 1 mole of barium hydroxide?

*****Always start with what you know*****

5) Worked Example: I have 1.95×10^{24} atoms of sulfur. (Sig figs matter!)
   a. How many moles of sulfur do I have?
   b. How many grams does my sample weigh?

6) Worked Example: Mark has 88.0 grams of solid KF.
   a. How many moles of KF does Mark have?
   b. How many atoms of KF does Mark have?

Conversion Mapping:

You can go from ______________ to ______________ (and vice-versa) by using the molar __________, found on the period table. ***You'll have to do some adding and multiplying for compounds.***

Review: What are the units for molar mass?

Review: How would you calculate the molar mass for K_2O, knowing there are ___ K atoms and ___ O atom?

You can go from ______________ to __________________ (and vice-versa) by using __________________'s Number.

Review: What is Avogadro's #?

Regardless of what you do in chemistry, **you MUST go <u>through</u> the mole**:

Mixed Practice:

7) How many molecules of water are in 60.0 grams of it?

$$\frac{60.0\,\cancel{g}\ H_2O}{1} \times \frac{1\ \cancel{mol}}{18.0\,\cancel{g}} \times \frac{6.02E23\ \text{molecules}}{1\ \cancel{mol}} = \_\_\_\_\_\_$$

8) How many grams does 4.77×10^{19} **formula units** of NaBr weight?

9) Which element has a molar mass of 196.97 g/mol? ______________________

10) Which element has a molar mass of about 83.8 g/mol? ______________________

11) Which **diatomic** element has a molar mass of 37.997 g/mol? ______________________

12) 6.02×10^{23} atoms of vanadium will have a mass of ______________ grams.

13) 7.5593×10^{38} atoms of vanadium will have a mass of ______________ grams.

14) 7.5593×10^{38} atoms of vanadium is equal to ____________ moles of vanadium.

15) 3.7 moles of H_2O will have a mass of ______________ grams.

16) 32,600 milligrams of carbon will contain __________________ C atoms.

17) 10.0 moles of phosphorus pentafluoride will have a mass of ________________ grams, and will contain ____________________ molecules of the compound.

18) Students calculated the molar mass of strontium iodide to be 307.13 g/mol. Calculate their % error.

$KE = \frac{1}{2}mv^2$

GASES and The Mole

Gases take up space. A lot of space!

Recall that the amount of matter in something give us its ***mass***, and the amount of ***space*** that something takes up is called volume.

Volume has a few different units.

We know that 1 cm^3 = 1 mL = 1 cc.

… and 1000 mL = 1 liter

The common unit for VOLUME in chemistry is the liter.

VERY IMPORTANT:

1 mole of a gas is gonna take up 22.4 liters.

In fact, 1 mole of any gas will take up 22.4 **L of space.** @ STP 0°C 1 atm

KNOW THIS: 1 mole (g) = 22.4 L

Review: According to Professor BrINClHOF ("Brinklehoff"), what are the 7 diatomic elements? Write them out in X_2 format, and circle the ones that are gases at standard (normal) temperature and pressure.

I–I

Br_2 liquid, I_2 solid, N_2 Gas, Cl_2 gas, H_2 gas, O_2 gas, F_2 gas

Practice with moles and gases: Use the right units! ALWAYS GO THROUGH THE MOLE!

19) How much volume does 8.40 moles of hydrogen (H_2) gas have?

3

$$\frac{8.40\ \cancel{mol\ H_2}}{1} \times \frac{22.4\ L\ H_2}{1\ \cancel{mol\ H_2}} = 118\ L\ H_2$$

20) How much space (in L) does 113 moles of O_2 take up?

$$\frac{113\ mol\ O_2}{1} \times \frac{22.4\ L\ O_2}{1\ mol\ O_2} = 2530\ L\ O_2$$

21) 5.4 x 10^{16} atoms of helium gas will take up _______________ liters of space.

$$\frac{5.4E16\ \cancel{atoms\ He}}{1} \times \frac{1\ \cancel{mol}}{6.02E23\ \cancel{atoms}} \times \frac{22.4\ L}{1\ \cancel{mol}}$$

$= 2.0E-6\ L$

22) How much mass does 59.4 L of chlorine gas have?

Cl_2

$$\frac{59.4\ \cancel{L\ Cl_2}}{1} \times \frac{1\ \cancel{mol\ Cl_2}}{22.4\ \cancel{L\ Cl_2}} \times \frac{70.9\ g\ Cl_2}{1\ \cancel{mol\ Cl_2}}$$

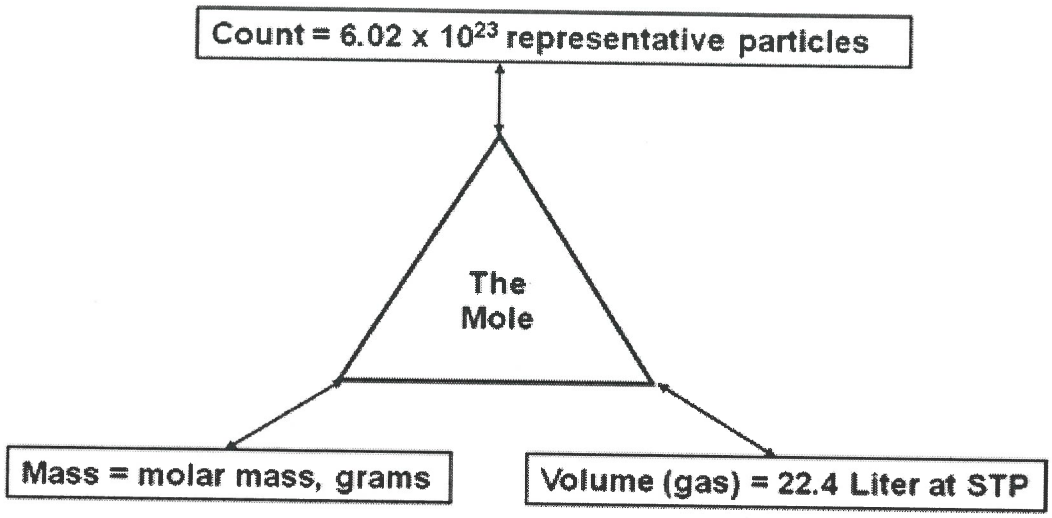
23) 44.8 L of nitrogen gas is _____________ moles of nitrogen.

24) 44.8 L of nitrogen gas is _____________ grams of nitrogen.

25) 44.8 L of nitrogen gas is _____________ molecules of nitrogen.

26) Determine the mass of carbon dioxide gas if the volume of the gas is 96.7 L.

Chapter Seven: The Mole is the Chemist and Physicist's Dozen



The mole = the number of atoms in exaxtly twelve grams of Carbon-12.

1958: 1 mole = 6.02×10^{23}
2006: 1 mole = $6.02214078 \times 10^{23}$

Avogadro's number = ______________________

Some mole facts—see if you can find the pattern

- 6.02×10^{23} atoms in twelve grams of C-12.
- 6.02×10^{23} atoms in 12.011 g of carbon (naturally occurring
- 6.02×10^{23} atoms in 1.008 g of hydrogen
- 6.02×10^{23} atoms in ________ g of oxygen

Converting between moles and counts of representative particles

1 mole = 6.02×10^{23} representative particles or

$$\frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ rep. part}} = \frac{6.02 \times 10^{23} \text{ rep. part.}}{1 \text{ mol}}$$

Representative particles:

Elements ____________________________________

Molecules ____________________________________

Ionic compounds ____________________________________

3 couples →

$\boxed{\text{1 mole} = 6.02 \times 10^{23} \text{ representative particles}}$

1. How many formula units of MgS are there in 0.482 mol of MgS?

2. How many moles are in 1.204×10^{25} molecules of nitrogen dioxide?

3. How many sodium atoms are there in 3.2 moles of sodium?(Ans = 1.9×10^{24} Na atoms)

4. How many moles are there in 6.32×10^{24} formula units of Iron(III) sulfide?(Ans = 10.5 mol)

Review of Counting Atoms in Formulas

One mole of NaCl = 1 mol of Na, 1 mol of Cl

One mole of Na_2S = 2 mol of Na, 1 mol of S

One mole of $Al_2(SO_4)_3$ = 2 mol Al, 3 mol S, 12 mol Oxy

You do:

CAMEL

CO_3^{2-}

- Li_3PO_4 = ______

$Pb(CO_3)_2$ = ______

$Fe_2(Cr_2O_7)_3$ = ______

Calculating Molar Mass (formula units, atoms, molecules)

Definitions:

- Molar Mass = the mass of $6.02x10^{23}$ representative particles of an element, molecule, or ionic compound. Molar Mass may also be called Formula Mass.

