Unit E

The Mole and more

Review:

Percent Composition by mass

parts per 100 parts

What is the composition of a molecule?

What is part of the total mass is one element, what part of the total mass is the other element, and so on?

$$\frac{part}{total}\times 100 = Percent\ Composition\ by\ Mass$$

Analytical Chemists

- Lab analysis to experimentally determine the percent composition of substances.
- **Mass** percentage of each element in a compound.

- A lab will experimentally "*certify*" the values as an indication of the accuracy of the analysis.

Lets review how to calculate the mass of a compound.

Using the Masses in the Periodic Table

- Let's determine the mass (in grams) of magnesium fluoride, MgF_2

Mass - Using the Masses in the Periodic Table

- The mass (in grams) of magnesium fluoride
- Determine the mass of MgF_2
  - ✓ 1 Mg = 24.31 g
  - ✓ 1 F = 19.00 g

Mass - Using the Mass in the Periodic Table

- The mass (in grams) of magnesium fluoride
- Determine the mass of MgF_2

✓ 1 Mg = 24.31 g

✓ 2 F = 2 (19.00 g) = 38 g

✓ 24.31 g + 38 g = 62.31 g of MgF_2 formula units

"ionicules"

Sometimes in chemistry it is useful to report the amount of each element in mass percentages for a compound.

Percent Composition (by mass)
aka Elemental Analysis

- Determine the *theoretical* % composition by **mass**
  - ✓(aka an elemental analysis of MgF_2)
- *Remember that % is always part out of total*
- What part of the mass is magnesium? and what part of the mass is fluorine?

Percent Composition (by mass) *aka* Elemental Analysis

- Determine the *theoretical* % composition by mass
  - ✓(aka an elemental analysis of MgF_2)
- Mg: 24.31 g, F = 19.00 g
- MgF_2 = 62.31 g
- *Remember that % is always part out of total*
- Mg % calc $\frac{24.31}{62.31} \times 100 = 39.01\%$
- F (all of it) = 60.99% by subtraction or $\frac{38}{62.31} \times 100 = 60.99\%$

So what can we do with the percentages from elemental analysis?

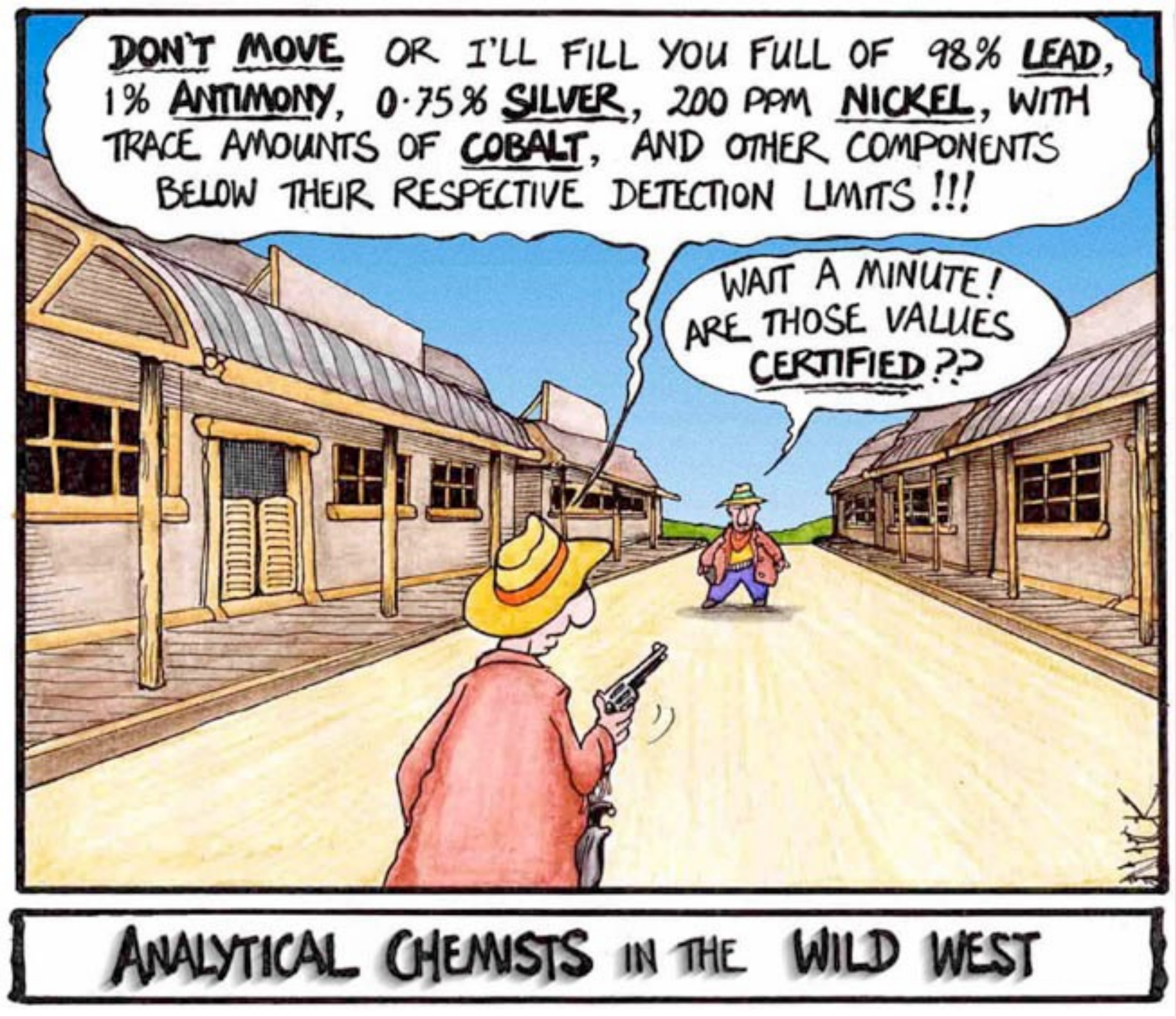
- Knowing MgF_2
  - ✓ is 61.0% F
  - ✓ and 39.0% Mg
- Given 0.483 g of magnesium fluoride, what is the mass of magnesium you could recover?

So what can we do with the percentages from elemental analysis?

- Knowing MgF_2
  - ✓ is 61.0% F
  - ✓ and 39.0% Mg
- Given 0.483 g of magnesium fluoride, what is the mass of magnesium you could recover?

$$39.0\%\ is \Rightarrow 0.310 \times 0.483 g = 0.188 gMg$$

Elemental Analysis



Mass of Large Compounds

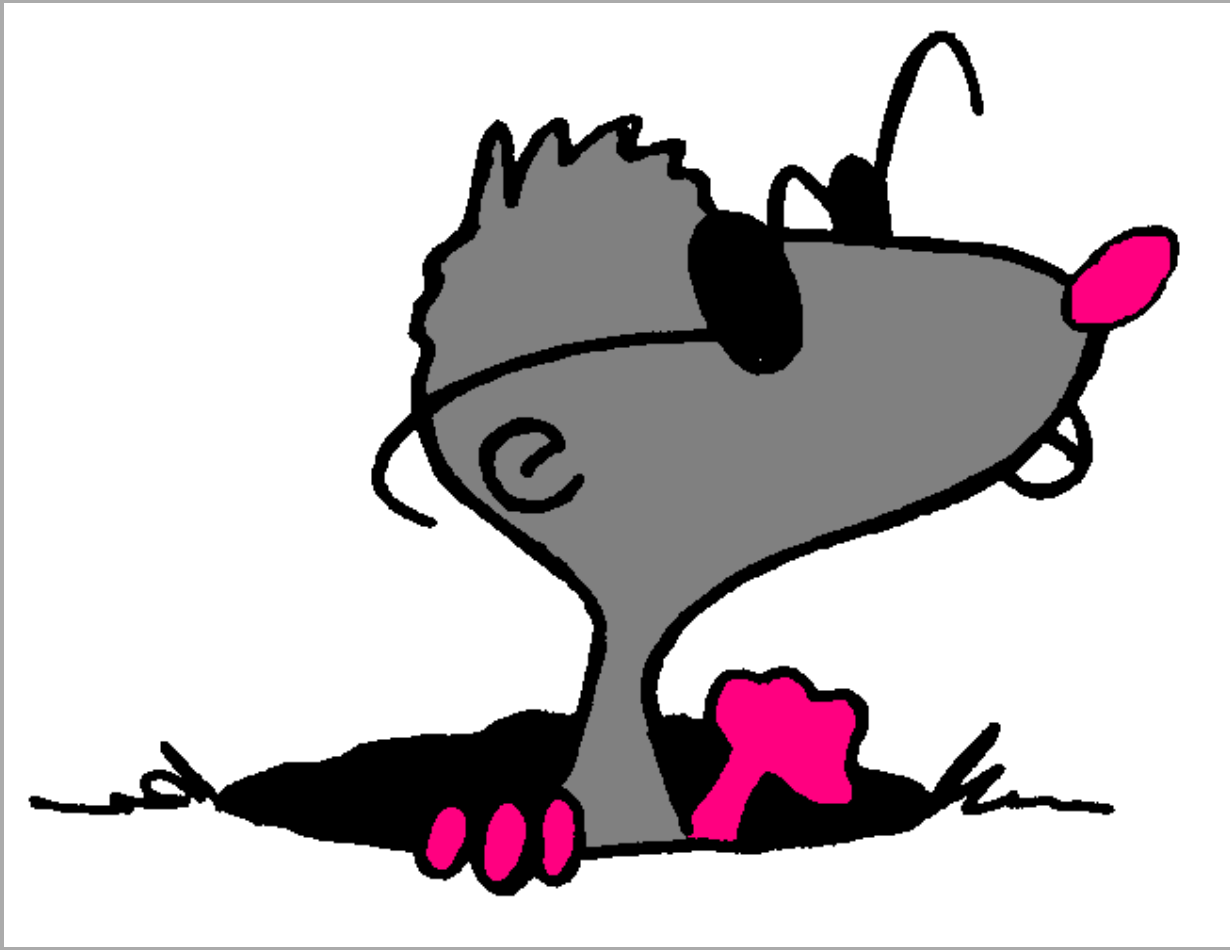
- Determine the MM (molar mass) of iron(III) sulfate: $Fe_2(SO_4)_3$

Mass of Large Compounds

- Determine the MM (molar mass) of iron(III) sulfate: $Fe_2(SO_4)_3$
  - ✓2 mole Fe = 2 (55.85 g) = 111.70 g
  - ✓3 mole S = 3 (32.07 g) = 96.21 g
  - ✓12 mole of O = 12 (16.00) = 192.00g
  - ✓111.70 g + 96.21 g + 192.00 g =
    - ✷399.91 g of iron(III) sulfate ionicules

- But what is the 400 g of iron(III) sulfate? Surely this is not the mass of one single ionicule? After all these atoms, ions and thus ionicules are VERY very very very small.

Enter....the mole



A quantity useful for counting very small things

I say, “I have a dozen....”

- You say....
  - ✓A dozen *what*?
  - ✓eggs, donuts, pencils, dogs, siblings....whatever.
- What is a dozen?
  - ✓12 items
- What unit label would you put on the #12 ?
  - ✓12 items per 1 dozen $\frac{12\ items}{1\ dozen}$

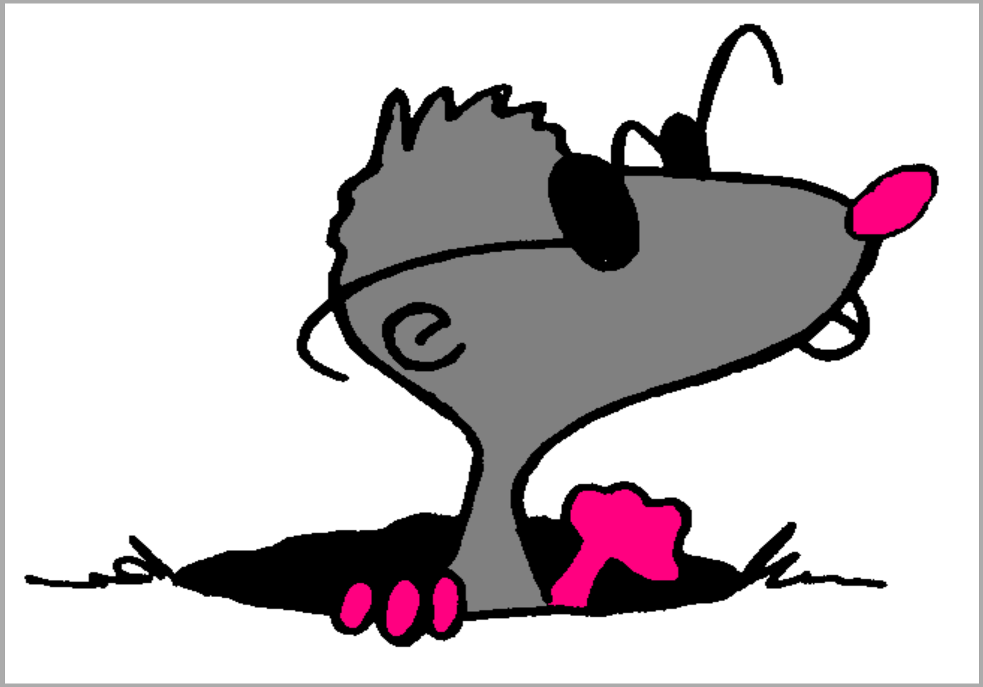
I'll give you a dozen rice…

- You might say, "Big deal."
- This would not be useful because rice grains are so small.
- You would still be hungry.

