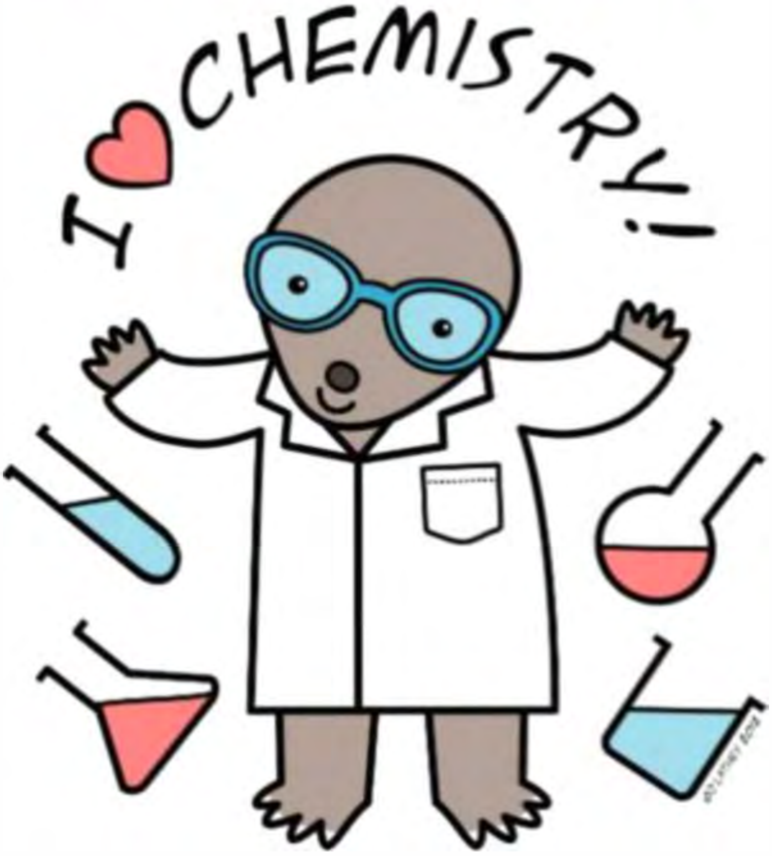
CHEMISTRY 11

UNIT 4: THE MOLE



BOOK 2: FORMULAE, HYDRATES & COMPLEX DILUTIONS

Name: Key

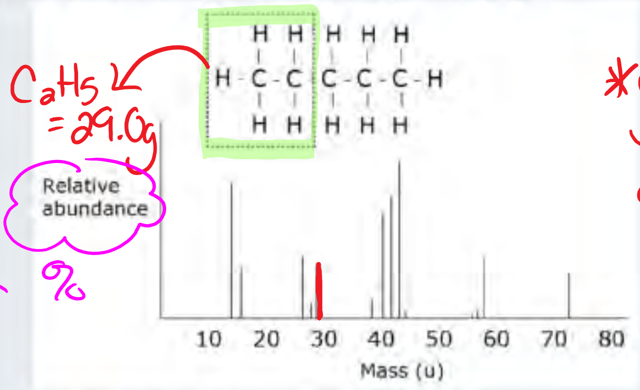
Block: ______

Composition Analysis — Determining Formulas

Fun Facts:

Forensic investigators collect samples from crime scenes. How do technicians identify the unknown samples? An instrument called a **mass spectrometer** can identify the vast majority of compounds. Each compound has a unique mass spectrum; much like each person has a unique fingerprint. A mass spectrometer breaks most of the molecules into fragments. In so doing, it creates a variety of particles from individual atoms to the intact molecule itself, and then marks the mass of each of these particles along a graph's horizontal axis. **The height of the line in the spectrum indicates the relative abundance of that particle.**

Below is a simplified mass spectrum of a compound called pentane (C_5H_{12}).



Mass spectrum analysis of pentane

*used to determine the composition of unknown samples.

Part A: Percent Composition

Percent Composition is the percentage of a compounds mass contributed by each type of atom in the compound.

A compound's percentage composition can be determined theoretically from its chemical formula.

EXAMPLE: **What is the percentage composition of CH_4?** (methane)

C: 12.0g H: 4×1.0g = 4.0g.

Assume there is 1 mol of the compound. molar mass = 16.0g

total mass of C in compound = 12.0g

total mass of H in compound = 4.0g

% of C in compound = $\frac{12.0g}{16.0g} \times 100\% = 75.0\%$ (3 sf)

% of H in compound = $\frac{4.0g}{16.0g} \times 100\% = 25.0\%$

should add to 100%

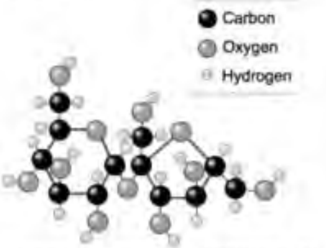
(*always calculate, don't sub. from 100%)

Sample Problem — Determining Percentage Composition

What is the percentage composition of a sugar with the formula $C_{12}H_{22}O_{11}$? ← assume 1 mole

What to Think about

1. Calculate the sugar's molar mass.
2. Thus one mole of this sugar contains 144 g C, 22 g H, and 176 g O.
3. Express each element's percentage of the molar mass.



A sugar molecule with 12 carbon atoms, 22 hydrogen atoms, and 11 oxygen atoms.

How to Do It

Total mass of C = 12 × (12.0g) = 144g C
Total mass of H = 22 × (1.0g) = 22g H
Total mass of O = 11 × (16.0g) = 176g O

(determine sig figs.)

$C_{12}H_{22}O_{11}$ = 342 g/mol

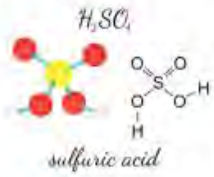
$\% C = \frac{144g\ C}{342g\ C_{12}H_{22}O_{11}} \times 100 = 42.1\%$ (3 sf.)

$\% H = \frac{22g}{342} \times 100 = 6.4\%$ (2 sf.)

$\% O = \frac{176g}{342} \times 100 = 51.5\%$

PERCENT COMPOSITION = $\frac{\text{(TOTAL MASS OF ELEMENT PRESENT)}}{\text{(MOLECULAR MASS)}} \times 100$

EXAMPLE: What is the percentage composition of H_2SO_4?



Assume there is 1 mol of the compound. molar mass = 98.1g
total mass of H in compound = 2 × 1.0g = 2.0g
total mass of S in compound = 1 × 32.1g = 32.1g
total mass of O in compound = 4 × 16.0g = 64.0g

% of H in compound = $\frac{2.0g}{98.1g} \times 100 = 2.0\%$ H

% of S in compound = $\frac{32.1g}{98.1g} \times 100 = 32.7\%$

% of O in compound = $\frac{64.0g}{98.1g} \times 100 = 65.2\%$

(% of each element in the compound.)

EXAMPLE: What is the percentage of water in $CuSO_4 \cdot 5H_2O$?

$CuSO_4 \cdot 5H_2O$

Assume there is 1 mol of the compound. molar mass = 249.6g of $CuSO_4 \cdot 5H_2O$
total mass of H_2O in compound = 5×18.0g = 90.0g of H_2O
% of H_2O in molecule =

$\frac{90.0g\ H_2O}{249.6g\ CuSO_4 \cdot 5H_2O} \times 100 = 36.1\%\ H_2O$ (3 sf.)

chemistry homework

Assignment #8- Hebden Questions page 91
#44-45 (every 2nd letter)
Complete ALL assignments on a separate piece of paper with a clear heading and title

Empirical, Molecular & Structural Formulas

Every molecular compound has three formulas; an **empirical formula**, a **molecular formula,** and a **structural formula**.

- The **empirical formula** is the simplest ratio of the different types of atoms in the compound.
- The **molecular formula** is the actual number of each type of atom in each molecule of the compound.
- The **structural formula** shows how the atoms in a molecule are arranged. It is a diagram that shows the pattern of the atoms' connections. (more common in organic chem.)

Glucose is an organic compound with a molecular formula of $C_6H_{12}O_6$.
The subscripts 6, 12, 6 ÷6 can be reduced (simplified) to 1, 2, 1. We don't show the number 1 as a subscript in a formula so the **empirical formula of glucose is** (simplified ratio) CH_2O

Many compounds have the same empirical formula but ***different molecular formulas****. Their molecular formulas all reduce to the same ratio. For example, all alkenes such as ethene (C_2H_4), propene (C_3H_6), and butene (C_4H_8), have an empirical formula of CH_2 because each of their molecular formulas can be reduced to a 1 to 2 ratio.*

PRACTICE

1. Complete the following table.

represents → simplified ratio

Structural Formula	Molecular Formula	Empirical Formula
H-C(H)(H)-C(=O)-O-H (double bond)	$C_2H_4O_2$ ÷2	CH_2O
H-O-C(=O)-C(=O)-O-H	$C_2H_2O_4$ ÷4	CHO_2

The **Empirical formula** is sometimes called the **simplest formula** and is the smallest **WHOLE** number ratio of atoms which represents the molecular composition of a chemical species.

