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1. Unit 1 and 2: Fundamental Skills Review

1.1 SI Units

1.1.1

SI Units

The following units are important for chemistry:

Length- meter (m)
Time- second (s)
Mass- kilogram (kg)
Temperature- kelvin (K)
Amount of a substance- mole (mol)

Sometimes we will see these prefixes being used when units are smaller/larger than the above base SI units:

Common Unit Prefixes			
Prefix	**Symbol**	**Factor**	**Example**
femto	f	10^{-15}	1 femtosecond (fs) = 1×10^{-15} s (0.000000000000001 s)
pico	p	10^{-12}	1 picometer (pm) = 1×10^{-12} m (0.000000000001 m)
nano	n	10^{-9}	4 nanograms (ng) = 4×10^{-9} g (0.000000004 g)
micro	μ	10^{-6}	1 microliter (μL) = 1×10^{-6} L (0.000001 L)
milli	m	10^{-3}	2 millimoles (mmol) = 2×10^{-3} mol (0.002 mol)
centi	c	10^{-2}	7 centimeters (cm) = 7×10^{-2} m (0.07 m)
deci	d	10^{-1}	1 deciliter (dL) = 1×10^{-1} L (0.1 L)
kilo	k	10^{3}	1 kilometer (km) = 1×10^{3} m (1000 m)
mega	M	10^{6}	3 megahertz (MHz) = 3×10^{6} Hz (3,000,000 Hz)
giga	G	10^{9}	8 gigayears (Gyr) = 8×10^{9} yr (8,000,000,000 yr)
tera	T	10^{12}	5 terawatts (TW) = 5×10^{12} W (5,000,000,000,000 W)

WIZE TIP

Memorizing these units can be very helpful on exams!

Example: you might see wavelength in *nm* instead of *m*. You should know that **1 nm=1x10^{-9}m**

Other Important Units for Chemistry

- Relationship between L and mL (1 L=1000 mL)

- The Angstrom (Å) is $1x10^{-10}$m (low yield)
  - Used in chemistry because it is a very tiny value, perfect for using when measuring the size of an atom or molecules!

1.2 Naming Compounds

1.2.1

Naming Compounds

Naming Ionic Compounds

Ionic compounds are written as a cation (+) followed by an anion (-)

Example: NaCl is made up of Na^+ and Cl^-

WIZE TIP

Name is written in this format:
[cation’s name] [space] [anion’s name with “-ide” ending]

Example:
NaCl is sodium chloride

WATCH OUT!

If the metal (cation) is a **transition metal**, it could have more than one possible oxidation state or "charge".

Indicate the charge in the name using this format:
[transition metal name] (oxidation state in roman numerals) [anion's name with "-ide" ending]

Example:
Fe_2O_3 is iron (III) oxide

Examples:

a) KCl

Ionic or covalent? ________

Name: ____________________

b) PbO

Ionic or Covalent? ________

Name: ________________

Naming Ionic Compounds With Polyatomic Ions

Polyatomic Ion Cheatsheet			
Ion	**Name**	**Ion**	**Name**
NH_4^+	Ammonium	MnO_4^-	Permanganate
NO_2^-	Nitrite	CrO_4^{2-}	Chromate
NO_3^-	Nitrate	$Cr_2O_7^{2-}$	Dichromate
SO_3^{2-}	Sulfite	O_2^{2-}	Peroxide
SO_4^{2-}	Sulfate	O_2^-	Superoxide
HSO_4^-	Hydrogen sulfate*	$C_2O_4^{2-}$	Oxalate
HSO_3^-	Hydrogen sulfite	$HC_2O_4^-$	Hydrogen Oxalate
OH^-	Hydroxide	HCO_2^-	Formate
CN^-	Cyanide	$S_2O_3^{2-}$	Thiosulfate
PO_4^{3-}	Phosphate	$HS_2O_3^-$	Hydrogen Thiosulfate
HPO_4^{2-}	Hydrogen Phosphate	BrO_4^-	Perbromate
H_2PO^{4-}	Dihydrogen Phosphate	BrO_3^-	Bromate
CO_3^{2-}	Carbonate	BrO_2^-	Bromite
HCO_3^-	Hydrogen Carbonate**	BrO^-	Hypobromite
ClO^-	Hypochlorite	IO_4^-	Periodate
ClO_2^-	Chlorite	IO_3^-	Iodate
ClO_3^-	Chlorate	IO_2^-	Iodite
ClO_4^-	Perchlorate	IO^-	Hypoiodate
$C_2H_3O_2^-$	Acetate		

WIZE TIP

Naming ionic compounds that contain polyatomic ions are very similar to the naming of other ionic compounds.

[Name of metal cation] [space] [name of polyatomic anion]

Example:

$NaNO_3$ is sodium nitrate

WATCH OUT!

Pay attention to the number of O atoms in a polyatomic ion when naming it!

Example:

ClO^- → **hypo**chlor**ite**

ClO^{2-} → chlor**ite**

ClO^{3-} → chlor**ate**

ClO^{4-} → **per**chlor**ate**

Note: The charge does not change, but the endings change depending on how many oxygen atoms there are!

****Tip:** If you memorize the number of O atoms and charge for each polyatomic ion with the "ate" ending it will be easy to determine the names of the rest of the polyatomic ions!

Examples:

a) $Sn(NO_3)_2$ is ____________________

b) KNO_2 is ____________________

Naming Covalent Compounds

Covalent bonds are formed between two non-metals.

WIZE TIP

Covalent compounds are usually written from least to most electronegative.
[prefix][name of first element] [space] [prefix] [name of 2nd element with "-ide" ending

Prefixes are used to describe the number of atoms. See the prefixes in the table below :)

Note: If there is only one atom on the left, we drop the mono prefix.

Examples:
CO is carbon **mon**oxide (the last vowel in the prefix is dropped if the next letter is also a vowel)
CO_2 is carbon **di**oxide

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

Example:

N_2O_4 is ____________________

1.2.2

Example: Providing Chemical Formulae and Names for Ionic Compounds

Give a chemical formula for the following compounds:

a) Potassium Fluoride

b) Manganese(II) oxide

c) Ammonium sulfate

Give a name to the following chemical compounds:

a) $TiBr_4$

b) Na_2S

c) $CuSO_4$

1.2.3

Example: Naming Covalent Compounds

a) NF_3

b) CF_4

c) N_2O

1.2.4

Example: Naming Ionic and Covalent Compounds

First state whether each compound is ionic or covalent then name it!

Cl_2O_7: ____________, ________________________

$CaCl_2$: ________, ____________________

$(NH_4)_2S$: _________________________, ______________________

KCN: _________________________, ______________________

PbI_2: ________, __________________

$KClO_4$: _________________________, __________________________

1.2.5

Practice: Naming

What is the name of the chemical compound: $Co_2(SO_3)_3$?

- Cobalt sulfate
- Cobalt(III) sulfate
- Cobalt(III) sulfite
- Cobalt(II) sulfate
- Cobalt(II) sulfite

1.3 Oxidation States

1.3.1

Oxidation States

Oxidation state: A way to describe the degree of oxidation of a chemical

WIZE TIP

Tips to Determine Oxidation State:

1. Free elements: oxidation state = 0

Examples: $Cl_2(g)$ has an oxidation state of 0, Na(s) has an oxidation state of 0

2. Monoatomic ions: oxidation state = ionic charge

Examples: $Cl^-(aq)$ has an oxidation state of -1, $Na^+(aq)$ has an oxidation state of +1.