6.02×10^{23} representative particles=1 Mole = molar mass, gram

Examples:

1. Find the molar mass of Sodium.

2. Find the molar mass of NaCl.

3. Find the molar mass of $CaCl_2$

4. Find the molar mass of $Cu(NO_3)_2$

5. Find the molar mass of $Mg(OH)_2$

6. Find the molar mass of $Ca_3(PO_4)_2$

Converting between Moles and Molar Mass

1 mole = 6.02×10^{23} rep. particles = molar mass, g

$$\frac{1\ \text{mol}}{\text{molar mass, g}} = \frac{\text{molar mass, g}}{1\ \text{mol}}$$

1. What is the mass of 2.3 moles of $MgBr_2$?(Ans = 420 g)

2. How many moles of potassium iodide are in 29.3 g of KI? (Ans = 0.177 mol)

3. How many grams of SO_3 are present in 2.3 moles of SO_3?(Ans = 180 g)

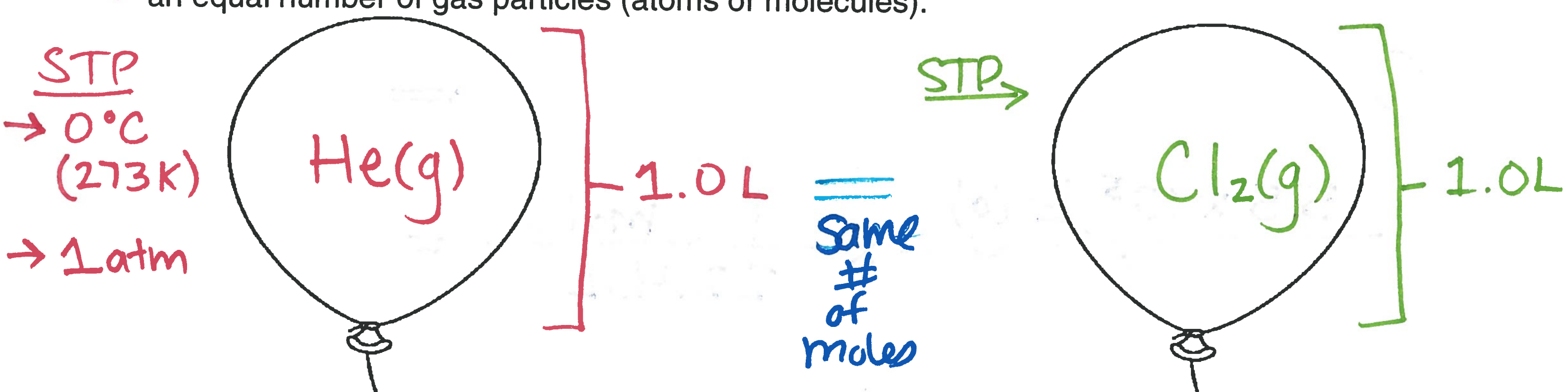
4. How many moles of CaF_2 are equivalent to 450 grams of CaF_2?(Ans = 5.8 mol)

5. How many grams of titanium(IV) sulfide, TiS_2, are present in 0.056 moles of titanium(IV) sulfide? (Ans = 6.3 g)

6. How many moles of ammonium sulfate $(NH_4)_2SO_4$ are present in 52.3 grams of ammonium sulfate?(Ans = 0.396 mol)

Molar Volumes

Avogadro's Hypothesis: Equal volumes of gas at the same temperature and pressure contain an equal number of gas particles (atoms or molecules).



Molar Volume: The volume of one mole of gas at standard temperature and pressure (STP)

STP = 0°C (273K) & 1.0 atm.

$\boxed{\text{1 mole gas = 22.4 L at STP for any gas}}$

Assumptions about ideal gases

- The gas particles have no volume (points in space)
- The gas particles have no intermolecular attractions
- The gas particles collide elastically like billiard balls.
- Ideal gases never condense no matter how cold it is.

Real Gases Condense!!!!

$\star$ 1 mol of ANY gas = 22.4 L

Converting between Moles and Molar Gas Volume

1 mole = 6.02×10^{23} rep. particles = molar mass, g = molar gas volume (22.4L at STP)

$$\frac{1 \text{ mol}}{22.4 \text{ L}} = \frac{22.4 \text{ L}}{1 \text{ mol}}$$

1. A neon light contains (neon is a noble gas) 0.51 liters of neon gas at STP. How many moles does the light contain?(Ans = 2.3×10^{-2} mol)

2. A helium balloon contains 0.325 moles of gas at STP. What is the balloon's volume in liters? (Ans = 7.28 L He)

$$\frac{0.325 \cancel{\text{mol}}}{1} \times \frac{22.4 \text{ L (g)}}{1 \cancel{\text{mol}} \text{ (g)}} = 7.28 \text{ L}$$

3. An underground cavern contains 5.5×10^5 liters of natural gas, CH_4, at STP. How many moles of gas are in the cavern? (Ans = 2.5×10^4 mol)

$$\frac{5.5 \times 10^5 \cancel{\text{L}} \text{ (g)}}{1} \times \frac{1 \text{ mol}}{22.4 \cancel{\text{L}}} =$$

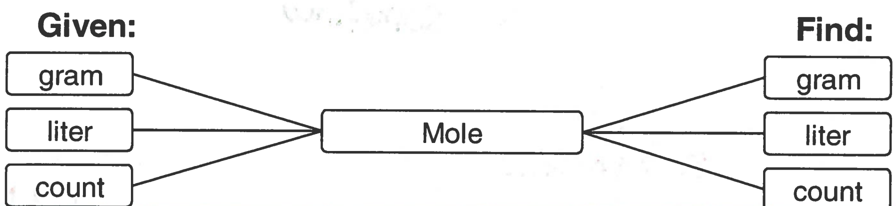
4. A blimp contains 35,000 liters of hydrogen gas (flammable!) at STP. How many moles of hydrogen does it contain? (Ans = 1.6×10^3 mol)

More Molar Conversions—You can go anywhere!

1 mole = 6.02 x 10^{23} rep. part = gram formula mass = 22.4 Liters at STP

Chapter Seven Objectives
- Memorize Avogadro's number
- Memorize STP volume of 1 mole of gas (0°C, 1 atm)
- Convert between grams, moles, representative particles and liters using factor label method.
- Calculate molar masses

unit to mole	mole to unit
$\frac{1 \text{ mole}}{6.02 \times 10^{23} \text{ rep. part.}}$	$\frac{6.02 \times 10^{23} \text{ rep. part.}}{1 \text{ mole}}$
$\frac{1 \text{ mole}}{\text{g. molar mass}}$	$\frac{\text{g. molar mass}}{1 \text{ mole}}$
$\frac{1 \text{ mole}}{22.4 \text{ L. gas}}$	$\frac{22.4 \text{ L. gas}}{1 \text{ mole}}$



1. How many atoms of gold are contained in a 505 gram bar of Au?(Ans = 1.54 x 10^{24} Au atoms)

2. Find the number of moles of Cl_2 gas in a 1.46 x 10^4 liter tank at STP.(Ans = 62 mol)

3. A balloon contains 1.2 grams of Helium. What is the balloon's volume at STP?~~(Ans = 7.3 L)~~

$$\frac{1.2\,\cancel{g\ He(g)}}{1} \times \frac{1\ \cancel{mol\ He}}{4.0\,\cancel{g\ He}} \times \frac{22.4\ L}{1\ \cancel{mol}}$$

4. How many molecules of fluorine gas are in an $\boxed{0.0030 \text{ liter}}$ ampule at STP?(Ans = 8.2 x 10^{19} molec.)

$$\frac{3.0 \times 10^{-3}\ L}{1} \times \frac{1\ mol}{22.4\ L} \times \frac{6.02E23\ molecules}{1\ mol}$$

Chapter Eight Skeleton Notes Part 1

3/7 I. Chemical Equations $3NaN_3(s) \rightarrow 4N_2(g) + 1Na_3N(s)$

Symbols	Use
$\rightarrow$	forward rxn (goes to completion) — reaction
$\leftrightarrows$ $\rightleftharpoons$	reversible rxn (equilibrium)
(s), (l), (g)	states of matter (solid, gas, liquid)
(aq)	dissolved in H_2O (aqueous)
$\xrightarrow{catalyst}$	speeds up a rxn!!!
$\xrightarrow{heat}$	add (or remove) sometimes from a rxn

lactose

~~lactose~~

B: Examples of chemical equations

synthesis rxn
$A + B \rightarrow AB$

$Mg(s) + O_2(g) \rightarrow MgO$

Reactants = magnesium, oxygen (gas)

Products = magnesium oxide

Decomposition Rxn
$CD \rightarrow C + D$

$OF_2(g) \rightarrow F_2(g) + O_2(g)$

Reactants = oxygen difluoride

Products = fluorine, oxygen

Writing chemical equations from word equations

1. Sodium metal reacts with chlorine gas to form sodium chloride.

$$2Na(s) + 1Cl_2(g) \rightarrow 2NaCl(s)$$

2. Iron metal reacts with oxygen gas to form rust, iron(III)oxide.

Fe^{3+} O^{2-}

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

3. Solid nitrogen triiodide decomposes to solid iodine and nitrogen gas.

$$2NI_3(s) \rightarrow 3I_2(s) + 1N_2(g)$$

Catalysts: Speeds up a reaction WITHOUT being consumed.

Skeleton equations do not show the amounts of products and reactants.

$\boxed{E=mc^2}$

similar to energy

C. Balancing chemical reactions using coefficients

Law of conservation of mass: "Matter cannot be created, nor destroyed... only transformed"

Balanced equations use coefficients in front of formulas to show the the balanced # of atoms / molecules

$2x^3 + 4y^5$

A chemical reaction is balanced if there are the same number of each kind of element on both sides of the chemical equation. If there are four oxygens on the reactant side, there will be _____ oxygens on the product side.

Count the atoms of each element of both sides.→ Indicate which equations are balanced.

$C(s) + O_2(g) \rightarrow CO_2(g)$

1 carbon	1 carbon
2 Oxygen	2 oxygens

✓ balanced!

$Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$

1 Fe	2 Fe
2 Oxy	3 oxy

X Not balanced!

$3NH_3 + H_3PO_4 \rightarrow (NH_4)_3PO_4$ coefficients apply to the entire compound

3 N	3 N
12 H	12 H
1 P	1 P
4 O	4 O

✓ Balanced!

$C_4H_{10} + 4O_2 \rightarrow 4CO_2 + 5H_2O$

4 C	4 C
10 H	10 H
8 O	13 O

X Not Bal. :(

$Ca(ClO_3)_2 \rightarrow CaCl_2(s) + 3O_2(g)$

OYO

balancing chemical equations

a. Write the skeleton chemical equation leaving blanks for the coefficients:
b. Count the number of each element in the reactant and product side
c. Balance the equation using whole number coefficients (NEVER SUBSCRIPTS)
d. Track your changes
   1. Balance the other compounds to the most complicated compound.
   2. Balance the binary compounds (H_2O, CO_2, NO_2)
   3. Balance diatomics and elements last
   4. If you end up with an odd number that won't balance (3 oxygens on one side, two on the other) double all the coefficients filled in so far.
   5. Double check when you're done.

