

Some of the most important lab math to know (in my opinion):

- **Metric conversions**

- Get really comfortable moving decimal places to convert between μ , m, etc. (esp. μL , mL & L)
 - Try to get to the point where these is second nature, and you can do it “automatically” in your head

- **Converting between molecular weight (MW) or formula weight (FW)*, molarity, and weight/volume**

- $\text{MW (or FW)} = \text{g/mol}$
- $\text{Molarity (M)} = \text{mol/L}$
- $\text{g/L} = (\text{MW})(\text{molarity})$
- $\text{M} = (\text{g/L}) / (\text{MW or FW})$

* although the terms molecular weight and formula weight are often used interchangeably, molecular weight is technically only for covalent compounds, whereas formula weight can refer to ionic compounds as well as hydrates & other complexes. Both are g/mol

- **For protein biochemists, converting between kDa (what the MW for proteins is often given as), mg/mL, & molarity**

- $1 \text{ kDa} = 1000 \text{ Da (Daltons)}$
- $1 \text{ Da} = 1 \text{ g/mol}$
- $1 \text{ kDa} = 1000 \text{ g/mol}$
- $1 \text{ mg/mL} = 1 \text{ g/L}$
- Going from **MW in kDa** and **molarity (M)** to **mg/mL**:
 - $(\text{MW in kDa})(\text{M} \times 1000) = \text{mg/mL}$
- Going from **MW in kDa** and **mg/mL** to molarity:
 - $\text{Molarity} = (\text{mg/mL}) / ((\text{MW in kDa})(1000))$

- **Understanding & using when necessary % w/v**

- $\% \text{ w/v} = \text{g/L} \times 100\%$
 - $1 \% \text{ w/v} = 1 \text{ g} / 100 \text{ mL}$

- **Making solutions from solids**

- If you have a desired volume & concentration:

$$(\text{Desired volume})(\text{needed mass}) = \text{desired concentration}$$

$$\text{Needed mass} = (\text{desired concentration}) / (\text{desired volume})$$

- Calculating the volume of solvent needed to get to a final desired concentration based on a measured amount (e.g. for when you're measuring small volumes and it's impractical to weigh out precisely the target mass)

$$(\text{measured mass}) / (\text{needed volume}) = \text{desired concentration}$$

$$\text{Needed volume} = (\text{measured mass}) / (\text{desired concentration})$$

- Salvaging solutions by adding extra solvent and/or mass to match what you actually have present in your current solution

- Too much mass? Calculate volume

$$(\text{Measured mass}) / (\text{total volume needed}) = \text{desired concentration}$$

$$\text{Total volume needed} = (\text{measured mass}) / (\text{desired concentration})$$

$$\text{Needed additional volume} = \text{total volume needed} - \text{current volume}$$

- Too much volume? Calculate mass

$$(\text{total mass needed}) / (\text{current volume}) = \text{desired concentration}$$

$$\text{Total mass needed} = (\text{desired concentration}) (\text{current volume})$$

$$\text{Extra mass needed} = \text{total mass needed} - \text{current mass}$$

Alternatively, calculate the mass you need to add directly (rather than first finding total then subtracting). To do that...

$$\text{Extra volume} = \text{current volume} - \text{intended volume}$$

$$\text{Extra mass needed} = (\text{extra volume}) (\text{desired concentration})$$

If adding that mass will not appreciably change the final volume, just add it. If adding that mass would change the final volume, calculate mass for a higher volume with enough wiggle room. Add the mass for that, then adjust the volume to that higher volume.

● Diluting solutions

- From any form of concentration (as long as same units on both sides)

- $C_1V_1 = C_2V_2$

- Initial concentration (C_1) x initial volume (V_1) = final concentration (C_2) x final volume (V_2)

- From relative concentrations (X factors)

- $(\text{Desired total volume}) / (\text{X factor}) = \text{needed volume of the stock solution}$

- $(\text{Volume of the stock solution you want to use}) (\text{X factor}) = \text{total volume to dilute to}$

- **Serial dilutions**

- Dilution factor (DF): how much you are diluting by each time
 - e.g. if you're halving it each time, DF would be 2
- Multiply the initial concentration by $(1/DF)$ raised to the number of dilutions to get the concentration of that dilution

- **Making master mixes**

- Pre-mix the parts of reaction mixtures that are the same between all reactions. Then add that to the unique parts
 - Make more than you need to account for volume loss during transfers

Links to more information

more on metric conversions: blog: https://bit.ly/metric_calcs ; YouTube: <https://youtu.be/A-J7kdLNb8c>

more on sig figs here: <http://bit.ly/significantfigures> & <https://youtu.be/q29Cv5i5PuI>

more on dimensional analysis: <http://bit.ly/dimensionalanalysing> & <https://youtu.be/KQMA0aAGfP4>

more about concentrations: <http://bit.ly/sciencestocksolutions> ; YouTube: <https://youtu.be/RjXYvOiLh2w> & <https://youtu.be/wC623-AGBOs>

and a couple cheat sheets for molecular weights:

