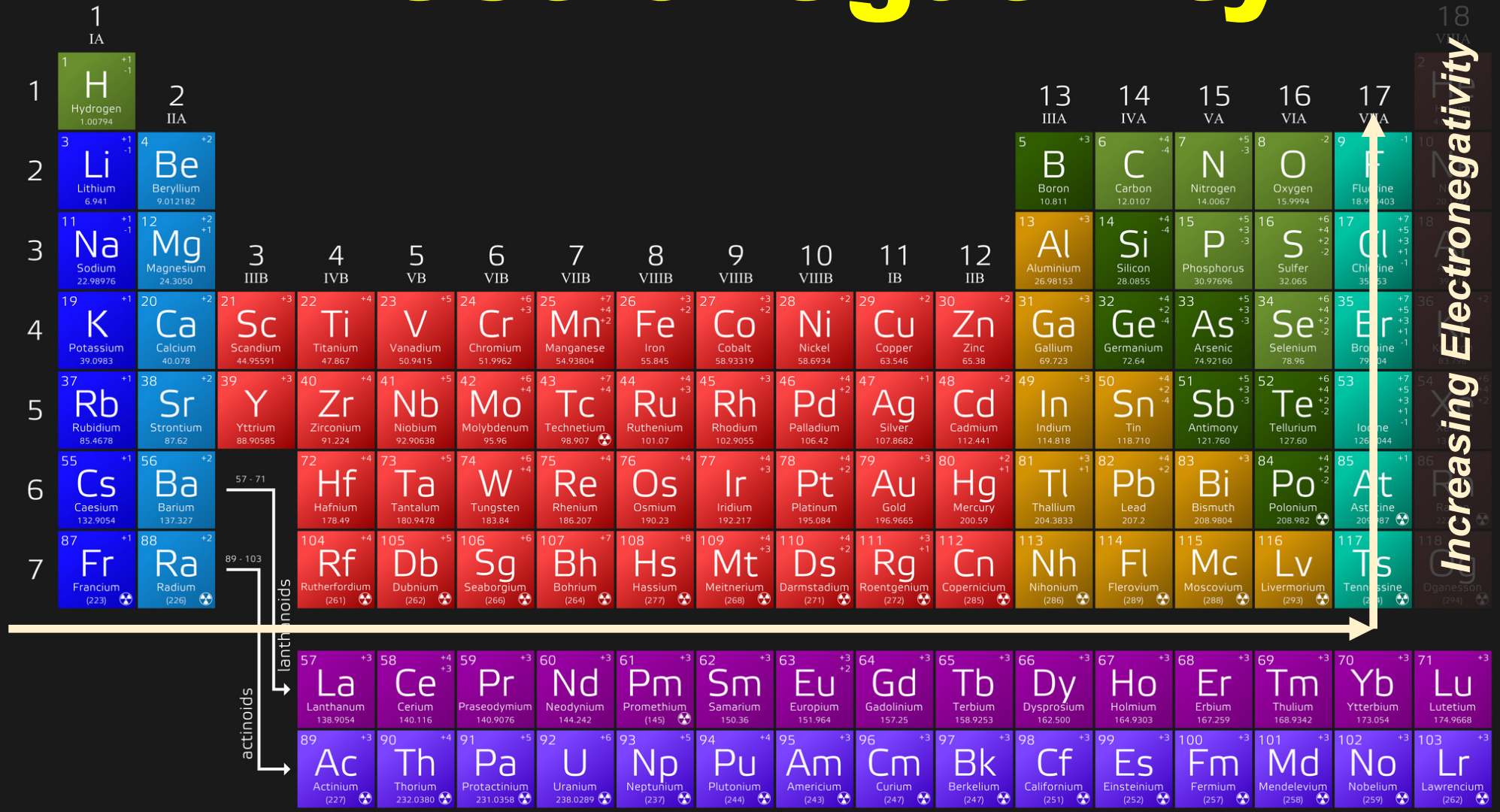
Solutions
(Solubility)

Polar or Nonpolar?

As I've taught previously, when bonded atoms have a significant difference in their electronegativities, they share their electrons unequally. This causes their bond to be polar. We should, of course, remember that . . .