OK, I'll give you a dozen gold atoms…

- You might again say, "So what."
- This would not be valuable because a dozen atoms is so small, so in chemistry we need a much LARGER dozen.
- Introducing...

Introducing the *really big* dozen



The Mole

The Chemist's Dozen - A Mole Avogadro's Number of "items"

Lorenzo Romano Amedeo Carlo Avogadro

Italian Chemist
1776-1856

$$\frac{6.02\times10^{23}\,items}{1\,mol} \quad OR \quad \frac{1\,mol}{6.02\times10^{23}\,items}$$

He never knew the number that has been given his name.

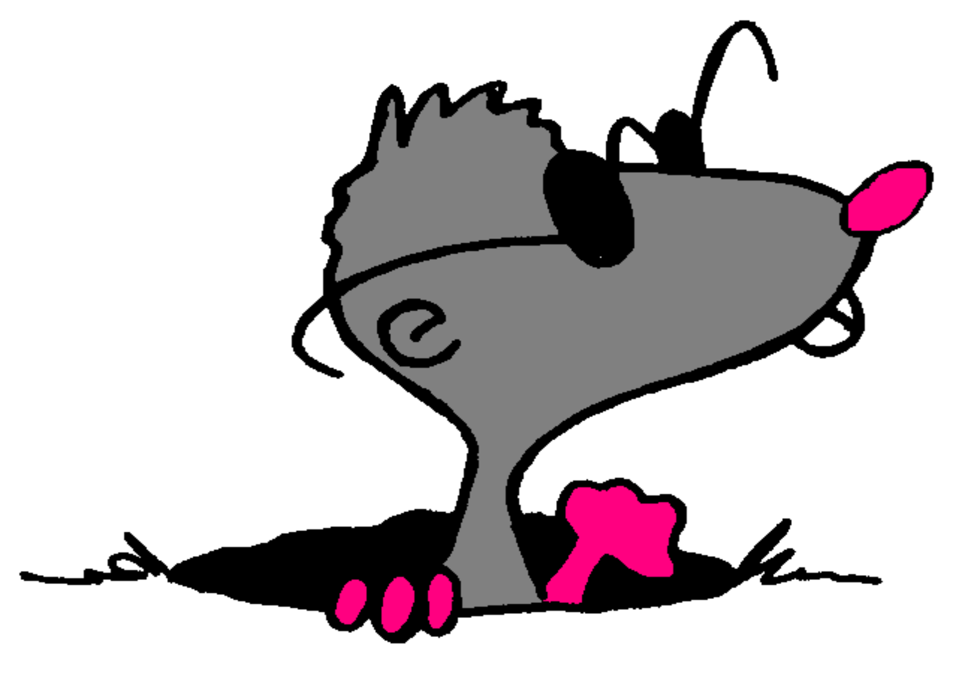
The Chemist's Dozen - A Mole
Avogadro's Number of "items"

- 6.022 141 99 x 10^{23} items in one mole
  - ✓Atoms, molecules, ions, formula units, or whatever (electrons, dogs, donuts, eggs, bicycles, etc)

sextillions quintillions quadrillions trillions billions millions thousands

- 602,214,199,000,000,000,000,000 items/mol
- The average atomic mass of any element is the mass of *a mole of atoms* of that element.
  - ✓ Hydrogen atoms, 1mole of atoms = 1.01 grams
  - ✓ Carbon atoms, 1 mole of atoms = 12.01 grams
  - ✓ Iron atoms, 1 mole of atoms = 55.85 grams

Mole calculations



Using the In-class Sheet

So what is the unit label on the molar masses that you look up in the periodic table?

- grams of one mole of substance
- grams per 1 mole

from the Periodic Table

$$\frac{\# g}{1mol} \quad OR \quad \frac{1mol}{\# g}$$

So what is the unit label on Avogadro's number? 6.02×10^{23}

- the number of items in one mole
- items per 1 mole

$$\frac{6.02\times10^{23}\,items}{1\,mol} \quad OR \quad \frac{1\,mol}{6.02\times10^{23}\,items}$$

- In chemistry, those items are most likely to be atoms, molecules, ions, formula units, or maybe even electrons.

Mass ↔ Moles

- How many moles of Mg in 75.0 g of Mg?

Mass ↔ Moles

- How many moles of Mg in 75.0 g of Mg?
  - ✓ from the periodic table.....
  - ✓ 1 mole Mg = 24.31 g
- Put the # 75.0 g on your paper
  (way over on the left)
- Use the equality above (in purple) to imagine the following conversion factors,

$$\frac{1mol}{24.31g} \quad \text{OR} \quad \frac{24.31g}{1mol}$$

- then multiply one of them by the 75 grams so that grams will cancel out.

Mass ↔ Moles

- How many moles of Mg in 75.0 g of Mg?
  - ✓ 1 mole Mg = 24.31 g
  - ✓ $75g \times \frac{1mol}{24.31g} = 3.085150144mol$
    - ✴ label and round off, how many sig figs?
      - three sig figs, the 1 in the conversion factor is as significant as you need it to be since the 1 is just the definition of the other quantity.
      - 3.09 moles of Mg atoms

Mass ↔ Moles

- What is the mass of 0.732 mole of Cl_2 molecules?

Mass ↔ Moles

- What is the mass of 0.732 mole of Cl_2 molecules?
  - ✓ look on the periodic table to find out that 1 mole Cl atoms = 35.45 g
  - ✓ 1 mole Cl_2 molecules = 70.90 g
- Put 0.732 mole on your paper (way over on the left)
- Use the equality above (in purple) to imagine the following conversion factors, $\frac{1 mol}{70.90 g}$ OR $\frac{70.90 g}{1 mol}$
- then multiply the 0.732 mole by one of these conversion factors so that mole will cancel out

Mass ↔ Moles

- What is the mass of 0.732 mole of Cl_2 molecules?
  - ✓ look on the periodic table to find out that 1 mole Cl atoms = 35.45 g
  - ✓ 1 mole Cl_2 molecules = 70.90 g
- Put 0.732 mole on your paper (way over on the left)

$$0.732 mol \times \frac{70.90 g}{1 mol} = 51.8988 g$$

- ✳ label and round off to how many sig figs?
  - three
  - 51.9 g of Cl_2 molecules

Moles ↔ Items

- How many atoms are in 4.7 moles of Mg?

Moles ↔ Items

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  - ✓ Just like a dozen is 12 items, we know that
  - ✓ 1 mol Mg = 6.02×10^{23} atoms

Moles ↔ Items

- How many atoms are in 4.7 moles of Mg?
  - ✓ Just like a dozen is 12 items, we know that
  - ✓ 1 mol Mg = 6.02 x 10^{23} atoms
- Put 4.7 moles on your page
  (way over on the left)
- then use the equality above (in purple) to imagine the conversion factors shown below,

$$\frac{1 mol}{6.02\times10^{23} atoms} \quad \text{OR} \quad \frac{6.02\times10^{23} atoms}{1 mol}$$

- then multiply one or the other by the 4.6 moles so that mole will cancel out

Moles ↔ Items

- How many atoms are in 4.7 moles of Mg?
  - ✓ 1 mol Mg = 6.02 x 10^{23} atoms
- Put 4.7 moles on your page
  (way over on the left)

✓ $$4.7mol \times \frac{6.02 \times 10^{23} atoms}{1mol} = 2.8294 \times 10^{24} atoms$$

- ✳ label and round off to how many sig figs?
  - two
  - 2.8 x 10^{24} atoms of Mg

Moles ↔ Items

- How many mole of molecules are 5.83×10^{22} oxygen molecules?
- Just like you might ask, “How many dozen eggs is 30 eggs?”

Moles ↔ Items

- How many mole of oxygen molecules in 5.83×10^{22} molecules?

✓We know that....

✓1 mole O_2 = 6.02×10^{23} molecules

- Put 5.83×10^{22} molecules on your page
  (way over on the left)
- then use the equality above (in purple) to imagine the conversion factors shown below,

$$\frac{1 mol}{6.02\times10^{23} molecules} \quad \text{OR} \quad \frac{6.02\times10^{23} molecules}{1 mol}$$

- then multiply one or the other by the
  5.83×10^{22} molecules so that molecules will cancel out

Moles ↔ Items

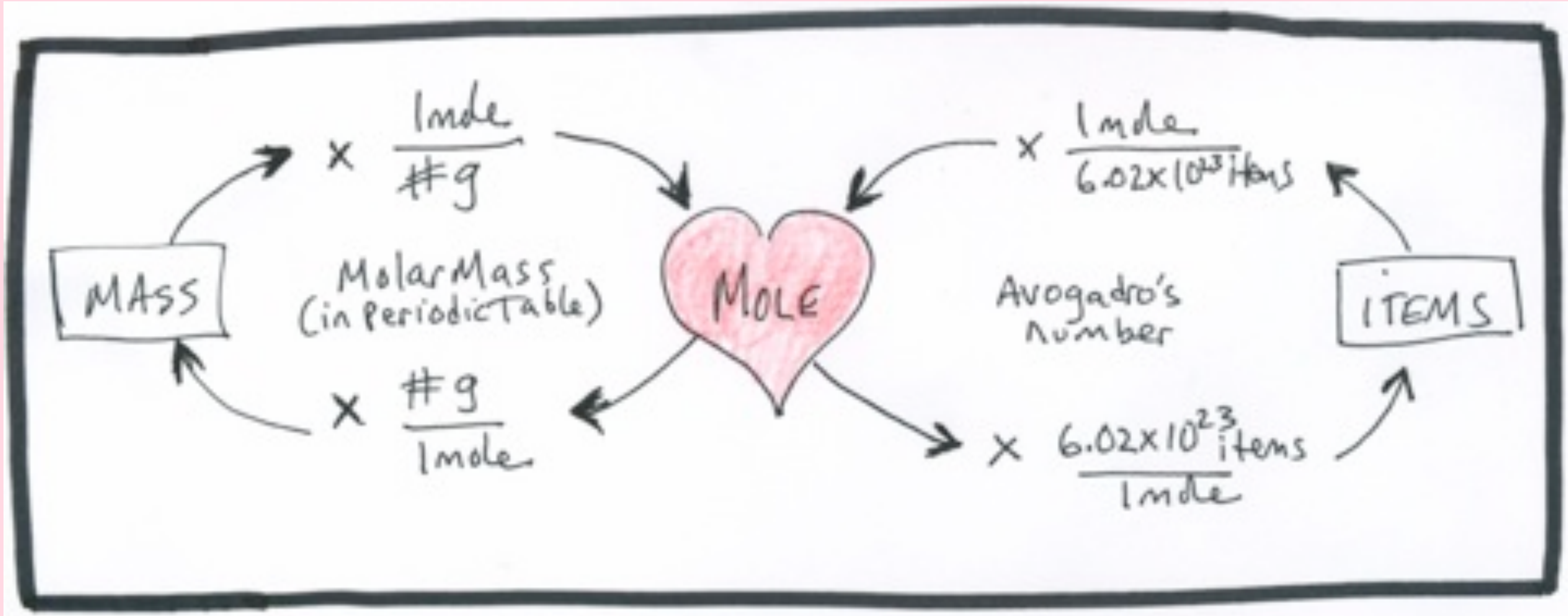
- How many mole of oxygen molecules in 5.83×10^{22} molecules?

✓ 1 mole O_2 = 6.02×10^{23} molecules

✓ $$5.83 \times 10^{22} molecules \times \frac{1 mol}{6.02 \times 10^{23} molecules} = 0.096843853 mol$$

- ✳ label and round off to how many sig figs?
  - three
  - 0.0968 mole of O_2 molecules

The mass-mole-items Road Map



Mass ↔ Moles ↔ Items

- What is the mass of 2.8×10^{24} atoms of Mg?
  - ✓ This is a two step process:
  - ✓ Change to moles, then change to grams
  - ✓ You're going to need TWO conversion factors

$$\frac{1 mol}{6.02 \times 10^{23} atoms} \quad \text{OR} \quad \frac{6.02 \times 10^{23} atoms}{1 mol}$$

AND

$$\frac{1 mol}{24.31 g} \quad \text{OR} \quad \frac{24.31 g}{1 mol}$$

Mass ↔ Moles ↔ Items

- What is the mass of 2.8×10^{24} atoms of Mg?
- Put 2.8×10^{24} molecules on your page
  (way over on the left)
- then multiply by the appropriate conversion factors so molecules cancel out...and then mole cancel out.

$$\frac{1mol}{6.02\times10^{23}\,atoms} \quad \text{OR} \quad \frac{6.02\times10^{23}\,atoms}{1mol}$$

AND

$$\frac{1mol}{24.31g} \quad \text{OR} \quad \frac{24.31g}{1mol}$$

Mass ↔ Moles ↔ Items

- What is the mass of 2.8×10^{24} atoms of Mg?
  - ✓ Two step process:
  - ✓ Change to moles, then change to grams
  - ✓ $$2.8 \times 10^{24}\,atoms \times \frac{1mol}{6.02 \times 10^{23}\,atoms} \times \frac{24.31g}{1mol} = 113.0697674g$$
    - ✳ label and round off to:
      - 110 g of Mg atoms (only 2 sig figs)

Mass ↔ Moles ↔ Items

- How many ammonia molecules are in 34 g of NH_3?

✓Again, This is a two step process:

✓Change to moles, then change to molecules

- Put 34 g of NH_3 on your page

(way over on the left)

- then imagine some useful conversion factors that will cancel out grams ... and cancel out mole.

$$\frac{1mol}{17.04g} \quad \text{OR} \quad \frac{17.04g}{1mol}$$

AND

$$\frac{1mol}{6.02\times10^{23}molecules} \quad \text{OR} \quad \frac{6.02\times10^{23}molecules}{1mol}$$

Mass ↔ Moles ↔ Items

- How many ammonia molecules are in 34 g of NH_3?