$\underline{1}\ N_2O_4(g) \xrightarrow{Pt} \underline{1}\ N_2(g) + \underline{2}\ O_2(g)$

$\underline{3}\ K_2S + \underline{2}\ FeCl_3 \rightarrow \underline{1}\ Fe_2S_3 + \underline{6}\ KCl$

6 K	6 K
3 S	3 S
2 Fe	2 Fe
6 Cl	6 Cl

For $\boxed{\text{combustion}}$ reactions, use the CHO rule (C first, H second, O last)

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

$\boxed{} + O_2 \rightarrow CO_2 + H_2O$

$\underline{2\ \cancel{1}}\ C_2H_6 + \underline{7\ \cancel{\tfrac{7}{2}}}\ O_2 \rightarrow \underline{4\ \cancel{2}}\ CO_2 + \underline{6\ \cancel{3}}\ H_2O$

$\_\_\_ C_2H_6O + \_\_\_ O_2 \rightarrow \_\_\_ CO_2 + \_\_\_ H_2O$

Chlorine gas was used in chemical warfare during WWI. The Germans used Chlorine gas on the Allied Forces in Ypres, France in 1915. Chlorine reacts with the moisture in lungs to produce hydrochloric acid, HCl.

You try balancing the reaction for chlorine in your lungs:

$\_\_\_Cl_2(g) + \_\_\_H_2O(l) \rightarrow \_\_\_HCl(l) + \_\_\_O_2(g)$

Now try balancing the reaction for phosgene (Cl_2CO) in your lungs. This is another poisonous gas used in warfare.

$\_\_\_Cl_2CO(g) + \_\_\_H_2O(l) \rightarrow \_\_\_HCl(l) + \_\_\_CO(g) + \_\_\_O_2(g)$

Chem F'17

3

Chemistry Unit 2

Primary reference: ***Chemistry: Matter and Change*** **[Glencoe, 2017]**

Topic	Essential Knowledge	Study and Practice
Scientific Investigation 1.2 **SOL 1b, 1g**	Understand and use **Material Safety Data Sheet** (MSDS) warnings including: handling chemicals, lethal dose (LD), disposal and chemical spill clean-up. The **percent by mass of an element in a compound** can be determined: $\% \text{ by mass of element} = \frac{\text{total mass of element in compound}}{\text{molar mass of the compound}} \times 100$	**Ch 10:** Read pp. 341-342 on percent composition
Atomic Structure and Periodic Relationships **2.2** **SOL 2a,2c,2d, 2e,2i**	**Democritus:**Greek philosopher who suggested the idea of atoms @ 400 BC. **John Dalton** atomic theory of 4 postulates was based on experimentation—early 1800s **J.J. Thomson** and **Millikan** discovered the electron and it's charge respectively. **Ernest Rutherford's** gold foil experiment showed the atom was mostly empty space with a small, dense, positively charge nucleus. Atoms are made of **protons, neutrons** in the nucleus. A cloud of electrons surrounds the nucleus. An atom's **atomic number** = the number of **protons**. All atoms of the same element have the same number of **protons**. A **proton** has a **positive** charge and a relative mass of one. The number of **electrons** equals the number of protons in a neutral atom. An **electron** has a **negative** charge and a relative mass of zero. A **neutron** has **no charge** and a relative mass of one. Isotopes are atoms of the same element with a different number of neutrons (Example C-12 and C-13). $\boxed{\textbf{mass number = \#protons + \# neutrons}}$ The **atomic masses** on the periodic tables are a weighted average of the isotope masses. **Dmitri Mendeleev** created a Periodic Table based on the elements' masses and physical and chemical properties. **Moseley** reordered the table slightly based on atomic number. Rows are called **periods** and columns are called **groups** or **families**. Named families are **alkali metals, alkaline earth metals, halogens**, and **noble gases**.	**Ch 4:** Read pp. 102-105 on early atomic models and atomic theory. Read pp. 115-119 about atomic numbers and mass numbers **Ch 6:** Read pp. 174-181 about the periodic table.
Nomenclature, Formulas, and Reactions **3.2** **SOL 3a, 3b, 3c**	**Subscripts** in a **chemical formula** represent the relative number of each type of atom. The subscript follows the element symbol. Example: a water molecule, H_2O, has 2 hydrogen atoms and one oxygen atom. **Parentheses** are used when a **subscript** affects a group of atoms. Example: $Mg(NO_3)_2$ has a ratio of one magnesium atom, 2 nitrogen atoms and 6 oxygen atoms in the compound. **Molecules** form from non-metals and **ionic compounds** form from a metal cation and a non-metal anion. Metals lose electrons to become **cations**. Non-metals gain electrons to form **anions**. For ionic compounds, the charges of the anions and cations must add to zero. In binary ionic compounds, we name the metal first followed by the anion ending with –ide. Roman numerals are used to show the charge/oxidation state of metals other than alkali or alkaline earth metals. In binary molecular compounds, we use prefixes in front of the element names and end with –ide. A **chemical equation** shows the formulas of all the **reactants** on the left hand side of the arrow, and the formulas for all the **products** on the right hand side. Chemical reactions follow the **Law of Conservation of Mass**—matter is neither created nor destroyed during a chemical reaction. We **balance** chemical equations using **coefficients in front of each substance** in the equation so that each side has the same number of atoms of each element.	**Ch 3:** Read p. 85 **Ch 7:** Read pp. 206-209 on ions. *Electron configurations will be learned later.* Read pp. 210-216 on ionic compounds. Read pp. 221-224 on polyatomic ions and formulas pp 149-151 and 158-159 **Ch 9:** Read pp. 282-288 on chemical equations
Molar Relationships 4.2 **SOL 4a, 4b, 4d**	**Molar mass** is the sum of all the atomic masses in a compound. The **mole** can be used to convert between mass, particles and gas volume using unit cancelation. $\boxed{1 \text{ mole} = 6.02 \times 10^{23} \text{ things} = \text{molar mass} = 22.4 \text{ L(gas at 0°C \& 1atm only)}}$ Ionic compounds **dissociate** in water to form **electrolyte** solutions (conduct electricity) whereas molecular compounds do not. An example of an ionic compound as it dissociates in water (into ions) is seen here: $MgBr_2(s) + H_2O(l) \rightarrow Mg^{2+}(aq) + 2Br^-(aq)$	**Ch 10:** Read pp. 325-340 on molar conversions; Read pp. 341-343 on percent composition (by mass). **Ch 7:** Read pp. 214-216 on electrolytes
Phases of Matter and Kinetic Molecular Theory **5.2** **SOL 5a, 5d**	**Kinetic Molecular Theory** describes the behavior of gases based on a model of an ideal gas. Ideal gases do not exist but help us understand how real gases behave. **Real gases** exist, have intermolecular forces, particle volume and can change states, whereas ideal gases do not. **Avogadro's hypothesis**: Equal volumes of gas at the same pressure and temperature will contain the same number of gas particles. $\boxed{1 \text{ mole gas} = 22.4 \text{ Liters at 0°C and 1 atm.}}$	**Ch 12 and Ch 13:** Read pp. 400-406 for an introduction to gases; Read p. 452. *Gases will be revisited in greater detail;* Read pp. 457-459 on real and ideal gases.

Unit 2 Objectives

Chemistry: Matter and Change (Glencoe, 2017)

I. Basic Atomic Structure
   A. Early Atomic Models through Rutherford
   B. Atomic number, mass number, atomic mass and isotopes
II. Introduction to the Periodic Table
   A. Parts of the periodic table
III. Chemical Names and Formulas
   A. Differentiating between molecular and ionic compounds
   B. Ionic charges of Elements
   C. Names ↔ Formulas Binary Ionic Compounds
   D. Names ↔ Formulas Binary Molecular Compounds
   E. Diatomic Elements (Review)
IV. Mole Calculations
   A. Molar Mass
      1. Review of counting atoms in formulas
      2. Calculating molar mass
      3. Converting between moles and molar mass
   B. Molar Volume of Gases at STP
      1. Avogadro's hypothesis
      2. Converting between moles and molar volume at STP (1 mole gas = 22.4 L)
   C. More Molar Conversions
      1. Conversions: mass↔ volume, mass↔ count, volume↔ count
V. Chemical Reactions
   A. Understanding chemical reaction symbols
   B. Balancing Chemical Reactions

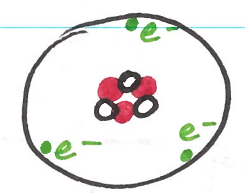
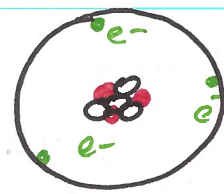
Objectives (SOL)

1. Identify the contributions of Democritus, Dalton, Thomson, Rutherford, and Millikan, to the development of the modern atomic model. (2i)
2. Describe the structure of an atom, including the location of protons, electrons and neutrons.(2c).
3. Define the charges and relative masses of electrons, protons and neutrons.
4. Determine the number of protons, neutrons and electrons in elements and isotopes. (2a)
5. Explain how isotopes differ, yet are still the same element.(2a)
6. Calculate the atomic mass for an element given the weighted averages of the isotopes.(2b)
7. Identify the contributions of Mendeleev and Mosely to the modern periodic table.(2i)
8. Identify the following areas on the periodic table: alkali metals, alkaline earth metals, halogens, noble or inert gases, representative elements, transition metals, non-metals, metals, and metalloids.(2d)
9. Distinguish between ionic and molecular compounds.(2g, 3a)
10. Count the number of atoms present in compound formulas(3c)
11. Explain how anions and cations are formed.(2g)
12. Predict monatomic ion charges using the periodic table (2g)
13. Use the roman numeral Stock System to identify and name transition metal ions.(3a)
14. Predict the ionic compound formed from any two monatomic ions.(3c)
15. Write the formulas for binary ionic and molecular compounds given their names and *visa versa*.(3a&3c)
16. Name the seven diatomic elements.(3a)
17. Explain Avogadro's Hypothesis.(4a)
18. Memorize molar volume = 22.4 Liter at 1 atmosphere and 0°C(4a)
19. Calculate the molar mass of a substance given the formula.(4a)
20. Calculate conversions between moles, molar masses, molar volumes, and particle counts.(4a)
21. Master reading and writing chemical equations using chemical formulas and symbols correctly. (3b)
22. Explain the Law of Conservation of Mass
23. Balance equations (3b)
24. Explain a catalyst's role in a chemical reaction. (3f)