3. Hydrogen: oxidation state = +1

a. **Exception: hydrogen in hydrides** (e.g., NaH) – **oxidation state = -1**

4. Oxygen: oxidation state = -2

a. **Exception: oxygen in peroxides** (e.g., H_2O_2) – **oxidation state = -1**

5. For **neutral molecules, oxidation numbers add up to zero.**
For **polyatomic ions, oxidation numbers add up to the ion charge.**

Examples: $Cl_2(g)$ has an oxidation state of 0, NaCl oxidation numbers add up to 0, SO_4^{2-} oxidation numbers add up to 2-.

1.3.2

Oxidation States - Additional Information

Other Low Yield Tips to Help Determine Oxidation States

- Alkali metals: oxidation state= +1
- Alkaline earth metals: oxidation state= +2
- Nitrogen Group: oxidation state= -3
- Oxygen family: oxidation state= -2
- Halogens: oxidation state= -1

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Note on Transition Metals

Transition metals can have different charges (ex. Pb^{2+} and Pb^{4+})
As a result, transition metals can also have different oxidation states!

Example:

What is the oxidation state of Cu in CuCl? ______

What is the oxidation state of copper in $Cu(NO_3)_2$? ______

When naming compounds involving transition metals, the **oxidation state of the transition metal is written in Roman numerals and put in brackets:**

- copper(I) chloride: CuCl
- copper(II) nitrate: $Cu(NO_3)_2$

Example: Oxidation State Practice

What is the oxidation state of:

S in S_2O_3: ______

O in S_2O_3: ______

What about Pb in Pb_2O_3?

This one is different because this compound is a mixed oxidation state compound, so although Pb should have 2+ or 4+, here, we get an oxidation state of ______ and that is ok because both Pb^{2+} and Pb^{4+} are present in the compound.

1.3.4

What is the oxidation state of Br in BrO_3^- ?

- [ ] +4
- [ ] +5
- [ ] +6
- [ ] -4
- [ ] -5
- [ ] -6

1.3.5

What is the oxidation state for the following: (enter answer as -1 or +2 for example)

a) P in P_2O_5:

b) Cr in $Na_2Cr_2O_7$:

c) Al in $Al_2(SO_4)_3$:

d) Na in Na_2O_2:

e) O in Na_2O_2:

f) N in NH_3:

1.4 Reduction-Oxidation (Redox) Reactions

1.4.1

Redox Reactions

OIL RIG! Oxidation Is Loss of electrons, Reduction Is Gain of electrons
OR
LEO GER! Loss of Electrons is Oxidation, Gain of Electrons is Reduction

OXIDATION (lose e-)	REDUCTION (gain e–)
Reducing agent reduces others and is itself oxidized.	**Oxidizing agent** oxidizes others and is itself reduced.
Oxidation number increases.	Oxidation number is reduced.

Keep in mind that in organic chemistry the definitions of oxidation and reduction are different:

- For **reduction**, we would look for **more bonds to H** and **less bonds to O**
- For **oxidation**, we would look for **less bonds to H** and **more bonds to O!**

How can you recognize if a redox reaction is occuring?

1. Can **check oxidation #s of each element** to see if they **change from reactants to products** (most important for now!)

2. Familiarize yourself with **common oxidizing and reducing agents** (you'll learn more about this in orgo!)

- **Oxidizing agents:** O_2, halogens, HNO_3, $Cr_2O_7$2-, MnO_4-
- **Reducing agents:** H_2, metals, C

1.4.2

Practice: Oxidation vs Reduction

Consider the following balanced redox equation:

$$16H^{+}_{(aq)} + 2MnO^{-}_{4(aq)} + 10Cl^{-}_{(aq)} \rightarrow 2Mn^{2+}_{(aq)} + 5Cl_{2(g)} + 8H_2O_{(l)}$$

Which species is being oxidized?

- H^+
- MnO_4^-
- Cl^-
- Mn^{2+}
- Cl_2

1.4.3

Practice: Oxidizing vs Reducing Agent

Consider the following balanced redox equation:

$$16H^{+}_{(aq)} + 2MnO^{-}_{4(aq)} + 10Cl^{-}_{(aq)} \rightarrow 2Mn^{2+}_{(aq)} + 5Cl_{2(g)} + 8H_2O_{(l)}$$

Which species is the reducing agent?

- H^+
- MnO_4^-
- Cl^-
- Mn^{2+}
- Cl_2

Practice: Oxidizing vs Reducing Agents

What is the reducing agent in the following chemical reaction?

$NO + Au^{+} \rightarrow NO_3^{-} + Au$

- Au^{+}
- NO
- NO_3^{-}
- Au

1.5 Significant Digits

1.5.1

Significant Digits

Significant digits (aka significant figures) indicate the precision of a measurement and determine how precisely a calculated number can be represented.
To count the significant figures in a value, follow these steps while reading from left to right through the number:

1) Start counting at the first non-zero digit

Example:

0.0095 → _____

2) If a digit is not zero, it is significant

Example:

456 → _____

3) If a digit is zero and there are nonzero digits or a decimal ahead, it is significant

Examples:

1290. → _____

12900 → _____

4) If a digit is zero and it occurs after a decimal point, it is significant

Example:

52.00 → _____

Adding and Subtracting

When adding or subtracting measurements, the **number with the least digits beyond the decimal point sets the limit for the answer**. Any extra digits must be rounded off.

Examples:
4.3568g - 0.14g

100mL + 150.47mL

Multiplying and Dividing

When multiplying or dividing measurements, the **number with the least significant figures sets the limit for the answer.**
Any extra digits must be rounded off.

Examples:
49.63m x 36m

57.6kPa/101.325kPa

0.0076 x 1.2455

Accuracy and Precision

Accuracy: how closely a measurement aligns with a correct value

Precision: how closely a measurement matches the same measurement when repeated

1.5.2

Example: Significant Digits

How many significant digits do the following numbers have?

0.00045 → ____

234000 → ____

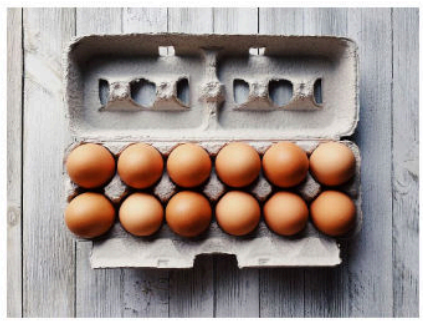
234000.0 → ____

1.6 Atoms, Molecules, and the Mole!

1.6.1

Moles

When someone says they want a dozen donuts, roses, or eggs we know they mean they want 12. A **mole** in chemistry also tells us the amount of something.



- Just like how a dozen of something = 12, 1 mole of something =$6.02x10^{23}$ molecules
- $6.02x10^{23}$ is referred to as **Avagadro's number (N_A)**

WIZE TIP

You should memorize Avagadro's number since it is not always provided on exams! We'll soon see that you need to know it to solve problems :)

$$\boxed{\mathbf{N_A = 6.02 \times 10^{23}}\ \text{molecules/mole}}$$

We use the unit "**moles**" to help make very small amounts more measurable. A mole is just like any other unit!