  ✓Again, This is a two step process:

  ✓Change to moles, then change to molecules

  ✓ $$34g \times \frac{1mol}{17.04g} \times \frac{6.02 \times 10^{23} molecules}{1mol} = 1.201173709 \times 10^{24} molecules$$

  ✳ label and round off to

  ✳ 1.2×10^{24} molecules of NH_3

- But wait... one more question...How many atoms is this??

- First you must ask yourself, how many atoms in NH_3?

Mass ↔ Moles ↔ Items

- So for.....1.2 x 10^{24} molecules of NH_3
- One more question...How many atoms is this??
- You must ask yourself, how many atoms in NH_3?
  - ✓ 4 atoms in NH_3
  - ✓ which can be written as yet two more conversion factors $\frac{1\ molecule}{4\ atoms}$ or $\frac{4\ atoms}{1\ molecule}$
  - ✓ Choose the appropriate conversion factor.....

$$1.2\times10^{24}\,molecules\times\frac{4atoms}{1molecule}=4.8\times10^{24}\,atoms$$

Moles ↔ molecules ↔ atoms

- How many atoms of oxygen in 13.0 moles of sulfuric acid?

Moles ↔ molecules ↔ atoms

- How many atoms of oxygen in 13.0 moles of sulfuric acid?

H_2SO_4

$$13moles \times \frac{6.02\times10^{23} molecules}{1mol} \times \frac{4 OxygenAtoms}{1molecule} = 3.13\times10^{25} atoms$$

Mass of Large Compounds

- Determine the MM (molar mass) of iron(III) sulfate: $Fe_2(SO_4)_3$
  - ✓2 mole Fe = 2 (55.85 g) = 111.70 g
  - ✓3 mole S = 3 (32.07 g) = 96.21 g
  - ✓12 mole of O = 12 (16.00) = 192.00g
  - ✓111.70 g + 96.21 g + 192.00 g =
    - ✳399.91 g of iron(III) sulfate ionicules
- But what is the 400 g of iron(III) sulfate? Surely this is not the mass of one single ionicule? After all these atoms, ions and thus ionicules are VERY very very very small.

Mass $\leftrightarrow$ Moles $\leftrightarrow$ molecules $\leftrightarrow$ atoms

- How many atoms of sulfur in 25.0 grams of $Fe_2(SO_4)_3$?

Mass ↔ Moles ↔ molecules ↔ atoms

- How many atoms of sulfur in 25.0 grams of $Fe_2(SO_4)_3$?
- This is a three step process....
- Put 25.0 g on your page
  (way over on the left)
- then pick out the appropriate conversion factors to cancel out so grams ... and cancel out moles ...and cancel out molecules.

Mass ↔ Moles ↔ Items

- How many atoms in 25.0 grams of $Fe_2(SO_4)_3$?
  - ✓ Change grams to moles, then to molecules, then change to atoms
  - ✓ 1 mole $Fe_2(SO_4)_3$ molecules has 3 moles of oxygen atoms

$$25.0 grams \times \frac{1 mol}{400 g} \times \frac{6.02 \times 10^{23} molecules}{1 mol} \times \frac{3 oxygenAtoms}{1 molecule} = 1.12875 \times 10^{23} atoms$$

- ✳label and round off to (sig figs?):
  - three
  - 1.13×10^{23} atoms of oxygen in 25.0 g of sulfuric acid

Mass Ratio (or percent)
REVIEW if needed

- If I had 3.60 g of aluminum and I reacted the aluminum completely with oxygen, what mass of aluminum oxide would I end up with?

Al_2O_3 102 g/mol

Calculating Molar Mass
REVIEW if needed

- What is the identity of some noble gas for which 5.00 g of this gas contains 3.59×10^{22} atoms

Determine the molar mass of this compound?
Molar mass is the mass per 1 mole of any substance.
We're halfway there….we have a mass, all we need is moles?

$$3.59 \times 10^{22}\, atoms \times \frac{1 mol}{6.02 \times 10^{23}\, atoms} = 0.0596 mol$$

$$\frac{5g}{0.0596 mol} = 83.8 g / mol \quad Kr$$

$$\frac{5g}{3.59 \times 10^{22}\, atoms} \times \frac{6.02 \times 10^{23}\, atoms}{1 mol} = 83.8 g / mol \quad Kr$$

Not just another pretty face....

The Mole

Determining an Empirical Formula

Empirical Formula:

The LOWEST whole number ratio of elements in a compound

Break out E2 - IN CLASS Practice

LAD D5 - Law of Constant Composition

- In this LAD we burned Mg in order to determine if the Mg/O ratio was constant.
- Lets suppose we don't know the formula for magnesium oxide. If Mg were a transition metal we would not know its charge.
- Lets calculate what the formula is; MgO, MgO_2, Mg_3O….etc
- First, how to calculate the mass of oxygen?

Eldon's Data	
mass magnesium (g)	0.483
mass magnesium oxide (g)	0.800
mass of oxygen (g)	0.317

LAD D5- Determining Empirical Formula

- Convert magnesium to moles $$0.483\times\frac{1mol}{24.31g}=0.01987mol$$
- Convert oxygen to moles $$0.317\times\frac{1mol}{16.0g}=0.0198mol$$
- Analyze the mole ratio to determine the simplest whole number ratio. If the numbers aren't so obvious, divide each mole value by the smaller of the two.

$$\frac{0.01987mol}{0.0198}=1.004 \qquad \frac{0.0198mol}{0.0198}=1$$

- thus the empirical formula: MgO
- Lowest whole number ratio 1:1

Eldon's Data	
mass magnesium (g)	0.483
mass magnesium oxide (g)	0.800
mass of oxygen (g)	0.317

What is an empirical formula?

- A chemical formula represented in the lowest whole number ratio
  - ✓ $BaCl_2$ is an empirical formula
  - ✓ PbN_2O_6 $Pb(NO_3)_2$ is an empirical formula
  - ✓ All ionic compounds are always written as empirical formulae

Law of Constant Composition

- In our LAD D5 we burned Mg in order to determine if the Mg/O ratio was constant.
- Lets suppose we repeated the lab burning nickel instead
- Lets calculate the formula of this nickel oxide compound.
  - ✓ NiO, NiO_2, Ni_3O, ….etc
- First calculate the mass of oxygen combined

Greta's Data	
mass nickel (g)	0.387
mass nickel oxide (g)	0.545
mass of oxygen (g)	

Determining Empirical Formula

- Convert nickel to moles $$0.387 \times \frac{1mol}{58.68g} = 0.006595mol$$

- Convert oxygen to moles $$0.158 \times \frac{1mol}{16.0g} = 0.009875mol$$

- Analyze the mole ratio to determine the simplest whole number ratio. If the numbers aren't so obvious, divide each mole value by the smaller of the two.

$$\frac{0.009875mol}{0.006595} = 1.5$$ x2 = 3

$$\frac{0.006595mol}{0.006595} = 1$$ x2 = 2

- Wait...this isn't a whole # ratio?!
- when ratio turns up with .5 multiply by 2
- thus the empirical formula: Ni_2O_3

Greta's Data	
mass nickel (g)	0.387
mass nickel oxide (g)	0.545
mass of oxygen (g)	0.158

If you were asked to determine an empirical formula, it's possible that someone may give you the quantities of each element as percentages.

Just as you can convert mass values into percentage quantities.

- Treat the % values as mass values.
  - ✓ Since % just means "out of 100"
  - ✓ We can "pretend" we have 100 g of the material, thus each item would be its % as a gram value.

An iron, oxygen, hydrogen compound was analyzed to be 52.2 % iron, 44.9 % oxygen and the rest hydrogen.

Determine the empirical formula for this ionic compound.

- First, determine the % of hydrogen.
  - ✓ 100% − 52.2 − 44.9 = 2.9%

- Remember, treat the % values as mass values.
  - ✓ Since % just means “out of 100”
  - ✓ We can “pretend” we have 100 g of the material, thus each item would be its % as a gram value.

An iron, oxygen, hydrogen compound was analyzed to be 52.2 % iron, 44.9 % oxygen and the rest hydrogen. Determine the empirical formula for this compound.

- Change each mass to moles

  - ✓ Fe: 52.2 g $52.2g \times \frac{1mol}{55.8g} = 0.935mol$

  - ✓ O: 44.9 g $44.9g \times \frac{1mol}{16.0g} = 2.81mol$

  - ✓ H: 2.9 g $2.9g \times \frac{1mol}{1.01g} = 2.87mol$

An iron, oxygen, hydrogen compound was analyzed to be 52.2 % iron, 44.9 % oxygen and the rest hydrogen. Determine the empirical formula for this compound.

- Divide by the mole value by the smallest mole value
  - ✓ Fe: 0.935 $\frac{0.935}{0.935} = 1$
  - ✓ O: 2.81 $\frac{2.81}{0.935} = 3.0$
  - ✓ H: 2.87 $\frac{2.87}{0.935} = 3.07$
- Round to the nearest whole number...
  - ✓ Voilà: FeO_3H_3

More Practice *if needed*

A compound is determined to be 38.76 % calcium, 19.97 % phosphorus and the rest oxygen.

Determine the empirical formula.

A compound is determined to be 38.76 % calcium, 19.97 % phosphorus and the rest oxygen. Determine the empirical formula. Then name this compound.

- Determine the % of oxygen.
  - ✓ 100% - 38.76% - 19.97% = 41.27%
- Treat the % values as mass values and change to moles
  - ✓ Ca: 38.76 g $$38.76g \times \frac{1mol}{40g} = 0.969mol$$
  - ✓ P: 19.97 g $$19.97g \times \frac{1mol}{31g} = 0.64mol$$
  - ✓ O: 41.27 g $$41.27g \times \frac{1mol}{16g} = 2.58mol$$

A compound is determined to be 38.76 % calcium, 19.97 % phosphorus and the rest oxygen. Determine the empirical formula.

- Divide by the smallest mole value

  - ✓ Ca: 0.969 mol $\dfrac{0.969\,mol}{0.64\,mol} = 1.5$
  - ✓ P: 0.64 mol $\dfrac{0.64\,mol}{0.64\,mol} = 1$
  - ✓ O: 2.58 mol $\dfrac{2.58\,mol}{0.64\,mol} = 4$

A compound is determined to be 38.76 % calcium, 19.97 % phosphorus and the rest oxygen. Determine the empirical formula.

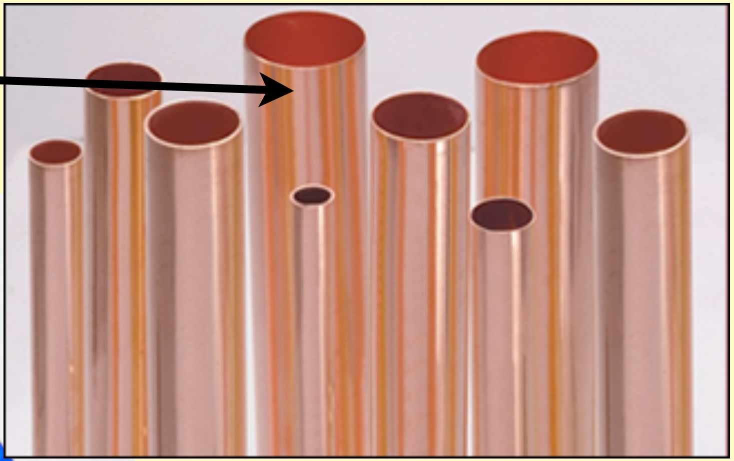
- Multiply them all by a factor of 2 to get whole numbers, then write the formula
  - ✓ Ca: $\frac{0.969 mol}{0.64 mol} = 1.5$
  - ✓ P: $\frac{0.64 mol}{0.64 mol} = 1$
  - ✓ O: $\frac{2.58 mol}{0.64 mol} = 4$
  - ✓ and Voilà: $Ca_3P_2O_8$

Hydrates

Percent Composition and Empirical Formula

Copper atoms & ions

- Color of copper?
  - ✓ copper colored
- Color of the compound?
  - ✓ blue
- Copper ions (when water is present) are various shades of blue.



- Why is the statue of liberty blue/green?
  - ✓ She's made out of copper and acid rain has produced copper ions on the surface.

Demonstrate the removal of water from copper(II) sulfate __ hydrate

- Color of the compound?
- Color after removal of water?
- Color after the water is put back?
- Copper + ions in the presence of water are blue color.

What is a hydrate?

- An ionic compound that contains water inside its crystal structure.
  - ✓ These compounds do not look or feel wet.
  - ✓ The water molecules are loosely attached and can be removed by heating.
- Generic Formula $M_xNm_y \bullet nH_2O$
  - ✓ $M_xNm_y \bullet nH_2O$ hydrate
  - ✓ M_xNm_y anhydrate
  - ✓ $M_xNm_y \bullet nH_2O$ water

"loosely attached" symbolized by the •

How do we analyze a hydrate?

- The water molecules are loosely attached and can be removed by heating.

This Δ (delta) over the $\longrightarrow$ means heat

- $M_xNm_y \bullet nH_2O \xrightarrow{\Delta} M_xNm_y + nH_2O$

  ✓ The remaining anhydrate (anhydrous crystals) may look the same, may have a different texture, or may even be a different color.