EXAMPLE:

Finding the empirical formula is essentially the opposite procedure to determining the percentage composition of a compound.

Part B: Determining an Empirical Formula from Percent Composition

EXAMPLE: **What is the empirical formula of a compound consisting of 80.0% C and 20.0% H?**

Note that neither the chemical formula nor molar mass is known.

Assume you have 100 g of the compound, so that:

% ⇒ g (mass)

mass of C = 80.0 % of 100 g = 80.0 g of C

mass of H = 20.0 % of 100 g = 20.0 g of H

Use the masses of each element present to determine the number of moles of each element.

mass → moles

moles of C = 80.0 g x $\frac{1 \text{ mol}}{12.0 \text{ g}}$ = 6.67 mol of C

moles of H = 20.0 g x $\frac{1 \text{ mol}}{1.0 \text{ g}}$ = 20.0 mol of H

$C_{6.67}H_{20}$ 6.67 : 20 (simple ratio)

Since "100 g" was an arbitrary (but convenient) mass, the numbers of moles calculated have no real significance. However, the RATIO which exists between the numbers of moles *is* significant.

To find the smallest whole-number ratio, divide both by the SMALLER number.

moles of C = 6.67 mol ÷ 6.67 = 1 mol C

moles of H = 20.0 mol ÷ 6.67 = 2.998 mol H ≅ 3 mol H

According to these "simplest ratio" values, you can now say that carbon and hydrogen atoms are present in the ratio:

C : H

1 : 3

and the simplest way to express the chemical formula is:

Empirical Formula CH_3

Sample Problem — Determining an Empirical Formula

Determine the empirical formula (simplest ratio) of a compound that is 48.65% carbon, 8.11% hydrogen, and 43.24% oxygen.

$C_{?}H_{?}O_{?}$

What to Think about

% → mass (g)

1. In 100.0 g of the substance, there would be 48.65 g C, 8.11 g H, and 43.24 g O. Convert these amounts into moles.
2. Divide each molar quantity by the smallest one and then multiply by whatever factor is necessary to find their integral ratio (as shown in a conventional formula).

   The mole ratio and the individual atom ratio are of course the same. This means the subscripts in a formula can be read either as mole ratios or as individual atom ratios. If this compound has 3 mol of carbon atoms for every 2 mol of oxygen atoms then it has 3 dozen carbon atoms for every 2 dozen oxygen atoms, and 3 carbon atoms for every 2 oxygen atoms.

How to Do It

mass → mol

mol C: 48.65 g × $\frac{1 \text{ mol}}{12.0 \text{ g}}$ = 4.0542 mol C

mol H: 8.11 g × $\frac{1 \text{ mol}}{1.0 \text{ g}}$ = 8.1100 mol H

mol O: 43.24 g × $\frac{1 \text{ mol}}{16.0 \text{ g}}$ = 2.7025 mol O

÷ by smallest # mol to get simplest ratio

4.0542 mol C, 8.1100 mol H, 2.7025 mol O ÷ 2.7025 → C = 1.5, H = 3, O = 1

×2 → C = 3, H = 6, O = 2 (*simplest WHOLE # ratio)

Empirical Formula $C_3H_6O_2$

IMPORTANT: You must be able to recognize the following fractions and their decimal equivalents.

0.20 = 1/5	0.40 = 2/5	0.67 = 2/3
0.25 = 1/4	0.50 = 1/2	0.75 = 3/4
0.33 = 1/3	0.60 = 3/5	0.80 = 4/5

*multiply by reciprocal to eliminate decimal

SNEAKY TRICK: You don't have to re–write fractions such as 2.67 in the form 8/3. All you have to do is to recognize that numbers such as 2.67, 1.33, 5.67 and 3.33 involve **THIRDS** and simply multiply the fraction by **3** to clear the fraction. Similarly, numbers like 1.75, 2.25 and 3.75 involve **QUARTERS**, so that multiplying by **4** will clear such fractions.

IMPORTANT

INCREDIBLY, VITALLY IMPORTANT NOTE: DO NOT round early or worry about sig figs.

Always carry out calculations to 3 or 4 digits and NEVER round off intermediate values. The numbers 3.60, 3.67, 3.75 and 3.80 are very close to one another and improper round–off of calculations will cause you to multiply by the wrong number when trying to "clear fractions".

EXAMPLE: **What is the empirical formula of a compound containing 81.8% C and 18.2% H?**

$C_{?}H_{?}$

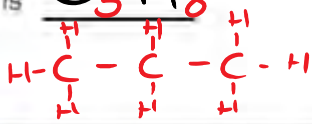
Assume 100.0 g of the compound is taken.

mass of C = 81.8g mass of H = 18.2g

moles C = 81.8 g × (1 mol / 12.0 g C) = 6.8167 mol (smallest #) ÷ 6.8167 = 1 × 3 = 3 mol C

moles H = 18.2 g × (1 mol / 1.0 g H) = 18.2 mol ÷ 6.8167 = 2.6699 ≅ 2.67 × 3 = 8 mol H (can NOT round up.)

Therefore the empirical formula is C_3H_8



chemistry homework

Assignment #9- Practice Problems #1-3 + Hebden Questions page 96 #46 (every 2nd letter)
Complete ALL assignments on a separate piece of paper with a clear heading and title

PRACTICE — Determining an Empirical Formula

1. A compound is 18.7% Li, 16.3% C, and 65.5% O. Determine its empirical formula.
2. A compound is 9.93% C, 58.6% Cl, and 31.4% F. Determine its empirical formula.
3. A sample of a compound contains 5.723 g Ag, 0.852 g S, and 1.695 g O. Determine its e

Answers:

1. Li_2CO_3
2. CCl_2F_2
3. Ag_2SO_4

6

% → mass → moles → simplify | mole ratio

(1) Li : $\frac{18.7g}{} \Big| \frac{1 mol}{6.9g}$ = 2.71 mol ÷ 1.3583 = 1.995 ≅ 2 : Li

C : $\frac{16.3g}{} \Big| \frac{1 mol}{12.0g}$ = 1.3583 mol ÷ 1.3583 = 1 = C

O : $\frac{65.5g}{} \Big| \frac{1 mol}{16.0g}$ = 4.09375 mol ÷ 1.3583 = 3.013 ≅ 3 = O

(16.0g = molar mass)

= Li_2CO_3

(2) % → g → mass → mol → simplify | Mole Ratio

C : $\frac{9.93g}{} \Big| \frac{1 mol}{12.0g}$ = 0.8275 mol ÷ 0.8275 = 1 = C

Cl : $\frac{58.6g}{} \Big| \frac{1 mol}{35.5g}$ = 1.6507 mol ÷ 0.8275 = 1.9948 ≅ 2 = Cl

F : $\frac{31.4g}{} \Big| \frac{1 mol}{19.0g}$ = 1.6526 mol ÷ 0.8275 = 1.99769... ≅ 2 = F

$C_1Cl_2F_2$ ⇒ CCl_2F_2

(3) Ag : $\frac{5.723g}{} \Big| \frac{1 mol}{107.9g}$ = 0.05304 mol ÷ 0.02654 = 1.998 ≅ 2

S : $\frac{0.852g}{} \Big| \frac{1 mol}{32.1g}$ = 0.02654 mol ÷ 0.02659 = 1

O : $\frac{1.695g}{} \Big| \frac{1 mol}{16.0g}$ = 0.10594 mol ÷ 0.02654 = 3.99 ≅ 4

Ag_2SO_4

Part C: Determining the Molecular Formula of a Compound

The difference between the **EMPIRICAL** formula and the **MOLECULAR** formula of a compound is:

- The **empirical formula** is the simplest ratio of the different types of atoms in the compound.
- The **molecular formula** is the ACTUAL number of each type of atom in each molecule.

If the empirical formula can be found, it is straightforward to calculate the molar mass of the empirical formula; that is, the empirical mass

emp. form.

C:H 1:2

The first example in this section pointed out that all of CH_2, C_2H_4, C_3H_6, C_4H_8 and C_5H_{10} have identical empirical formulae. Since all of these compounds have formulae which are whole–number multiples of CH_2, then the molar mass of all of the compounds must be a whole–number multiple of the empirical mass of CH_2.

Let N = the WHOLE NUMBER multiple of the empirical mass

$$\text{multiple} = N = \frac{\text{molar mass}}{\text{empirical mass}}$$

Since the molar mass is a multiple of the empirical mass, then the molecular formula must be the same multiple of the empirical formula:

molecular formula = N x (empirical formula) .