Electronegativity

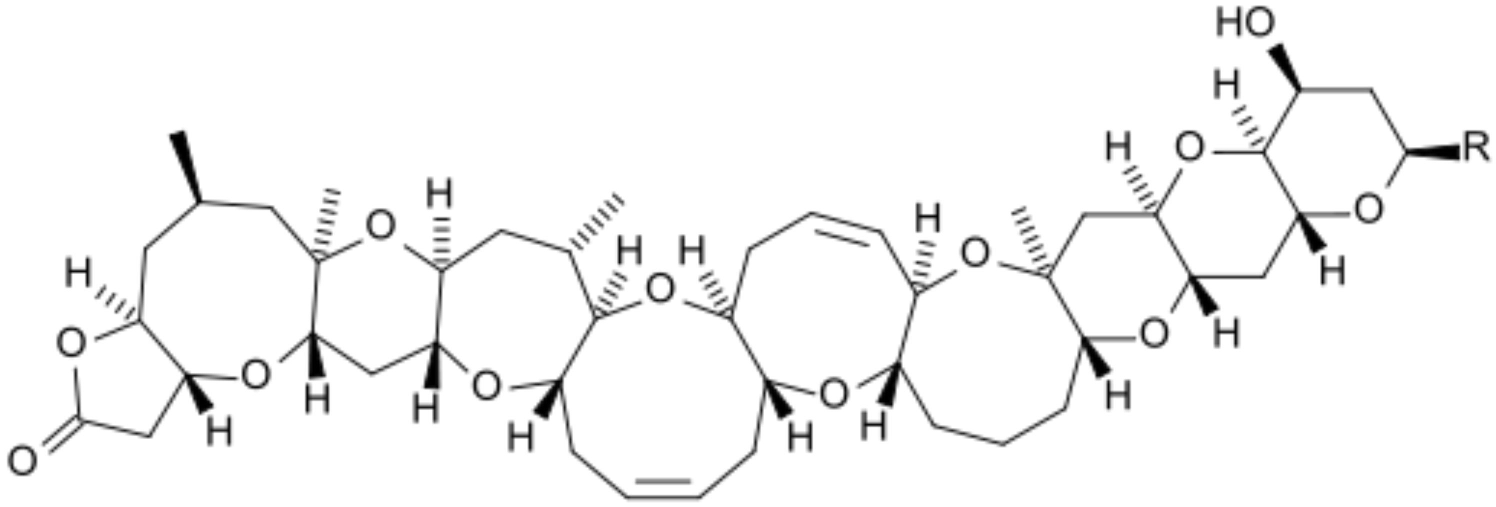


. . . An atom's ***electronegativity*** (its "thirst" for electrons) increases as you go up and to the right on the periodic table.

Polar or Nonpolar?

As I've taught previously, when bonded atoms have a significant difference in their electronegativities, they share their electrons unequally. This causes their bond to be polar. We should, of course, remember that . . .

Determining the ***polarity*** of entire complex molecules is complicated, as their ***polarities*** are the collective product of the combined polarities and geometries of their individual bonds.