  ✓ You analyzed a hydrate in LAD C5 - SG

Let’s Calculate Hydrates

Break out E3 IN CLASS Practice

LAD D2 - Law of Constant Composition

- In this Lab we measured SG
- We heated SG to drive off G
- We measured the mass of S
- We calculated the mass of G…how?
  - ✓ Calculate it now on your E3 IN-CLASS Practice
- We calculated the **experimental mass ratio** of S to G…how?
- Calculate it now on your E3 IN-CLASS Practice

Gertrude & Alice's Data	
mass SG (g)	5.658
mass S (g)	3.061
mass G (g)	2.597
mass ratio S/G	1.18

LAD D2 - Law of Constant Composition

- This data could be analyzed in another way.
- Calculate the **percentage of G** in SG.... how?
  - ✓ Try it now on your E3 IN CLASS Practice

Gertrude & Alice's Data	
mass SG (g)	5.658
mass S (g)	3.061
mass G (g)	2.597
% G in SG	45.9

LAD D2 - Law of Constant Composition

- You may remember that S was actually
  - ✓ Potassium aluminum sulfate
  - ✓ $KAl(SO_4)_2$

Gertrude & Alice's Data	
mass SG (g)	5.658
mass S (g)	3.061
mass G (g)	2.597
% G in SG	45.9

- We can calculate the molar mass of this anhydrate.

$39.1+27+2(32.07)+8(16)=258.2$

LAD D2 - Law of Constant Composition

- Let's analyze the data in one more way by calculating the mole ratio of water to anhydrate
- Calculate mass of water then moles of water and then moles of anhydrate
  - ✓ *Do this on your E3-IN CLASS Practice*

$$3.061g \times \frac{1mol}{258.2g} = 0.01186 mol Anhydrate$$

$$2.597g \times \frac{1mol}{18.02g} = 0.1441 mol H_2O$$

Gertrude & Alice's Data	
mass $KAl(SO_4)_2 \bullet ?H_2O$ (g)	5.658
mass $KAl(SO_4)_2$ (g)	3.061
mass water (g)	2.597
moles of anhydrate	0.01186
moles of water	0.1441
mole ratio: water/anhydrate	

LAD D2 - Law of Constant Composition

- Now use our simple math trick to calculate a whole number ratio.
  - ✓ *Do this on your E3-IN CLASS Practice*

$$\frac{0.01186}{0.01186} = 1 \qquad \frac{0.1441}{0.01186} = 12.15$$

Gertrude & Alice's Data	
mass $KAl(SO_4)_2 \cdot ?H_2O$ (g)	5.658
mass $KAl(SO_4)_2$ (g)	3.061
mass water (g)	2.597
moles of anhydrate	0.01186
moles of water	0.1441
mole ratio water/anhydrate	12.15

Given that most hydrates are only 1 salt, to multiple moles of water, we will round to the nearest whole number, and NOT multiply by any factor.

Voila: $KAl(SO_4)_2 \cdot 12\ H_2O$

LAD F1

Formula of a hydrate

Practice in preparation of PreLab Questions

Page 2 of E3 IN CLASS Practice

2. Calculate the *theoretical* percentage of water in sodium acetate trihydrate.

- The formula for this hydrate is….
  - ✓ $NaC_2H_3O_2 \cdot 3\ H_2O$
- Using masses from the periodic table, add the mass of all the water
  - ✓ $3\ H_2O$ 3 [2(1.01) + 16)] = 54.06 g/mol
- Then add the mass of the entire compound
  - ✓ 22.99 + 2(12.01) + 3(1.01) + 2(16) + 3(18.02) = 136.1 g/mol
- Then calculate the *theoretical* water/hydrate percentage

$$\frac{54.06 g / mol}{136.1 g / mol} \times 100 = 39.7\%$$

3. If a student measured 5.49 g of cobalt(III) bromide hydrate, and 3.86 g of cobalt(III) anhydrate,

a. Calculate the mass of water removed

b. Calculate the moles of water

c. Calculate the moles of anhydrate

d. Calculate the whole number mole ratio of water to anhydrate

e. Write the correct formula, and give a complete name for this hydrate.

3. If a student measured 5.49 g of cobalt(III) bromide hydrate, and 3.86 g of cobalt(III) anhydrate,

a. Calculate the mass of water removed

- ✓ 5.49 - 3.86 = 1.42 g water removed

b. Calculate the moles of water $$1.42g \times \frac{1mol}{18g} = 0.0789 mol H_2O$$

c. Calculate the moles of anhydrate

- ✓ Formula and molar mass?
- ✓ $CoBr_3$ 58.93 + 3(79.9) = 298.63

$$3.86g \times \frac{1mol}{298.63g} = 0.0129 mol CoBr_3$$

d. Calculate the whole number mole ratio of water to anhydrate $$\frac{0.0789}{0.0129} = 6 mol H_2O$$

e. Write the correct formula, and give a complete name for this hydrate.

- ✓ Thus $CoBr_3 \cdot 6\ H_2O$ Cobalt(III) bromide hexahydrate

4. In the lab a student measured 8.98 g of a aluminum chlorate hydrate and after heating, the anhydrate weighed 7.05 g

a. Calculate the experimental percent of water in this hydrate

b. Later you were told that this compound is a tetrahydrate, calculate the theoretical percent of in this hydrate.

c. Do you suppose the error source for the experimental trial was caused by anhydrate falling out of the dish on the way to the balance, or drooling into the dish after heating the hydrate to dryness? Justify.

4. In the lab a student measured 8.98 g of a aluminum chlorate hydrate and after heating, the anhydrate weighed 7.05 g

a. Calculate the experimental percent of water in this hydrate

- ✓ 8.77 - 7.04 = 1.73 g water removed

$$\frac{1.73g}{8.77g} \times 100 = 19.7\%$$

b. Later you were told that this compound is a tetrahydrate, calculate the theoretical percent of in this hydrate.

- ✓ $Al(ClO_3)_3 \bullet 4H_2O$
- ✓ $4H_2O$ 4(18.02) = 72.08
- ✓ $Al(ClO_3)_3$ 26.98 + 3(35.45) + 9(16) + 4(18.02) = 349.41

$$\frac{72.08}{349.41g} \times 100 = 20.6\%$$

c. Do you suppose the error source for the experimental trial was caused by anhydrate falling out of the dish on the way to the balance, or drooling into the dish after heating the hydrate to dryness? Justify.

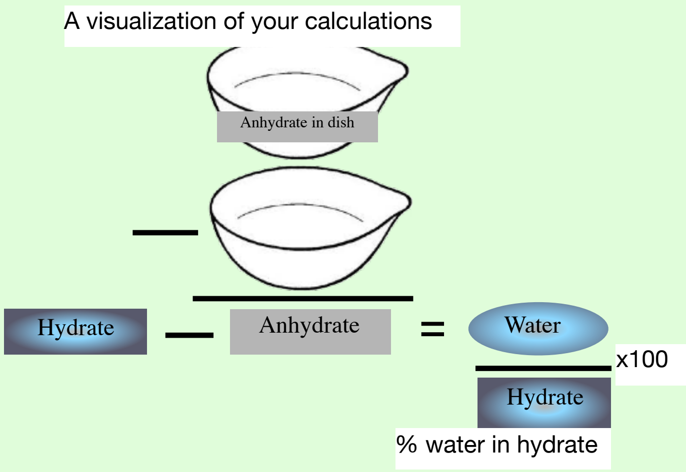
4. In the lab a student measured 8.77 g of a aluminum chlorate hydrate and after heating, the anhydrate weighed 7.04 g

c. Do you suppose the error source for the experimental trial was caused by anhydrate falling out of the dish on the way to the balance, or drooling into the dish after heating the hydrate to dryness? Justify.

$$Experimental : \frac{1.73g}{8.77g} \times 100 = 19.7\% \qquad Theoretical : \frac{72.08}{349.41g} \times 100 = 20.6\%$$

It is important to notice the the Exp is Larger than Theor

•



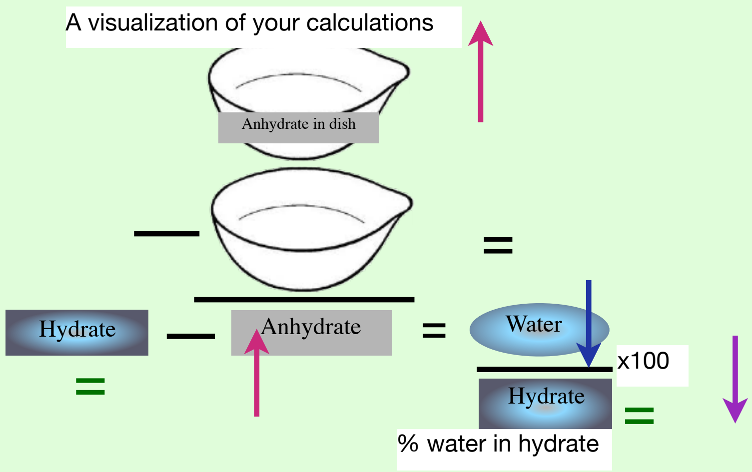
4. In the lab a student measured 8.77 g of a aluminum chlorate hydrate and after heating, the anhydrate weighed 7.04 g

c. Do you suppose the error source for the experimental trial was caused by anhydrate falling out of the dish on the way to the balance, or drooling into the dish after heating the hydrate to dryness? Justify.

$$Experimental: \frac{1.73g}{8.77g} \times 100 = 19.7\% \qquad Theoretical: \frac{72.08}{349.41g} \times 100 = 20.6\%$$

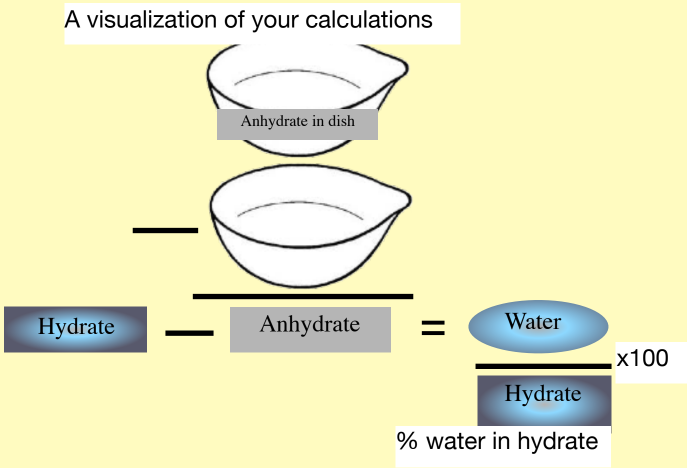
Exp is Larger than Theor

- If the mass of the anhydrate were be larger than it should be, when subtracted from the unchanged hydrate, the water would appear too small, resulting in an experimental percentage of water that appears smaller than the theoretical value.



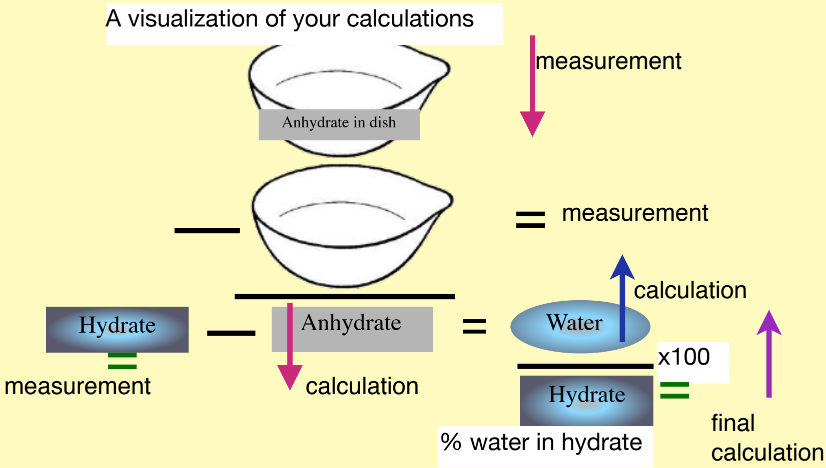
Post Lab Error Analysis Practice

- Suppose the student tripped and spilled some anhydrate on the way to the balance after heating the anhydrate to a constant mass?
- Would this source of error cause your percentage of water in the hydrate appear too small, too large, or have no effect?
- Follow those changes completely through the calculations.



Post Lab Error Analysis Practice

- Suppose the student tripped and spilled some anhydrate on the way to the balance after heating the anhydrate to a constant mass?

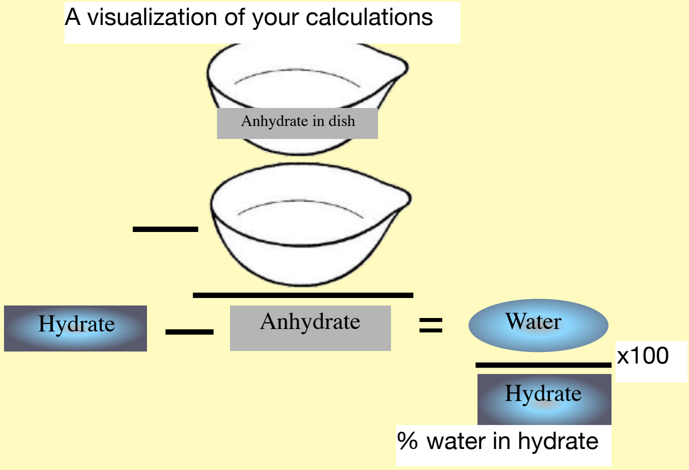


- Would this source of error cause your percentage of water in the hydrate appear too small, too large, or have no effect?

- The mass of the anhydrate would be smaller than it should be, and when subtracted from the unchanged hydrate, the water would appear too large, resulting in a percentage of water that would appear greater than expected.