<https://drive.google.com/file/d/1TqCRh23TO8s5T23DjBcJeCUWI23ps9sl/view?usp=sharing>

and concentrations: <https://drive.google.com/file/d/1v7pff7cEpJTvr-lADLwdwwFOIEDdz8Ny/view?usp=sharing>

more lab-mathy things: <https://thebumblingbiochemist.com/lets-talk-science/math-y-stuff/> & <https://www.youtube.com/playlist?list=PLUWsCDtjESrGfML5aO-J8vgzm2OYGPwhL>

More on making solutions: blog form: <http://bit.ly/sciencestocksolutions> ; YouTube: https://youtu.be/4R_5qxmMzas & <https://youtu.be/9OtxOHLyaOQ>

More on serial dilutions: <http://bit.ly/serialdilutions>

more about all sorts of things: All (with topics listed) <http://bit.ly/2OIIAB0> or search blog: <http://thebumblingbiochemist.com>

common SI prefixes

peta	P	10^{15}	quadrillion
tera	T	10^{12}	trillion
giga	G	10^9	billion
mega	M	10^6	million
kilo	k	10^3	thousand
hecto	h	10^2	hundred
deka	da	10^1	ten
		10^0	one
deci	d	10^{-1}	tenth
centi	c	10^{-2}	hundredth
milli	m	10^{-3}	thousandth
micro	μ	10^{-6}	millionth
nano	n	10^{-9}	billionth
pico	p	10^{-12}	trillionth
femto	f	10^{-15}	quadrillionth

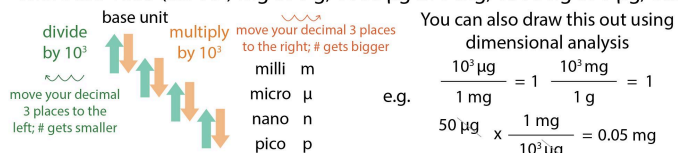
Metric prefixes let us work with large & tiny numbers

You can stick these prefixes in front of any base unit (g, mol, M, etc.)

milli	m	10^{-3}	1/1,000	thousandth of a...
micro	μ	10^{-6}	1/1,000,000	millionth of a...
nano	n	10^{-9}	1/1,000,000,000	billionth of a...
pico	p	10^{-12}	1/1,000,000,000,000	trillionth of a...

Multiply by factors of 10 (just move your decimal point!) to easily convert between units with different metric prefixes

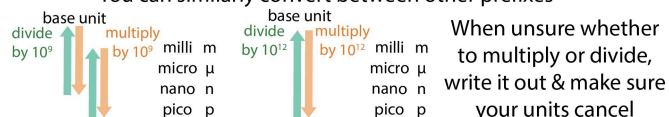
There are 1000 (i.e. 10^3) mg in 1 g; 1000 μ g in 1 mg; 1000 ng in 1 μ g, etc.



There are 1,000,000 (i.e. 10^6) μ g in 1 g; 1,000,000 ng in 1 mg, etc.



You can similarly convert between other prefixes



Shortcuts for metric concentration conversions

When converting to thinking in terms of **less stuff** (e.g. a smaller volume or mass), the **number should get smaller** (e.g. there are fewer molecules in a μ L or a μ g than there are in a L or a g).

But when you're converting to thinking of **same amount of stuff, just in terms of smaller units**, the **number should be bigger** (e.g. there are more μ g in a g than there are mg in a g and more μ L in a L than mL in a L).

When the top & bottom prefixes are the same, no matter what they are, you will be dividing/multiplying by the same factor of 10, so any such concentrations are equivalent.

For example, often concentrations are given in mg/mL.

This is the same as $\mu\text{g}/\mu\text{L}$, and ng/nL, and g/L!

Using similar reasoning to remember, M (mol/L) is the same as $\mu\text{mol}/\mu\text{L}$, and mmol/mL, and nmol/nL.

& nmol/ μ L is the same as mM

Since nano is one thousandth of μ , you need to use the one thousandth prefix, with M, hence "mM"

When unsure what to do, write it out & make sure units cancel out.

e.g. If your concentration is in mM (mmol/L), how many mol per μ L?

$$\frac{Y \text{ mmol}}{1 \text{ L}} \times \frac{1 \text{ mol}}{10^3 \text{ mmol}} \times \frac{1 \text{ L}}{10^6 \mu\text{L}} = Y \times 10^{-9} \text{ mol}/\mu\text{L}$$

10⁻⁹ - that's the nano prefix! So this is the same as Y nmol/ μ L!
e.g. a 50 mM NaCl solution would have 50 nmol NaCl per μ L

dimensional analysis!