$E=mc^2$

atomic number → 6 C 12.01

Chapter 5 Atomic Structure Skeleton Notes

What is an atomic number of an element and where do we find it?	atomic # : # of protons in nucleus (★ : It identifies the element)
What is a mass number?	mass # : # of p^+ and n^0 (whole # integer) > carbon-12 $n^0=6$ > Carbon-14 $n^0=8$
What is an isotope?	★ Atoms of same element w/ same # of p^+, BUT diff. # of n^0
How do you read isotope symbols? $^{6}_{3}$Li or ^{6}Li or Li-6 (Δ = # of n^0) vs $^{7}_{3}$Li or ^{7}Li or Li-7	top # = mass # (sum p^+ & n^0) (Isotopes ID) bottom # = atomic # (# of p^+)  Li-6 ● = p^+ o = n^0 • = e^-  Li-7
How many protons, neutrons and electrons are in $^{35}_{17}$Cl ?	protons 17 (atomic #), neutrons 18, electrons 17 (neutral) • ^{35}Cl • Chlorine-35 • Cl-35 $^{35}_{17}Cl^-$
How many protons, neutrons and electrons are in $^{37}_{17}$Cl ?	p^+ 17, n^0 20, e^- 17 • ^{37}Cl • Chlorine-37 • Cl-37
How many protons, neutrons and electrons are in Calcium-42?	p^+ 20, n^0 22, e^- 20 $^{42}_{20}$Ca
What is the atomic mass of an element and where do we find it?	Avg mass of all isotopes of an element. (Cl-35, Cl-37) 75% 25%

4% → 84%

Calculating Atomic mass

Elements contain a mix of isotopes. If we are given the percent composition of each isotope, we can calculate the atomic mass using weighted averages.

Analogy: Weighted Grades Becky

Type	% Weight	x Score	= Contribution
Tests	50% (0.50)	75 pts	= 37.5 pts
Quizzes	25% (0.25)	92 pts	= 23 pts
Homework	25% (0.25)	95 pts	= 23.75 pts

SUM = 84.25

We use mass number x % abundance (composition) to calculate the approximate atomic mass
Example: Find the atomic mass of chlorine using the data below. (Ans = 35.4846 amu)

Isotope	% Abundance	x mass number	= Contribution
Cl-35	~~75.77~~ 0.7577	35	+
Cl-37	~~24.23~~ 0.2423	37	

OYO

= $\boxed{}$

We can calculate a more accurate value by using % abundance and isotope mass in atomic mass units, amu, to calculate the value. (Ans = 35.4528)

Isotope	% Abundance	x amu	= Contribution
Cl-35	~~75.77~~ 0.7577	x 34.969	= +
Cl-37	~~24.23~~ 0.2423	x 36.966	=

SUM = 35.4528 amu

★ Use the "better" # if available

OYO Practice: Naturally occurring oxygen contains 99.757 % Oxygen-16 0.038% Oxygen-17 and 0.205% Oxygen-18. Calculate the approximate atomic mass.(Ans = 16.00448 amu) = 16.00

Practice: Use the atomic mass unit data in the following table to calculate oxygen's atomic mass more accurately. (Ans = 15.999 amu) → 16.0

Isotope	% abundance	amu
O-16	99.757	15.995
O-17	0.038	16.999
O-18	0.205	17.999

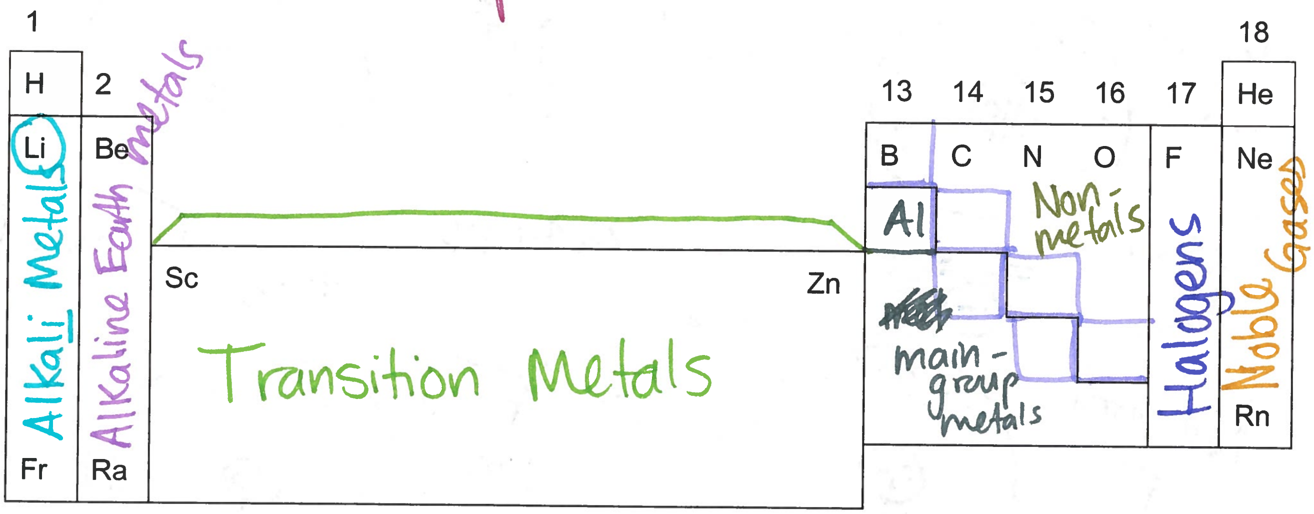
Animation for mass spec and isotopes at:
http://wps.prenhall.com/wps/media/objects/4974/5093961/emedia/ch02/MassSpectrometer/c2s4item20/MassSpectrometer.html

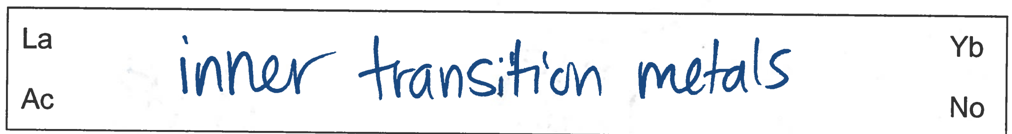
Chapter 6: Chemical Names and Formulas

Areas of the Periodic Table (Chapter 5)

Columns are called groups or families

Rows are called periods





Nonmetals are located in the ABOVE & to the RIGHT of

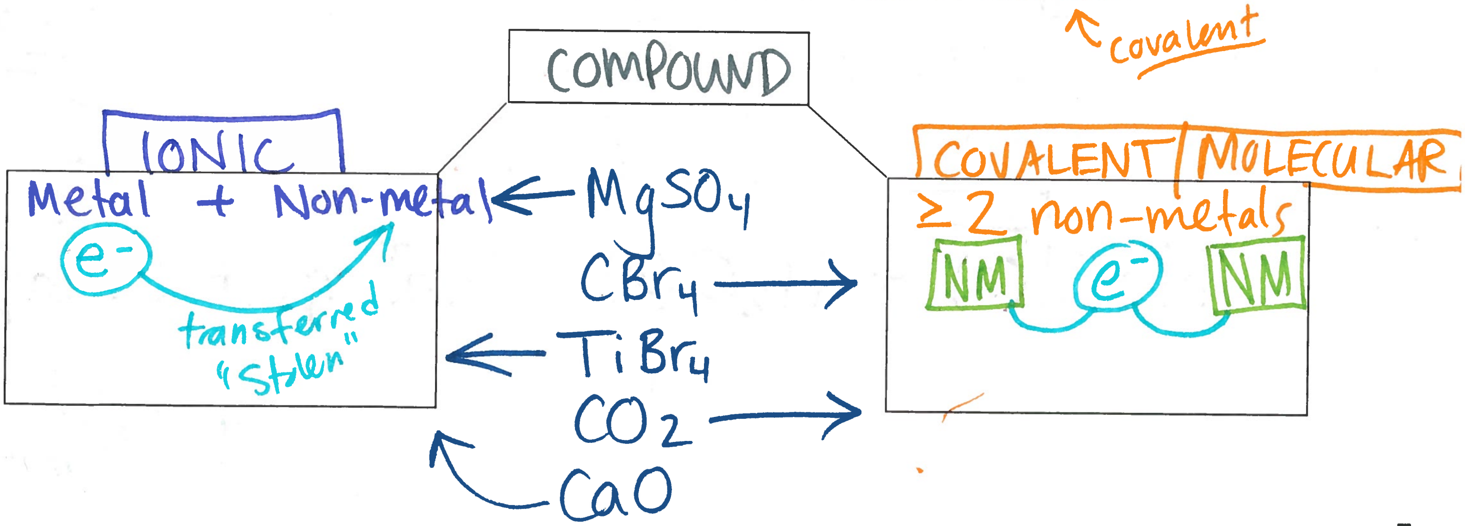
Metals are located in the BELOW & to the LEFT of

A. Types of Compounds

1) Compounds are chemically-combined elements

2) Elements may combine to form IONIC or MOLECULAR compounds

covalent



CaF_2

$-(-1) = +1$

3) IONIC COMPOUNDS form when a metal $\boxed{\text{cation}}$ ⊕ combines with a non-metal $\boxed{\text{anion}}$ ⊖ → "a neg. ion"