"N"

molecular formula = empirical formula × (molar mass / empirical mass)

A SUMMARY OF METHODS FOR FINDING THE MOLAR MASS

molar mass = g / mol

(a) ***Finding the Molar Mass from the Density of a Gas at STP***

If: density of gas X = 1.43 $\frac{g}{L}$ (at STP)

then: mass of 1 mol of X = 1.43 $\frac{g}{L}$ x $\frac{22.4\ L}{1\ mol}$ = **32.0** $\frac{\mathbf{g}}{\mathbf{mol}}$.

(b) Finding the Molar Mass from the Mass and Volume of a Gas at STP

If you are told: "0.0425 L of gas X at STP has a mass of 0.135 g"

then: density of gas X = $\frac{0.135\ g}{0.0425\ L}$ = 3.176 g/L

and: molar mass of X = 3.176 $\frac{g}{L}$ x $\frac{22.4\ L}{1\ mol}$ = **71.2 g/mol**

(c) ***Finding the Molar Mass if the Mass of a Certain Number of Moles is Given***

If you are told: "0.0250 mol of X has a mass of 1.775 g"

then: molar mass = $\frac{1.775\ g}{0.0250\ mol}$ = **71.0** $\frac{\mathbf{g}}{\mathbf{mol}}$

(d) ***Finding the Molar Mass if the Molar Mass is Given as a Multiple of a Known Molar Mass***

If you are told: X has a molar mass which is 1.64 times that of CO_2

then: molar mass CO_2 = 44.0 g/mol

and: molar mass of X = 1.64 x 44.0 g/mol = **72.2 g/mol**

EXAMPLE: A molecule has an empirical formula of HO and a molar mass of 34.0 g. (of the molecule)

What is the molecular formula?

H + O
1.0g + 16.0g

empirical mass of HO = 17.0g

N = molar mass / empirical mass = 34.0g / 17.0g = 2

and: molecular formula = N × (empirical formula)

= 2 × (HO) = H_2O_2

check: 2·(1.0g), 2·(16.0g) } 34.0g

EXAMPLE: A gas has the empirical formula POF_3. If 0.350 L of the gas at STP has a mass of 1.62 g, what is the molecular formula of the compound?

1 mol = 22.4L

molec. formula = emp. formula · (molar mass / empirical mass) = POF_3 · (103.712 / 104.0)

= POF_3 · (0.997...) ≅ 1

empirical mass of POF_3:
P = 31.0g
O = 16.0g
3×F = 3·(19.0g)
= 104.0g

density of gas = 1.62g / 0.350L

molar mass = 4.63g / 1L · 22.4L / 1 mol = 103.712 g/mol

molecular formula = POF_3

EXAMPLE: The empirical formula of a compound is SiH_3. If 0.0275 mol of compound has a mass of 1.71 g, what is the compound's molecular formula?

molec. formula = emp. formula · (molar mass / empirical mass) — "N"

Empirical mass: SiH_3
Si = 28.1g
3·H = 3·(1.0g)
31.1g

molar mass = g / mol
m.m = 1.71g / 0.0275mol
m.m = 62.18 g/mol

molec. formula = SiH_3 · (62.18 g/mol / 31.1g)

= SiH_3 · (1.999...) ≅ 2

= Si_2H_6

PRACTICE — Determining a Molecular Formula

A compound has an empirical formula of CH_2 and a molar mass of 42.0 g/mol. Determine its molecular formula.

"N" = multiple

What to Think about	How to Do It
1. Calculate the molar mass of the empirical formula.	molecular formula = empirical formula · (molar mass / empirical mass); emp. mass CH_2: C = 12.0g, H = 2·(1.0g), CH_2 = 14.0g; = CH_2 · (42.0 g/mol / 14.0g)
2. Divide the molar mass of the molecular formula by the molar mass of the empirical formula.	= CH_2 · (3)
3. Multiply the empirical formula itself by this factor.	molecular formula = 3 · (CH_2) = C_3H_6

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Part D: Determining the Formula of a Hydrate

water "trapped" in the solid, crystalline salt.

Recall that a **Hydrate** is a compound that contains a definitite amount of water.

The water can be removed by heating it up and evaporating all water to leave only the crystalline ionic compound ⟸ Anhydrate (or "anhydrous compound")

The difference between the **EMPIRICAL** formula and the **MOLECULAR** formula of a compound is: → shows # of H_2O

* **Unknown Hydrates** are written with an "*n*" before the water to show that the number of moles of water is unknown

always simplify so the anhydrate ratio = 1

Magnesium Sulfate Hydrate ← Hydrate

$1\,MgSO_4\;nH_2O$

Anhydrate — water

Solve for "how many waters?"

This is an example!

A 5.018 gram sample of a certain hydrate of magnesium sulfate, $MgSO_4\bullet xH_2O$, is heated until all the water is driven off. The resulting anhydrous compound weighs 2.449 grams. What is the formula of the hydrate? — anhydrate $MgSO_4$

Steps to finding the formula of a hydrate:

1. Determine the mass of the water that has left the compound

   mass of hydrate - mass of anhydrate = mass of water

① hydrate (salt + water) − anhydrate (salt) = water only

5.018g − 2.449g = 2.569g

2. Convert the mass of water to moles

$$\frac{\text{mass of water}}{\text{molar mass of water}} = \text{moles of water}$$

molar mass water H_2O = 18.0g

② mass → mol (H_2O)

$$2.569\text{g} \times \frac{1\text{ mol}}{18.0\text{g}} = 0.1427 \text{ mol } H_2O$$

3. convert the mass of anhydrate into moles of anhydrate:

$$\frac{\text{mass of anhydrate}}{\text{molar mass of anhydrate}}$$

molar mass $MgSO_4$ = 24.3 + 32.1 + 4(16.0) = 120.4g

③ mass → mol (anhydrate ⇒ ionic salt)

$$2.449\text{g} \times \frac{1\text{ mol}}{120.4\text{g}} = 0.02034 \text{ mol } MgSO_4$$

* should ALWAYS be smaller than H_2O

4. Find the water to anhydrate ratio:

   generally you will have more waters than anhydrates

   (water) > anhydrate

$$\frac{\text{moles of water}}{\text{moles of anhydrate}} = \text{mole ratio}$$

coefficients in the molecular formula

④ Mole ratio:

H_2O 0.1427 mol ÷ 0.02034 = 6.999 ≅ 7

anhydrate 0.02034 mol ÷ 0.02034 = 1

simplify

5. Use the **mole ratio** to determine the formula of the hydrate

⑤ $\boxed{\;}MgSO_4 \cdot \boxed{n}H_2O$

anhydrate : water

1 : 7 ⇒ molecular Formula $MgSO_4 \cdot 7H_2O$

chemistry homework

Assignment #10 Hebden Questions page 95 #47-55
+ "Formula of Hydrates" Questions #1-5
Complete ALL assignments on a separate piece of paper with a clear heading and title

Formula of Hydrates

$Cu(NO_3)_2 \cdot nH_2O$ — molar mass 187.5g — heating — H_2O is evaporated.

1. Find the formula when 6.00g of hydrated Cu(NO3)2 is decomposed and the mass decreases by 1.34g.

H_2O removed.

mass hydrate: 6.00g
mass H_2O: 1.34g → moles
mass anhyd: 6.00 − 1.34 = 4.66g → mol

mol H_2O: 1.34g | 1mol / 18.0g = 0.0744 mol ÷ 0.024853 = 2.99 ≅ 3 (H_2O)

mol anhyd: 4.66g | 1mol / 187.5g = 0.024853 ÷ 0.02485... = 1 ⇒ $Cu(NO_3)_2$

= $Cu(NO_3)_2 \cdot 3H_2O$

2. Determine the formula for hydrated Mg(NO3)2, given the following data:

Mass of container + hydrated $Mg(NO_3)_2$	12.22 g
Mass of container	5.44 g
Mass of container + anhydrous $Mg(NO_3)_2$	9.36 g

3. When 5.17 g of anhydrous Na2CO3 is left open to the air it becomes hydrated and the mass increases to 6.05 g. Determine the formula.