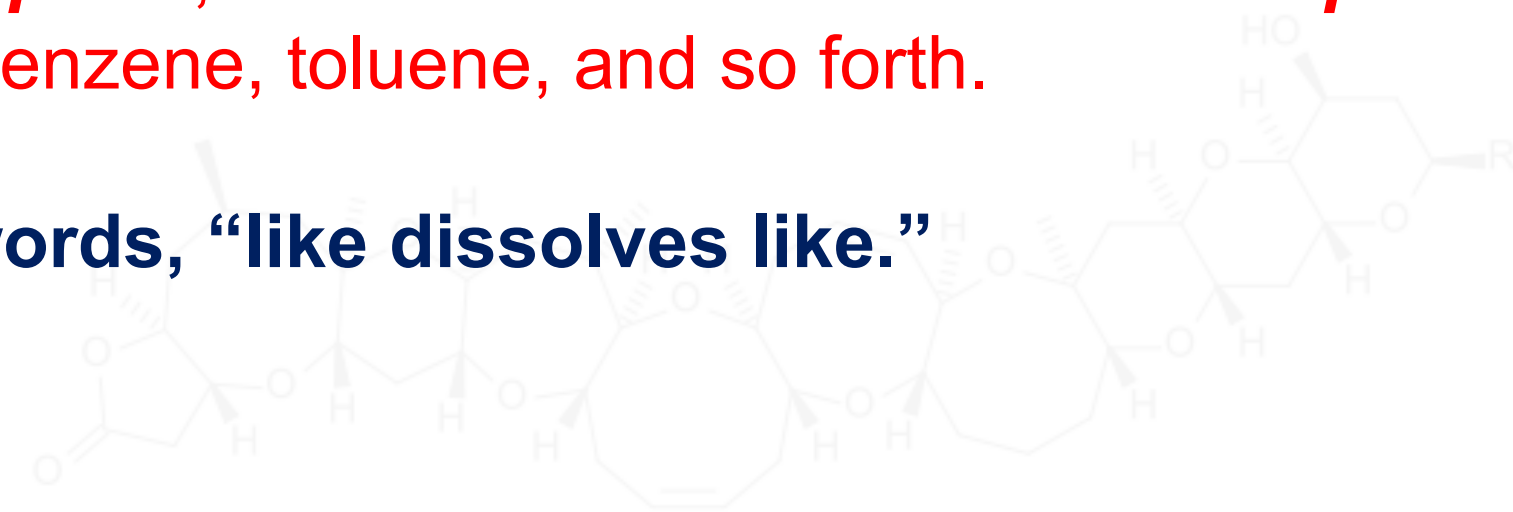


Brevetoxin A

Polar or Nonpolar?

We care about ***polarity*** because ***polarity*** determines solubility. In other words, if something is ***polar*** then it will be soluble in water (and other polar solvents). If it's ***nonpolar***, then it will be soluble in ***nonpolar*** solvents, such as hexane, benzene, toluene, and so forth.

In other words, "like dissolves like."



Polar or Nonpolar?

For simpler molecules, we can more easily determine if they are ***polar*** or ***nonpolar***. Here are some tips:

- O–H bonds are polar. The more OH's your molecule has, the more polar it is.
- C–H bonds are nonpolar. The more CH's your molecule has, the more nonpolar it is.

For other simple molecules, I'll now give you some quick rules.

Polar or Nonpolar?

So how do we determine if a molecule is polar or not?

1. Draw your molecule's ***Lewis structure*** flat on your paper and then spread out all of its groups (including its lone pairs) as far as possible, around its central atom.

2. Draw arrows between every atom in the molecule, going from the less electronegative atom (**A**) to the more electronegative atom (**B**) in each bond, like this:

A – B

3. Answer the following question: "If my central atom were a truck stuck in the mud being pulled in the directions indicated by the arrows, would the truck move?" If so, then your molecule is ***polar***. If not, then it's ***nonpolar***.

Example Question

Which of the following, when added to carbon tetrachloride, will form a miscible solution?

I. hexane
II. water
III. propane
IV. ethyl acetate

A. **I** and **III** only
B. **II** and **IV** only
C. **I**, **III**, and **IV**
D. **I** and **IV** only
E. **III** and **IV** only

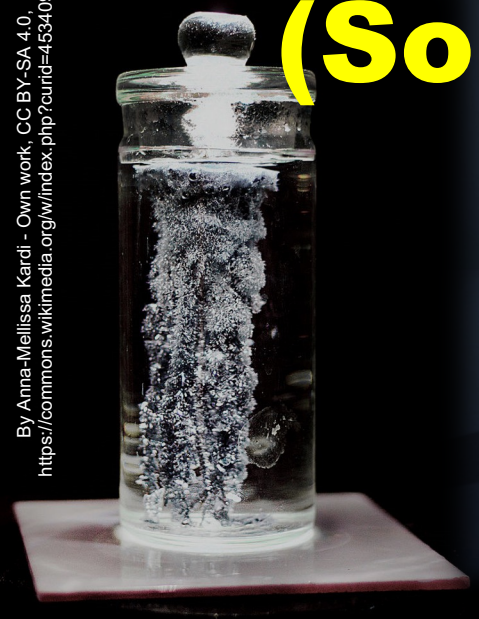
Example Question

Arrange the following alcohols in order of increasing solubility in water

I. methanol, CH_3OH
II. ethanol, CH_3CH_2OH
III. n-propanol, $CH_3CH_2CH_2OH$

A. I < II < III
B. I = II = III
C. I < III < II
D. III < II < I
E. II < III < I

Solutions
(Solutions and Concentrations)



Important Vocab

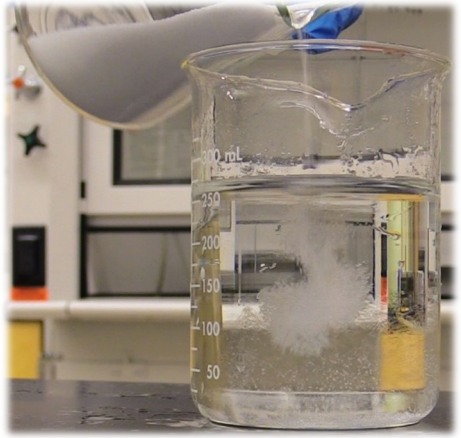
Solution: a uniform mixture of two or more substances