There's many a way to same-thing-say! And this lets us convert units!

when you divide something by itself, even if that something is described in different ways, you get 1

so you can harmlessly change units by multiplying or dividing anything by 1

for examples...

starting info: I have 2 pizzas. Each has 12 slices. There are 24 people

How many slices do I have?

start point destination

2 pizzas = ? slices

$\frac{2 \text{ pizzas}}{1 \text{ pizza}} \times \frac{12 \text{ slices}}{1 \text{ pizza}} = 24 \text{ slices}$

How many slices does each person get?

start point destination

24 people = ? slices per person

$\frac{2 \text{ pizzas}}{24 \text{ people}} \times \frac{12 \text{ slices}}{1 \text{ pizza}} = 1 \text{ slice per person}$

What're your cookie's concentrations?

molarity is a way to report number of somethings (in mol) in some volume (in L)

it's like saying

50 chocolate chips per cup

10 almonds per cup

but in moles of those per L

a mole is just 6×10^{23} of something (Avogadro's Number)
it's kinda like a biochemist's "dozen"

and a cup is ~ 0.24 L

so we can rewrite these as (6×10^{23})

50/
(6×10^{23}) $\checkmark$

8.3×10^{-23} mol chocolate chips

$\sim 3.5 \times 10^{-22}$ M chocolate chip

10/
(6×10^{23}) $\checkmark$

1.7×10^{-23} mol almonds

$\sim 6.9 \times 10^{-23}$ M almond

0.24 L batter

0.24 L batter

A 1 molar (1M) solution of something has 6.02×10^{23} individual somethings per liter of solution - that's a lot of somethings! So we often deal with metric-prefixed molarities

these generic parts mean...

milli	m	10^{-3}	thousandth of a...	molar	1 mM = 6.02×10^{20} mol/L
micro	μ	10^{-6}	millionth of a...	molar	1 μ M = 6.02×10^{17} mol/L
nano	n	10^{-9}	billionth of a...	molar	1 nM = 6.02×10^{14} mol/L
pico	p	10^{-12}	trillionth of a...	molar	1 pM = 6.02×10^{11} mol/L

stick them in front of molar...

chemical concentrations cheat sheet

molarity (M) is mol solute/L solution

this is what we use most commonly in biochemistry

a mole (mol) is just 6.02×10^{23} copies of the molecule, and 1M is 1 mol/L

molality (m) is mol solute/kg solvent

good for when you're drastically changing temperature (since volume is temperature-dependent) or when components interact so volumes don't add

mole fraction, X, is mol something/total mol

note: mol fraction of one solute will decrease if you stuff in more other solutes, even though the molarity of that solute stays constant

partial pressure is the mole fraction times total pressure

this is used to report gas concentrations, and tells you the amount of pressure contributed by a component of the gas (which is directly related to how much of it there is (Dalton's law of partial pressures))

% w/w

weight-weight percentage
weight solute
weight solution

% v/v

volume-volume percentage
volume solute
volume solution

% w/v

weight-volume "percentage"
g solute
mL solution

ppm (parts per million) &

ppb (parts per billion)

multiply one of those ratios by 1 million (10^6) or 1 billion (10^9), respectively
1 ppm = 1 mg/kg or 1 mg/L

good for reporting amounts of super dilute things like contaminants

relative concentrations of

stock solutions

concentration relative to the final, "working" concentration - e.g. 10X is 10 times more concentrated than you want to use it at

working with protein concentrations

proteins vary greatly in size, so mass/volume can be misleading
The molecular weight (MW) of proteins is typically given in **Daltons (Da)** or **kiloDaltons (kDa)**

1 Da = 1 g/mol 1 kDa = 1000 Da = 1000 g/mol = 1 kg/mol

To go from Da to kDa, divide by 1000. To go from kDa to Da, multiply by 1000
you can think of the Da as shorthand so you don't have to write "g/mol"

quick estimates: the average amino acid is ~ 0.1 kDa (100 Da) so a 1000 amino-acid-long protein would have a MW of ~ 100 kDa

estimated protein length from MW (in kDa) = kDa $\times 10$

estimated MW (in kDa) from protein length = # amino acids / 10

to go from mass/volume* to molarity:

$$\text{molarity} = \left(\frac{\text{mass}}{\text{volume}} \right) / \frac{\text{MW}}{\text{mol}} \quad \text{make sure this is in Da (convert from kDa if needed)}$$

*shortcut: protein concentration is generally measured/reported in mg/mL, but mg/mL = g/L = $\mu\text{g}/\mu\text{L}$ because in each case you're dividing both of these by the same factor of 10 in terms of kDa

$$\text{molarity} = (\text{mg/mL}) / \text{MW (in Da)} \quad \text{molarity} = (\text{mg/mL}) / (\text{MW (in kDa)} \times 1000)$$

mass/volume from molarity and MW

$$\text{mg/mL} = \text{molarity} \times \text{MW (in Da)} \quad \text{mg/mL} = \text{molarity} \times (\text{MW (in kDa)} \times 1000)$$

mass from molarity and volume

$$\text{mg} = \text{molarity} \times \text{MW (in Da)} \times \text{mL} \quad \text{or} \quad \mu\text{g} = \text{molarity} \times \text{MW (in Da)} \times \mu\text{L}$$

% weight/volume (%w/v) is a relative concentration

it is sometimes used for dilute aqueous solutions, especially in medicine

$$1\% \text{ w/v} = 1\text{g}/100\text{mL}$$

Why?

it comes from the fact that water has a density of 1g/mL (at room temp) and adding a little solute won't change that

So the density of most biological solutions is close to the density of pure water, so $\sim 1\text{g/mL}$.

