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12:12 PM

The Mole

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The Mole
Unit Conversions

It's all about laying out your problem in a stepwise, logical manner, so you can solve it. LEARN HOW!

Conversions in the every day

There are terms we use every day like "*dozen*," which you all know means 12, to describe how much of something you have. You can convert between different units as long as they measure the same thing (like feet, and meters both measure distance)

Ex. Timmy has 18 eggs, how many dozen eggs does he have?

This is how to set up a unit conversion problem properly!!!

KQ CF UQ

$$18 \text{ eggs} \times \frac{1 \text{ dozen}}{12 \text{ eggs}} = 1.5 \text{ dozen}$$

$\frac{12 \text{ eggs}}{1 \text{ dozen}}$ or $\frac{1 \text{ dozen}}{12 \text{ eggs}}$

Metric System conversions

The metric system is based around 10's (like scientific notation) with prefixes denoting the power on the 10. Common prefixes are centi (10^{-2}), milli(10^{-3}), kilo(10^{3}), micro(10^{-6}), mega(10^{6}), nano(10^{-9}), giga(10^{9}),

Ex. Convert 1.8 km into meters

$$1.8 \text{ km} \times \frac{1000 \text{ m}}{1 \text{ km}} = 1800 \text{ m}$$

Ex. Convert 1.8 km into centimeters

$$1.8 \text{ km} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{100 \text{ cm}}{1 \text{ m}} = 180{,}000 \text{ cm}$$

Ex. Convert 5.4 Gm into nm

$$5.4 \text{ Gm} \times \frac{10^{9} \text{ m}}{1 \text{ Gm}} \times \frac{1 \text{ nm}}{10^{-9} \text{ m}} = 10^{18} \text{ nm}$$

5.4×10^{18} nm

Or...

$$5.4 \text{ Gm} \times \frac{10^{9} \text{ m}}{1 \text{ Gm}} \times \frac{10^{9} \text{ nm}}{1 \text{ m}} = 5.4 \times 10^{18} \text{ nm}$$

Formula Masses

Chemical formulas represent the number and type of each atom present in a compound or element (recall law of definite proportions)

Ex. NaCl: 1 sodium atom and 1 chlorine atom
H_2SO_4: 2 hydrogen atoms, 1 sulphur atom, and 4 oxygen atoms

The *law of conservation of mass* states that mass is neither created nor destroyed in a chemical reaction. Therefore the mass of a formula unit can be determined by adding together the masses of each atom present.

The term "*formula mass*" applies to atoms, molecules and ionic compounds; "*molecular mass*" only applies to covalent compounds

Relative Atomic Masses

Relative mass: comparing the mass of one object to the mass of another

Ex. A certain number of oranges have a mass of 3000g and an equal number of grapefruits have a mass of 5000g. What fraction of the mass of the grapefruits do the oranges have?

$$\frac{3000\text{g}}{5000\text{g}} = 0.6 \quad \Leftarrow \text{ratio of orange to grapefruit mass } (\textbf{relative mass})$$

(Note that the mass of an individual orange or grapefruit is not known)

In the same way, scientists have determined the relative masses of atoms. The units used to measure mass of atoms is the **atomic mass unit** or **amu** which has the SI unit symbol **u**.

The carbon-12 atom is assigned a mass of **12u** so one **u** is equal to $\frac{1}{12}$ the mass of a single carbon-12 atom

→ HW? STUDY!

Calculating Formula Masses (a strategy)

1) Determine the number of atoms of each element present:

Ex. $Fe(CH_3COO)_3$

1 Fe   6 C   9 H   6 O

2) Look up the atomic masses on the periodic table (keep at least one decimal place):

Fe = 55.8u   C = 12.0u   H = 1.01u   O = 16.0u

3) Multiply the masses by the number of each atom present

Fe = 1 (55.8) = 55.8u
C = 6 (12.0) = 72.0u
H = 9 (1.01) = 9.09u
O = 6 (16.0) = 96.0u

4) Add the masses together:

55.8u + 72.0u + 9.09u + 96.0u = 232.9u

Thus the **formula mass** of $Fe(CH_3COO)_3$ is 232.9u

Try these: NaCl, CO_2, CaC_2O_4

NaCl = 23.0u + 35.5u = 58.5u

CO_2 = 12.0u + 2(16.0u) = 44.0u

CaC_2O_4 = 40.1u + 2(12.0u) + 4(16.0u) = 128.1u

Avogadro's Number and the Mole

The mole or 'moles' is just a way of measuring the amount of things, just like a dozen. Except where a dozen means 12 things, a mole means: N_A 602,000,000,000,000,000,000,000 (yes, 6.02×10^{23}) things!!!

An amu is far too small to measure in the lab, so chemicals are measured in grams. To be of use we need to know how many ${}^{12}_{6}C$ atoms have a mass of exactly 12 grams. This number was found to be about 6.02214×10^{23} atoms.

We call this number Avogadro's number (N_A) named after an Italian scientist.

So just like we can do a calculation to figure out how many dozen are 54 eggs. We can do calculations to figure out how many moles of an atom we have if we have… say 5.0 x 10^{22} atoms.

KQ CF UQ

$$5.0 \times 10^{22} \text{atoms} \times \frac{1 \text{ mole}}{6.022 \times 10^{23} \text{atoms}} = 0.083 \text{ moles}$$

If you look at the periodic table carbon has an atomic mass of 12.011u not 12 exactly. This is due to the presence of isotopes (recall from science 10)

Isotopes are atoms of the same element (same # protons) but with a difference in # neutrons.

12 x 0.9889 = 11.8668
+ 13 x 0.0111 0.1443
12.0111

Most elements occur naturally as mixtures of different isotopes.
Ex. Carbon contains 98.89% $^{12}_{6}C$ and 1.11% $^{13}_{6}C$ (contains an extra neutron)

Atomic mass is a **weighted average** based on abundance of each isotope.

Moles ↔ Molecules

small huge!

To convert from the number of moles of particles to the number of particles:

$$\text{number of moles} \times \frac{6.022 \times 10^{23} \text{atoms}}{1 \text{ mole}} = \text{number of particles}$$

Ex. How many sodium atoms are in 2.50 moles of *sodium?*

KQ CF UQ

$$2.50 \text{ mol} \times \frac{6.022 \times 10^{23} \text{atoms}}{1 \text{ mole}} = 1.51 \times 10^{24} \text{sodium atoms}$$

To convert from the number of particles to the number of moles of particles:

$$\text{particles} \times \frac{1 \text{ mole}}{6.022 \times 10^{23} \text{atoms}} = \text{moles}$$

Ex. How many moles of hydrogen molecules are there in 3.01 x 10^{24} molecules of H_2?

$$3.01 \times 10^{24} \text{molecules of } H_2 \times \frac{1 \text{ mole}}{6.022 \times 10^{23} \text{atoms}} = 5.00 \text{ moles of } H_2$$

particles

Molar Mass

One mole of particles has a mass in **grams** equal to the mass of one particle in **amu**

Ex. Carbon-12 = 12u, so one mole would be 12g
Oxygen = 32u, so one mole would be 32g
NaCl = 58.5u, so one mole is 58.5g etc…

The molar mass in grams has the same numerical value as the formula mass in amu, but now we're dealing with values we can measure in the lab

Molar mass can also be used as a unit conversion (show as a ratio)

If NaCl has a mass of 58.5g per mole we can write it as: $\frac{58.5\text{g}}{1\text{ mole}}$ or $\frac{1\text{ mole}}{58.5\text{ g}}$

Ex. How many grams would 0.250 mol of water weigh?