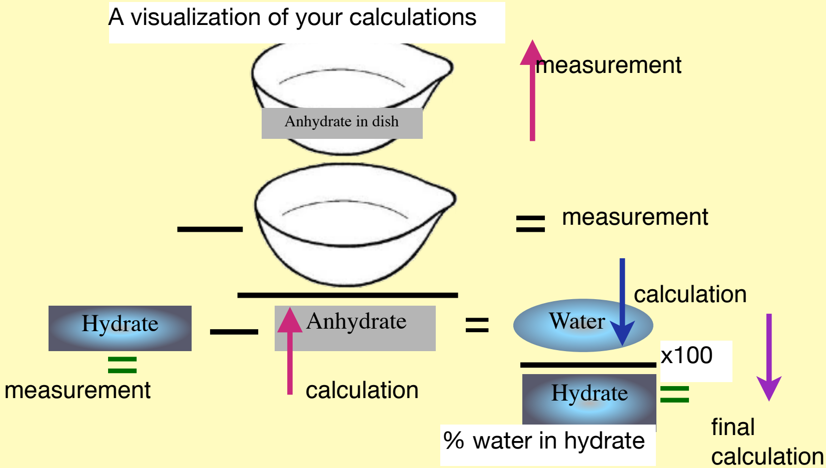
Post Lab Error Analysis Practice

- Suppose the student drooled into the dish on the way to the balance after heating the anhydrate to a constant mass?
- Would this source of error cause your percentage of water in the hydrate appear too small, too large, or have no effect?
- Follow those changes completely through the calculations.



Post Lab Error Analysis Practice

- Suppose the student drooled into the dish on the way to the balance after heating the anhydrate to a constant mass?



- Would this source of error cause your percentage of water in the hydrate appear too small, too large, or have no effect?

- The mass of the anhydrate would be larger than it should be, and when subtracted from the unchanged hydrate, the water would appear too small, resulting in a percentage of water that would appear smaller than expected.

Formula of a Hydrate

When we don't know
the empirical formula

#6 of E3 IN CLASS Practice

more F4 IN CLASS Practice (as needed)

5. Lets test a copper(II) sulfate hydrate.

- Given 8.986 g hydrate, heat the compound and the remaining anhydrate weighs 5.746g.
- Determine the mole ratio of water to anhydrate
- write the chemical formula
- give the complete name.

Formula of copper(II) sulfate hydrate

- Calculate the mass of water removed.
  - ✓ 8.986 - 5.746 = 3.243 g water removed
- Calculate moles of water $$3.243g\frac{1mol}{18g}=0.180molH_2O$$
- Calculate moles of anhydrate
  - ✓ Formula and molar mass?
  - ✓ $CuSO_4$ $$5.764g\frac{1mol}{159.6g}=0.036molCuSO_4$$
  - ✓ 63.55 + 32.07 + 4(16) = 159.6
- Determine the mole ratio water / anhydrate
  - ✓ $\frac{0.180}{0.036}=5molH_2O$
  - ✓ $CuSO_4 \cdot 5\ H_2O$
  - ✓ This is copper(II) **penta**hydrate

A hydrate was analyzed and determined to be 34% water.

Further the anhydrate was analyzed to be 23.96% nickel, 17.16% nitrogen, and 58.88% oxygen.

Determine the formula for this hydrate, and name it.

- First you must notice that this problem is a little bit different than some of the others that you may have tried.
- You are not told the chemical name of the anhydrate.
- Instead you are given information that will allow you to determine this, which must be done before moving on to the moles of water in the hydrate.

A hydrate was analyzed and determined to be 34% water. Further the anhydrate was analyzed to be 23.96% nickel, 17.16% nitrogen, and 58.88% oxygen.
Determine the formula for this hydrate, and name it.

- Treat the percentages mass and convert to moles
  - ✓ Ni: 23.96 g $$23.96g \times \frac{1mol}{58.69g} = 0.408mol$$
  - ✓ N: 17.16 g $$17.16g \times \frac{1mol}{14.01g} = 1.225mol$$
  - ✓ O: 58.88 g $$58.88g \times \frac{1mol}{16g} = 3.68mol$$

A hydrate was analyzed and determined to be 34% water. Further the anhydrate was analyzed to be 23.96% nickel, 17.16% nitrogen, and 58.88% oxygen.
Determine the formula for this hydrate, and name it.

- Divide each mole value by the smallest mole value
  - ✓ Ni: 0.408 $\frac{0.408}{0.408} = 1$
  - ✓ N: 1.225 $\frac{1.225}{0.408} = 3$
  - ✓ O: 3.68 $\frac{3.68}{0.408} = 9.02$
- Round to the nearest whole number and Voilà: NiN_3O_9

A hydrate was analyzed and determined to be 34% water. Further the anhydrate was analyzed to be 23.96% nickel, 17.16% nitrogen, and 58.88% oxygen.
Determine the formula for this hydrate, and name it.

- Now we are ready to find the mole ratio between the water and the anhydrate. Note that the Ni, N, and O add up to 100%.
- A completely separate analysis tell us that the hydrate is 34% water and thus 66% anhydrate
- Calculate moles of water $34g\frac{1mol}{18g}=1.89molH_2O$
- Calculate moles of anhydrate $66g\times\frac{1mol}{244.69g}=0.270molNiN_3O_9$
  - ✓ Determine the mole ratio water / anhydrate
  - ✓ $\frac{1.89}{0.270}=7molH_2O$
  - ✓ Thus $NiN_3O_9 \bullet 7\ H_2O$

Determining Molecular Formula *Molecular Compounds* *(all nonmetals)*

Break out E4 IN CLASS Practice

Empirical Formulae and Molecular Formulae

- All ionic compounds are always written as empirical formulae, including hydrates.
- Molecular compounds may not always be represented in the lowest whole number ratio.
  - ✓ Water is both an empirical formula and it is the molecular formula of this compound: H_2O
  - ✓ The chemical formula for sugar is not an empirical formula: $C_6H_{12}O_6$
    - The empirical formula of sugar is CH_2O
    - however this molecule is formaldehyde

A carbon, hydrogen, oxygen compound was analyzed to be 40.0 % carbon, 53.0 % oxygen and the rest hydrogen. Determine the empirical formula for this compound.

The compound was further analyzed and has molar mass of 180 g/mole, determine the molecular formula.

- Determine the % of hydrogen.
- Treat the % values as mass values. and change to moles
- Divide by the smallest mole value

A carbon, hydrogen, oxygen compound was analyzed to be 40.0 % carbon, 53.0 % oxygen and the rest hydrogen. Determine the empirical formula for this compound.

The compound was further analyzed and has molar mass of 180 g/mole, determine the molecular formula.

- Determine the % of hydrogen.
  - ✓ 100% - 53% - 40% = 7%
- Treat the % values as mass values. and change to moles
  - ✓ C: 40 g $$40g \times \frac{1mol}{12g} = 3.33mol$$
  - ✓ O: 53g $$53g \times \frac{1mol}{16.0g} = 3.31mol$$
  - ✓ H: 7g $$7g \times \frac{1mol}{1.01g} = 6.93mol$$

A carbon, hydrogen, oxygen compound was analyzed to be 40 % carbon, 53 % oxygen and the rest hydrogen. *Determine the empirical formula for this compound.*

The compound was further analyzed and has molar mass of 180 g/mole, determine the molecular formula.

- Divide by the smallest mole value
  - ✓ C: 3.33 $\frac{3.33}{3.31} = 1.01$
  - ✓ O: 3.31 $\frac{3.31}{3.31} = 1$
  - ✓ H: 6.93 $\frac{6.93}{3.31} = 2.09$
- Rounding to the nearest whole number and Voilà - Empirical Formula: CH_2O

A carbon, hydrogen, oxygen compound was analyzed to be 40 % carbon, 53 % oxygen and the rest hydrogen. Determine the empirical formula for this compound.

The compound was further analyzed and has molar mass of 180 g/mole, determine the molecular formula.

- Calculate the molar mass of the empirical formula
  - ✓ CH_2O 12 + 2(1) + 16 = 30
- Divide the molar mass of the molecule by the molar mass of the empirical formula.
  - ✓ $\frac{180}{30} = 6$
- Multiply the factor of 6 through the empirical formula to get the molecular formula
  - ✓ $C_6H_{12}O_6$ is the molecular formula. This is actually the molecule. It's sugar.

3.4 g phosphorus was burned in oxygen and the resulting compound weighed 7.8 g.

Determine the *empirical* and *molecular* formulas, then name the compound. The molar mass is 284 g/mole.

Try this on your E4 InClass Practice

3.4 g phosphorus was burned in oxygen and the resulting compound weighed 7.8 g. Determine the *empirical* and *molecular* formulas, then name the compound. The molar mass is 284 g/mole.

- Calculate the mass of oxygen
  - ✓ 7.8 g P?O? - 3.4 g P = 4.4 g O
- Determine the moles
  - ✓ 4.4 g O = 0.275 mol
  - ✓ 3.4 g P = 0.1098 mol
- Divide by the smallest
  - ✓ Oxy 0.275 mol $\frac{0.275 mol}{0.1098 mol} = 2.5$
  - ✓ P 0.1097 mol $\frac{0.1097 mol}{0.1097 mol} = 1$

3.4 g phosphorus was burned in oxygen and the resulting compound weighed 7.8 g. Determine the *empirical* and *molecular* formulas, then name the compound. The molar mass is 284 g/mole.

- Calculate the mass of oxygen
  - ✓ 7.8 g P?O? - 3.4 g P = 4.4 g O
- Determine the moles
  - ✓ 4.4 g O = 0.275 mol
  - ✓ 3.4 g P = 0.1098 mol
- Divide by the smallest
  - ✓ Oxy 0.275 mol $\frac{0.275 mol}{0.1098 mol} = 2.5$ | 5
  - x2
  - ✓ P 0.1097 mol $\frac{0.1097 mol}{0.1097 mol} = 1$ | 2
- Thus P_2O_5

if you carry enough sig figs, you will only ever see 0.5 OR 0.3 0.67 Thus x2 and x3 factors

3.4 g phosphorus was burned in oxygen and the resulting compound weighed 7.8 g. Determine the *empirical* and *molecular* formulas, then name the compound. The molar mass is 284 g/mole.

- Empirical formula: P_2O_5
- But what about the molecular formula.
- Add the molar mass of the empirical formula
  - ✓ 2(31) + 5(16) = 142
- Divide into the molar mass given
  - ✓ $\frac{284}{142} = 2$
- Multiply the 2 factor through the empirical formula
  - ✓ P_4O_{10} name ????
  - ✓ **tetra**phosphorus **dec**oxide

7.835 g methoxsalen resulted in 3.371 g carbon, 0.407 g hydrogen, 0.627 g nitrogen, and 1.431 g oxygen. The molecular mass is approximately 350 g/mole

Determine the *empirical* and *molecular* formulas.

Try this on your E4 InClass Practice

Analysis of 7.835 g of methoxsalen, used to treat psoriasis resulted in 5.223 g carbon, 0.293 g hydrogen, and 2.319 g oxygen. The molecular mass is approximately 216 g/mole Determine the empirical and molecular formulas.

- Determine the moles of each element
- Divide by the smallest
- Add the molar mass of the empirical formula
- Determine the Molar Mass/Emp mass factor

$$5.223g \times \frac{1mol}{12.01g} = 0.435molC \quad \frac{0.435}{0.145} = 3$$

$$0.293g \times \frac{1mol}{1.01g} = 0.290molH \quad \frac{0.290}{0.145} = 2 \qquad C_3H_2O$$

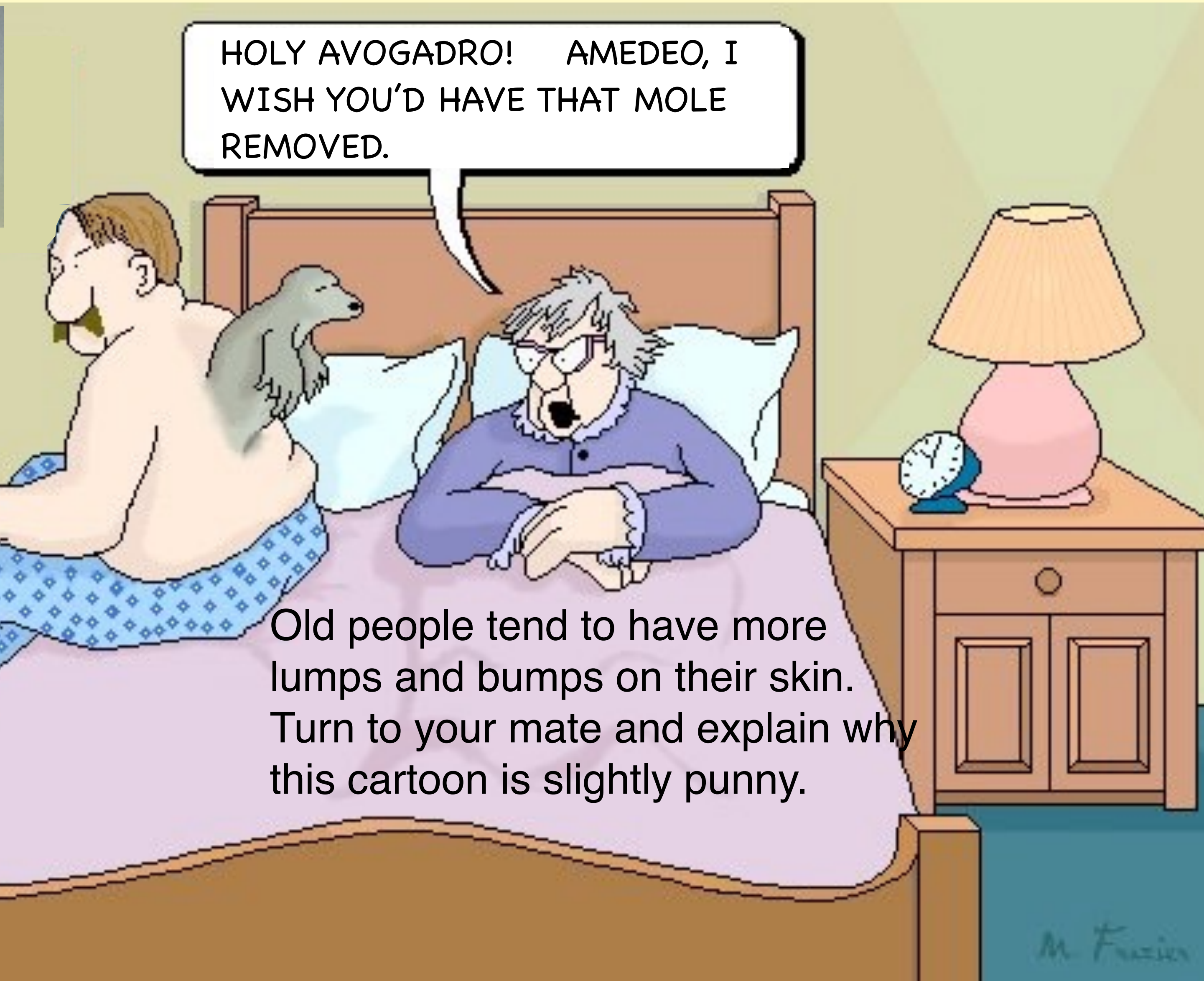
$$2.319g \times \frac{1mol}{16g} = 0.145molO \quad \frac{0.145}{0.145} = 1$$