So if you say that there's 1g of something else in there, and the solution still had a density of 1g/mL it'd have to mean that the solution was 100% your thing. So 1g/mL is 100% w/v.

warning: for this to hold true, the thing you're adding has to be similar to the density of water and/or there's just so little of it compared to water that it doesn't make too much of a dent

Solution-making from solids

calculate



Use the formula weight or molecular weight (g/mol) to figure out how much you need to weigh out to get the desired molarity (mol/L)

weigh

do it



dissolve

Start with a lower volume than your final volume because the solid will take up space



add the stir bar

add the solid gradually



stir to dissolve



Always start by adding some liquid (but not a final volume-worth!) to the vessel before solid, or you'll get a hard lump, perhaps with a stir bar stuck in it.



Transfer to a graduated cylinder

Adjust volume in a graduated cylinder - NOT the beaker - & beware of the stir bar!



salty stock solution example

NaCl has a formula weight of ~58.44 g/mol

If I want a 5 molar (5M) solution in molarity:

$$\text{MOLARITY} = \frac{\text{moles of something}}{\text{liter}}$$

a mole is just 6×10^{23} of something (Avogadro's Number)

$$\frac{5 \text{ mol NaCl}}{1 \text{ L solution}} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = \frac{292.2 \text{ g NaCl}}{1 \text{ L solution}}$$

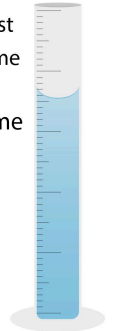
weigh



dissolve in lower volume



adjust volume



Solution from solid

$$\frac{\text{mol}}{\text{L}} \times \frac{\text{g}}{\text{mol}} = \frac{\text{g}}{\text{L}} \times \text{L} = \text{g}$$

↑ molarity ↑ MW ↑ same as mg/mL

if you don't weigh exactly that...

$$\frac{\text{weighed}}{\text{X mL}} = \frac{\text{mg}}{\text{mL}}$$

← final desired conc.

$$X = \frac{\text{weighed mg}}{\text{desired mg/mL}}$$

example: want 5mg/mL BSA

only need 100μL

0.1 mL

$$\frac{5 \text{ mg}}{1 \text{ mL}} \times 0.1 \text{ mL} = 0.5 \text{ mg}$$

TOO SMALL

weigh ~5 mg

actual weighed: 9 mg

$$\frac{9 \text{ mg}}{\text{X mL}} = \frac{5 \text{ mg}}{1 \text{ mL}}$$

$$X = \frac{9 \text{ mg}}{5 \text{ mg/mL}} = 1.8 \text{ mL}$$

"fixing" things

weighed too much?

↓

adjust vol. to match

$$\frac{\text{measured g}}{\text{needed total vol.}} = \text{desired } \frac{\text{g}}{\text{L}}$$

$$\text{needed total vol} = \frac{\text{measured g}}{\text{desired g/L}}$$

$$\text{needed added vol} = \text{needed total vol} - \text{current vol}$$

alternatively, calculate just for "extra" directly rather than subtracting at the end

too much solvent?

↓

adjust mass to match

will needed mass appreciably change vol?

YES → same as if NO, but calculate for an even greater volume then adjust both mass & volume

NO →

$$\frac{\text{needed total g}}{\text{measured vol}} = \text{desired g/L}$$

$$\text{needed total g} = \text{desired g/mol} \times \text{measured vol}$$

$$\text{needed added g} = \text{needed total g} - \text{current g}$$

fixing solutions

if you want 100 mL of 100 g/L

you put in 11g

$$\frac{11 \text{ g}}{x} = \frac{100 \text{ g}}{1 \text{ L}}$$

$$x = \frac{11 \text{ g}}{100 \text{ g/L}}$$

$$x = 0.11 \text{ L}$$

$$= 110 \text{ mL}$$

adjust vol to 110 mL

110 - 100 = 10 more mL needed (best to adjust vol in grad cylinder so this is your final vol.)

your volume got to 110 mL

$$\frac{x}{110 \text{ L}} = \frac{100 \text{ g}}{1 \text{ L}}$$

$$x = 100 \text{ g} \times \frac{110 \text{ L}}{1 \text{ L}}$$

$$x = 11 \text{ g}$$

11g - 10g = 1g

were already had

if addition wouldn't appreciably change the volume, add directly

if it would... calculate for higher vol. w/ wiggle room to adjust. ex. 125 mL

$$\frac{x}{125 \text{ L}} = \frac{100 \text{ g}}{1 \text{ L}}$$

$$x = 12.5 \text{ g} \times \frac{100 \text{ g}}{1 \text{ L}}$$

$$x = 12.5 \text{ g}$$

add 12.5 - 11 = 1.5 g & adjust vol to 125

2 Alternatively, calculate vol. just for extra: $11 \text{ g} - 10 \text{ g} = 1 \text{ g}$ $\frac{1 \text{ g}}{100 \text{ g/L}} = 0.01 \text{ L}$ $0.11 \text{ L} + 0.01 \text{ L} = 0.12 \text{ L}$ 120 mL

Stuff takes up space!