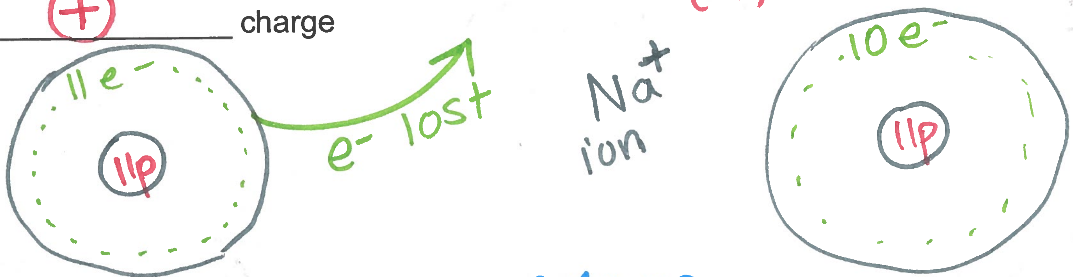
"paws"itive

ca+ion

a) cations are metal atoms that have LOST electrons, so they acquire a ⊕ charge

$-(-1)$

Neutral Na



Na^+ ion

-10

+

0

Δ

ISO-

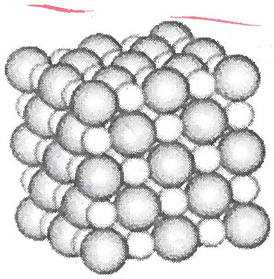
same

b) anions are nonmetal atoms that have GAINED electrons so they acquire a ⊖ charge

$+(-1)$

Neutral F

F^- ion



⊕ ⊖ NaCl

⊕ ⊖ $MgCl_2$

⊕ ⊖ K_2O

LiF

• NaCl

• $C_{12}H_{22}O_{11}$

Formula Unit: lowest whole-number ratio of CATIONS to ANIONS

$\boxed{Ca_5Cl_{10}}$ → ~~$CaCl_2$~~ → ~~Ca + 2Cl~~

Properties of $\boxed{\text{Ionic}}$ Compounds ← w/ a metal & non-metal

→ Melting points are HIGH AF !!!

Physical State at room temperature always solid

9/27 → Ionic Compounds $\boxed{\text{dissociate}}$ into ions in water when they dissolve.

$NaCl(s) + H_2O(l) \rightarrow Na^+(aq) + Cl^-(aq)$

aqueous

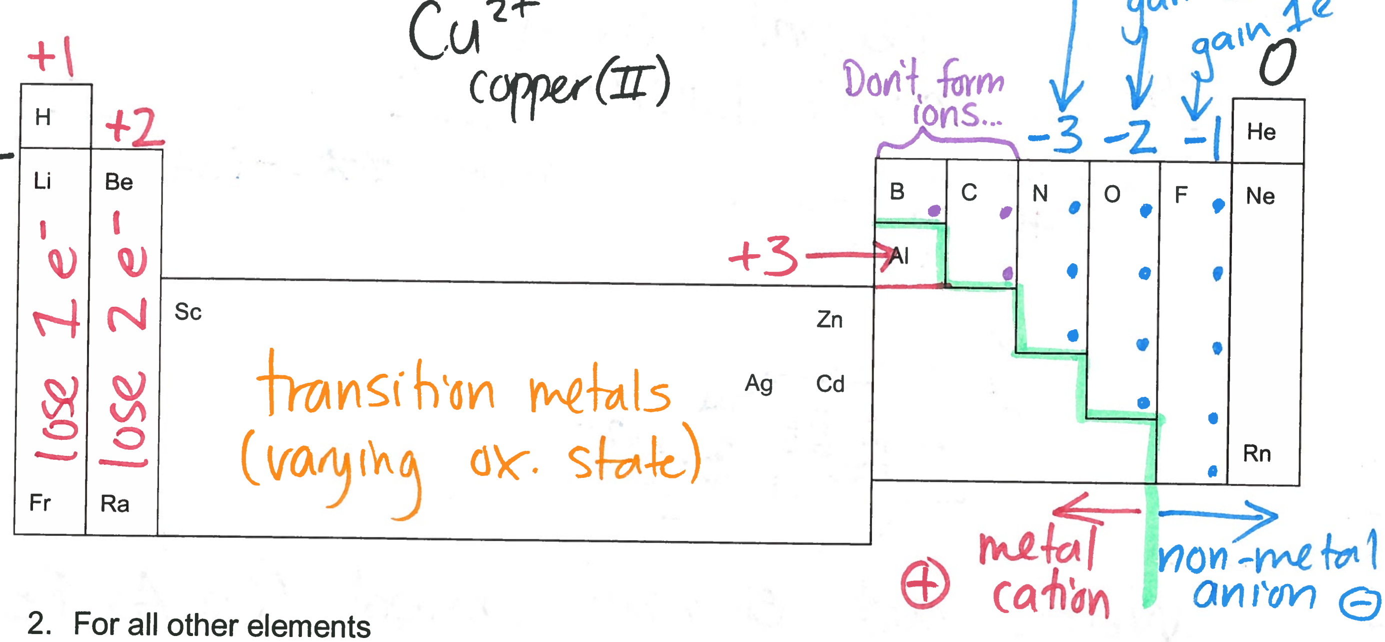
"dissolved in H_2O"

$-(+1) = +1$

B. Common Oxidation States (Charges) of Elements in Ionic Compounds

1. For alkali metals, alkaline earth metals, non-metals and noble gases.

Cu^{2+} copper(II)



2. For all other elements

- Transition metals may have more than one common Oxidation state (charge)
- The oxidation state is indicated by a Roman numeral (I, II, III, IV, V, VI, VII)
- Example: iron(II) = Fe^{2+} iron(III) = Fe^{3+}
- Exceptions: silver Ag^{+}, Zinc Zn^{2+}, Cadmium Cd^{2+}, Aluminum Al^{3+}

Review: formulas tell us two things: $Mg_3(PO_4)_2$

1) Which elements are present
2) Subscript: How many atoms are there

Examples:

NaCl: 1 Na, 1 Cl

Al_2O_3: 2 Al, 3 Oxy

$Al_2(SO_4)_3$: 2 Al, 3 S, 12 Oxy

$CH_3(CH_2)_{12}CH_3$: 14 C, 30 H

$(NH_4)_3PO_4$: 3 N, 12 H, 1 P, 4 Oxy

2 ← M + Nm

C. Binary Ionic Compounds

1) Binary compound: a compound w/ 2 diff elem.

2) Ionic compounds are neutral (overall). The positive charges from the cation must be cancelled / balanced by the negative charges from the anion.

3) The cation (metal) is always written first, and the anion ends in -ide.

Examples: sodium chloride is NaCl

Iron(III)oxide is Fe_2O_3

$\underbrace{2\,Fe^{3+}}_{+6}$ $\underbrace{3 \cdot O^{2-}}_{-6}$

2 3 5 8 13 21 34

4) Steps for writing Binary Ionic Formulas from names:

a) Write out element symbol (Ca, Ag, Fe...) with charge!

b) Figure out how many cations (M^+) & anions (Nm^-) are needed to cancel.

Examples:

metal

Calcium bromide

Ca^{2+} Br^-

$CaBr_2$

copper(II)chloride

Cu^{2+} Cl^-

$CuCl_2$

Magnesium nitride

Mg^{2+} N^{3-}

Mg_3N_2

lead(IV)sulfide

Pb^{4+} S^{2-} → Pb_2S_4

PbS_2

11

Zinc phosphide

Zn^{2+} P^{3-}

Zn_3P_2

Lithium oxide

Li^+ O^{2-} → Li_2O

$K_2S \rightarrow$ potassium sulfide

potassium sulfide $\rightarrow K_2S$

5) Steps for naming binary ionic compounds from formulas with transition metals

a) Write out cation's ⊕ name: ____

b) If it's a tran. metal, "open up par." (__)

c) Write out anion's ⊖ name: ____ -ide

d) Fill in (__) w/ the metal's ox. state ~~&~~ (charge)

K_3N

potassium nitride

$\underbrace{3K^+}_{\oplus\oplus\oplus}$ $\underbrace{1N^{3-}}_{\ominus\ominus\ominus}$

SnF_2

tin(II) fluoride

$\underbrace{1Sn^{\boxed{2+}}}_{+2}$ $\underbrace{2F^-}_{-2}$

Fe_2O_3

$\underbrace{2Fe^{\boxed{3+}}}_{+6}$ $\underbrace{3 \cdot O^{2-}}_{-6}$

iron(III) oxide

TiO_2

$\underbrace{1Ti^{\boxed{4+}}}_{+4}$ $\underbrace{2 \cdot O^{2-}}_{-4}$

titanium (IV) oxide

CaO

calcium oxide

D. Binary Molecular Compounds

2 NM & NM

1) Composed of two non-metals

2) Because they are molecules, no charges are involved.

3) Prefixes are used to show the number of each type of atom present:

1 =	mono-	6 =	hexa-
2 =	di-	7 =	hepta-
3 =	tri-	8 =	octa-
4 =	tetra-	9 =	nona-
5 =	penta-	10 =	deca-

CO_2

4) Exceptions: Do not use mono for the first element

Drop the prefixes ending –a or –o with oxygen compounds.

NO_2

CO carbon monoxide

$SiBr_4$ S_2Cl_6 P_3I_5

Examples:

N_2O

Phosphorus trichloride PCl_3

Nitrogen dioxide NO_2

Dinitrogen pentoxide N_2O_5

Carbon monoxide ______

XeF_7

CF_4 carbon tetrafluoride

P_2O_5 diphosphorus pentoxide

SiO_2 Silicon dioxide

$AsCl_3$ arsenic trichloride

Na_2

Mixed Ionic and Molecular Naming

Molecular	Ionic
NM + NM	M + NM
• DO use numerical prefixes • 2nd element: ___-ide • No "mono" for 1st	• Do NOT use prefixes • Cation is 1st (name it) → If it's a tran. metal, open (__) • Anion: ends in ___-ide (usually)

Circle the molecular compounds, then name all compounds correctly.