FULLY WORKED SOLUTIONS ON THE FOLLOW PAGES

4. When 5.00 g of ZnSO4•7H2O is heated, what mass of anhydrous ZnSO4 remain?

5. How much will the mass change when 10.46 g of CuSO4•5H2O becomes anhydrous?

Complete Lab Activity 4D: Formula of a Hydrate

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Hydrate Calculations Worksheet

November 1, 2017 1:32 PM

1. mass of H_2O = 1.34g
   mass of anhydrous salt ($Cu(NO_3)_2$) = 6.00g − 1.34g
   = 4.66g

mols H_2O = 1.34g x $\frac{1 \text{ mol}}{18.0 g}$ = ~~0.07444 mols~~ / 0.024853 mols | 2.9936 ≈ 3

mols $Cu(NO_3)_2$ = 4.66g x $\frac{1 \text{ mol}}{63.5 + 2(14.0) + 6(16.0) g}$ = ~~0.024853 mols~~ / 0.024853 mols | 1

⇒ Formula is $Cu(NO_3)_2 \cdot 3H_2O$

2. Mass of anhydrous $Mg(NO_3)_2$ = 9.36g − 5.44g = 3.92g
   Mass of hydrated $Mg(NO_3)_2$ = 12.22g − 5.41g = 6.81g
   ∴ Mass of H_2O = 6.81g − 3.92g = 2.81g

mols H_2O = 2.81g x $\frac{1 \text{ mol}}{18.0 g}$ = ~~0.15611 mols~~ / 0.02643 mols | 5.91 ≈ 6

mols $Mg(NO_3)_2$ = 3.92g x $\frac{1 \text{ mol}}{24.3 + 2(14) + 6(16) g}$ = ~~0.02643 mols~~ / 0.02643 mols | 1

⇒ Formula is $Mg(NO_3)_2 \cdot 6H_2O$

3. Mass of anhydrous Na_2CO_3 = 5.17g
   Mass of hydrated Na_2CO_3 = 6.05g
   ∴ Mass of water = 6.05g − 5.17g = 0.88g

mols of H_2O = 0.88g x $\frac{1 \text{ mol}}{18.0 g}$ = ~~0.048889 mols~~ / 0.048774 mols | 1.002 ≈ 1

mols of Na_2CO_3 = 5.17g x $\frac{1 \text{ mol}}{2(23.0) + 12.0 + 3(16.0) g}$ = ~~0.048774 mols~~ / 0.048774 mols | 1

⇒ Formula is $Na_2CO_3 \cdot H_2O$

4. ① Find % $ZnSO_4$ in $ZnSO_4 \cdot 7H_2O$ (Assume 1 mol)

$M.M.$ of $ZnSO_4 \cdot 7H_2O = 65.4g + 32.1g + 4(16.0g) + 7(18.0g)$
$= 287.5g$

$\%\ ZnSO_4 = \dfrac{65.4g + 32.1g + 4(16.0g)}{287.5g} \times 100\%$

$= \dfrac{161.5g}{287.5g} \times 100\%$

$= 56.17\%$

② Mass of $ZnSO_4 = 5.00g \times 0.5617$

3 sf $= 2.81g$

5. ① Find % H_2O in $CuSO_4 \cdot 5H_2O$ (Assume 1 mol)

$MM = 63.5g + 32.1g + 4(16.0g) + 5(18.0g)$
$= 249.6g$

$\%\ H_2O = \dfrac{5(18.0g)}{249.6g} \times 100\%$

$= \dfrac{90.0g}{249.6g} \times 100\%$

$= 36.06\%$

② Mass of $H_2O = 10.46g \times 0.3606$

4 sf $= 3.772g$

Part E: Molar Concentration

Examples of common household liquids

Warm Up

1. List three products in your refrigerator that are solutions.
   milk (colloid), orange juice, soy sauce
2. Name some substances that are dissolved in these solutions.
   water, salt, natural juices
3. Where else in your home are solutions kept?
   under sink, bathroom, laundry, pantry

" Molarity " The *Unit* of Concentration

→ Homogeneous (consistent throughout)

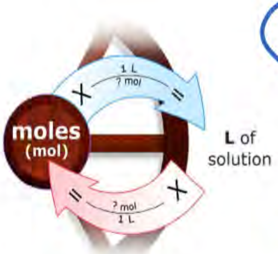
A solution is a type of mixture in which the chemical species are completely (mixed) dissolved. A solute is a minor component of the mixture, generally what has been dissolved. The solvent is the major component of the mixture, generally what the solute was dissolved in. (most often water)

Concentration is any expression of the proportion of a chemical in a solution. Chemists need to know the amount of solute present in any volume of solution they might dispense.

Concentration is most usefully expressed as an amount of solute per volume of solution rather than per volume of solvent.

There are many units of concentration. Common units of concentration express the amount of solute in grams. These include grams per litre of solution, percent m/v, etc...... → isopropyl alcohol 50%, 70% (90%)

But the ***most useful*** unit of concentration for chemists is the quantity of solute in MOLES!



"concentration" [conc.]

Molarity (M) is the number of moles of solutes (chemical), per Litre (volume) of solution

For example, 1.8 M HCl means 1.8 mol HCl per litre of solution. ***Molar concentrations allow chemists to directly compare the number of particles in the same volume of different solutions.*** *For example, 10 mL of 2 M Li^+ contains twice as many ions as 10 mL of 1 M Na^+.*

Name	Equivalence Statement	Conversion Factors	
Molar concentration	1 L solution = ? mol solute	$\frac{\text{? mol solute}}{\text{1 L solution}}$	$\frac{\text{1 L solution}}{\text{? mol solute}}$
Example: 3 M HCN	1 L solution = 3 mol HCN	$\frac{\text{3 mol HCN}}{\text{1 L solution}}$	$\frac{\text{1 L solution}}{\text{3 mol HCN}}$

"three molar solution"

3 mol → ? g (mass) ↑ how much to dissolve in water

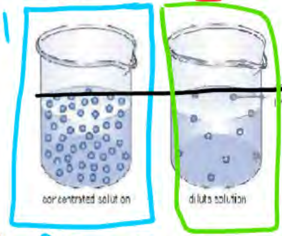
Molar Concentration

Knowing the concentration of a solution provides a way to find out how much of a particular substances exists in a given volume of the solution. "molar" = unit

Definitions: The CONCENTRATION = molarity = $\frac{mol}{L}$ = [? M]

A **CONCENTRATED** solution a high amount of mol dissolved per unit volume

A **DILUTE** solution a lower amount of mol dissolved per unit volume



same volume

↑[conc.] ↓[conc.]

A **SATURATED SOLUTION** the max amount of solute has been dissolved for that particular volume → anymore will sediment on the bottom.

Chemists frequently use the "mole" to describe the amount of a substance in a solution. ← don't usually use mass

The **MOLAR CONCENTRATION** or **MOLARITY** of a substance in solution is the number of moles of the substance contained in 1 L of ***solution***.

$M = \frac{\#\ of\ mol}{volume}$

EXAMPLE: **If 2.0 L of solution contain 5.0 mol of NaCl, what is the molarity of the NaCl?**

molar concentration = molarity = $\frac{mol}{L} = \frac{5.0\ mol}{2.0\ L} = 2.5\ \frac{mol}{L} = 2.5\ M$

[NaCl] = 2.5 M "concentration of"

NOTES:
1. The unit symbol for "molarity" is "M". = mol/L
2. When expressed in words, the unit symbol "**M**" is written and read as "molar".
3. The short-hand symbol for "**molar concentration of ...**" is a set of brackets: [ ]
   → molarity = 5.0 M = [5.0]

EXAMPLES: If a 1.0 L of solution contains 2.5 mol of NaCl, the molar concentration can be expressed in several equivalent ways

$M = \frac{2.5\ mol}{1.0\ L} = 2.5\ M$

$= [2.5]$

$= 2.5\ \frac{mol}{L}$

(most common) → [NaCl] = 2.5 M "the concentration of NaCl, is 2.5 molar"

molar concentration = $\frac{\textbf{moles}}{\textbf{volume}}$	where:	c = molar concentration, in mol/L
or: $c = \frac{n}{V}$ (same) $M = \frac{n}{V}$		n = number of moles
		V = volume, in litres (L)

EXAMPLE: **What is the [NaCl] in a solution containing 5.12 g of NaCl in 250.0 mL of solution?**

"concentration of NaCl" 3 s.f. volume; mass → mol; → L

m.m. NaCl: Na = 23.0 g, Cl = 35.5 g, 58.5 g

$[NaCl] = M = \frac{n}{V}$

Mol NaCl: $5.12\ g \times \frac{1\ mol}{58.5\ g} = 0.0875\ mol$

vol mL → $250.0\ mL \times \frac{10^{-3}\ L}{1\ mL} = 0.2500\ L$

$M = \frac{0.0875\ mol}{0.2500\ L}$

$M = 0.350\ \frac{mol}{L}$

$\therefore [NaCl] = 0.350\ \frac{mol}{L} = 0.350\ M$ (same)

$C = M = \frac{mol}{L}$

EXAMPLE: What mass of NaOH is contained in 3.50 L of 0.200 M NaOH?

Plan: The molarity (c) and volume (V) are given so moles (n) can be found. Moles can then be converted to mass.