Molarity (M) is mol solute / L solution

This is **total** solution volume

So when you prepare a solution of a specific molarity, you need to take into account the volume of the solute.

When making solutions from solids...

Calculate & weigh out solid.



Dissolve in partial volume*



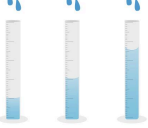
Once dissolved, add solvent (e.g. water) to desired final volume.



*The more solid you need to stuff in there, the more space you need to leave.

When making solutions from liquids...

Measure out & add each component (stock solutions of salts, buffering agents, etc.).



Add solvent* (e.g. water) to desired final volume.



*If the liquids interact or otherwise affect one another (e.g. EtOH & water), the solvent volume might be different than what you'd calculate.

$C_1 V_1 = C_2 V_2$

initial concentration X initial volume = final concentration X final volume

rearrange

isolate the unknown

$$C_1 = \frac{C_2 V_2}{V_1} \text{ or } V_1 = \frac{C_2 V_2}{C_1} \text{ or } C_2 = \frac{C_1 V_1}{V_2} \text{ or } V_2 = \frac{C_1 V_1}{C_2}$$

then exchange

stick in the known
replace those placeholders with numbers
and math it out
solve for the unknown

Relative concentrations

Often, concentrated solutions will be described in terms of relative concentrations (relative to the final, "working" concentration).

1X is the **working concentration** - the concentration you want it to be at in your final solution.

In $C_1V_1=C_2V_2$ terms, 1X is your C_2
The concentration of the stock is your C_1

$$1X = \frac{C_1V_1}{C_2V_2} = \frac{\text{initial concentration} \times \text{initial volume}}{\text{final concentration} \times \text{final volume}}$$

The relative concentration (x-factor) = $\frac{C_1}{C_2}$

The number before the X is the dilution factor of the working concentration

10 X

10X more concentrated than you want

50 X

50X more concentrated than you want

1000 X

1000X more concentrated than you want

Divide final desired volume by x-factor to figure out how much concentrate you need.

$$V_1 = \frac{V_2}{\text{x-factor}}$$

e.g. if you want a final volume of 100mL, add 100/10 = 10mL of 10X; 100/50 = 2mL of 50X or 100/100 = 1mL of 100X stock

Multiply starting volume of concentrate by x-factor to figure out final volume to dilute it to*.

$$V_2 = \text{x-factor} \times V_1$$

e.g. if you start with 10mL concentrate, dilute 10X to 10*10 = 100mL, 50X to 10*50 = 500mL; and 100X to 10*100 = 1000mL (1L)

*Before you dilute it directly to that final volume, keep in mind that, in addition to the concentrate, there might be other components contributing to that final total volume.

calculations with relative concentrations

you can treat relative concentrations just like other concentrations

in terms of $C_1V_1=C_2V_2$
the working concentration, 1X, is your C_2
the stock concentration, something X, is your C_1
the X's cancel out (just like units would) so we can ignore them and focus on the number before the X

$$\frac{C_1V_1}{X \text{ initial volume}} = \frac{C_2V_2}{X \text{ final volume}}$$

the relative concentration (x-factor) = $\frac{C_1}{C_2}$

How much concentrate do I need to make a certain final volume?

$$V_1 = \frac{C_2V_2}{C_1} = \frac{V_2}{\text{x-factor}}$$

if you know the final volume you want, divide by the # before the x to figure out how much of the concentrated stock you need

e.g. if you want a final volume of 100mL, add 100/10 = 10mL of 10X stock

What should my final volume be given starting amount of stock?

$$V_2 = \frac{C_1V_1}{C_2} = \text{x-factor} \times V_1$$

if you know the initial volume you want to use, multiply by the # before the x to figure out how much final volume you want (what should you dilute it to?)

e.g. if you start with 10mL concentrate, dilute 10X to 10*10 = 100mL

tip: if you're making stock solutions, where possible make them with dilution factors of multiples of 10 for easy mental math

for example, if your stock is 1000X, you can quickly calculate that you need 1mL per L

similarly, when possible, choose a final volume that's divisible by the x-factor
for example, if you have a 6X gel loading buffer, prepare 6, 12, 24, 36... μ L

solution-making calculating example - Part 1

$$C_1V_1=C_2V_2$$

initial concentration X initial volume = final concentration X final volume

what stock solutions do you have?

initial concentrations

C_1 1M Tris pH 8
5M NaCl

initial volumes

V_1 this is what you need to find

what final solution do you want?

desired concentrations

C_2 10mM Tris pH 8
100mM NaCl

desired volumes

V_2 250 mL

$C_1V_1=C_2V_2$ it for each component, then add water to get to the final desired volume

try it out (at least the calculating part!) and I'll meet you with the answer in the next figure!