$mol^{-1} = \frac{1}{mol}$

1) calculate the molar mass of H_2O: 2(1.0) + 16.0 = 18.0g·mol^{-1}
2) multiply by number of moles:

$$0.250\text{ moles of }H_2O \times \frac{18.0\text{ g}}{1\text{ mole}} = 4.50\text{g of }H_2O$$

Ex how many moles of CO_2 is 75g?

$$75g \times \frac{1\text{ mol}}{44.0g} = 1.7\text{ mol}$$

Ex how many grams is 0.30 mol of oxygen gas? O_2

$$0.30\text{mol} \times \frac{32.0g}{1\text{ mol}} = 9.6\text{ g}$$

Hw # 6-10, 13-14 (pg. 77-83)

Percent Composition and Empirical Formulae

Percent composition gives you the percentage amount of an element (or both elements) in a compound

Ex. Percent composition of a sample of containing 4.00g of hydrogen and 32.0g of oxygen

$$\% \text{ Oxygen } \times \frac{32.0\text{ g}}{32.0\text{g} + 4.00\text{g}} \times 100\% = 88.9\% \text{ O}$$

$$\% \text{ Hydrogen } \times \frac{4.00\text{ g}}{32.0\text{g} + 4.00\text{g}} \times 100\% = 11.1\% \text{ H}$$

Ex. Percent composition of sulphuric acid (H_2SO_4) (Try this one!)

98.1g/mol

$2(1.01) + 32.1 + 4(16.0) = 98.1$ g/mol

$$\% \text{ Sulphur } \times \frac{32.1\text{ g}}{98.1\text{g}} \times 100\% = 32.7\% \text{ S}$$

$$\% \text{ Hydrogen } \times \frac{2.02\text{ g}}{98.1\text{g}} \times 100\% = 2.0\% \text{ H}$$

$$\% \text{ Oxygen } \times \frac{64.0\text{ g}}{98.1\text{g}} \times 100\% = 65.2\% \text{ O}$$

Calculate % Composition of K, Al, S, O in $KAl(SO_4)_2$ ... thanks :)

258.3g/mol

$$K = \frac{39.1}{258.3} \times 100\% = 15.1\%$$

$$Al = \frac{27.0}{258.3} \times 100\% = 10.5\%$$

$$S = \frac{64.2}{258.3} \times 100\% = 24.9\%$$

$$O = \frac{128}{258.3} \times 100\% = 49.6\%$$

Empirical formula gives you the lowest possible ratio of elements in a compound. For example $C_6H_{12}O_6$ (glucose) would have an empirical formula of CH_2O, even though it's 6 times that in reality.

Ex. Some compound has 50.4g of nitrogen, and 115.8g of oxygen (from elemental analysis):

$$\text{moles of N} = 50.4\text{g} \times \frac{1\text{ mole}}{14.0\text{g}} = 3.60\text{ mol N}$$

$$\text{moles of O} = 115.8\text{g} \times \frac{1\text{ mole}}{16.0\text{g}} = 7.24\text{ mol O}$$

Ratio of moles of N : moles of O(put the smaller # moles on the bottom)

$$= \frac{3.60\text{ mol N}}{3.60\text{ mol N}} : \frac{7.24\text{ mol O}}{3.60\text{ mol N}} = 1.00 : 2.01 \rightarrow \mathbf{1:2}$$

So the empirical formula is $\mathbf{NO_2}$ (which has a molar mass of **46.0g**) ← empirical mass

Molecular formula has to be a multiple of the empirical formula: $(NO_2)_n$

Where $n = \dfrac{\text{molar mass of compound (has to be given – from mass spec)}}{\text{molar mass of empirical formula}}$ (in Q.)

$n = \dfrac{92.0\text{g (given value)}}{46.0\text{g (calculated)}} = 2$, so this compound is N_2O_4 $(NO_2)_2$

Ex. Empirical formula of a compound with 58.5% C, 7.4% H and 34.1% N

TRICK: Imagine you have 100.0g of this compound, then there would be 58.5g C, 7.4g H and 34.1g N

	Divide by 2.44 (lowest)
$\text{moles of C} = 58.5\text{g} \times \frac{1\text{ mole}}{12.0\text{g}} = 4.88\text{ mol C}$	2.00
$\text{moles of H} = 7.4\text{g} \times \frac{1\text{ mole}}{1.01\text{g}} = 7.4\text{ mol H}$	3.0
$\text{moles of N} = 34.1\text{g} \times \frac{1\text{ mole}}{14.0\text{g}} = 2.44\text{ mol N}$ *	1.00

ratio is 2C's to 3H's to 1N, or $\mathbf{C_2H_3N}$

HW Q's 44-46 pg's 91-93

Prep lab 4B

Ex. Empirical formula of Rust

Mass of rust analyzed 15.53g
Mass of iron in sample - 10.87g
Mass of oxygen = 4.66g