$$3(12.01) + 2(1.01) + 16 = 54.05$$

$$\frac{216g}{54.05} = 4 \qquad Thus \quad C_{12}H_8O_4$$

More great chemistry humor…
Lorenzo Romano Amedeo Carlo Avogadro

The Count of Quaregna and Cerreto

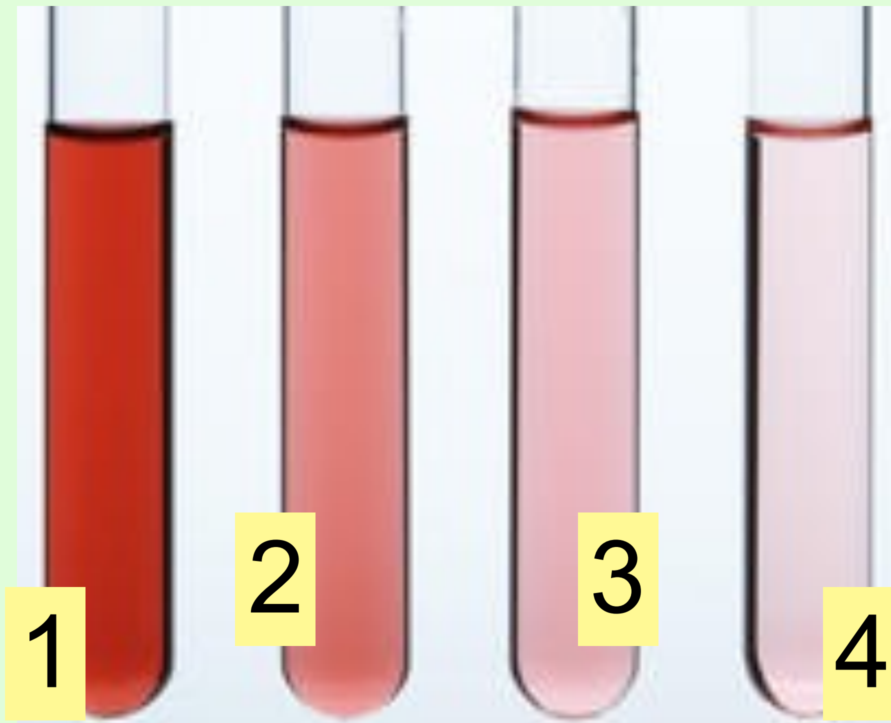


Old people tend to have more lumps and bumps on their skin. Turn to your mate and explain why this cartoon is slightly punny.

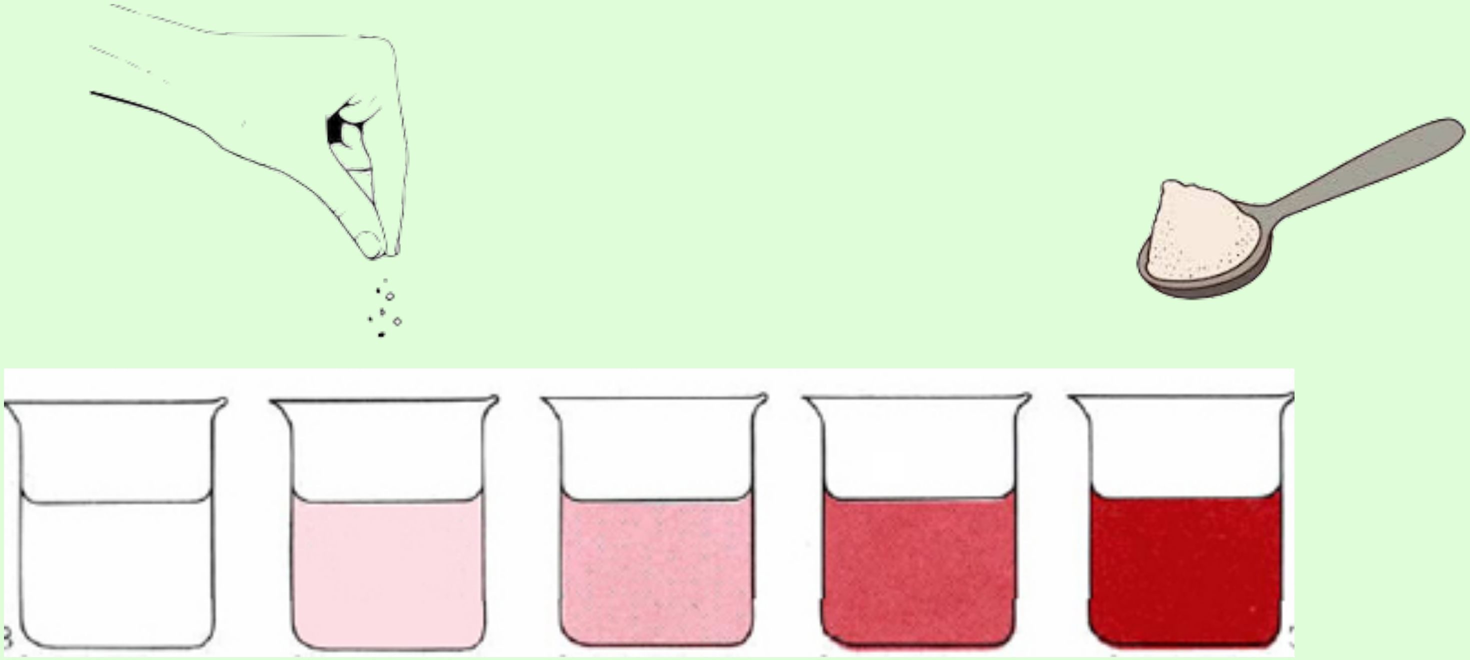
Molarity

A measure of concentration

Which cobalt(II) chloride solution is most concentrated?



Concentration:
We can look at the intensity of the light *absorbed* (the darkness) to tell us more or less concentrated.



Molarity, a measure of concentration

- Solid is the solute (dissolvee)
- Water is the solvent (dissolver)
- Salt water is the solution
- Molarity is the number of moles of solute per Liter of solution

- $$Molarity(M) = \frac{moles\ of\ Solute}{Volume(L)\ of\ Solution}$$

Molarity,
a measure of concentration

$$Molarity = \frac{molesSolute}{LiterSolution}$$

- If there were dissolved 5.0 g of sodium chloride, NaCl in 250. mL of solution, what is the concentration of this salt solution? MM 58.5

- Convert to moles, change to Liters, calculate molarity….enter your answer.
How many sig figs?

✓*Try it on your E5 IN CLASS Practice Sheet.*

Molarity, a measure of concentration

$$Molarity = \frac{molesSolute}{LiterSolution}$$

- If there were dissolved 5.0 g of sodium chloride in 250. mL of solution, what is the concentration of this salt solution?
- NaCl 58.45 g/mol

$$5g \times \frac{1mol}{58.45g} = 0.0855mol$$

$$\frac{0.0855mol}{0.25L} = 0.34M$$

Molarity

Calculations

Molarity, Volume, Moles

$$Molarity = \frac{molesSolute}{LiterSolution}$$

MM 36.5 g/ml

- If you had 35.0 ml of 1.5 M HCl solution, how many moles of HCl would you have?

✓ *Work this on your E5 IN CLASS Practice Sheet.*

Molarity, Volume, Moles

$$Molarity = \frac{molesSolute}{LiterSolution}$$

MM 36.5 g/ml

- If you had 35.0 ml of 1.5 M HCl solution, how many moles of HCl would you have?
- $\frac{1.5 moles}{1L} \times 0.035L = 0.052 moles$

$$Molarity = \frac{molesSolute}{LiterSolution}$$

cross multiply

-

Molarity, Volume, Moles

$$Molarity = \frac{molesSolute}{LiterSolution}$$

MM 36.46 g/mol

- If you had 35.0 ml of 1.5 M HCl solution, how many moles of HCl would you have?

- $\frac{1.5 moles}{1 L} \times 0.035 L = 0.052 moles$

- What mass of HCl is in this solution?

✓ *Work it on your E5 IN CLASS Practice Sheet.*

Molarity, Volume, Moles

$$Molarity = \frac{molesSolute}{LiterSolution}$$

MM 36.46 g/mol

- If you had 35.0 ml of 1.5 M HCl solution, how many moles of HCl would you have?
- $\frac{1.5moles}{1L} \times 0.035L = 0.052moles$
- What mass of HCl?
- $0.052mol \times \frac{36.46g}{1mol} = 1.9g$

Molarity, Volume, Moles

$$Molarity = \frac{moles\ of\ Solute}{Liters\ of\ Solution}$$

- If you wanted to make 25 ml of a 0.876 M silver nitrate, $AgNO_3$ (MM 169.91) solution, what would you need?

- Calculate the mass of silver nitrate needed.

Molarity, Volume, Moles

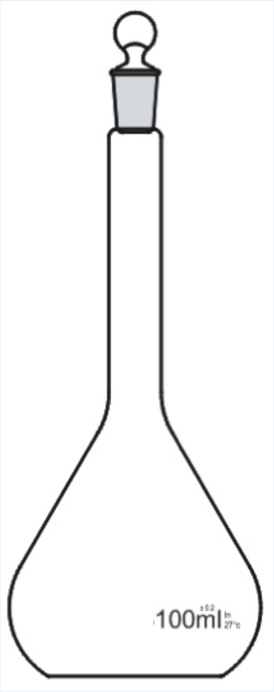
- If you wanted to make 25 ml of a 0.876 M silver nitrate solution, what would you need?

- the correct mass of $AgNO_3$

- $$\frac{0.876mol}{1L} \times 0.025L \times \frac{169.87g}{1mol} = 3.7gAgNO_3$$

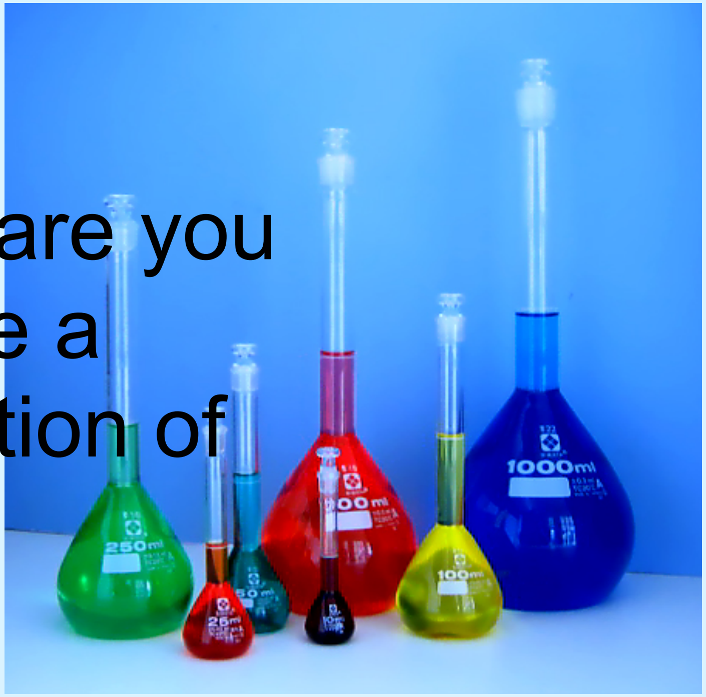
- An accurate 25 ml container

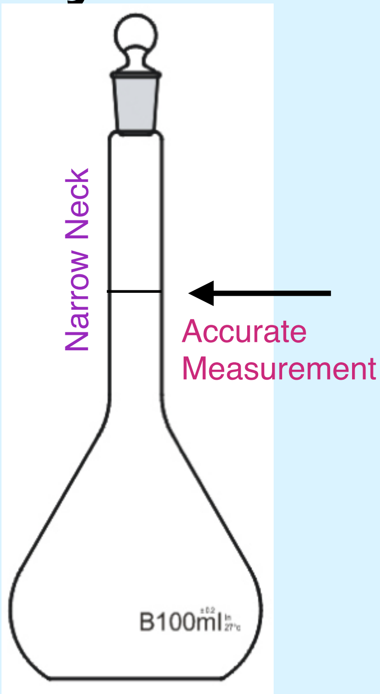
- Enter...the volumetric flask



what is a Volumetric Flask?