solution-making calculating example

Use 1M Tris & 5M NaCl stock solutions to make 250mL of a solution that is 10mM Tris, 100mM NaCl

convert units so they're consistent

1000 mM = 1 M
so...
10 mM = 0.01 M
100 mM = 0.1 M

Calculate amount of Tris needed

$$C_{1(\text{Tris})}V_{1(\text{Tris})} = C_{2(\text{Tris})}250 \text{ mL}$$

$$1M_{(\text{Tris})}V_{1(\text{Tris})} = 0.01M_{(\text{Tris})}250 \text{ mL}$$

$$V_{1(\text{Tris})} = \frac{0.01M_{(\text{Tris})}250 \text{ mL}}{1M_{(\text{Tris})}}$$

$$V_{1(\text{Tris})} = 2.5 \text{ mL}$$

Calculate amount of NaCl needed

$$C_{1(\text{NaCl})}V_{1(\text{NaCl})} = C_{2(\text{NaCl})}250 \text{ mL}$$

$$5M_{(\text{NaCl})}V_{1(\text{NaCl})} = 0.1M_{(\text{NaCl})}250 \text{ mL}$$

$$V_{1(\text{NaCl})} = \frac{0.1M_{(\text{NaCl})}250 \text{ mL}}{5M_{(\text{NaCl})}}$$

$$V_{1(\text{NaCl})} = 5 \text{ mL}$$

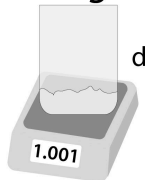
add water to get to the final volume

calculate pH buffer stock making



use the formula weight to figure out how much you need to weigh out to get the desired molarity

weigh



do it

dissolve/dilute partway



leave space for the acid or base



adjust pH

add concentrated acid/base dropwise while monitoring pH change

once it (stably) gets to the desired pH (at the desired temp)

adjust volume

do this in a graduated cylinder - NOT the beaker - and beware of the stir bar!



Tris base has a formula weight of ~121.14 g/mol

We want a 1 molar (1M) solution

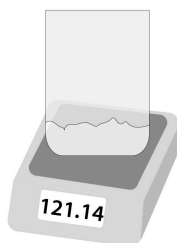
in molarity:

$$\text{MOLARITY} = \frac{\text{moles of something}}{\text{liter}}$$

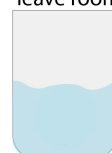
a mole is just 6×10^{23} of something (Avogadro's Number)

$$\frac{1 \text{ mol Tris}}{1 \text{ L solution}} \times \frac{121.14 \text{ g Tris}}{1 \text{ mol Tris}} = \frac{121.14 \text{ g Tris}}{1 \text{ L solution}}$$

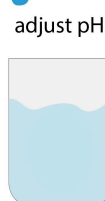
weigh



dissolve
leave room



HCl



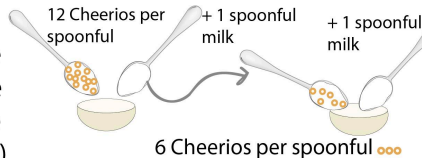
adjust pH

adjust volume



serial dilutions

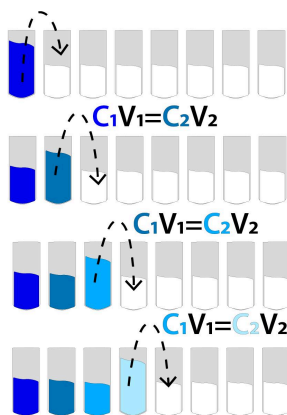
a series of stepwise dilutions where you're diluting a thing by the same factor each time (the dilution factor, DF)



$$C_1 V_1 = C_2 V_2$$

initial concentration X initial volume = final concentration X final volume

in a serial dilution, the final concentration of 1 thing (C_2) becomes initial concentration (C_1) for the next



You can easily figure out the concentration of each dilution in a standard curve without having to math out the whole series.

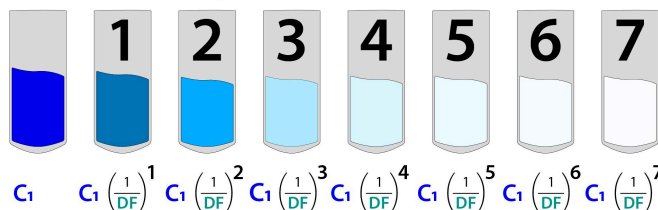
you just need to know the initial concentration, the # of dilutions since then and the **Dilution Factor**: how much are you diluting by each time?

for example, if you halve it each time (i.e. each dilution has 1/2 the concentration of the prior one), you'd have a dilution factor of 2.

if each dilution has 1/3 the concentration of the prior one), you'd have a dilution factor of 3.

if each dilution has 1/4 the concentration of the prior one), you'd have a dilution factor of 4. etc. etc. etc.

multiply the initial concentration by $(1/DF)$ raised to the # of dilutions to get the concentration of that dilution