	Divide by 0.195 mol
$\text{moles of Fe} = 10.87\text{g} \times \frac{1\text{ mole}}{55.8\text{g}} = 0.195\text{ mol Fe}$	1.00
$\text{moles of O} = 4.66\text{g} \times \frac{1\text{ mole}}{16.0\text{g}} = 0.291\text{ mol O}$	1.49

Ratio of Fe:O approximately 1:1.5... can't have half an atom, so multiply by 2, and we get 2:3 (**a whole number ratio**) 2 Fe's and 3 O's

Empirical formula for Rust: Fe_2O_3

HW Q's 47-55 pg 95
AND Lab 4B!

Avogadro's Hypothesis and STP

STP: Standard Temperature and Pressure

- Gas volumes vary with temperature (PV = nRT)
- Standard conditions chosen so gases can be compared
- Standard T = 0°C (or 273 K)
- Standard P = 1 atm (101.3 kPa)

History of gas and the mole

Early 19th century: Chemists reacted gases together and compared ratios of gas volumes and masses
Dalton: Atoms react in fixed simple whole number ratios
Gay-Lussac: Law of combining volumes: gases combine in volumes of fixed whole number ratios
Avogadro: saw that the atom and gas-volume ratios were identical; and that some gases consisted of atoms and some were molecules

★**Avogadro's hypothesis (1811):** equal volumes of different gases at the same T and P contain the same number of particles! ← ∴ same # of moles

1 mole of a gas (any gas, but ONLY gases!) has a volume of 22.4L at STP

Thus the molar mass of any unknown substance (converted to a gas) can be determined

22.4 L/mol

Recall that several elements exist as diatomic molecules: H O, N, F, Cl, Br, I so that 22.4L of oxygen at STP has 6.02 x 10^{23} molecules, but 1.20 x 10^{24} atoms of O

Ex. A 34.0g sample of an unknown gas has a volume of 12.0L at STP. What is its ~~molecular~~ molar mass?

$$12.0\text{L} \times \frac{1\text{ mole}}{22.4\text{L}} = 0.536\text{ mol} \quad \text{and} \quad \frac{34.0\text{g}}{0.536\text{ mol}} = 63.4\text{g/mol}$$

Ex. What is the mass of 3.5L of NO_2 gas (at STP)?

MW of NO_2 = 14.0 + 2(16.0) = 46.0 g/mol

$$3.5\text{ L} \times \frac{1\text{ mole}}{22.4\text{ L}} \times \frac{46.0\text{g}}{\text{mol}} = 7.2\text{g of NO}_2$$

Ex. What mass does a molar volume of CO_2 gas have at 100°C?

← whatever volume, = 1 mol

The mass of 1 molar volume of any gas is it's molar mass, so for CO_2 it's just: 12.0 + 2(16.0) = 44.0 g/mol = **44.0g**

LAB 4B & Q's 15-17 pg 84
23-24 pg 87

Moles and Solutions

Molarity

The number of moles in a given volume of solution describes the **concentration** of that solution.

The molar concentration we call the **Molarity (M)** of a solution. It is measured by the number of dissolved moles of solute per liter of solution.