There are 2 Equations Related to Moles:

$$\boxed{n = \frac{m}{M}} \qquad \boxed{n = \frac{N}{N_A}}$$

n=# of ***moles***

m=mass (***g***)

M=molar mass (***g/mol***)

N=# of ***molecules or atoms***

N_A=Avagadro's number=***6.02x10^{23} molecules***

Example #1:

If we are told that a sample of CO_2(s) weighs 11g, how many moles are present?

Example #2

We have 2 moles of CO_2 present in our sample, how many molecules are there?

Example #3

How many oxygen atoms there in 1 mole of CO_2?

1) Find Number of Molecules In the Sample

2) Find the Number of O atoms in the Sample

In each molecule there are ______ O atoms.
Therefore to find the number of O atoms:
# of molecules in sample x ______ O atoms/molecule =# of O atoms

# of O atoms= ________________________________

1.6.2

Example: Calculate the Number of Molecules of Ethyl Mercaptan

The volatile liquid ethyl mercaptan, C_2H_6S, is one of the most odoriferous substances known. It is added to natural gas to make gas leaks detectable. (***d*** = 0.84 g/mL; **MW** = 62.1 g/mol)

1. How many C_2H_6S, molecules are contained in a 3.0 μL sample?

2. In the same 3.0 μL sample, how many C and H atoms are there?

1.6.3

Practice: Converting Mass to Number of Atoms

Calculate the number of nitrogen atoms in 2.25 g of bismuth(III) nitrate.

a) 1.03×10^{22} atoms

b) 1.03×10^{21} atoms

c) 3.43×10^{21} atoms

d) 3.43×10^{22} atoms

1.6.4

Practice: Finding the Number of Moles of Iron

Calculate the number of moles of iron atoms in 14.1 g of iron oxide, Fe_2O_3

A) 0.177 mol

B) 0.0821 mol

C) 0.0906 mol

D) 0.0451 mol

1.7 Dimensional Analysis

1.7.1

Dimensional Analysis

This is helpful when converting units or figuring out a way to calculate one quantity without a formula.

Unit Conversion Examples

Convert 4.56 km/h into m/s.

Convert 1.63 Ms to μs.

Calculate the speed of an object which moves 12m in 8s in m/s.

1.8 Isotopes and Atomic Weight

1.8.1 Isotopes and Atomic Mass

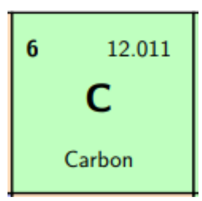
Isotopes and Atomic Mass

Isotopes

- When two atoms have the **same number of protons and electrons**, but a **different number of neutrons**, we call these **isotopes**.
- Isotopes have the **same atomic number**, but a **different mass number.**

^{12}C	^{13}C	^{14}C
protons	protons	protons
neutrons	neutrons	neutrons
98 %	1.07 %	Trace

How is the Atomic Mass in the Periodic Table Calculated?



- The atomic mass is written as 12amu for C for example, but this does not mean that each C atom weighs exactly 12amu!

- **The atomic mass # is actually a weighted average** based on the **relative abundance of isotopes**
    - Isotopes have similar reactivity to one another, that's why we can form C bonds with either C-12 or 13 etc

To determine the average mass of an element, use this equation:

$$\boxed{A.W. = \sum_{i=1}^{n} (mass\ of\ each\ isotope_i)(abundance\ of\ each\ isotope_i)}$$

WATCH OUT!

Plug in the **mass of each isotope in amu** and plug in the **relative abundance of each isotope in the form of a decimal** not percentage!

For example, above we are told the relative abundance of C-12 is 98%. You would want to plug in 0.98 for this isotope's relative abundance.

1.8.2

Example: Solving for the Atomic Weight

Chlorine can be found in nature as $^{35}\mathrm{Cl}$ (mass 34.969u, 75.78% abundance) and $^{37}\mathrm{Cl}$ (mass 36.966u, 24.22% abundance). What is the average atomic mass of Cl?

1.8.3

Practice: Solving for the Weight of an Isotope

Naturally occurring potassium contains two stable isotopes. The lighter isotope, ^{39}K (38.9637 amu) is the more abundant isotope, accounting for 93.26% of the nuclei. What is the weight of the heavier isotope, ^{41}K?

- 41.00 amu
- 40.96 amu
- 39.09 amu
- 41.08 amu

1.8.4

Bromine is a noxious fuming red liquid in its elemental form. Naturally occurring bromine has two stable isotopes ^{79}Br (78.918*u*) and ^{81}Br (80.916*u*). What are the relative abundances of these two isotopes?

Enter your answer as a percentage (ex. If the answer is 16.26%, enter 16.26 as the answer)

Relative Abundance of Br-79

Relative Abundance of Br-81

1.9 Definitions & Solutions Calculations

1.9.1

Everything You Need to Know About Concentrations

Solutions

- Homogeneous mixture containing a solvent (liquid) and one or more solutes (can be gas, liquid or solid)
- Concentrations can be measured in different ways:

a. **Molarity (M):** $\frac{\text{moles of solute}}{\text{litre of solution}} = \frac{mol}{L}$

b. **Mass Percent (%w/w):** $\frac{\text{g of solute}}{\text{g solution}} \times 100$

c. **Percent weight by volume (%w/v):** $\frac{\text{g of solute}}{\text{100mL of solution}} \times 100$

d. **Percent volume by volume (%v/v):** $\frac{\text{mL of solute}}{\text{100mL of solution}} \times 100$

e. Concentration can be expressed in **parts per million (ppm) or parts per billion (ppb)**, usually when working with trace impurities (ex: the label on bottled water)

$$\frac{[mass\ of\ solute\ (g)]}{mass\ of\ solution\ (g)} x10^6\ \ (for\ ppm)\ \ or\ 10^9\ \ (for\ ppb)$$

- To convert from **mL** of a liquid to **grams**, use the density.

$$volume\ \ (mL)\ \cdot\ density\ \ \left(\frac{g}{mL}\right)\ = mass\ \ (g)$$

Dilutions

- When diluting a solution from an initial concentration, we can find the final volume for a desired final concentration or vice versa:

- The number of moles of solute does not change, but the volume of the solvent increases.

$$n_1 = n_2, therefore$$

$$C_1V_1 = C_2V_2$$

Note that a diluted/standard/and stock solution are all different!

Standard solution: solution that we know the concentration for accurately
Stock solution: standard solutions kept in concentrated form
Dilution: Stock and standard solutions are diluted to prepare solutions of a lower concentration

1.9.2

Concentrations can be measured in four ways:

1. Molarity (M) $\longrightarrow \frac{\text{moles of solute}}{\text{liter of solution}} = \frac{\text{mol}}{L}$
2. Mole fraction (X) $\rightarrow X_A = \frac{n_A}{n_{total}}$
3. Mass percentage (%) $\rightarrow \frac{\text{grams of solute}}{\text{grams of solution}} \times 100\%$
4. Molality (m) $\rightarrow \frac{\text{moles of solute}}{\text{kilograms of solvent}} = \frac{\text{mol}}{kg}$

1.9.3

Practice: Concentration

Calculate the concentration of Ca^{2+} ions in a solution containing 100.0 mL of 0.100 M $CaSO_4$ and 50.0 mL of 0.120 M $Ca_3(PO_4)_2$.