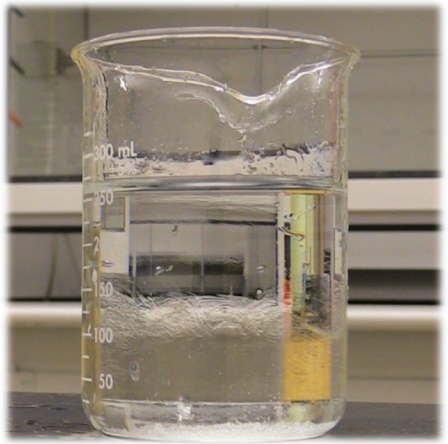
Solvent: The substance in a solution that is present in a larger amount. For the EXAM, this is usually a liquid. In our table salt/water example, water is the ***solvent***.

Solute: The substance in a solution that is present in a smaller amount. In our table salt/water example, salt is the ***solute***.

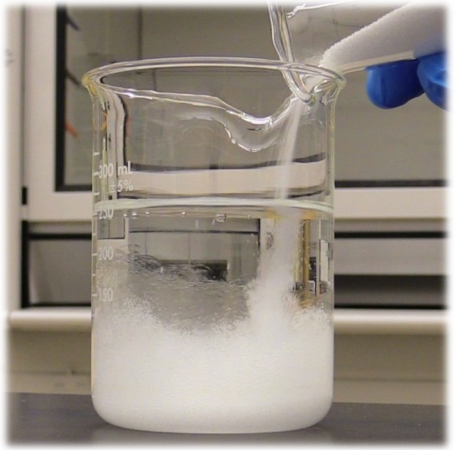
Important Vocab

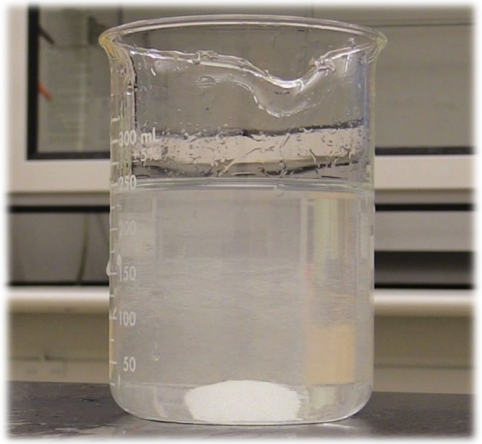
Water is, of course, able to dissolve table salt up to a certain point. If we keep adding salt, however, then we eventually get to the point where the water can no longer dissolve any more salt.





unsaturated





saturated

This is called water's ***saturation point***. Prior to reaching this point, you have an ***unsaturated solution*** of aqueous NaCl. After this point, you have a ***saturated solution*** of aqueous NaCl.

Concentration

The amount of ***solute*** in a solution is called the solution's ***concentration***. For the EXAM, concentration can be expressed in two different ways: ***molality*** and ***molarity***.

$$\text{molality } (m) = \frac{\text{moles solute}}{\text{kg of solvent}}$$

$$\text{molarity (M)} = \frac{\text{moles solute}}{\text{L of solution}}$$

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Molarity is useful for determining a solution's concentration. Or, if you know the concentration and volume, then you can figure out how many moles or grams of solute you have.

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Note the difference between ***molality*** and ***molarity***. They both have moles of solute in the numerator. However, ***molality*** is divided by "kg solvent," while ***molarity*** is divided by "L of solution." ***Solution*** includes both ***solute*** and ***solvent***.

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We use ***molarity*** most of the time, but we use ***molality*** when dealing with ***colligative properties***, which we'll discuss in a later video in this chapter.

Questions

You have 2.0 L of 0.5 M NaCl. How many grams of NaCl do you have?

Questions

What mass (in grams) of $MnSO_4$ (150 g/mol) is required to prepare a 5.0 *m* manganese sulfate solution with 500 g of water?