(I) CaF_2 calcium fluoride

(I) CaO calcium oxide

(M) NF_3 nitrogen trifluoride

(M) P_2O_5

(M) SiO_2

(I) ~~Fe_2N_3~~ $MgBr_2$ magnesium bromide

H_2O_2

E. DIATOMIC ELEMENTS: Professor Br_2 I_2 N_2 Cl_2 H_2 O_2 F_2

$H_2 + O_2 \rightarrow H_2O$

"Brinklehoff"

The Mole & Avogadro's Number

1 mole of ANYTHING (cars, people, atoms, molecules, books, protons...) is equal to 6.02×10^{23}. You definitely need to memorize this.

1 mole is equal to ↑, just like 1 dozen is equal to 12 things.

Moles are usually referring to atoms and compounds (**1 mole of Na atoms** = 6.**022 × 10^{23} Na atoms!**) because atoms and compounds are very, very, very small.

6.02×10^{23} is Avogadro's number. It's equal to 1 mole. (you only need to know it to "6.02"...). GET OUT YOUR PERIODIC TABLE.

Video Guide: "How Big is a Mole?" (TED-Ed)

1) Who was the first guy to propose numerical "counting" of particles like atoms and molecules?
   Avogadro
2) If you have 6.02×10^{23} (that's about 602,000,000,000,000,000,000,000) molecules of water (H_2O)...
   a. ...how much will it weigh in grams? 18.01 g ~~H + O + O~~ ~~1.01 + 16 + 18~~ H + H + O 1.01 + 1.01 + 16.0
   b. ...and since the density of H_2O is 1 g/mL, 18.01 g of H_2O should also have a volume of 18.01 mL.

KNOW THIS:

1 mole of He atoms = 6.02×10^{23} He atoms.
1 mole of P atoms = 6.02×10^{23} P atoms
1 mole of Cu atoms = 6.02×10^{23} Cu atoms.
1 mole of Na atoms = 6.02×10^{23} atoms of Na.
1 mole of H_2 molecules = 6.02×10^{23} molecules of H_2.
1 mole of CO_2 molecules = 6.02×10^{23} molecules of CO_2. EASY!

Check out the next pattern. Have your PT ready:

1 mole of He atoms = 6.02×10^{23} He atoms = **4.003 g of He**
1 mole of P atoms = 6.02×10^{23} P atoms = **30.97 g of P**
1 mole of Cu atoms = 6.02×10^{23} Cu atoms = 63.55 **g of Cu**
1 mole of Na atoms = 6.02×10^{23} atoms of Na = 22.99 **g of Na**
1 mole of H_2 molecules = 6.02×10^{23} molecules of H_2 = **2.02 g of H_2**

The first tricky thing: How many H atoms are in each molecule of H_2? 2

So... you can easily find the **molar mass** of H on the PT. Why did we have to multiply that PT number (for a *single* H) by **2** in order to get the **molar mass** of H_2?

1 mole of CO_2 molecules = 6.02×10^{23} molecules of CO_2 = 44.01 g of CO_2.
(1 carbon is 12.01 **g/mol**. 2 oxygens = 2× 16.00 **g/mol**)

KNOW THIS before you practice: When you look on the PT and see that the **molar mass** for magnesium (**Mg**) is ____________, you've GOT to have units!

amu g g/mol

- The unit for a regular pack of eggs is 12 **eggs/dozen**.
- The unit for an average pack of paper is 500 **sheets/ream**.
- The unit for a molar mass is **grams/mole**. (seen usually as **g/mol** for short.)

Worked Example for Calculating the Molar Mass of H_2O:

~~H + 2(O)~~ 2(H) + O

18.01 = 1.01 + 1.01 + 16.00

Worked Example for Calculating the Molar Mass of **nitrogen trifluoride**:

NF_3

(14.0) + 3(19.0) = 71.0 g/mol

OYO Worked Example for Calculating the Molar Mass of **ammonium chloride**:

NH_4^+ Cl^- → NH_4Cl

Practice. Find the molar mass of each compound, using your PT. Must use appropriate units!

1) How much molar mass does 1 mole of zirconium have?
2) Determine the molar mass of 1 mole of sodium chloride.
3) Find the molar mass of 1 mole of CF_4.
4) What is the molar mass of 1 mole of barium hydroxide?

*****Always start with what you know*****

5) Worked Example: I have 1.95 × 10^{24} atoms of sulfur. (Sig figs matter!)

a. How many moles of sulfur do I have?

particles → moles

$$\frac{1.95E24\ \cancel{S\ atoms}}{1} \times \frac{1\ mol\ S}{6.02E23\ \cancel{S\ atoms}} = \boxed{3.2392\ mol\ S}$$

b. How many grams does my sample weigh?

moles → mass

$$\frac{3.2392\ \cancel{mol\ S}}{1} \times \frac{32.07\ g\ S}{1\ \cancel{mol\ S}} = \boxed{103}\ g\ S$$

6) Worked Example: Mark has 88.0 grams of solid KF. ← white powder

a. How many moles of KF does Mark have?

mass → moles (use MM)

$$\frac{88\ \cancel{g\ KF}}{1} \times \frac{1\ mol\ KF}{58.10\ \cancel{g\ KF}} = 1.514\ mol\ KF$$

b. How many formula units (particle-pairs/set) of KF does Mark have?

moles → units

$$\frac{1.514\ \cancel{mol\ KF}}{1} \times \frac{6.02E23\ u.\ KF}{1\ \cancel{mol\ KF}} = \boxed{}\ formula\ units\ KF$$

Conversion Mapping:

You can go from mass **to** moles **(and vice-versa) by using the molar __________, found on the period table.** ***You'll have to do some adding and multiplying for compounds.***

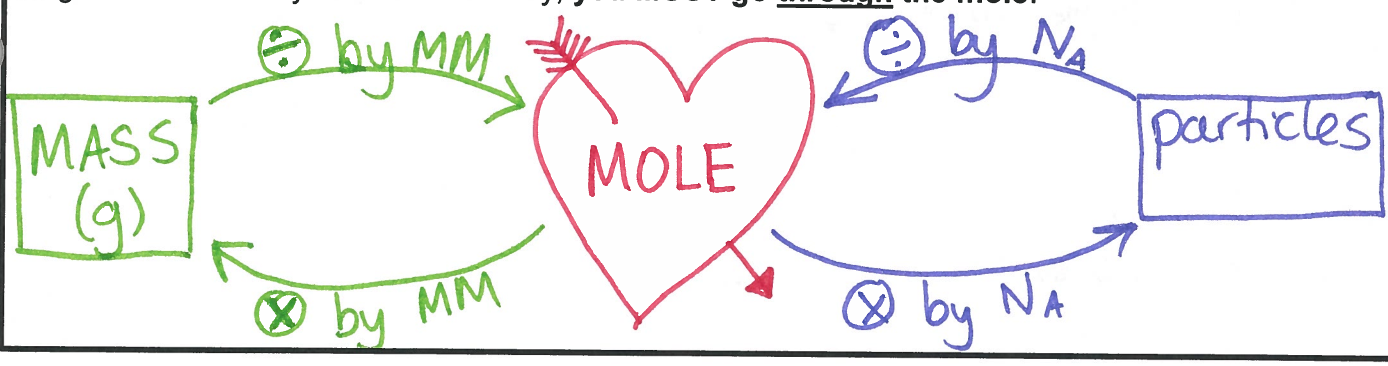
Review: What are the units for molar mass? g/mol amu or g

Review: How would you calculate the molar mass for K_2O, knowing there are 2 K atoms and 1 O atom? $K + K + O = 2(K) + O$ ✓

You can go from moles to particles (and vice-versa) by using ________________'s Number.

Review: What is Avogadro's #? 6.02×10^{23}

Regardless of what you do in chemistry, **you MUST go <u>through</u> the mole:**



Mixed Practice:

7) How many molecules of water are in 60.0 grams of it?

$$\frac{60.0\ \cancel{g\ H_2O}}{1} \times \frac{1\ \cancel{mol\ H_2O}}{18.0\ \cancel{g\ H_2O}} \times \frac{6.02E23\ H_2O\ molecules}{1\ \cancel{mol\ H_2O}} = 2.01 \times 10^{24}\ H_2O\ molecules$$

8) How many grams does 4.77×10^{19} **formula units** of NaBr weight?

$$\frac{4.77E19\ u.\ NaBr}{1} \times \frac{1\ mol\ NaBr}{6.02E23\ u.NaBr} \times \frac{102.9\ g\ NaBr}{1\ mol\ NaBr}$$

$$= \boxed{\quad}\ g\ NaBr$$

37.997 g/mol

X X

9) Which element has a molar mass of 196.97 g/mol? Au (gold)

10) Which element has a molar mass of about 83.8 g/mol? Kr (krypton)

11) Which **diatomic** element has a molar mass of 37.997 g/mol? F

12) 6.02 × 10^{23} atoms of vanadium will have a mass of ______________ grams.

13) 7.5593 × 10^{38} atoms of vanadium will have a mass of ______________ grams.

14) 7.5593 × 10^{38} atoms of vanadium is equal to ____________ moles of vanadium.

15) 3.7 moles of H_2O will have a mass of ______________ grams.

16) 32,600 milligrams of carbon will contain __________________ C atoms.

17) 10.0 moles of phosphorus pentafluoride will have a mass of _______________ grams, and will contain ____________________ molecules of the compound.

18) Students calculated the molar mass of strontium iodide to be 307.13 g/mol. Calculate their % error.