Answer

1.9.4

Practice: Molarity

Calculate the molarity of a KNO_3 solution with a mole fraction of 0.225 and a density of 1.19 g/mL.

Answer

1.9.5

For a 3.94 M aqueous solution of strontium nitrate with density of 1.11 g/mL, calculate the mole fraction of strontium nitrate.

Answer

1.10 Introduction to Chemical Equations and Balancing

1.10.1

Balancing Reactions

Reactions must be balanced before the stoichiometry can be analyzed.

What makes a reaction balanced?

The number of ________ of each type on the ______________ must equal the number of atoms of each type on the ______________ of the reaction.

Example:

$$\text{Unbalanced:} \qquad C_3H_{8(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$$

$$\text{Balanced:} \qquad C_3H_{8(g)} + 5\ O_{2(g)} \rightarrow 3\ CO_{2(g)} + 4\ H_2O_{(g)}$$

1.10.2

Balancing Chemical Reactions

Unbalanced Reaction

$$\mathbf{CH_4 + O_2 \longrightarrow CO_2 + H_2O}$$

The **subscripts** tell us the ratio of different atoms in one molecule.

Example: One CH_4 molecule has 4 hydrogen atoms and 1 carbon atom

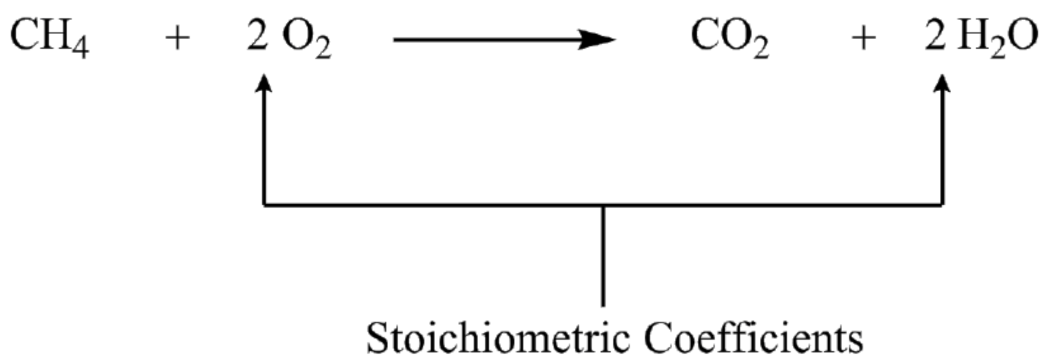
! **WATCH OUT!**

Sometimes you could be given unbalanced equations and you are expected to balance them! We **need to use the balanced equation for calculations** in order to get the correct answer!

Steps to Balance Chemical Equations

1. Balance **elements that only appear once** on the reactant side and once on the product side (typically elements not including C, N, H, O)
2. **Group polyatomic ions** and balance them as a group (not as separate elements, *e.g.* balance NO_3 as 1 NO_3 group, not 1 N and 3 O)
3. **Balance all other elements**, starting with the least common elements to the most common elements
4. Make sure all **coefficients are whole numbers** AND are represented in the **simplest integer ratio possible**
5. **Check** to make sure all elements are balanced on each side of the chemical equation

Balanced Reaction



- The **stoichiometric coefficients** of the balanced equation tell us that we need _____ oxygen molecules for each methane (CH_4) molecule and we produce _____ waters for every methane consumed.
- It also tells us that for every 1 mole of methane used we produce _____ mole of CO_2.

WATCH OUT!

We want the ***lowest* possible coefficients.**

Example:

If I balanced an equation and got:

$$2MgCl_2\,(aq) + 4NaOH\,(aq)\ \rightarrow 2Mg\,(OH)_2\,(s)\ +\ 4NaCl\,(aq)$$

The **greatest common factor** that can go into each coefficient is 2 so we need to divide each coefficient by 2 to get the lowest possible numbers!

$$MgCl_2\,(aq) + 2NaOH\,(aq)\ \rightarrow Mg\,(OH)_2\,(s)\ +\ 2NaCl\,(aq)$$

Example: Balancing a Chemical Reaction

Balance the following equation:

$$NO_{(aq)} \quad + \quad O_{2(g)} \quad + \quad H_2O\ _{(l)} \quad \rightarrow \quad HNO_{2(aq)}$$

WIZE CONCEPT

The physical state of each chemical compound is often included in chemical equations:

(s) = solid
(l) = liquid
(g) = gas
(aq) = aqueous

Polyatomic Ions

Note: There is no need to memorize the following table. It is just included here so you have an idea about what a polyatomic ion is!

Polyatomic Ion Cheatsheet

Ion	Name	Ion	Name
NH_4^+	Ammonium	MnO_4^-	Permanganate
NO_2^-	Nitrite	CrO_4^{2-}	Chromate
NO_3^-	Nitrate	$Cr_2O_7^{2-}$	Dichromate
SO_3^{2-}	Sulfite	O_2^{2-}	Peroxide
SO_4^{2-}	Sulfate	O_2^-	Superoxide
HSO_4^-	Hydrogen sulfate*	$C_2O_4^{2-}$	Oxalate
HSO_3^-	Hydrogen sulfite	$HC_2O_4^-$	Hydrogen Oxalate
OH^-	Hydroxide	HCO_2^-	Formate
CN^-	Cyanide	$S_2O_3^{2-}$	Thiosulfate
PO_4^{3-}	Phosphate	$HS_2O_3^-$	Hydrogen Thiosulfate
HPO_4^{2-}	Hydrogen Phosphate	BrO_4^-	Perbromate
H_2PO^{4-}	Dihydrogen Phosphate	BrO_3^-	Bromate
CO_3^{2-}	Carbonate	BrO_2^-	Bromite
HCO_3^-	Hydrogen Carbonate**	BrO^-	Hypobromite
ClO^-	Hypochlorite	IO_4^-	Periodate
ClO_2^-	Chlorite	IO_3^-	Iodate
ClO_3^-	Chlorate	IO_2^-	Iodite
ClO_4^-	Perchlorate	IO^-	Hypoiodate
$C_2H_3O_2^-$	Acetate		

1.10.5

Example: Balance a Chemical Equation with Polyatomic Ions

$$(NH_4)_2\,SO_4\ (aq) \quad + \quad Ba\,(NO_3)_2\ (aq) \quad \rightarrow \quad NH_4NO_3\ (aq) \ + \ BaSO_4\ (s)$$

1.10.6

Practice: Balancing the Combustion of Furan

Choose the answer that balances the reaction shown below:

$$\mathrm{A}\ C_4H_4O(l) + \mathrm{B}\ O_2(g) \rightarrow \mathrm{C}\ CO_2(g) + \mathrm{D}\ H_2O(g)$$