- Very exact **single** volume measuring device.
- You already used the 25 ml flask when you were measuring the density of water.
- Volumetric flasks come in a wide variety of sizes.
- This is the glassware you would use to make a precise concentration of some solution.





Embrace the Millimole

$$Molarity = \frac{molesSolute}{LiterSolution}$$

- Just what can M for molarity mean?

4. Let's change 2.5 M (moles/L) to millimoles/L

Embrace the Millimole

$$Molarity = \frac{molesSolute}{LiterSolution}$$

- Just what can M for molarity mean?

4. Let's change 2.5 M (moles/L) to millimoles/L

$$2.5M = \frac{2.5moles}{1L} \times \frac{1L}{1000ml} \times \frac{1000\,millimol}{1mol} = \frac{2.5\,millimol}{1ml} = 2.5M$$

$$Molarity(M) = \frac{1molesSolute}{1LiterSolution} \times \frac{1L}{1000ml} \times \frac{1000\,millimol}{1mol} \xrightarrow{\text{Voilà}} Molarity(M) = \frac{1millimolesSolute}{1milliliterSolution}$$

Embrace the Millimole

$$Molarity = \frac{\#\, millimolesSolute}{1 milliliterSolution}$$

5. If you had 20. ml of 0.15 M $AlCl_3$, calculate the millimoles of $AlCl_3$ units.

Embrace the Millimole

$$Molarity = \frac{\# millimolesSolute}{1 milliliterSolution}$$

5. If you had 20. ml of 0.15 M $AlCl_3$, calculate the millimoles of $AlCl_3$ units.

$$0.15M\left(\frac{mmol}{ml}\right)\times 20ml = 3.0 mmolAlCl_3$$

OR.... $$0.15M = \frac{x\ mmolAlCl_3}{20ml} \quad x = 3.0 mmolAlCl_3$$

- What are the millmole of chloride ions Cl^-

Embrace the Millimole

$$Molarity = \frac{\# millimolesSolute}{1 milliliterSolution}$$

5. If you had 20. ml of 0.15 M $AlCl_3$, calculate the millimoles of $AlCl_3$ units.

$$0.15M \times 20ml = 3.0mmolAlCl_3$$

- What are the millmole of chloride ions Cl^-

$$3.0mmolAlCl_3 \times \frac{3Cl^-}{1AlCl_3} = 9mmolCl^-$$

Dilution

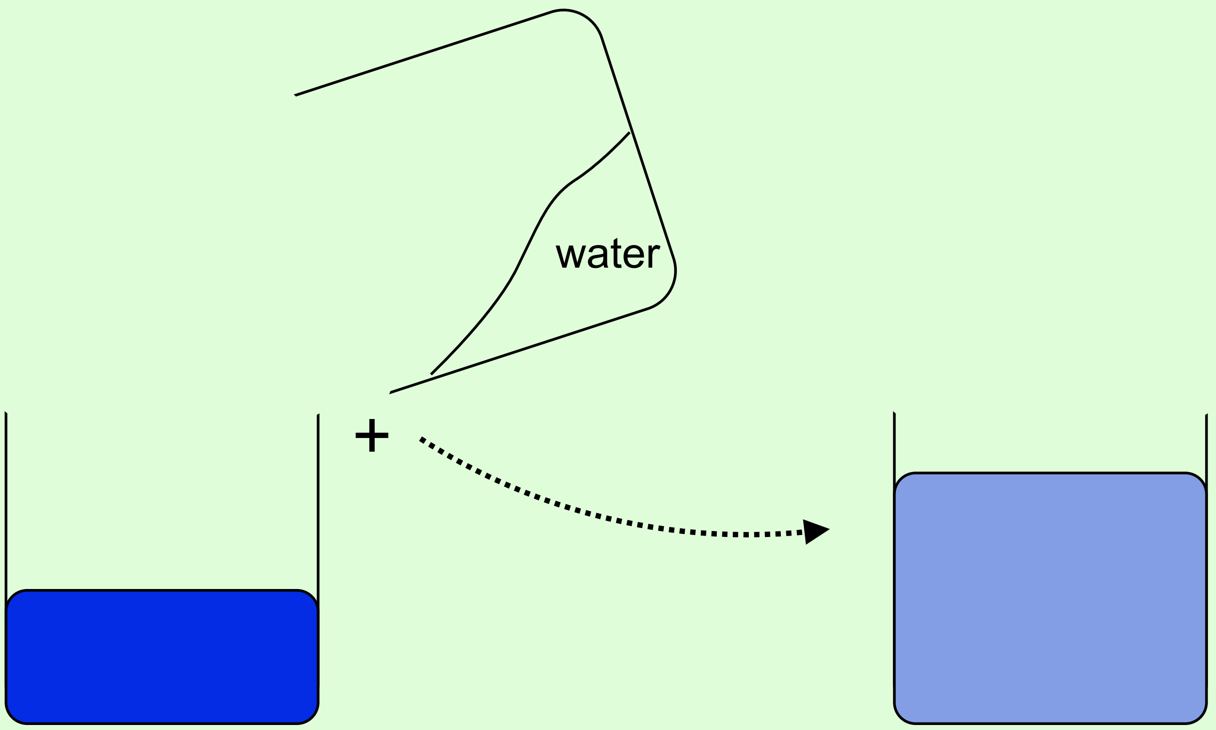
Making a solution
less concentrated

$$Molarity(M) = \frac{moles\ of\ Solute}{Volume(L)\ of\ Solution}$$

Dilution

$$Molarity = \frac{moles\ of\ Solute}{Liters\ of\ Solution}$$

The number of moles of copper ions in the **beaker on right** compared to the number of moles in the beaker on left are



1. more
2. less
3. same

$M \times V = moles$

$$Molarity = \frac{moles\ of\ Solute}{Liters\ of\ Solution}$$

- When you dilute a salt water solution by adding water the number of moles of chemical in the water does not change.
- $moles_{before} = moles_{after}$
- $M_c \times V_c = moles_{before} = moles_{after} = M_d \times V_d$
- Introducing the *dilution equation* ... just a modification of the molarity equation

$$M_c V_c = M_d V_d$$

Molarity, Volume, Moles, millimoles!

6. If you wanted to make 300. ml of 0.25 M NaOH solution and you had access to 2.0 M NaOH stock solution.

- What volume of *stock* *(concentrated)* solution should you dilute?
  - *Work on your E5 IN CLASS Practice Sheet.*

$$M_c V_c = M_d V_d$$

$$Molarity = \frac{moles\ of\ Solute}{Liters\ of\ Solution}$$

Molarity, Volume, Moles, millimoles!

6. If you wanted to make 300. ml of 0.25 M NaOH solution and you had access to 2.0 M NaOH stock solution.

- What volume of *stock* solution should you dilute?

$$2.0M \times V_c = 0.25M \times 300ml$$

$$V_c = 37.5ml$$

OR….

$$2.0M \times V_c = 0.25M \times 0.3L$$

$$0.0375L$$

$$Molarity = \frac{moles\ of\ Solute}{Liters\ of\ Solution}\ OR\ \frac{millimoles\ of\ Solute}{milliLiters\ of\ Solution}$$

Molarity, Volume, Moles, millimoles!

7. If you had 25 ml of 1.5 M barium chloride, and you **added** 175 ml of water (assume volumes are additive), calculate the molarity of the resulting solution.

✓*Work it on your E5 IN CLASS Practice Sheet.*

$$M_cV_c = M_dV_d$$

$$Molarity = \frac{moles\ of\ Solute}{Liters\ of\ Solution}\ OR\ \frac{millimoles\ of\ Solute}{milliliters\ of\ Solution}$$

Molarity, Volume, Moles, millimoles!

7. If you had 25 ml of 1.5 M barium chloride, and you **added** 175 ml of water (assume volumes are additive), calculate the molarity of the resulting solution.

$$1.5M \times 25 = M_d \times 200ml$$

$$M_d = 0.19M\ BaCl_2$$

OR....

$$1.5M \times 0.025L = M_d \times 0.200L$$

$$M_d = 0.19M\ BaCl_2$$

- What is the concentration of the chloride ions?

$$Molarity = \frac{moles\ of\ Solute}{Liters\ of\ Solution}\ OR\ \frac{millimoles\ of\ Solute}{milliLiters\ of\ Solution}$$

Molarity, Volume, Moles, millimoles!

7. If you had 25 ml of 1.5 M Barium chloride, and you added 175 ml of water (assume volumes are additive, calculate the molarity of the resulting solution. What volume of *stock* solution should you dilute?

$$1.5M \times 25 = M_d \times 200ml$$

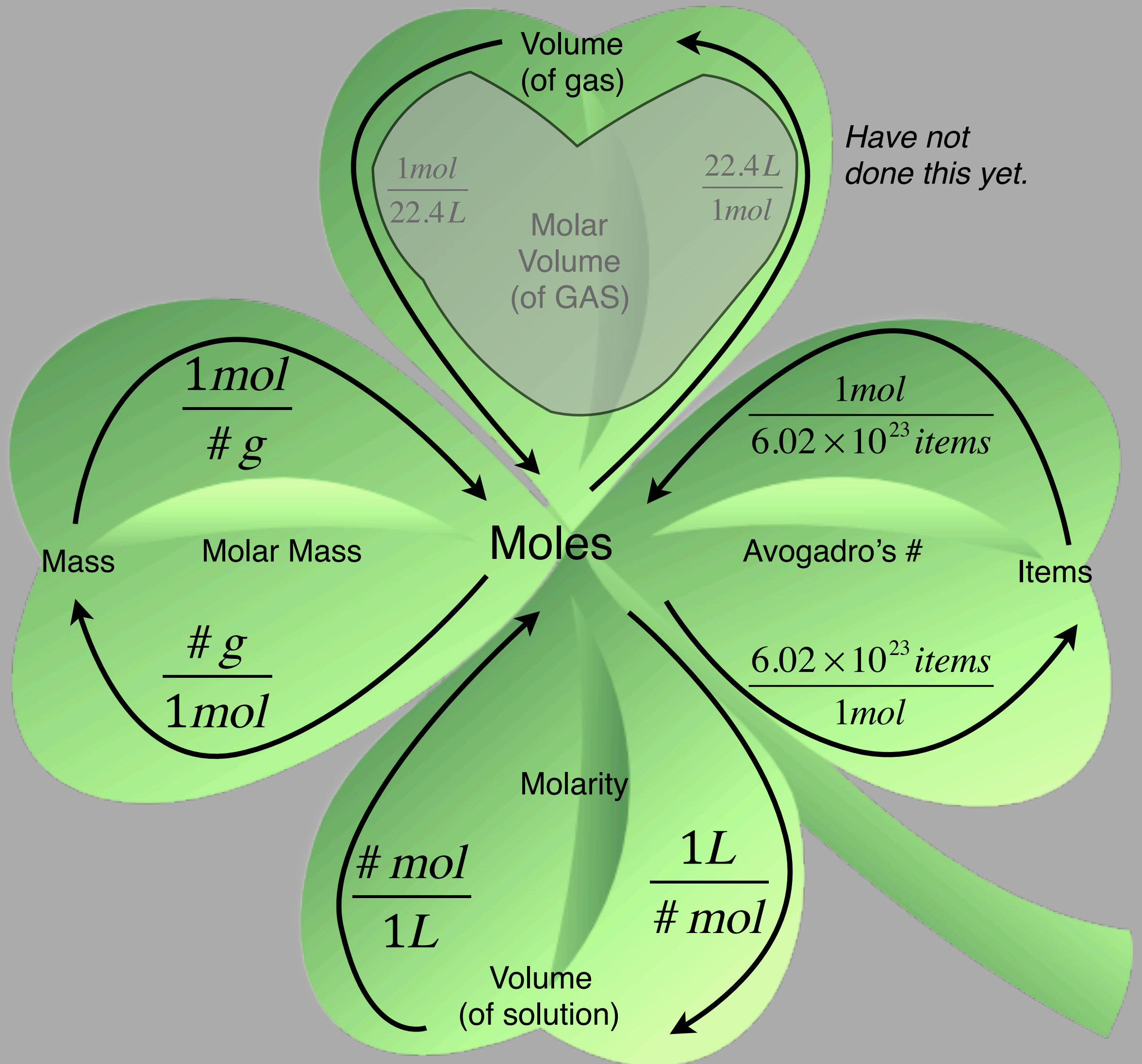
$$M_d = 0.19M\ BaCl_2$$

- What is the concentration of the chloride ions?

$$0.19MBaCl_2 \times \frac{2Cl^-}{1BaCl_2} = 0.38M\ Cl^-$$

$$Molarity = \frac{moles\ of\ Solute}{Liters\ of\ Solution}\ OR\ \frac{millimoles\ of\ Solute}{milliLiters\ of\ Solution}$$

The Mole Shamrock

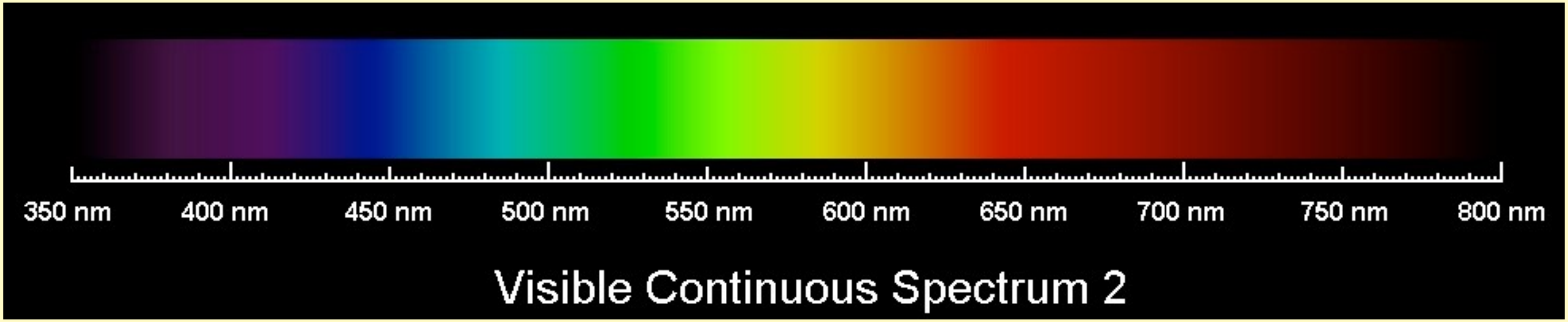


Know Your Units !!