Adjust your volume to match your measured mass

(Rather than trying to measure out exact tiny masses)

$$\text{desired concentration} = \frac{\text{target mass}}{\text{target volume}} = \frac{\text{measured mass}}{\text{needed volume}}$$

Calculate a **target mass** based on your **target volume** & **desired concentration**

Note: If you start in molarity, convert to mg/mL (= g/L) by multiplying M by the MW

Figure out how much you need, then calculate to make that.

$$\text{desired concentration} = \frac{\text{target mass}}{\text{target volume}}$$

$$\text{target mass} = \text{desired concentration} \times \text{target volume}$$

Measure out a mass close to the target mass

Calculate **needed volume** to reach **desired concentration**

$$\text{desired concentration} = \frac{\text{measured mass}}{\text{needed volume}}$$

$$\text{needed volume} = \frac{\text{measured mass}}{\text{desired concentration}}$$

Be sure your units are consistent so they cancel out appropriately.

Example: You want ~20 mL of 10 mM NADH disodium salt

Convert desired concentration to mg/mL (= g/L) if needed

If you start in molarity, convert to mg/mL (= g/L) by multiplying M by the MW

NADH disodium salt MW (technically a FW) = 709.4 g/mol

$$\frac{10 \text{ mmol}}{\text{L}} \times \frac{1 \text{ mol}}{1000 \text{ mmol}} = \frac{0.010 \text{ mol}}{\text{L}} \times \frac{709.4 \text{ g}}{\text{mol}} = \frac{7.094 \text{ g}}{\text{L}} = \frac{7.094 \text{ g}}{10^3} = \frac{7.094 \text{ mg}}{\text{mL}}$$

Calculate a target mass based on your target volume & desired concentration

$$\text{target mass} = \text{desired concentration} \times \text{target volume}$$

$$\text{target mass} = \frac{7.094 \text{ mg}}{\text{mL}} \times 20 \text{ mL} = 141.9 \text{ mg}$$

Measure out a mass close to the target mass

In this example, say you measured 161 mg...

Calculate needed volume to reach desired concentration

$$\text{needed volume} = \frac{\text{measured mass}}{\text{desired concentration}}$$

$$\text{needed volume} = \frac{161.0 \text{ mg}}{7.094 \text{ mg/mL}} = 22.7 \text{ mL}$$

Reasonableness check: The value is close to 20 mL, but slightly greater than it since the measured mass was slightly greater than the target mass. ✓

Preparing solutions from solids

$$\text{Mass needed} = \text{MW} \times \text{Molarity} \times \text{Volume}$$

$$\text{mass} = \frac{\text{g}}{\text{mol}} \times \frac{\text{mol}}{\text{L}} \times \text{L}$$

Molecular weight (MW) or formula weight (FW) Desired molarity Desired total volume

Calculate & weigh out solid.



Dissolve in partial volume.



Once dissolved, add solvent (e.g. water) to desired final volume.



Preparing solutions from liquids

$$C_1 V_1 = C_2 V_2$$

initial concentration × initial volume = final concentration × final volume

Unlike with solids, volumes & concentrations can be in any unit (even relative things like % or X), but they need to be consistent.

$$V_1 = \frac{C_2 V_2}{C_1}$$

Volume needed Desired concentration Desired total volume

Stock concentration

Measure out & add calculated volume.



When applicable, do this for each additional component (stock solutions of salts, etc.).



Add solvent* (e.g. water) to desired final volume.



Solution-making from solids

calculate



Use the formula weight (F.W.) or molecular weight (M.W.) (g/mol) to figure out how much you need to weigh out to get the desired molarity (mol/L)

weigh

do it

$$\text{Mass needed} = \text{MW} \times \text{Molarity} \times \text{Volume}$$

dissolve

Start with a lower volume than your final volume because the solid will take up space



Add the stir bar

Add the solid gradually

adjust volume

Stir to dissolve

Always start by adding some liquid (but not a final volumes-worth!) to the vessel before solid, or you'll get a hard lump, perhaps with a stir bar stuck in it.



Transfer to a graduated cylinder

Adjust volume in a graduated cylinder - NOT the beaker - & beware of the stir bar!



Reasonableness Checks (RCs)

Silly errors are easy to make when doing calculations. But often they're easy to catch before it's too late. **Make it a habit to incorporate a reasonableness check into your calculation routine.**

Ask yourself: Does my answer at least make sense?

The 3 components of a formal RC

- 1) Expectations:
- 2) Work:
- 3) Reasonableness check:

It's good to explicitly write these out when you're learning new problem types, but then you can do much of it in your head in the future

1) **Before doing calcs**, think qualitatively/general sense what you expect the answer to be.

For example...

- If you add an acid, $[H^+]$ should increase and pH should decrease
- If you add base, $[H^+]$ should decrease and pH should increase
- If thermodynamically-favorable, K_{eq} should be > 1 & ΔG should be $-$
- If thermodynamically-unfavorable, K_{eq} should be less < 1 & ΔG should be $+$

2) Now **do those calcs** (and show your work)

3) **After doing the calcs**, ask yourself: Does your answer make sense based on expectations?