- A = 2; B = 9 ;C = 8; D = 4
- A = 1; B = 3 ;C = 5; D = 2
- A = 1; B = 1 ;C = 1; D = 1
- A = 4; B = 25 ;C = 20; D = 10

1.11 Stoichiometry

1.11.1

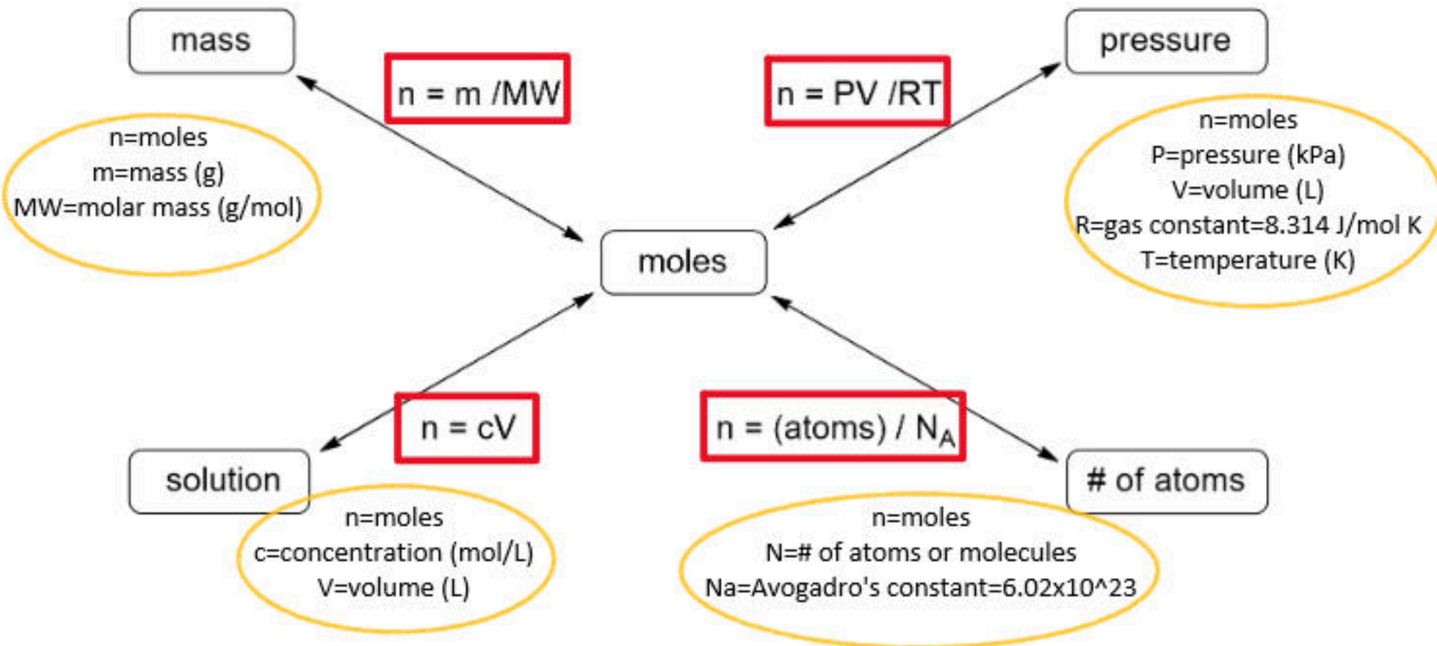
Intro to Stoichiometry

Stoichiometry allows us to predict the quantities of products or reagents across a chemical reaction.

Typically this is done in a few steps:

1. Calculating the number of moles of the reagents
2. Calculating the number of moles of the products using the stoichiometric ratio
3. Calculating the mass or pressure of the products.

Moles are the central unit!



In addition we can convert between mass and volume of a pure substance by looking at it's **density**,

$$\boxed{\delta = m/V}$$

Stoichiometry of a Reaction

We use the **coefficients** of the **balanced reaction** along with the our equations that convert mass, volume, and concentration into moles to predict the quantities of reactants and products in a chemical reaction.

To answer any stoichiometry problem, focus on converting to and from moles!

General Steps to Solving a Stoichiometry Problem:

1. Convert the values given in the problem about a reactant or product to a **number of moles**
2. Use the **stoichiometric coefficients** from the **balanced** reaction to **find the number of moles of the unknown** you are being asked for
3. Convert the number of moles of your unknown to a mass, or whatever quantity you are being asked for

Example

2.6g of sodium metal (Na) reacts with water to form NaOH and H_2 according to the unbalanced reaction below. Once the reaction is complete how many grams of NaOH are formed?

$$Na_{(s)} + \ H_2O_{(g)} \rightarrow \ NaOH_{(s)} + \ H_{2(g)}$$

WIZE TIP

On an exam, your prof will NOT specify if an equation is balanced or unbalanced.

Always double check that the equation is balanced. If it's not, balance the equation before continuing!

The equation must be balanced in order to get the correct answer!

Step 1: Balance the reaction

$$2\ Na_{(s)} + 2\ H_2O_{(g)} \rightarrow 2\ NaOH_{(s)} + H_{2(g)}$$

Step 2: Find the number of moles of sodium

$$n_{Na} = \frac{m}{MW} = \frac{2.6\ g}{22.990\ g/mol} = 0.113\ mol$$

Step 3: Find the number of moles of NaOH

$$n_{NaOH} = n_{Na} \times \frac{coefficient\ NaOH}{coefficient\ Na} = 0.113\ mol \times \frac{2}{2} = 0.113\ mol$$

Step 4: Convert the number of moles of NaOH to a mass

$$n = \frac{m}{M}$$

$$m_{NaOH} = n \times MW = 0.113\ mol \times 39.997\ g/mol = 4.520\ g = 4.5\ g$$

1.11.3

Example: Converting Mass of Reactant to Mass of Product

Calculate the mass of hydrogen gas produced if 0.550 g of iron powder is reacted with an excess amount of sulfuric acid.

$$Fe\ (s) \quad + \quad H_2SO_4\ (aq) \quad \rightarrow \quad Fe_2\left(SO_4\right)_3\ (aq) \quad + \quad H_2\ (g)$$

1.11.4

Example: Solution Stoichiometry to Determine Mass of Product

Lithium metal was added to a 25mL of a 1.3M solution of Ag_2SO_4. The unbalanced chemical reaction is shown below.

$$Ag_2SO_{4(aq)} + Li_{(s)} \rightarrow Li_2SO_{4(aq)} + Ag_{(s)}$$

Once the reaction has gone to completion, what mass of silver metal is produced?

a) 0.0070 g
b) 7.0 g
c) 3.5 g
d) 0.0035 g

Practice: Calculate the Mass of Product

16.0 mL of Hg (mp = -38.8^{o}C, density = 13.56 g/mL) reacts with fluorine gas and the reaction goes to completion as shown below.

$$Hg_{(l)} + F_{2(g)} \rightarrow HgF_{2(s)}$$

Once the reaction has gone to completion, what mass of mercury fluoride, HgF_2, is produced?

A) 1.48 g

B) 514 g

C) 142 g

D) 258 g

1.11.6

Practice: Stoichiometry Calculations Application

A solution that is 183mM (millimolar) NaCl(aq) is isoosmotic with plasma. This means that cells don't swell or shrink when in this solution. How many grams of sodium chloride are required to make 150mL of isoosmotic NaCl(aq)? The molar mass of NaCl is 58.44g/mol.