$c = \frac{n}{V}$ ∴ $n = c \cdot V$

$n = (0.200 \frac{mol}{L})(3.50 L)$

$n = 0.700$ mol NaOH

mol → mass

$\frac{0.700 \text{ mol}}{} \times \frac{40.0 \text{ g NaOH}}{1 \text{ mol}} = 28.0$ g NaOH

EXAMPLE: What is the molarity of pure sulphuric acid, H_2SO_4, having a density of 1.839 g/mL?

g → mol

Notice that density and molarity both have units of amount/volume

where: density = $\frac{\text{amount (as mass)}}{\text{volume}}$ and molarity = $\frac{\text{amount (as moles)}}{\text{volume}}$

$[H_2SO_4] = \frac{1.839 \text{ g}}{1 \text{ mL}} \times \frac{1 \text{ mol}}{98.1 \text{ g}} = 0.0187 \frac{mol}{mL} \times \frac{1 \text{ mL}}{10^{-3} \text{ L}} = 18.7$ M

(molar mass)

EXAMPLE: What is the molarity of the $CaCl_2$ in a solution made by dissolving and diluting 15.00 g of $CaCl_2 \cdot 6H_2O$ to 500.0 mL?

when $CaCl_2 \cdot 6H_2O$ dissolves in water:

$1CaCl_2 \cdot 6H_2O \rightarrow 1CaCl_2 + 6H_2O$

$C = M = \frac{n}{V} = \frac{mol}{L}$

$\frac{15.00 \text{ g}}{} \times \frac{1 \text{ mol}}{219.1 \text{ g}} = 0.06846$ mol $CaCl_2 \cdot 6H_2O$

$M = \frac{n}{L} = \frac{0.06846 \text{ mol}}{0.5000 \text{ L}} = 0.1369 \frac{mol}{L} = 0.1369$ M

Assignment #11 Hebden Questions page 98-99

#59-71 *Complete ALL assignments on a separate piece of paper with a clear heading and title*

ADD H_2O

concentration = molarity
$C = M = \frac{mol}{L}$

Part F: Simple Dilutions

Assume: (start)
- **initial concentration** of solution (in more concentrated form) = C_{conc} = C_i
- **initial volume** of solution (in more concentrated form) = V_{conc} = V_i

(after adding H_2O)
- **diluted concentration** (after water is added) = C_{DIL} = C_f
- **diluted volume** (after water is added) = V_{DIL} = V_f

The "diluted volume" can also be thought of as the "total volume" (after H_2O is added)

Since $c = \frac{n}{V} = \frac{\#mol}{L}$	then $n = C \cdot V$

Then $n_i = n_f$, $C_i \cdot V_i = C_f \cdot V_f$

which means moles of chemical in concentrated solution = $n_i = c_i \times V_i$ $\qquad n_{CONC} = c_{CONC} \times V_{CONC}$

and moles of chemical in diluted solution = $n_f = c_f \times V_f$ $\qquad n_{DIL} = c_{DIL} \times V_{DIL}$.

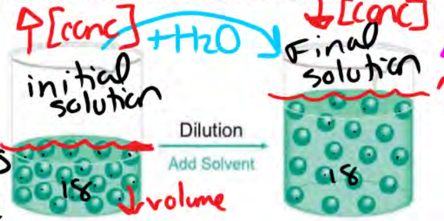
The amount of chemical is not changed when the **solution is diluted**,

only the concentration is changed. (due to an ↑ volume) → add H_2O only, NOT adding more solute.

moles of concentrated chemical (start) = moles of diluted chemical (end)

$n_i = n_f$

Dilution is the procedure for preparing a less concentrated solution from a more concentrated solution.



#particles does not change

Moles of solute before dilution (i) = Moles of solute after dilution (f)

$n_i = C_i V_i = C_f V_f = n_f$

EXAMPLE: If 200.0 mL of 0.500 M NaCl is added to 300.0 mL of water, what is the resulting [NaCl] in the mixture? ? M (3 s.f.)

0.500 M, 200 mL [NaCl] — +300 mL, 200 mL, 500 mL

$n_i = n_f$

$C_i \cdot V_i = C_f \cdot V_f$

$\therefore C_f = \frac{C_i \cdot V_i}{V_f} = \frac{(0.500M)(200.0mL)}{(200.0+300.0mL)}$

$C_f = [NaCl] = 0.200M$

MAKING DILUTE SOLUTIONS FROM CONCENTRATED SOLUTIONS

Again, this calculation is based on the fact that the moles of chemical in the diluted solution equals the moles of chemical poured from the concentrated solution. That is, $n_{CONC} = n_{DIL}$.

$c_{CONC} \times V_{CONC} = c_{DIL} \times V_{DIL}$

} $n_i = n_f$, $C_i \cdot V_i = C_f \cdot V_f$

EXAMPLE: **What volume of 6.00 M HCl is used in making up 2.00 L of 0.125 M HCl?** (final)

i: 6.00 M, ? L → +H_2O → f: 0.125 M, 2.00 L

$C_i \cdot V_i = C_f \cdot V_f$

$V_i = \frac{C_f \cdot V_f}{C_i}$

$V_i = \frac{(0.125M)(2.00L)}{(6.00M)}$

$V_i = 0.041\overline{6}L$ (3 s.f.)

$= 0.0417L$ or $41.7mL$ ($4.17 \times 10^{-2} L$)

$n_i = n_f$

$C_i \cdot V_i = C_f \cdot V_f$

EXAMPLE: A student mixes 100.0 mL of water with 25.0 mL of a sodium chloride solution having an unknown concentration. If the student finds the molarity of the sodium chloride in the diluted solution is 0.0876 M, what is the molarity of the original sodium chloride solution?

+H_2O V_i C_i C_f

→ what is the initial [conc.]? $C_i = ?$

what is V_f? Remember in a dilution (+H_2O) volumes are always added.

✱ $V_f = V_i + H_2O$

$= 25.0\text{ mL} + 100.0\text{ mL}$

$V_f = 125.0\text{ mL}$

If, $C_i V_i = C_f \cdot V_f$

Then, $C_i = \dfrac{C_f \cdot V_f}{V_i} = \dfrac{(0.0876\text{ M})(125.0\text{ mL})}{(25.0\text{ mL})}$

$C_i = [NaCl]_i = 0.438\text{ M}$

chemistry homework

Assignment #12 Hebden Questions page 102 #78, 80, 82-83, 85-87, 89 + 91
Complete ALL assignments on a separate piece of paper with a clear heading and title

Simple Dilutions: ① solution + adding water

Part G: Complex Dilutions:

Mixing 2 solutions having ***DIFFERENT*** concentrations the **SAME** chemical

One solution dilutes the other, and vice versa.
In our calculations, one solution is called #1 (A) and the other is #2 (B) to keep track.

Since these involve combining both moles and solvent from both solutions, the *moles before* and *moles after* will ~~be~~ NOT the same.

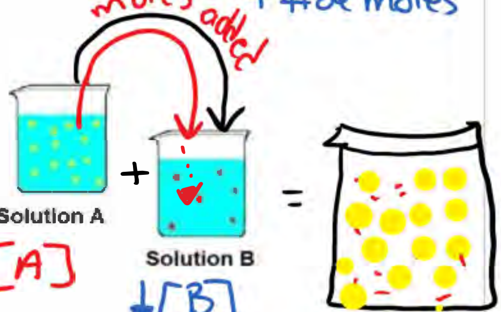
+volumes
+ # of moles

$C = \frac{n}{V}$

$n = C \cdot V$

Let n_A = # of moles in Solution A

Let n_B = # of moles in Solution B



Combine Solutions A + B:

moles = $n_A + n_B$ = $C_AV_A + C_BV_B$ = $(n_A) + (n_B)$

total volume = $V_A + V_B$

2 solution mixing together, ADD volumes!

Recall that:

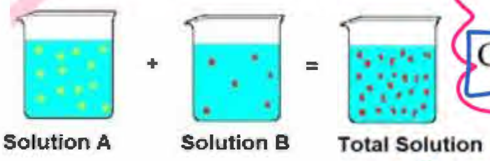
C = moles / volume

Mixing *two or more solutions* having the same solute, but different concentrations: adding Sol$_A$ + Sol$_B$

$$C_{Final} = \frac{\text{Total moles}}{\text{Total volume}}$$

$$C_{Final} = \frac{(C_AV_A + C_BV_B + \ldots)}{(V_A + V_B + \ldots)}$$

($n_A + n_B$)



*start by identifying Sol$_A$ and Sol$_B$ and what info you have.

*Start By Identifying "Solution A" + "Solution B"

Example: If 450.0 mL of 0.150 M HCl solution (solution A) is mixed with (+) 125.0 mL of a 0.220 M HCl solution, what is the resulting HCl concentration?

Solution A: [HCl] 450.0 mL, 0.150 M

(+)

Solution B (same): [HCl] 125.0 mL, 0.220 M

= Solution C (A + B): 450+125 = 575.0 mL; [HCl] = ? M

$C = \frac{mol}{L}$ ∴ mol = C·V

Solution C = Sol$_A$ + Sol$_B$

$= \frac{mol_A + mol_B}{\text{total volume}}$

$[Sol_C] = \frac{(C_A \cdot V_A) + (C_B \cdot V_B)}{V_A + V_B}$

$[Sol_C] = \frac{(0.150M)(450.0mL) + (0.220M)(125.0mL)}{(575.0mL)}$

$[Sol_C] = [HCl]_f = 0.165M$ (after mixing)

$Sol_c = HCl_f = C$ (after mixing)

Example: 500.0 mL (→ 0.5000 L) of 0.6750 M hydrofluoric acid (A) was combined with (+) 1.875 L of 0.375 M HF (B). Determine the concentration of the new solution.