A. 10 g

B. 375 g

C. 750 g

D. 1500 g

E. 3000 g

Solutions
(Solubility Rules)

Physical States

As you probably already know, we use little subscript notations to indicate a substance's physical state, as follows:

$Cu_{(s)}$ ← solid or a solid precipitate

$H_{2\,(g)}$ ← gas

$H_2O_{(l)}$ ← liquid

$Na^+_{(aq)}$ ← dissolved in water

Solubility Rules

There are certain ionic compounds that dissolve in water and other ionic compounds that do not. The "rules" about which ones dissolve in water and which ones do not are called ***solubility rules***.

These ***solubility rules*** can get pretty complicated. To keep things simple, you should just memorize the following two rules, which will cover 99% of the solubility questions you're likely to see on the EXAM.

Rule #1: Most ionic compounds that have any of the following in them are soluble in water:

Group I metal cations, **nitrate** (NO_3^-) **perchlorate** (ClO_4^-), **acetate** ($C_2H_3O_2^-$), and **ammonium** (NH_4^+)

Solubility Rules

There are certain ionic compounds that dissolve in water and other ionic compounds that do not. The "rules" about which ones dissolve in water and which ones do not are called ***solubility rules***.

These ***solubility rules*** can get pretty complicated. To keep things simple, you should just memorize the following two rules, which will cover 99% of the solubility questions you're likely to see on the EXAM.

Rule #2: Most ionic compounds that have any of the following in them are insoluble in water:

silver (Ag^{+}), **lead** (Pb^{2+}), **sulfide** (S^{2-}), **hydroxide** (OH^{-}),
dimercury (Hg_2^{2+}), **carbonate** (CO_3^{2-}), **phosphate** (PO_4^{3-})

Solubility Rules

Solubles

Group I metal cations
nitrate (NO_3^-)
perchlorate (ClO_4^-)
acetate ($C_2H_3O_2^-$)
ammonium (NH_4^+)

Insolubles

silver (Ag^+)
lead (Pb^{2+})
sulfide (S^{2-})
hydroxide (OH^-)
dimercury (Hg_2^{2+})
carbonate (CO_3^{2-})
phosphate (PO_4^{3-})

The solubles generally trump the insolubles.

Questions

Predict whether each of the following ionic compounds is soluble or insoluble in water:

a. $AgNO_3$

b. $AgCl$

c. PbI_2

d. Na_2CO_3

e. $(NH_4)_2CO_3$

g. $HgCl_2$

h. $AlPO_4$

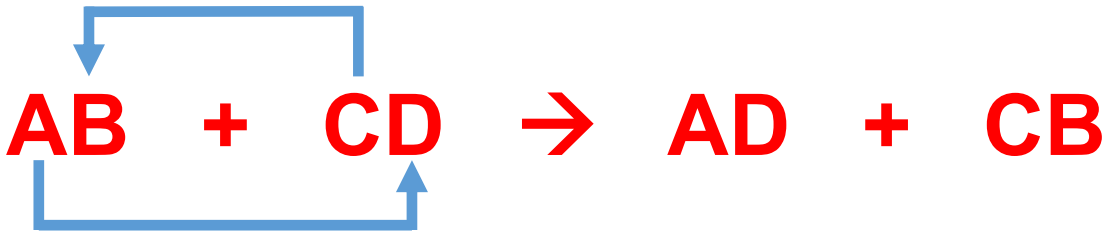
i. MgS

j. $Zn(CH_3COO)_2$

Solutions
(Net Ionic Equations)

Ionic Stuff

When two different ionic compounds are dissolved together in water, they sometimes react with each other. The way they react is by doing a ***metathesis*** (also called a ***double-displacement***) reaction. It's essentially a "partner swap":



This partner swap will lead to an actual reaction if either **AD** or **CB** are insoluble in water.

Ionic Stuff

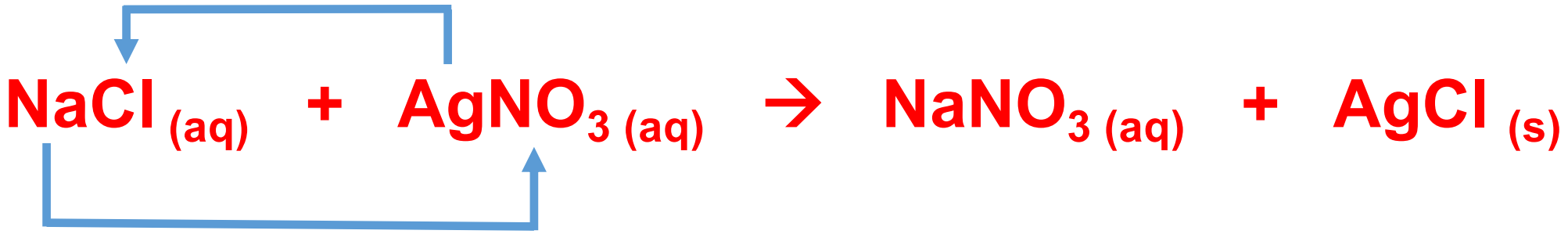
If not, then the reactants just dissolve . . .