- 0.2g
- 1.6g
- 2.2g
- 4.9g

1.12 Limiting Reagents

1.12.1

Introduction to Limiting Reagents

Anytime **reactant species are in limited supply** and not present in perfectly proportional amounts, a chemical reaction will have a **limiting reagent** which will be totally **consumed before any other reactant**

The quantity of the limiting reagent available directly **determines the maximum number of product molecules** that can be formed!

Let's Consider an Example:

When making smores (yum!!) the "reaction" looks something like:

2 graham crackers + 1 piece of chocolate + 1 marshmellow → 1 smore

If I had 10 graham crackers, 6 pieces of chocolate, and 6 marshmellows, what would be the limiting reagent? In other words, what would I run out of first? And how many smores could I make?

- The "chemistry way" to figure this out would be to **take the # of moles of each reactant and divide by its stoichiometric coefficient:**

Graham crackers: 10/2=5
Pieces of chocolate: 6/1=6
Marshmellows: 6/1=6

- Now, **to figure out the limiting reagent, look at which of the above numbers are the smallest!!**
    - 5 is smallest, therefore Graham crackers are the limiting reagent!!

Now how many smores could we make?

We know graham crackers will determine how much product we get since we'll run out of the crackers first.

$$moles\ of\ graham\ cracker \times \frac{1\ mol\ smores}{2\ moles\ graham\ cracker} = moles\ of\ smores$$

$10 \times (1/2) = 5$ moles of smores created!

WIZE CONCEPT

It is necessary to **determine the limiting reagent** whenever we are **given the amounts of two or more reactants** in a chemical reaction.

The limiting reagent is **totally consumed in the reaction**. **Any other reactant is in excess.**

Limiting reagent stops the reaction by **running out first.**

The **quantity of the limiting reagent** available directly **determines the maximum number of products** molecules that can be formed

1.12.2

Example: Determine the Mass of Product in a Limiting Reagent Problem

Iron and chlorine gas react to form iron (III) trichloride. If 110 g of iron and 105 g of chlorine gas are reacted, which species is the limiting reagent? What is the maximum mass of $FeCl_3$ that can be formed?

$$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$$

WIZE TIP

Steps for Solving Limiting Reagent Problems:

Step 1 – Write & **balance** the equation.

Step 2 – Calculate **moles of each reactant.**

Step 3 – Use the molar ratios of the present reactants (from the balanced equation) to determine the limiting reactant (LR).

- Take the **# of moles of each reactant** from step 2 and **divide by the stoichiometric coefficient** for that reactant
  - **Smallest value** from this calculation tells us the **LR**

Step 4 – From the limiting reactant, use the **molar ratio** (from the balanced equation) to **calculate moles of the desired product.**

Step 5 – Convert moles to the desired units (density, molarity, grams, etc...)

$$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$$

Step 1 – Write & balance the equation.

Is the equation given balanced? ______

Step 2 – Calculate moles of each reactant.

i) Find moles of Fe

ii) Find moles of Cl_2

Step 3 – Use the molar ratios of the present reactants (from the balanced equation) to determine the limiting reactant.

Therefore the limiting reagent is ______ ! (because it had the smallest number)

Step 4 – From the limiting reactant, use the molar ratio (from the balanced equation) to calculated moles of the desired product.

- The maximum mass of $FeCl_3$ that is formed depends on how much ______ we have (LR).

Step 5 – Convert moles to the desired units (density, molarity, grams, etc...)

1.12.3

Example: What is the Limiting Reagent?

2.0 g of aqueous Barium hydroxide ($Ba(OH)_2(aq)$, 171.32 g/mol) and 1.5 g of liquid hydrogen bromide ($HBr(l)$, 80.91g/mol) react together to give aqueous barium bromide ($BaBr_2(aq)$) and liquid water ($H_2O(l)$). Which species would be the limiting reagent of this reaction?

1.12.4

Example: How Many Grams of the Excess Reagent Will Remain?

The reaction between P_4 and Br_2 is very exothermic and produces PBr_5 as the only product. If 7.0 g of P_4 react with 12.0 g of Br_2 how many grams of the excess reagent will remain?

a) 0.4 g
b) 4.7 g
c) 6.1g
d)11.1 g

1.12.5

Practice: What is the Limiting Reagent

If 12 g of sodium is mixed with 0.64 g of H_2 and 17 g of O_2 which reagent will be limiting in the reaction shown below?

$$2\ Na_{(s)} + O_{2(g)} + H_{2(g)} \rightarrow 2\ NaOH$$

A) Na

B) O_2

C) H_2

D) None, they will all be completely consumed

1.13 Percent Yield

1.13.1

Percent Yield

So far we have been assuming that the reaction proceeds 100% to products. In reality, this is rarely the case!

When we are doing experiments in the lab we often lose some of our product by spilling it or residues of it get left on glassware and spatulas.

So far we have been calculating the **theoretical yield.**

Theoretical Yield - the maximum amount of product produced based on the quantity of the limiting reagent (*e.g.* the amount of product produced in a calculation).

Actual Yield - the actual amount of product produced in the reaction (e.g. the amount of product obtained in a laboratory experiment).

$$\boxed{\%\ yield = \frac{actual\ yield}{theoretical\ yield}\ x100}$$

Test Your Understanding

1) In a lab, you measure your product and record 5.222g. This is the ________ yield.

2) If everything went 100% perfectly according to the reaction we would get the ____________ yield

3) If I determined the limiting reagent, and then calculated the number of moles of product and the mass of product based on this, I've just calculated the ____________ yield

Example: What is the Percent Yield

When 49.00 g of a hydrocarbon fuel with formula $C_7H_{10}O_2$ is reacted with excess oxygen, a total of 21.56 g of water is collected. What was the percent yield of the reaction?

$$\_\_\_\_ C_7H_{10}O_2 + \_\_\_\_ O_2 \rightarrow \_\_\_\_ H_2O + \_\_\_\_ CO_2$$

Practice: Calculate Moles Given Percent Yield

If the yield for the following reaction is 45%, how many moles of $KClO_3$ are needed to produce 1 mol of O_2?

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

A) 0.4 moles

B) 1.1 moles

C) 1.5 moles

D) 4.6 moles

Practice: Limiting Reagent & Percent Yield

Iron oxidizes when exposed to oxygen to form iron oxide according to the following equation:

$$Fe + O_2 \rightarrow Fe_2O_3$$

352g of pure iron is exposed to 12.0 mols of O_2, and after a period of time 46.7 g of iron oxide (rust) is collected.

Part 1

What is the limiting reagent in this reaction?

- A) Fe
- B) O_2
- C) Fe_2O_3
- D) There is no limiting reagent

Practice: Limiting Reagent & Percent Yield

Iron oxidizes when exposed to oxygen to form iron oxide according to the following equation:

$$Fe + O_2 \rightarrow Fe_2O_3$$

352g of pure iron is exposed to 12.0 mols of O_2, and after a period of time 46.7 g of iron oxide (rust) is collected.

Part 2

What is the percent yield?