"what is the C_f? or $[Sol_c]$? $[HF]_f$?" all mean the same thing.

$$[Sol_c] = C_f = [HF]_f = \frac{(C_A \cdot V_A) + (C_B \cdot V_B)}{(V_A + V_B)}$$

$$[HF]_f = \frac{[(0.6750M)(0.5000L)] + [(0.375M)(1.875L)]}{(0.5000L + 1.875L)}$$

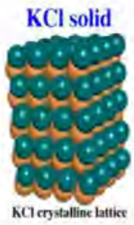
(0.3375; 0.703125; 0.9375)

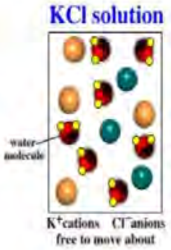
$$[HF]_f = 0.438M$$

Assignment #13 Hebden Questions page 102 #79, 81, 84, 88, 90 + 92
Complete ALL assignments on a separate piece of paper with a clear heading and title

Part H: Ions in Solution

Ionic compounds have no net charge (→ overall charge = ∅). The ions associate together in the ratio that results in their charges canceling.





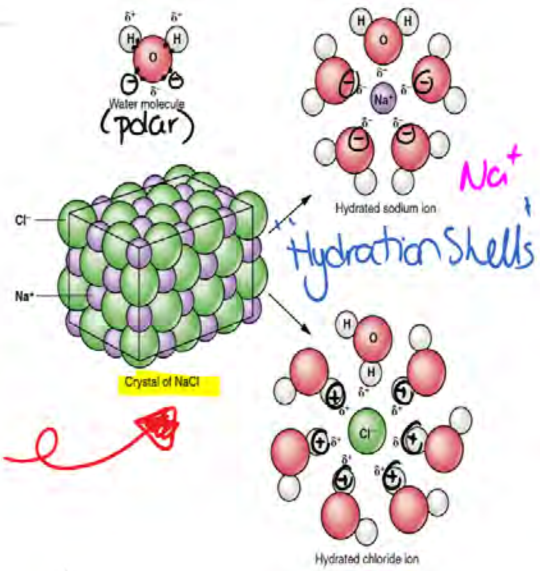
For example: $2Al^{3+}(aq) + 3SO_4^{2-}(aq) \rightarrow Al_2(SO_4)_3\ (s)$

The ionic compound is neutral; the ions have a net charge of zero: $2(3+) + 3(2-) = 0$.

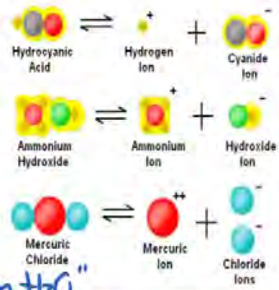
When an ionic compound dissolves, the same ions that associated together to form the compound now dissociate (dis-associate) and travel independently through the solution.

For example,

$Al_2(SO_4)_3\ (s) \rightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$



Being able to relate the concentration of dissolved ions to the concentration of their parent compound is extremely important in chemistry.



The **dissociation equation** provides the ratio of the released ions to each other and to their parent compound;

1:3

$AlCl_3(aq)$ is an ionic compound which releases 3 Cl^- ions in water for every molecule of $AlCl_3$ which dissolves. (→ 1 Al^{3+} ion)

$AlCl_3(s) \longrightarrow Al^{3+}(aq) + 3\,Cl^-(aq)$ (aq means "dissolved in H_2O")

Sample Problem — Relating the Concentration of Dissolved Ions to the Concentration of Their Parent Compound

What concentrations of ions are present in 3.0 M $CaCl_2(aq)$?

What to Think about	How to Do It
The ratios in the dissociation equation show that 1 mol Ca^{2+} and 2 mol Cl^- are formed for each mole of $CaCl_2$ dissolved.	

$CaCl_2(s)$ (solid)	→ (dissociate)	$Ca^{2+}(aq)$	+ $2Cl^-(aq)$
1	:	1	: 2 (molar ratio)
3.0M	ions dissociate	3.0M	6.0M = 6 mol/L

(1:1, 1:2, ×2)

EXAMPLE: What is the molar concentration of the chloride ions in 0.25 M $AlCl_3(aq)$?

$AlCl_3(aq)$ is an ionic compound which releases 3 Cl^- ions in water for every molecule of $AlCl_3$ which dissolves.

dissociation reaction ① $AlCl_3(s) \longrightarrow Al^{3+}(aq) + 3Cl^-(aq)$ ((aq) = solution "dissolved in H_2O")

1 : 1 : 3 mole ratio

Therefore: $[Cl^-] = 0.25 \times 3 = [Cl^-] = 0.75M$

An alternate way to carry out this calculation is as follows. Since 1 mol of $AlCl_3$ produces 1 mol of Al^{3+} and 3 moles of Cl^-: (1 mol, 3 mol *mole ratio)

$AlCl_3(s) \longrightarrow Al^{3+}(aq) + 3Cl^-(aq)$

1 mol

you can simply state that the $[Cl^-]$ in solution is three times the dissolved $[AlCl_3]$.

$[Cl^-] = 3 \times [AlCl_3]$

$[Cl^-] = 3 \times (0.25M) = 0.75M$

Remember:

M, moles
Standard number of molecules.
1 mole = 6.02×10^{23} molecules.

C, concentration
Is equivalent to moles per litres

V, volume
The amount of water (usually)
Measured in = litres

> The **molar concentration** of a chemical is indicated by putting **square brackets [ ]** around the chemical's formula.

For example, [Na^+] means the molar concentration of Na^+.
A couple of precautions:

- "M" already means "mol per L" therefore *don't write* "M per L" because that would mean "moles per litre per litre," which doesn't make sense.
- You can write "2 M Na^+" or "[Na^+] = 2 M" but *don't write* "2 M [Na^+]" because that would mean "two molar the molar concentration of Na^+," which doesn't make sense.

> The **dissociation equation** provides the **conversion factor** for calculations. ⟹ mole ratio

Sample Problem — Three-Step Conversion: Volume of Solution to Number of Ions

	~~Aluminum chloride can be used to produce aluminum chlorohydrate, an active ingredient in antiperspirants.~~ How many chloride ions are in 0.025 L of 0.30 M $AlCl_3$? (dissociation reaction)
What to Think about	Convert: L soln ($AlCl_3$) → mol $AlCl_3$ → mol Cl^- → ions Cl^- $AlCl_{3(s)} \rightarrow Al^{3+}_{(aq)} + 3Cl^-_{(aq)}$ 1 mol $AlCl_3$ = 3 mol Cl^-
How to Do It	$0.025\,L \times \frac{0.30\,mol\,AlCl_3}{1\,L} \times \frac{3\,mol\,Cl^-}{1\,mol\,AlCl_3} \times \frac{6.02\times10^{23}\,ions}{1\,mol\,Cl^-} = 1.4\times10^{22}$ ions of Cl^- (molarity) (mole ratio) (avagadro's #)

0.025 L, [Cl^-] = ?, 0.30 M $AlCl_3$ = 0.30 mol/L

EXAMPLE: What is the concentration of each type of ion in a solution made by mixing 50.0 mL of 0.240 M $AlBr_3$ and 25.0 mL of 0.300 M $CaBr_2$?

① When $AlBr_3$ and $CaBr_2$ dissolve in water, they undergo dissociation.

$AlBr_{3(s)} \rightarrow Al^{3+}_{(aq)} + 3Br^-_{(aq)}$ (1 : 1 : 3) $\quad CaBr_{2(s)} \rightarrow Ca^{2+}_{(aq)} + 2Br^-_{(aq)}$ (1 : 1 : 2)

② First perform a dilution calculation on each of the starting solutions.

Simple DiL x2: $C_f = \frac{C_i V_i}{V_f}$

$[AlBr_3]_f = \frac{(0.240M)(50.0mL)}{(50.0+25.0)} = 0.160M \qquad [CaBr_2]_f = \frac{(0.300M)(25.0mL)}{(50.0+25.0mL)} = 0.100M$

③ Below each dissociation equation, indicate each ion's concentration.