$Sr^{2+}\ I^{-} \rightarrow \boxed{SrI_2}$

$Sr + 2(I) = 341.42$ g/mol TV

$$\%\ err = \frac{|EV - TV|}{TV} \times 100$$

$$= \frac{(307.13 - 341.42)}{341.42} \times 100 = 10.0\%$$

35 psi

GASES and The Mole

Gases take up space. A lot of space!

Recall that the amount of matter in something give us its ***mass***, and the amount of ***space*** that something takes up is called volume.

Volume has a few different units.

15 cc of medicine

15 mL = 15 cm^3

We know that 1 cm^3 = 1 mL = 1 cc.

… and 1000 mL = 1 L

The common unit for VOLUME in chemistry is the liter.

VERY IMPORTANT:

1 mole of a gas is gonna take up lots of space

In fact, 1 mole of any gas will take up 22.4 L **L of space.**

@ STP

Standard

temp. (0°C)

pressure (1 atm)

KNOW THIS: 1 mole (g) = 22.4 L

Review: According to Professor BrINClHOF ("Brinklehoff"), what are the 7 diatomic elements? Write them out in X_2 format, and circle the ones that are gases at standard (normal) temperature and pressure. ● = Solid ● = liquid ● = gas

Br_2 I_2 N_2 Cl_2 H_2 O_2 F_2

Practice with moles and gases: Use the right units! ALWAYS GO THROUGH THE MOLE!

19) How much volume (L) does 8.40 moles of hydrogen (H_2) gas have? (3)

$$\frac{8.40 \cancel{\text{mol } H_2(g)}}{1} \times \frac{22.4 \text{ L } H_2}{1 \cancel{\text{mol } H_2}} = 188 \text{ L } H_2(g)$$

20) How much space (in L) does 113 moles of O_2 take up?

$$\frac{113 \cancel{\text{mol } O_2(g)}}{1} \times \frac{22.4 \text{ L } O_2}{1 \cancel{\text{mol } O_2}} = 2530 \text{ L } O_2(g)$$

O–O

Cl–Cl

$\underbrace{Cl-Cl}_{70.9\ g/mol}$

21) 5.4×10^{16} atoms of helium gas will take up ______________ liters of space.

$$\frac{5.4E16\ \cancel{He\ atoms}}{1} \times \frac{1\ \cancel{mol\ He}}{6.02E23\ \cancel{He\ atoms}} \times \frac{22.4\ L\ He}{1\ \cancel{mol\ He}}$$

22) How much mass does 59.4 L of chlorine gas have?

$\boxed{volume} \rightarrow \boxed{moles} \rightarrow \boxed{mass}$

$$\frac{59.4\ L\ \cancel{Cl_2(g)}}{1} \times \frac{1\ \cancel{mol\ Cl_2}}{22.4\ \cancel{L\ Cl_2}} \times \frac{70.9\ g\ Cl_2}{1\ \cancel{mol\ Cl_2}}$$

$= 188\ g\ Cl_2(g)$

23) 44.8 L of nitrogen gas is ____________ moles of nitrogen.

24) 44.8 L of nitrogen gas is ____________ grams of nitrogen.

25) 44.8 L of nitrogen gas is ____________ molecules of nitrogen.

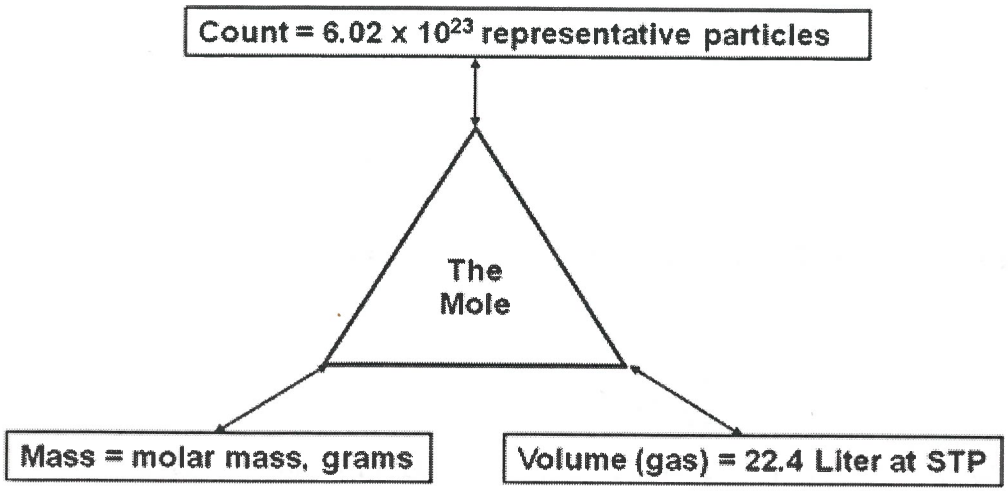
26) Determine the mass of carbon dioxide gas if the volume of the gas is 96.7 L. (STP)

$CO_2(g)$

$$V = \frac{96.7\ \cancel{L\ CO_2}}{1} \times \frac{1\ \cancel{mol\ CO_2}}{22.4\ \cancel{L\ CO_2}} \times \frac{44.0\ g\ CO_2}{1\ \cancel{mol\ CO_2}}$$

$= \cancel{188\ g\ CO_2}\ 190.\ g\ CO_2$

Chapter Seven: The Mole is the Chemist and Physicist's Dozen



The mole = the number of atoms in exaxtly twelve grams of Carbon-12.

1958: 1 mole = 6.02×10^{23}
2006: 1 mole = $6.02214078 \times 10^{23}$

Avogadro's number = ______________________________

Some mole facts—see if you can find the pattern

- 6.02×10^{23} atoms in twelve grams of C-12.
- 6.02×10^{23} atoms in 12.011 g of carbon (naturally occurring
- 6.02×10^{23} atoms in 1.008 g of hydrogen
- 6.02×10^{23} atoms in __________ g of oxygen

Converting between moles and counts of representative particles

1 mole = 6.02 x10^{23} representative particles or

$$\frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ rep. part}} = \frac{6.02 \times 10^{23} \text{ rep. part.}}{1 \text{ mol}}$$

Representative particles:

Elements____

Molecules ____

Ionic compounds_____________________________________

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

1. How many formula units of MgS are there in 0.482 mol of MgS?

2. How many moles are in 1.204×10^{25} molecules of nitrogen dioxide?

3. How many sodium atoms are there in 3.2 moles of sodium?(Ans = 1.9×10^{24} Na atoms)

4. How many moles are there in 6.32×10^{24} formula units of Iron(III) sulfide?(Ans = 10.5 mol)

Review of Counting Atoms in Formulas

One mole of NaCl = 1 mol Na, & 1 mol Cl

One mole of Na_2S = 2 mol Na, & 1 mol S

One mole of $Al_2(SO_4)_3$ = 2 mol Al; 3 mol S, 12 mol O

You do:

Li_3PO_4 = 3 mol Li, 1 mol P, 4 mol O

$Pb(CO_3)_2$ = 1 mol Pb, 2 mol C, 6 mol O

OYO $Fe_2(Cr_2O_7)_3$ = ______

lead(IV) carbonate

Calculating Molar Mass (formula units, atoms, molecules)
Definitions:

- Molar Mass = the mass of 6.02×10^{23} representative particles of an element, molecule, or ionic compound. Molar Mass may also be called Formula Mass.

6.02×10^{23} representative particles=1 Mole = molar mass, gram

Examples:

1. Find the molar mass of Sodium.

2. Find the molar mass of NaCl.

3. Find the molar mass of $CaCl_2$

4. Find the molar mass of $Cu(NO_3)_2$

5. Find the molar mass of $Mg(OH)_2$

6. Find the molar mass of $Ca_3(PO_4)_2$

Converting between Moles and Molar Mass

1 mole = 6.02 x 10^{23} rep. particles = molar mass, g

$$\frac{1\text{ mol}}{\text{molar mass, g}} = \frac{\text{molar mass, g}}{1\text{ mol}}$$

1. What is the mass of 2.3 moles of $MgBr_2$?(Ans = 420 g)

2. How many moles of potassium iodide are in 29.3 g of KI? (Ans = 0.177 mol)

3. How many grams of SO_3 are present in 2.3 moles of SO_3?(Ans = 180 g)

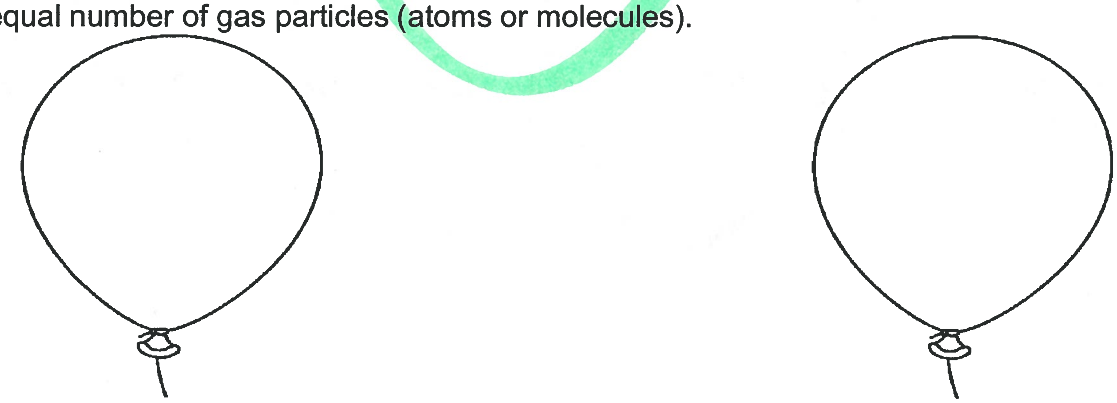
4. How many moles of CaF_2 are equivalent to 450 grams of CaF_2?(Ans = 5.8 mol)

5. How many grams of titanium(IV) sulfide, TiS_2, are present in 0.056 moles of titanium(IV) sulfide? (Ans = 6.3 g)

6. How many moles of ammonium sulfate $(NH_4)_2SO_4$ are present in 52.3 grams of ammonium sulfate?(Ans = 0.396 mol)

Molar Volumes

Avogadro's Hypothesis: Equal volumes of gas at the same temperature and pressure contain an equal number of gas particles (atoms or molecules).