- A) 2.2%
- B) 8.8%
- C) 9.4%
- D) 33.1 %

Practice: Limiting Reagent & Percent Yield

Iron oxidizes when exposed to oxygen to form iron oxide according to the following equation:

$$Fe + O_2 \rightarrow Fe_2O_3$$

352g of pure iron is exposed to 12.0 mols of O_2, and after a period of time 46.7 g of iron oxide (rust) is collected.

Part 3

How many Fe atoms are present in the rust produced?

A) 0.89×10^{17} atoms

B) 1.77×10^{23} atoms

C) 3.56×10^{23} atoms

D) 4.33×10^{23} atoms

1.14 Ideal Gas Law

1.14.1

The Ideal Gas Law

The equation below **describes ideal gases**. Real gases act differently, but act most similarly to ideal gases at **high temperatures and low pressures!**

$$\boxed{PV = nRT}$$

P *is* ***pressure*** *measured in kPa (or maybe atm Torr, mmHg...)*
V *is* ***volume*** *measured in* ***L*** *or m^3*
n *is* ***number of moles*** *of the gas*
R *is the* ***ideal gas constant*** *(will be either* ***8.314J/mol K*** *OR* ***0.082 L atm/mol K****)*
T *is* ***temperature*** *measured in* ***K***

WIZE TIP

Pressure=Force/unit area=N/m^2=1 Pa

How to convert different units of pressure:

$$\boxed{101\ 300\ Pa = 101.3\ kPa = 1\ bar = 1\ atm = 760\ mmHg = 760\ torr}$$

We mentioned that R (the ideal gas constant) could be either 8.314J/mol K or 0.082 L atm/mol K. **How do we know which value to use?**

WATCH OUT!

Using the correct value for R is more important than you might think! Students often make the mistake of using the wrong value of R on an exam and they get the question wrong. Let's see how we can easily prevent that :)

The ideal gas constant appears frequently in chemistry, for this reason, it comes in many forms depending on the context of its use and the **units involved:**

Common Values for R
8.3145 kPa · L / mol · K
8.3145 N · m / mol · K
8.3145 J / mol · K
8.3145 kg · m2/s2 mol · K
0.08206 atm · L / mol · K

Example:

Suppose the variables in the ideal gas law equation had the following units:

$$P(kPa) \times V(L) = n(mol) \times R(?) \times T(K)$$

What should R be in this case? (These terms need to combine correctly to ensure that this equation makes sense!)

1.14.2

Example: Ideal Gas Law

1.1 mols of Argon are stored in a 2.0 L container kept at 10.0 °C. What is the pressure of this gas? (ans. in kPa)

1.14.3

Example: Ideal Gas Law

a) A 4 L cylinder containing 3 moles of He(g) arrives from a chemical supply company. Given that the lab you are working in has a room temperature of 298 K, calculate the pressure inside the flask.

b) At what temperature will the cylinder have an internal pressure of 9 atm?

c) Your lab mate steals some of the gas from your cylinder for a reaction he is doing. He fills a 1 L bulb with 1 atm of your gas at 298 K. If the cylinder is at room temperature, what is the internal pressure of the cylinder once the gas has been removed?

1.14.4

Practice: Stoichiometry and the Ideal Gas Law

The combustion of methanol (CH_3OH) is shown below. If 16g of methanol is burned in oxygen, what volume of CO_2 is produced if the pressure is 101.3kPa, T=25°C, and R=8.314 kPa L/mol K? Round to the nearest whole number.

$$CH_3OH \quad + O_2 \rightarrow CO_2 \quad + \quad H_2O$$

- 6L
- 12L
- 24L
- 27L

1.15 Other Applications of the Ideal Gas Law

1.15.1

Applications of the Ideal Gas Law

Determining the Molecular Weight (M) of an Unknown Gas

$$PV = nRT \; plug \; in \; n = \frac{m}{M}$$

$$PV = \frac{m}{M}RT$$

where, m=mass (g) and M=molar mass (g/mol)

Rearrange the equation to solve for M (molecular weight) of the unknown gas:

$$M = \frac{mRT}{PV}$$

Finding the Density of a Gas

The equation for density (δ) is:

$$\boxed{\delta = \frac{m}{V}}$$

We can rearrange the ideal gas law to find density:

$$PV = nRT \; plug \; in \; n = \frac{m}{M}$$

$$PV = \frac{mRT}{M}$$

Solve for density:

$$\frac{m}{V} = \delta = \frac{MP}{RT}$$

1.15.2

Standard Molar Volume

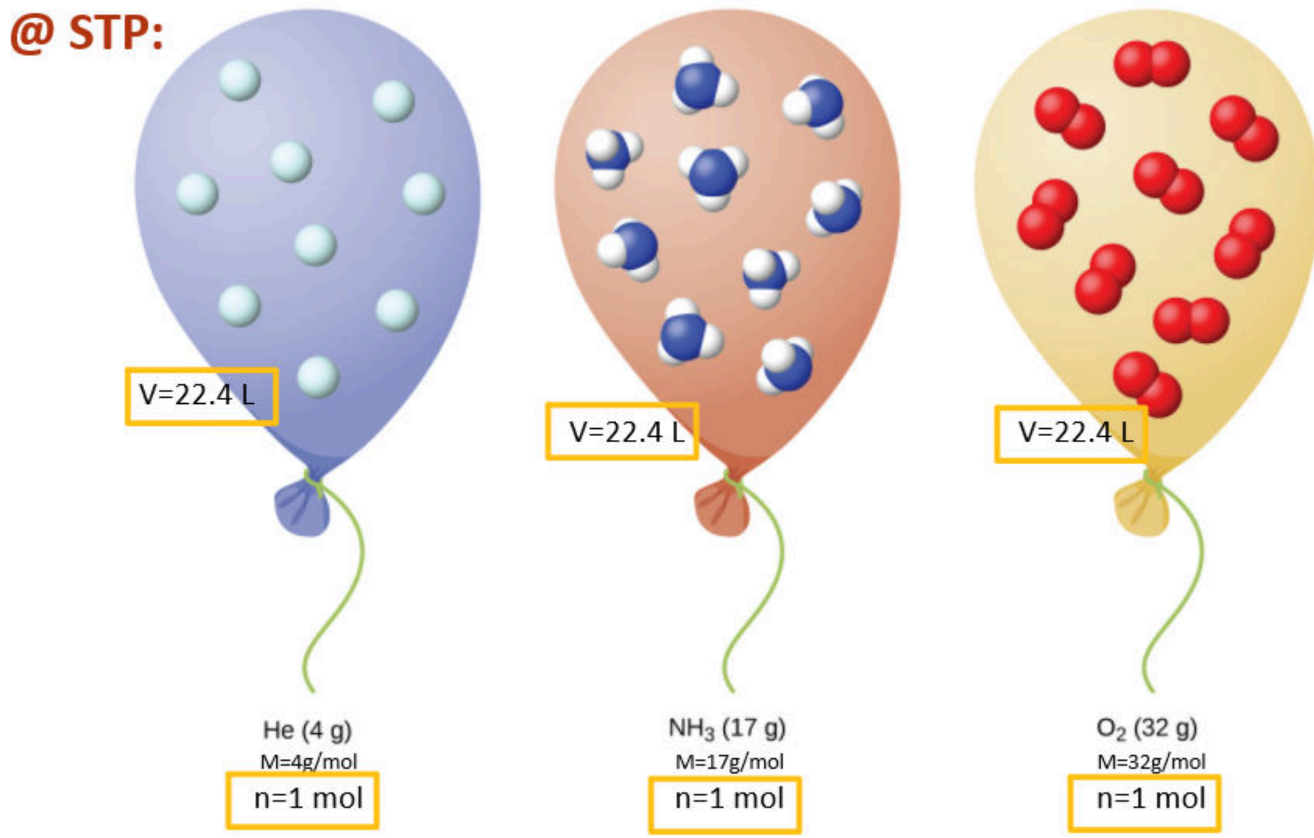
Recall: we talked about Avogadro's law where P and T are kept constant.
PV=nRT
V/n=RT/P where RT/P=constant
V/n=constant

Based on this, imagine if we were looking at two different gases in two different containers, each with the same P, T, and number of moles of gas particles.