Yes? - it at least has a chance of being right



No? - check your calcs!

Get the reference?

Quick laboratory reference sheets & calculation tips

Concentrations, unit conversions, UV-Vis spectroscopy, error statistics etc.

laboratory_reference_sheets

statistics

serial_dilutions

uv_protein_beer_conversion.pdf

uv_nucleicacid_ratios.pdf

scientific_notation_011925.pdf

relative_concentrations.pdf

protein_concentrations.pdf

molecular_weights.pdf

metric_shortcuts.pdf

metric_conversions.pdf

Important_Lab_Math.pdf

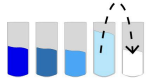
concentrations_cheat_sheet.pdf

Free (as always, no stringed attached) downloadable PDFs in the BB_resources drive. Hope they help!

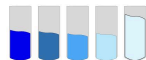
<https://bit.ly/labreferencesheets>

Remember you need to leave a desired volumes-worth in each tube

You calculated for this volume



But are only left with this volume



Each dilution, you're taking out some of what you made
So, in order to still have a desired final volume's worth left in the tube, you'll need to make extra

$$\frac{\text{transferred volume}}{\text{total volume}} = \frac{1}{\text{DF (dilution factor)}}$$

$$\text{total (intermediate) volume} - \text{transferred volume} = \text{diluent (e.g. buffer or water) volume}$$

$$\text{volume left in tube} = \text{diluent volume}$$

You need to calculate so that the diluent volume = the amount you want left in each tube (desired volume)

$$\text{total (intermediate) volume} = \frac{\text{desired volume}}{1 - (1/\text{DF})}$$

Make $(1/(1 - 1/\text{DF}))$ -times as much as you need of each intermediate dilution

$$\text{transferred volume} = \text{desired volume} \left(\frac{1}{1 - (1/\text{DF})} - 1 \right)$$

Serial dilution calculations to make sure you're left with how much you want

$$\frac{\text{transferred volume}}{\text{total volume}} = \frac{1}{\text{DF}}$$

$$\text{total (intermediate) volume} - \text{transferred volume} = \text{diluent (e.g. buffer or water) volume}$$

$$\text{volume left in tube} = \text{diluent volume}$$

You need to calculate so that the diluent volume = the amount you want left in each tube (desired volume)

$$\text{total (intermediate) volume} = \text{DF} \times \text{transferred volume}$$

$$\text{total volume} = \text{desired volume} + \text{transferred volume}$$

$$\text{total volume} = \text{desired volume} + \frac{(1/\text{DF}) \times \text{total volume}}{\text{volume}}$$

$$\text{total volume} - ((1/\text{DF}) \times \text{total volume}) = \text{desired volume}$$

$$\text{total volume} (1 - (1/\text{DF})) = \text{desired volume}$$

$$\text{total volume} = \frac{\text{desired volume}}{1 - (1/\text{DF})}$$

$$\text{transferred volume} = \text{total volume} - \text{desired volume}$$

$$\text{transferred volume} = \frac{\text{desired volume}}{1 - (1/\text{DF})} - \text{desired volume}$$

$$\text{transferred volume} = \text{desired volume} \left(\frac{1}{1 - (1/\text{DF})} - 1 \right)$$



Lab tip: Make Master Mixes

Determine which reagents are the same in each reaction mixture and mix those together (making enough for all reactions + a little extra to account for volume loss in transfers, etc.), then add that to the part(s) that are different.

Master Mixes

make sure you make enough happiness to share!
make enough as you need for # of people and a little to spare!

# of people	10		
prepare as if there were	11		
per person		total to make	
reagent	quantity	reagent	quantity
puppies	5	puppies	=S9*\$T\$3
kittens	1	kittens	
rainbows	2	rainbows	
Parafilm	1	Parafilm	
just add person!			

# of people	10		
prepare as if there were	11		
per person		total to make	
reagent	quantity	reagent	quantity
puppies	5	puppies	55
kittens	1	kittens	11
rainbows	2	rainbows	22
Parafilm	1	Parafilm	11
just add person!			

Master Mixes

MORE SERIOUSLY...

# of reactions	8		
prepare as if there were	8.5		
per reaction		total to make of mastermix	
reagent	quantity (uL)	reagent	quantity
RNA (at 50uM)	2		
10X PNK buffer	5	10X PNK buffer	42.5
water	37	water	314.5
hot ATP	5	hot ATP	42.5
PNK	1	PNK	8.5
final volume	50	volume mastermix to make	
			408
amount of master mix per sample	48	amount of wiggle room	
		amount you "need": $8 \times 48 = 384$	
		$408 - 384 = 24\text{uL extra}$	

Always prepare more than you think that you need...

You have to account for volume loss due to liquid sticking on the pipets, tubes, etc.
&/or evaporating out.

and, if possible, prepare enough to redo if at first you don't succeed!

2 + extra rule of thumb

Ideally, make/save enough sample to redo an experiment if it doesn't work without having to go all the way back to actually getting &/or preparing that sample.