$M = \frac{n}{V}$ Where, n = # of moles

V = volume of solution

M = molarity (concentration)

The shorthand symbol for "molar concentration of X" is a set of square brackets: [X]

[NaCl] ← "the concentration of Sodium chloride"

Ex. What is the molarity of a solution made by dissolving 3.0 mol of salt (NaCl) in 1.5 L of water

$$M = \frac{n}{V} = \frac{3.0\ \text{mol}}{1.5\ \text{L}} = 2.0\frac{\text{mol}}{\text{L}}\ \text{or}\ 2.0\ \text{M}$$

[NaCl] = 2.0 M

Technique for preparing a solution

1) Calculate the mass of solid required
2) Weigh out on a balance
3) Transfer the solid *quantitatively* (all of it!) to a volumetric flask
4) Add water to dissolve, swirl until mostly dissolved
5) add water up to mark and mix by inversion

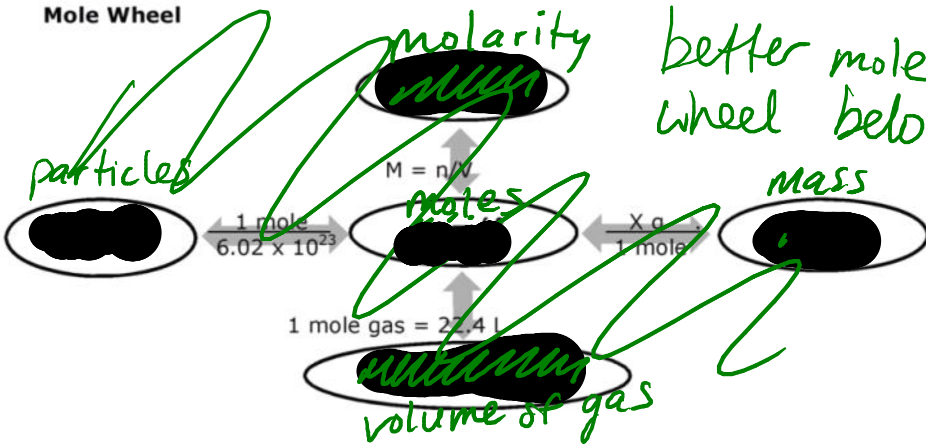
Ex. Describe how you would prepare 1.00L of a 0.100M solution of NaCl

$M = \frac{n}{V}$ so n = M x V = (0.100 M)(1.00L) = 0.100 mol NaCl

$$0.100\ \text{mol NaCl} \times \frac{58.5\ \text{g}}{1\ \text{mol}} = 5.85\ \text{g NaCl}$$

Weigh out 5.85g of NaCl on a balance, quantitatively transfer the solid to a volumetric flask and dissolve with sufficient water, add enough water to make up to the line on the volumetric flask. Done!

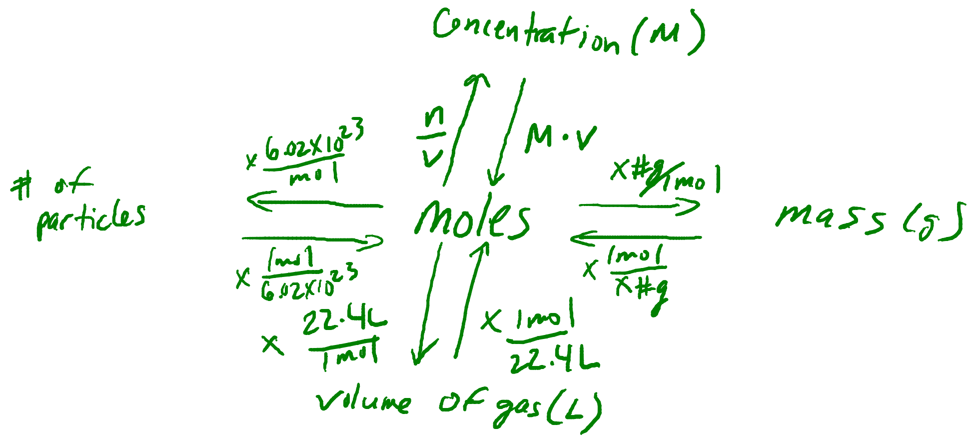
Mole Wheel



better mole wheel below!

Hw: pg 103/104 #95-101
(102) optional

Prep lab 7B


Concentration (M)
n/V
M·V
# of particles
x 6.02x10^23 / mol
x 1mol / 6.02x10^23
moles
x #g / 1mol
x 1mol / x #g
mass (g)
x 22.4L / 1mol
x 1mol / 22.4L
volume of gas (L)