We have made n (moles) a constant too. If that is the case, then the volume for all the gases should be the same, and they are!

WIZE CONCEPT

Specifically **at STP** **(P=1atm and T=0°C or 273K), 1 mol of any gas occupies 22.4L! **Memorize this!****
STP=Standard Temperature and Pressure


@ STP:
V=22.4 L
V=22.4 L
V=22.4 L
He (4 g)
M=4g/mol
n=1 mol
NH_3 (17 g)
M=17g/mol
n=1 mol
O_2 (32 g)
M=32g/mol
n=1 mol

1.15.3

Example: Standard Molar Volume

In the following reaction, if you start with 4L H_2(g) and 1.5L N_2(g) at STP, what would the volume of each of the 3 gases be when the reaction is complete?

$$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$$

1.15.4

Practice: Solving for Molar Mass of a Gas

The density of an unknown gas at 100°C and 746 torr is 1.994 g/L. What is the molecular mass of this unknown gas?

- 8.3 g/mol
- 31.0 g/mol
- 62.1 g/mol
- 124.2 g/mol
- 1689 g/mol

1.15.5

Practice: Determine the Gas

0.238 g of a gas were placed in a 250 mL vessel. If the gas is heated to 85 ^{o}C, the gas has a pressure of 7 atm. Which gas is in the vessel?

- H_2 (g)
- He (g)
- N_2 (g)
- F_2 (g)

1.15.6

Practice: Solve for Molecular Weight

A 6.00 L flask is evacuated and weighed, 21.64g. The vessel is then filled with a gas at STP, after filling the vessel weighs 27.31g. What is the molecular weight of the gas?

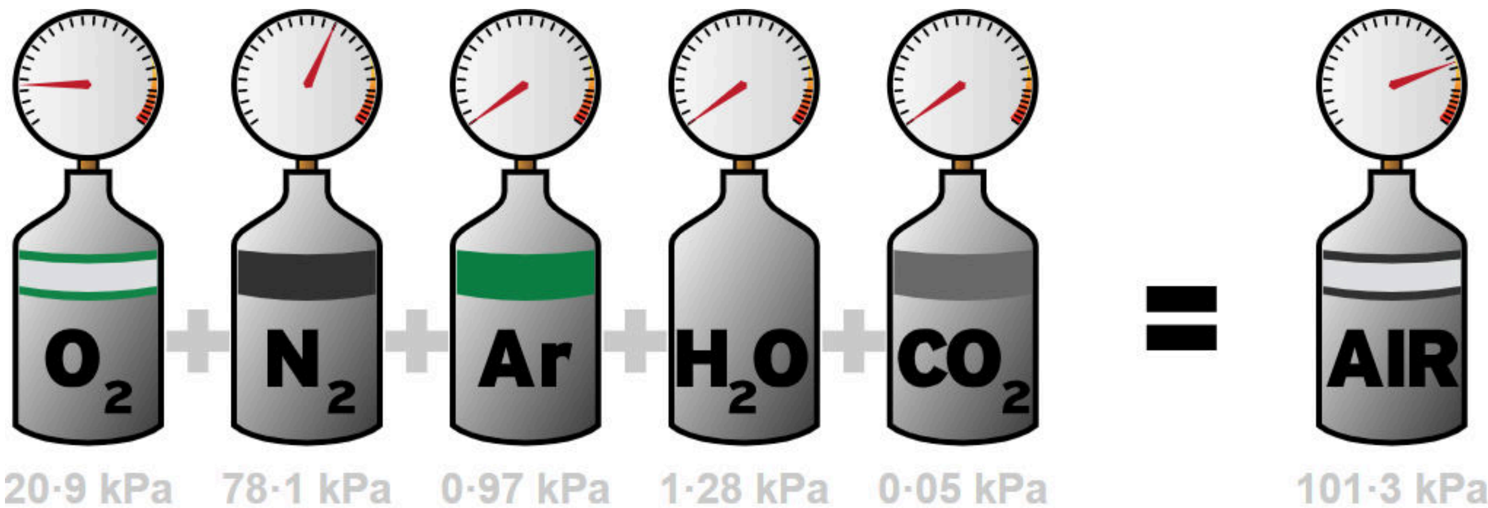
- 0.2446 g/mol
- 1.0 g/mol
- 2.446 g/mol
- 21.46 g/mol
- 227.2 g/mol

1.16 Gas Mixtures & Partial Pressures

1.16.1

Gas Mixtures & Partial Pressures

When dealing with mixtures of gases, the total pressure of the gases is equal to the sum of the **partial pressure** of each component.



Dalton's Law of Partial Pressures:

$$\boxed{P_{Total} = P_1 + P_2 + P_3 ... = \Sigma P_i}$$

P_{Total} is the single **measured pressure of the mixture**

P_1 is the **partial pressure of gas 1**, and so on

Partial pressure: is the pressure that would be exerted by one of the gases in the mixture if it occupied the same volume on its own.

Note: The total pressure depends on the total number of moles (not on the identity of the gases)

$PV = nRT$

$P_{total} = \frac{n_{total}RT}{V}$

We can also relate the partial pressure of an individual gas (ex. Gas "A") to the total pressure by the **mole fraction (X)** of that gas.
The **mole fraction (X)** is a **proportion** of that gas.

$$\boxed{P_A = \frac{n_A}{n_{Total}} P_{Total} = X_A P_{Total}}$$

1.16.2

Example: Partial Pressures

If we have a container with 75% nitrogen, 20% oxygen, and 0.04% carbon dioxide, the total pressure in the container is measured to be 760 Torr (recall this is equal to 1 atm!). What is the partial pressure of oxygen gas? (Ans in torr)

1.16.3

Practice: Partial Pressures

A vessel contains He, Ne, and Ar gas. There's twice as much He as Ne and half as much Ar as He. The total pressure in the vessel is 400torr. What is the partial pressure of Ar(g)?

- 75 torr
- 100 torr
- 175 torr
- 225 torr

1.16.4

Practice: Partial Pressures

Two gases were mixed together in a 5 L vessel at 243 K to create a total pressure of 9.3 atm. If gas 1 has 2 moles, what is the partial pressure of gas 2? (R = 0.08206 L*atm/mole*K)

- 0.7 atm
- 1.3 atm
- 3.1 atm
- 5.4 atm