Molar Volume: The volume of one mole of gas at standard temperature and pressure (STP)

STP = $0°C = 273 K$; 1 atm

$\boxed{\text{1 mole gas = 22.4 L at STP for any gas}}$

Assumptions about ideal gases

- The gas particles have no volume (points in space)
- The gas particles have no intermolecular attractions
- The gas particles collide elastically like billiard balls.
- Ideal gases never condense no matter how cold it is.

Real Gases Condense!!!!

Converting between Moles and Molar Gas Volume

1 mole = 6.02 x 10^{23} rep. particles = molar mass, g = molar gas volume (22.4L at STP)

$$\frac{1 \text{ mol}}{22.4 \text{ L}} = \frac{22.4 \text{ L}}{1 \text{ mol}}$$

1. A neon light contains (neon is a noble gas) 0.51 liters of neon gas at STP. How many moles does the light contain?(Ans = 2.3×10^{-2} mol)

2. A helium balloon contains 0.325 moles of gas at STP. What is the balloon's volume in liters? (Ans = 7.28 L He)

3. An underground cavern contains 5.5 x 10^5 liters of natural gas, CH_4, at STP. How many moles of gas are in the cavern? (Ans = 2.5×10^4 mol)

4. A blimp contains 35,000 liters of hydrogen gas (flammable!) at STP. How many moles of hydrogen does it contain? (Ans = 1.6×10^3 mol)

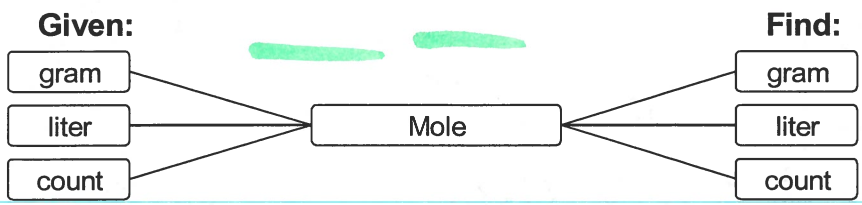
More Molar Conversions—You can go anywhere!

1 mole = 6.02×10^{23} rep. part = gram formula mass = 22.4 Liters at STP

Chapter Seven Objectives

- Memorize Avogadro's number
- Memorize STP volume of 1 mole of gas (0°C, 1 atm)
- Convert between grams, moles, representative particles and liters using factor label method.
- Calculate molar masses

unit to mole	mole to unit
$\frac{1\ \text{mole}}{6.02 \times 10^{23}\ \text{rep. part.}}$	$\frac{6.02 \times 10^{23}\ \text{rep. part.}}{1\ \text{mole}}$
$\frac{1\ \text{mole}}{\text{g. molar mass}}$	$\frac{\text{g. molar mass}}{1\ \text{mole}}$
$\frac{1\ \text{mole}}{22.4\ \text{L. gas}}$	$\frac{22.4\ \text{L. gas}}{1\ \text{mole}}$



1. How many atoms of gold are contained in a 505 gram bar of Au?(Ans = 1.54×10^{24} Au atoms)

2. Find the number of moles of Cl_2 gas in a 1.46×10^4 liter tank at STP.(Ans = 62 mol)

3. A balloon contains 1.2 grams of Helium. What is the balloon's volume at STP?(Ans = 7.3 L)

4. How many molecules of fluorine gas are in an 0.0030 liter ampule at STP?(Ans = 8.2×10^{19} molec.)

Chapter Eight Skeleton Notes Part 1

I. Chemical Equations $A + B \rightarrow C + D$

fwd rxn

Symbols	Use
$\rightarrow$	"goes forward" to completion
$\leftrightarrows$ $\rightleftharpoons$	reversible rxn (equilibrium)
(s), (l), (g)	solid, liquid, & gas (states)
(aq)	aqueous (solution; dissolved in H_2O)
$\xrightarrow{catalyst}$	speeds up a rxn
$\xrightarrow{heat}$	add heat

$A \xrightarrow{X} B + C$

$\xrightarrow{\Delta}$

$\xrightarrow{-\Delta}$

B: Examples of chemical equations

$Mg^{2+}(aq)$

Synthesis • $\boxed{Mg(s) + O_2(g)} \rightarrow \boxed{MgO}$ ← SOLID

Reactants = $Mg(s)$ & $O_2(g)$ Products = MgO

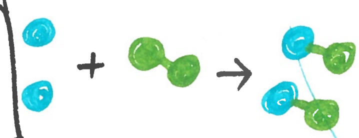
Decomposition • $\boxed{OF_2(g)} \rightarrow \boxed{F_2(g) + O_2(g)}$

Reactants = Oxygen difluoride Products =

Writing chemical equations from word equations

1. Sodium metal reacts with chlorine gas to form sodium chloride

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$



2. Iron metal reacts with oxygen gas to form rust, iron(III)oxide.

"2-3" switch

$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

3. Solid nitrogen triiodide decomposes to solid iodine and nitrogen gas.

$2NI_3(s) \rightarrow 3I_2(s) + N_2(g)$

Catalysts: Speed up rxns

Skeleton equations do not show the amounts of products and reactants.

C. Balancing chemical reactions using coefficients

Law of conservation of mass: mass is neither created nor destroyed... only transformed

Balanced equations use coefficients in front of formulas to show the # of moles of each compound / element

2×3

$5N_2$

A chemical reaction is balanced if there are the same number of each kind of element on both sides of the chemical equation. If there are four oxygens on the reactant side, there will be 4 oxygens on the product side.

Count the atoms of each element of both sides.→ Indicate which equations are balanced.

SYN $C(s) + O_2(g) \rightarrow CO_2(g)$

1 C	1 C
2 O	2 O

BAL ü

SYN $Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$

1 Fe	2 Fe
2 O	3 O

NB

SYN $3NH_3 + H_3PO_4 \rightarrow (NH_4)_3PO_4$ coefficients apply to the entire compound

3 N	3 N
12 H	12 H
1 P	1 P
4 O	4 O

Bal ü

Combustion $C_4H_{10} + 4O_2 \rightarrow \boxed{4CO_2} + \boxed{5H_2O}$

4 C	4 C
10 H	10 H
8 O	8+5=13 O

NB

DEC $Ca(ClO_3)_2 \rightarrow CaCl_2(s) + 3O_2(g)$

Ca	Ca
Cl	Cl
O	O

Bal

balancing chemical equations

a. Write the skeleton chemical equation leaving blanks for the coefficients:
b. Count the number of each element in the reactant and product side
c. Balance the equation using whole number coefficients (NEVER SUBSCRIPTS)
d. Track your changes
   1. Balance the other compounds to the most complicated compound.
   2. Balance the binary compounds (H_2O, CO_2, NO_2)
   3. Balance diatomics and elements last
   4. If you end up with an odd number that won't balance (3 oxygens on one side, two on the other) double all the coefficients filled in so far.
   5. Double check when you're done.

$\underline{1}\,N_2O_4(g) \xrightarrow{Pt} \underline{1}\,N_2(g) + \underline{2}\,O_2(g)$

1 | 1 2

✓ $\underline{3}\,K_2S + \underline{2}\,FeCl_3 \rightarrow \underline{1}\,Fe_2S_3 + \underline{6}\,KCl$

3 2 | 1 $\cancel{2}\cancel{3}$6

For combustion reactions, use the CHO rule (C first, H second, O last)

$\underline{1}\,CH_4 + \underline{2}\,O_2 \rightarrow \underline{1}\,CO_2 + \underline{2}\,H_2O$

1 2 $\underbrace{1 \quad 2}_{2\,+\,2}$

$\underline{2\,\cancel{1}}\,C_2H_6 + \underline{7\,\cancel{3}}\,O_2 \rightarrow \underline{\cancel{2}\,4}\,CO_2 + \underline{\cancel{3}\,6}\,H_2O$

— — $\underbrace{2 \quad 3}_{\underbrace{4\,+\,3}_{7\cdot\text{Oxy}}}$

$\underline{1}\,C_2H_6O + \underline{3}\,O_2 \rightarrow \underline{2}\,CO_2 + \underline{3}\,H_2O$

1 — $\underbrace{2 \quad 3}_{\underbrace{4\,+\,3}_{7\ \text{oxy}}}$

Chlorine gas was used in chemical warfare during WWI. The Germans used Chlorine gas on the Allied Forces in Ypres, France in 1915. Chlorine reacts with the moisture in lungs to produce hydrochloric acid, HCl.

You try balancing the reaction for chlorine in your lungs:

$\underline{2}\,Cl_2(g) + \underline{2}\,H_2O(l) \rightarrow \underline{4}\,HCl(l) + \underline{1}\,O_2(g)$

4 ~~2~~ Cl	4 ~~2~~ ~~1~~ Cl
4 ~~2~~ H	4 ~~2~~ ~~1~~ H
2 ~~1~~ O	2 ~~1~~ ~~2~~ O

Now try balancing the reaction for phosgene (Cl_2CO) in your lungs. This is another poisonous gas used in warfare.

$\underline{2}\,Cl_2CO(g) + \underline{2}\,H_2O(l) \rightarrow \underline{4}\,HCl(l) + \underline{2}\,CO(g) + \underline{1}\,O_2(g)$

4 ~~2~~ Cl	4 ~~2~~ ~~1~~ Cl
4 ~~2~~ H	4 ~~2~~ ~~1~~ H
2 ~~1~~ C	2 ~~1~~ C
4 ~~3~~ ~~2~~ O	4 ~~3~~ O