After mixing: * each solution has effectively been diluted by the other solution

$1\,AlBr_3 \rightarrow 1\,Al^{3+} + 3Br^-$: 0.160M → 0.160M (×1), 0.480M (×3)

$1\,CaBr_2 \rightarrow 1\,Ca^{2+} + 2Br^-$: 0.100M → 0.100M (×1), 0.200M (×2)

when solutions are mixed, there are 2 sources of ions that make up the new solution

$[Al^{3+}] = 0.160M$

$[Ca^{2+}] = 0.100M$

$[Br^-] = 0.480M + 0.200M = 0.680M$

Before: 50.0 mL, $[AlBr_3] = 0.240M$, $[Br^-] = 3(0.240M) = 0.720M$, $[Al^{3+}] = 0.240M$; 25.0 mL, $[CaBr_2] = 0.300M$, $[Ca^{2+}] = 0.300M$, $[Br^-] = 2(0.300M) = 0.600M$

Simple Dilution: $C_iV_i = C_fV_f$, $C_f = \frac{C_iV_i}{V_f}$

n, moles
Standard number of molecules.
1 mole = 6.02×10^{23} molecules.

C, concentration
is equivalent to moles per litres.

V, volume
The amount of water (usually)
Measured in → litres

The **molar concentration** of a chemical is indicated by putting **square brackets [ ]** around the chemical's formula.

For example, $[Na^+]$ means the molar concentration of Na^+.
A couple of precautions:

- "M" already means "mol per L" therefore *don't write* "M per L" because that would mean "moles per litre per litre," which doesn't make sense.
- You can write "2 M Na^+" or "$[Na^+]$ = 2 M" but *don't write* "2 M $[Na^+]$" because that would mean "two molar the molar concentration of Na^+," which doesn't make sense.

The **dissociation equation** provides the **conversion factor** for calculations.

Sample Problem — Three-Step Conversion: Volume of Solution to Number of Ions

Aluminum chloride can be used to produce aluminum chlorohydrate, an active ingredient in antiperspirants. How many chloride ions are in 0.025 L of 0.30 M $AlCl_3$? (0.30 mol / L)

What to Think about

Convert:
L soln → mol $AlCl_3$ → mol Cl^- → ions Cl^-

molarity as conversion factor. Avogadro's # to represent # of ions.

$AlCl_{3(s)} \rightarrow Al^{3+}_{(aq)} + 3Cl^-_{(aq)}$
1 : 1 : 3 molar ratio

How to Do It

$$0.025\,L\,AlCl_3 \times \frac{0.30\,mol\,AlCl_3}{1\,L} \times \frac{3\,mol\,Cl^-}{1\,mol\,AlCl_3} \times \frac{6.02\times10^{23}\,ions\,Cl^-}{1\,mol\,Cl^-} = 1.4\times10^{22}\,ions\,Cl^-$$

(molarity) (mole ratio)

EXAMPLE: What is the concentration of each type of ion in a solution made by mixing 50.0 mL of 0.240 M $AlBr_3$ and 25.0 mL of 0.300 M $CaBr_2$? ← 2 different solutions!!

Step ① dissociate

When $AlBr_3$ and $CaBr_2$ dissolve in water, they undergo dissociation.

$1AlBr_{3(s)} \rightarrow 1Al^{3+}_{(aq)} + 3Br^-_{(aq)}$ $\quad$ $1CaBr_{2(s)} \rightarrow 1Ca^{2+}_{(aq)} + 2Br^-_{(aq)}$

Step ② simple Dilution

First perform a dilution calculation on each of the starting solutions. (new total volume ⇒ ADD 50+25 = 75.0 mL)

$C_iV_i = C_fV_f$

$[AlBr_3]_f = \frac{(0.240M)(50.0mL)}{(75.0mL)} = 0.160M$ $\quad$ $[CaBr_2]_f = \frac{(0.300M)(25.0mL)}{(75.0mL)}$

$[CaBr_2]_f = 0.100M$

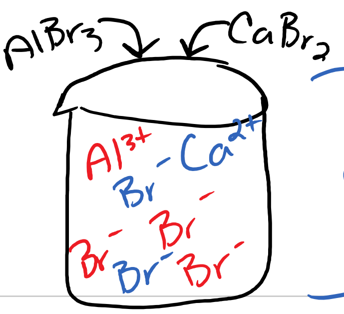
Step ③ [ion]

Below each dissociation equation, indicate each ion's concentration.

$AlBr_{3(s)} \rightarrow Al^{3+}_{(aq)} + 3Br^-_{(aq)}$ $\quad$ $CaBr_{2(s)} \rightarrow Ca^{2+}_{(aq)} + 2Br^-_{(aq)}$

1 : 1 : 3 $\quad$ 1 : 1 : 2

0.160M 0.160M 0.480M (×3) $\quad$ 0.100M 0.100M 0.200M (×2)



$[Al^{3+}] = 0.160M$

$[Ca^{2+}] = 0.100M$

$[Br^-] = (0.480M + 0.200M) = 0.680M$ (from each solution)

chemistry homework

Assignment #14 Hebden Questions page 212 #30-36
NOTE: THESE QUESTIONS ARE IN THE "SOLUTIONS" SECTION OF YOUR TEXTBOOK...not The Mole
Complete ALL assignments on a separate piece of paper with a clear heading and title

THE MOLE - UNIT REVIEW ASSIGNMENT

Write the Unit Conversion and/or Equivalence Statements you need to know & use in this box

All work must be shown to receive credit on thet test….so practice showing ALL your working our NOW!

1. Define each of the following words or phrases using complete sentences.

a) Molar Mass: the molar mass of a chemical compound is defined as the mass of a sample of that compound divided by the amount of substance in that sample, measured in moles. Or the mass of 1 mol of a substance

b) Mole: The mole is the unit of measurement for amount of substance in the International System of Units. It is defined as exactly 6.02214076×10^{23} particles, which may be atoms, molecules, ions, or electrons.

c) Avogadro's Law: States that under the same conditions of temperature and pressure, equal volumes of different gases contain an equal number of molecules. *(ie: 1 mol of any gas = 22.4 L)*

d) Empirical Formula: the empirical formula of a chemical compound is the simplest positive integer ratio of atoms present in a compound

e) STP: is the abbreviation for Standard Temperature and Pressure. STP most commonly is used when performing calculations on gases, such as gas density. The standard temperature is 273 K (0° Celsius or 32° Fahrenheit) and the standard pressure is 1 atm pressure.

f) Molarity: Molar concentration or molarity is most commonly expressed in units of moles of solute per litre of solution.

2. Calculate the molar mass of $Mg_3(PO_4)_2 \bullet 9H_2O$

(3 x Mg) + (2 x P) + 2(4 x O) + 9(2 x H) + (9 x O)

alternatively
17 (Oxygen)
17 (16.0g)

3(24.3) + 2(31.0) + 8(16.0) + 18(1.0) + 9(16.0)

72.9 + 62.0 + 128.0 + 18.0 + 144.0

= 424.9

all values to 1 d.p of precision ∴ answer to 1 d.p. for sig figs

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atoms → mol → mass

3. What is the mass of 1.23×10^{25} atoms of tin?

$$\frac{1.23\times10^{25}\text{ atoms}}{} \left| \frac{1\text{ mol}}{6.02\times10^{23}\text{ atoms}} \right| \frac{118.7\text{g}}{1\text{ mol}} = 2425.26578....$$

3 s.f.

$= 2.43\times10^{3}$ g of Sn

3 s.f.

4. Determine the number of moles of fluorine molecules contained in 12.35 L of fluorine gas at STP. 1 mol = 22.4L

volume → mol 4 s.f.

$$\frac{12.35\text{L}}{} \left| \frac{1\text{ mol}}{22.4\text{L}} \right. = 0.5513 \text{ mol of F gas}$$

4 s.f.

5. What is the density of dinitrogen tetroxide gas at STP? 1 mol = 22.4L.

2·N 4·O

$D = \frac{\text{mass}}{\text{volume}}$

$N_2O_4 = 2(14.0)+4(16.0)$

molar mass = 92.0g/mol

$$\frac{92.0\text{ g}}{1\text{ mol}} \left| \frac{1\text{ mol}}{22.4\text{L}} \right. = 4.11\text{ g/L}$$

6. Given two 42.0 g samples of Fe and Cu, show why one sample must contain more atoms than the other. convert mass → mol → # atoms for each.

Fe: $$\frac{42.0\text{g}}{} \left| \frac{1\text{ mol}}{55.8\text{g}} \right| \frac{6.02\times10^{23}\text{ atoms}}{1\text{ mol}} = 4.53\times10^{23}\text{ atoms Fe}$$

Cu: $$\frac{42.0\text{g}}{} \left| \frac{1\text{ mol}}{63.5\text{g}} \right| \frac{6.02\times10^{23}\text{ atoms}}{1\text{ mol}} = 3.98\times10^{23}\text{ atoms Cu}$$

Because Fe and Copper have different molar masses (ie: a Cu atom has a larger mass) there are less Cu atoms in a 42.0g sample

7. At STP, what volume would 37.6 g of NO_2 gas occupy?

1mol = 22.4L 3 s.f.

mass → mol → volume

molar mass NO_2

N: 14.0g

2·O = 2(16.0g)

46.0g/mol

$$\frac{37.6\text{g}}{} \left| \frac{1\text{ mol}}{46.0\text{g}} \right| \frac{22.4\text{L}}{1\text{ mol}} = 18.30956....$$

3 s.f.