LAD E2

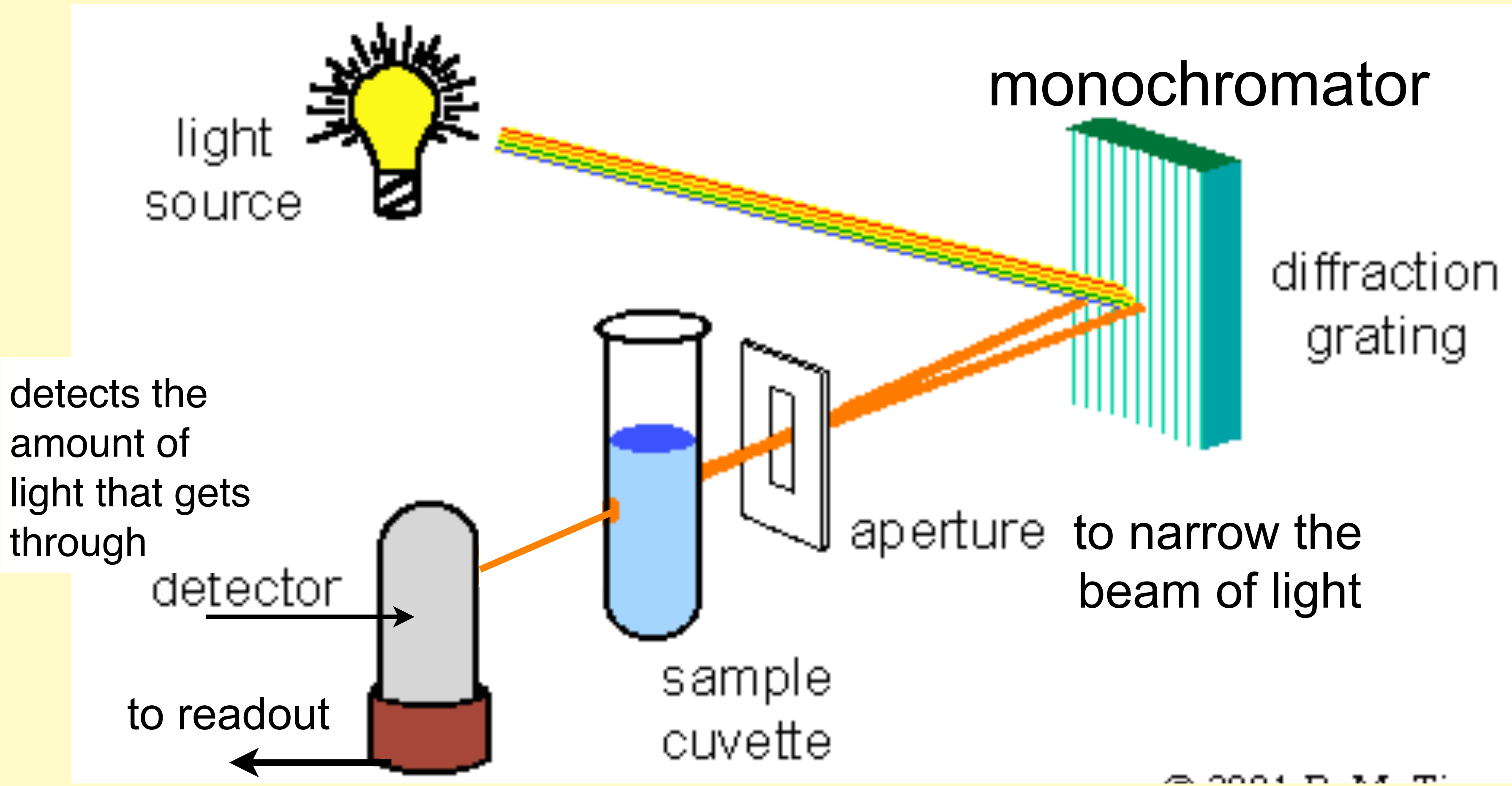
Confirming Concentration

Color as a function of wavelengths of Light

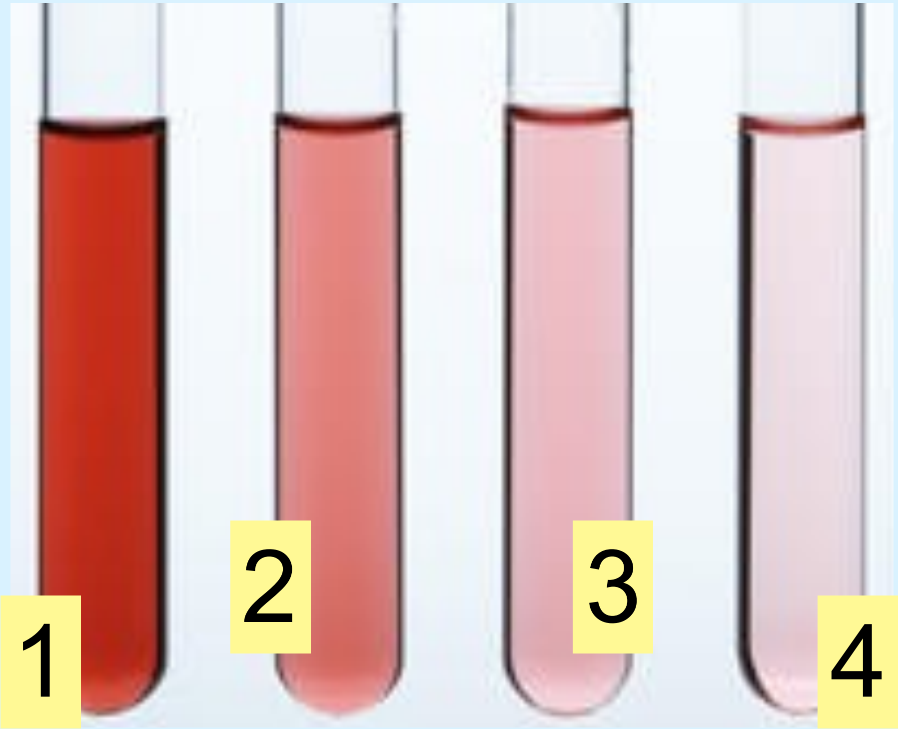


- Color in a visible spectrum and color wheel

Spectrophotometer schematic

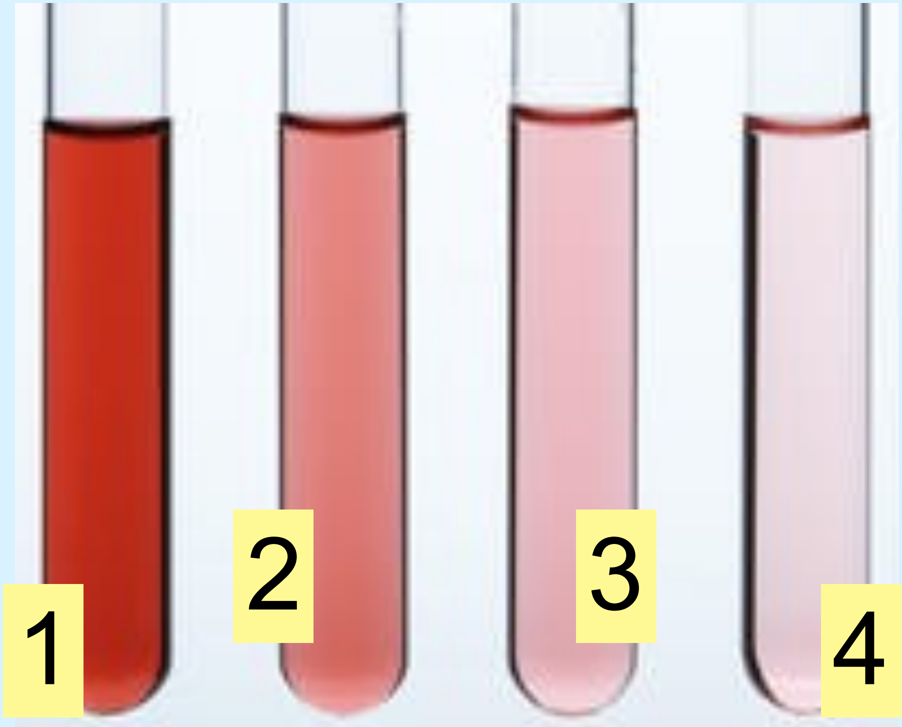


You know which cobalt(II) chloride solution is the most concentrated.



How is color intensity related to **absorbtion** of light?

Which cobalt(II) chloride solution absorbs the most light?



LAD F2

Molarity and Error Analysis

The Chemist's Dozen

Extra Clicker Questions

As needed or time allows

A dozen eggs

A mole of atoms

No calculators for ALL of these questions....

Which of the following contains 6.00×10^{16} atoms?

No Calculators

Select as many as apply. Be prepared to defend your choice(s).

1. 6.00×10^{16} He atoms
2. 6.00×10^{16} H_2O molecules
3. 3.00×10^{16} Cl_2 molecules
4. 3.00×10^{16} O_3 molecules
5. 1.00×10^{16} N_2O_4 molecules

Which of the following contains 6.00 x 10^{16} atoms?

Select as many as apply. Be prepared to defend your choice(s).

1. 6.00 x 10^{16} He atoms $3 \times 10^{16} \times \frac{2 atoms}{1 molecule} = 6 \times 10^{16} atoms$
2. 6.00 x 10^{16} H_2O molecules
3. 3.00 x 10^{16} Cl_2 molecules $1 \times 10^{16} \times \frac{6 atoms}{1 molecule} = 6 \times 10^{16} atoms$
4. 3.00 x 10^{16} O_3 molecules
5. 1.00 x 10^{16} N_2O_4 molecules

Which contains the least amount of moles of particles? *No Calculators*

Be prepared to defend your choice.

1. 4 g of helium
2. 16 g of oxygen gas
3. 18 g of water
4. they all contain the same amount of particles.

Which contains the least amount of moles of particles?

Select as many as apply. Be prepared to defend your choice(s).

1. 4 g of helium (this is one mole of He atoms)
2. 16 g of oxygen gas (diatomic)
   - this is 0.5 mole of O_2 molecules, thus the least amount of particles.
3. 18 g of water (this is one mole of water molecules)
4. they all contain the same amount of particles.

Which weighs the least?

Select as many as apply. Be prepared to defend your choice(s).

No Calculators

1. 71 g of lead
2. 71 g of Cl_2
3. 1 mole of chlorine gas
4. 1.2×10^{24} chlorine atoms
5. They all weigh about the same.

Which weighs the least?

Select as many as apply. Be prepared to defend your choice(s).

1. 71 g of lead
2. 71 g of Cl_2
3. 1.0 mole of chlorine gas (1 mole of Cl_2 weighs approximately 71 g
4. 1.2 x 10^{24} chlorine atoms (This is 2 moles of Cl atoms which weigh approximately 71 g)
5. They all weigh the same.

- You could try to argue that #3 actually weighs only 70.9, and that #4 actually weighs only 70.7 g calculation below:
- $$1.2\times10^{24}\times\frac{1mol}{6.02\times10^{23}atoms}\times\frac{35.45g}{1mol}=70.7g$$
- However, when you consider only 2 sig figs, you arrive at a value of 71 g for all of them.

Concerning water, choose the false statement(s) below.

Select as many as apply. Be prepared to defend your choice(s).

1. 1 mole water = 6.02×10^{23} g
2. 18 g = 6.02×10^{23} molecules
3. 2 moles = 36 g
4. Water is a triatomic molecule
5. 1 mole of water contains 6.02×10^{23} atoms.
6. Select this only if all statements are true.

No Calculators

Concerning water, choose the false statement(s) below. *Select as many as apply. Be prepared to defend your choice(s).*

1. 1 mole water = 6.02×10^{23} g
   - (this would be correct if you replaced the g with molecules)
2. 18 g = 6.02×10^{23} molecules
3. 2 moles = 36 g
4. Water is a triatomic molecule (3-atom molecule)
5. 1 mole of water contains 6.02×10^{23} atoms.
   - 1 mole contains 6.02×10^{23} molecules,
   - but 3x as many atoms.

- You could argue that 2 and 3 are false if you were to use more exact masses from the chart.
  - 18.02 g = 6.02×10^{23}
  - 2 moles = 36.04 g

Which of the following contain the most number of atoms? *No Calculators*

1. 1 g of He
2. 1 g of O_2
3. 1 g of Pb
4. They all contain the same amount.

Which of the following contain the most number of atoms?

1. 1 g of He - The He is the most amount of mole
    - (1g * 1mole/4 g)
2. 1 g of O_2 (1 g * 1mole/32 g)
3. 1 g of Pb (1 g * 1mole/207 g)
4. They all contain the same amount.

That's it for now