= 18.3 L

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3 Cl atoms / 1 molecule $CHCl_3$

8. How many chlorine atoms are present in 6.67 g of $CHCl_3$?

mass → mol → molecules → atoms Cl

molar mass $CHCl_3$: C = 12.0g, H = 1.0g, 3·Cl = 3(35.5g); 119.5g

$$\frac{6.67\,g}{} \times \frac{1\,mol}{119.5\,g} \times \frac{6.02\times10^{23}\,molec.}{1\,mol} \times \frac{3\ Cl\ atoms}{1\ molec.\ CHCl_3} = 1.01\times10^{23}\ atoms\ of\ Cl$$

9. Determine the percentage composition of ammonium bicarbonate. = NH_4HCO_3

N = 14.0g
5×H = 5.0g
C = 12.0g
3×O = 48.0g

Molar Mass = 79.0g

$\%N = \frac{14.0}{79.0}\times100$ N = 17.7%

$\%H = \frac{5.0}{79.0}\times100$ H = 6.3%

$\%C = \frac{12.0g}{79.0}$ C = 15.2%

$\%O = \frac{48.0g}{79.0}$ O = 60.8%

10. The percentage composition of a certain compound is 24.4% carbon, 3.39% H, and the remainder is chlorine. (100 − 21.4 − 3.39 = 72.21% chlorine)

(*assume 100g, % → g)

a) Determine the empirical formula

mol C: $\frac{24.4g}{} \times \frac{1mol}{12.0g}$ = 2.0333 mol C → ÷2.033 (smallest #) = 1 ×3 = 3 = C_3

mol H: $\frac{3.39g}{} \times \frac{1mol}{1.0g}$ = 3.39 mol H → 1.67 (1 2/3) ×3 = 5.01 = H_5

mol Cl: $\frac{72.21g}{} \times \frac{1mol}{35.5g}$ = 2.034 mol Cl → 1 ×3 = 3 = Cl_3

= $C_3H_5Cl_3$

b) If the molar mass of the substance is 442.5 g, then what is its molecular formula?

molecular formula: N × (empirical formula)

$$N = \frac{molar\ mass}{empirical\ mass} = \frac{442.5g}{147.5g} = 3\ (the\ multiple)$$

$C_3H_5Cl_3$
3(12.0g) = 36.0g
5(1.0g) = 5.0g
3(35.5g) = 106.5
147.5

∴ 3 × $C_3H_5Cl_3$

molecular formula = $C_9H_{15}Cl_9$

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11. What mass of water will be driven off when 5.65 g of hydrated barium chloride, $BaCl_2 \cdot 2H_2O$, is heated?

① Find % H_2O in $BaCl_2 \cdot 2H_2O$ (3sf.)

molar mass Ba = 137.3g
Cl = 2(35.5g)
H_2O = 2(18.0g)
= 244.3g/mol

$\% H_2O = \frac{2(18.0g)}{244.3g/mol} \times 100\% = 14.73598...$

② Mass of H_2O in a 5.65g sample

$(0.1473598)(5.65g) = 0.83258$ (3sf.) = **0.833g**

12. Heating drives off water from the structure of hydrated copper (II) chloride. The anhydrous salt is left behind. Using the following data, determine the formula of the hydrated salt.

$CuCl_2 \cdot [?] H_2O$ — How many mol of H_2O?

Mass of crucible (g)	38.44 g	
Mass of crucible & hydrate (g)	45.26 g	- 38.44g = 6.82g mass of hydrate
Mass of crucible & contents after first heating (g)	43.90 g	
Mass of crucible & contents after second heating (g)	43.82 g	- 38.44g = 5.38g mass of anhydrate

(anhydrate ... all H_2O gone)

$CuCl_2$: Cu = 63.5, Cl = 2(35.5), = 134.5g/mol

② mass of H_2O = mass of hydrate - mass of anhydrate
= 6.82g - 5.38g = 1.44g

③ mol of anhydrate: 5.38g × 1mol / 134.5g = 0.04 mol ÷0.04 (simplify) = 1

mol of water: 1.44g × 1 mol / 18.0g = 0.08mol ÷0.04 = 2

mole ratio $CuCl_2 : H_2O$ = 1 : 2

Formula: $CuCl_2 \cdot 2H_2O$

Simple Dilution: $M_iV_i = M_fV_f$ (adding water)

13. What is the concentration of the solution prepared by dissolving 71.6 g of $Ca(OH)_2$ in water up to a final volume of 750.0 mL?

$C = M = \frac{n}{V}$

① mass → (molar mass) → moles

② mL → L: 750.0mL × 10^{-3}L / 1mL = 0.75L

solution + water: vol. ↑ conc. ↓

① 71.6g × 1mol / 74.1g = 0.96626 mol $Ca(OH)_2$

molar mass Ca = 40.1, O = 2(16.0), H = 2(1.0), = 74.1g

③ $C = \frac{n}{V} = \frac{0.96626mol}{0.75L} = 1.288346...$

$[Ca(OH)_2] = 1.29$ M

mol → mass(g)

14. How many grams of NaCl must be used to produce 1.50 L of a 0.850 M solution?

① $C = \frac{n}{V}$ ∴ $n = C \cdot V = (0.850\,M)(1.50\,L) = 1.275$ mol NaCl

② 1.275 mol NaCl → mass?

$\frac{1.275\,mol \mid 58.5\,g}{\mid 1\,mol} = 74.5875 = $ **74.6 g NaCl** (3 sf.)

15. What is the final concentration when 150.0 mL of water is mixed with 250.0 mL of 1.250 M KCl solution?

(+ water; M_f; V_i, M_i; 1 solution)

Simple Dilution: $M_iV_i = M_fV_f \Rightarrow M_f = \frac{M_iV_i}{V_f}$

Total Volume: 250.0 mL + 150.0 mL

16. Calculate the concentration of the solution that results when 200.0 mL of 0.500 M $Mg(NO_3)_2$ is mixed with 175.0 mL of 0.230 M $Mg(NO_3)_2$.

(solution A; solution B; "solution C")

Complex Dilution:
- 2 volumes
- 2 concentrations
- 1 type of chemical species (both solutions are $Mg(NO_3)_2$)

$M_f = \frac{(M_AV_A) + (M_BV_B)}{(V_A + V_B)}$

17. Determine the ion concentrations when 325.50 mL of 0.625 M $MgCl_2$ solution is mixed with 416.11 mL of 0.356 M RbCl solution. *(Be sure to start with the two dissociation equations)*

$[Mg^{2+}]$ $[Cl^-]$

* 2 diff types of chemical species — NOT a complex dilution.

$[Rb^+]$ $[Cl^-]$

$MgCl_2$ 325.50; RbCl 416.11 mL; Total Volume = 741.61 mL; conc.?

① $MgCl_{2(aq)} \rightarrow Mg^{2+}_{(aq)} + 2Cl^-_{(aq)}$ (1 : 1 : 2) — 0.274 M, 0.274 M, 0.549 M

$RbCl_{(aq)} \rightarrow Rb^+_{(aq)} + Cl^-_{(aq)}$ (1 : 1 : 1) — 0.200 M, 0.200 M, 0.200 M

② $[MgCl_2]_f = \frac{(0.625\,M)(325.50\,mL)}{(741.61\,mL)}$

$[MgCl_2]_f = 0.274\,M$

$[RbCl]_f = \frac{(0.356\,M)(416.11\,mL)}{(741.61\,mL)}$

$[RbCl]_f = 0.19974 = 0.200\,M$

③ Final ion conc. $[Mg^{2+}] = 0.274\,M$

$[Cl^-] = 0.549\,M + 0.200\,M = 0.749\,M$

$[Rb^+] = 0.200\,M$

1. Check your test definitions carefully against those given in the notes and/or text; be very critical of the quality of your answers.
2. 424.9 g
3. 2430 g
4. 0.551 mol
5. 4.11 g/L
6. Given that both samples have masses of 42.0 g, calculate the number of atoms in each by using the strategy: mass → moles of element → number of atoms of the element. Make a statement to summarize what you have demonstrated.
7. 18.3 L
8. 1.01×10^{23} atoms
9. 17.7% N, 6.3% H, 15.2% C and 60.8% O
10. a) $C_3H_5Cl_3$ b) $C_6H_{[illegible]}Cl_5$
11. 0.833g
12. $CuCl_2 \cdot 2H_2O$
13. 1.29 M
14. 74.6 g
15. 0.7813 M
16. 0.374
17. $[Mg^{+2}]$ = 0.274M
    $[Rb^+]$ = 0.200M
    $[Cl^-]$ = 0.748M