$$AB + CD \rightarrow A^+_{(aq)} + B^-_{(aq)} + C^+_{(aq)} + D^-_{(aq)}$$

. . . Unless **AB** or **CD** are insoluble in water, in which case this kind of thing happens:

$$AB_{(s)} + CD_{(aq)} \rightarrow AB_{(s)} + C^+_{(aq)} + D^-_{(aq)}$$

Example



Because one of the products (**AgCl**) *is* insoluble in water, this reaction will occur.

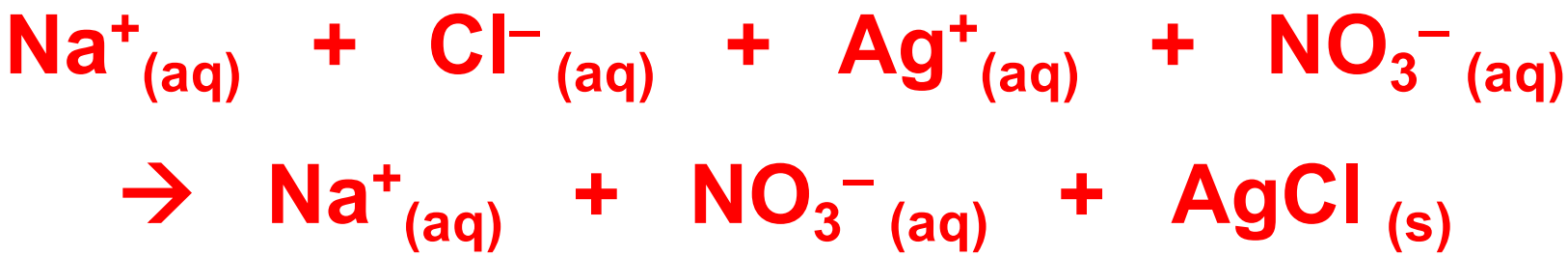
This equation is called a ***complete ionic equation***.

$$NaCl_{(aq)} + AgNO_{3\,(aq)} \rightarrow NaNO_{3\,(aq)} + AgCl_{(s)}$$

Example

In water, everything with an **(aq)** next to it dissolves. Thus, in real life, all the **(aq)** species in this reaction . . .

. . . Actually dissolve to separate their cations from their anions, like this:



Example

This equation . . .

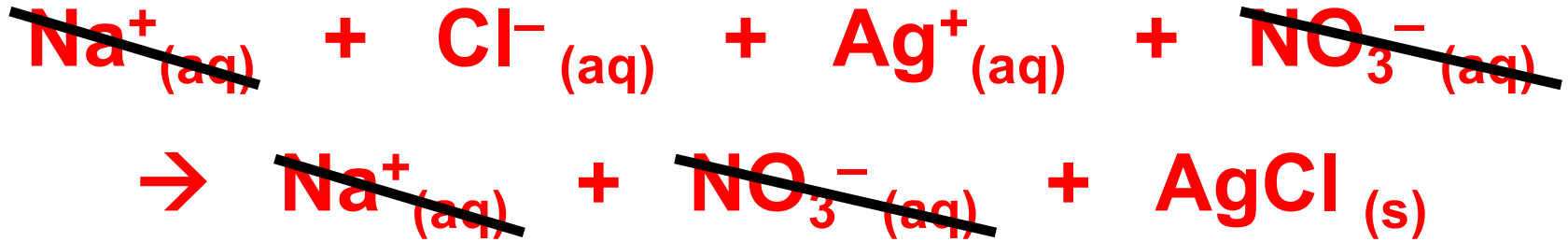
$$Na^+_{(aq)} + Cl^-_{(aq)} + Ag^+_{(aq)} + NO_3^-{}_{(aq)}$$

$$\rightarrow Na^+_{(aq)} + NO_3^-{}_{(aq)} + AgCl_{(s)}$$

. . . Is called the ***total ionic equation***.

Example

We can cancel out everything that is exactly the same on both sides:



The stuff that remains . . .

. . . Is called the ***net ionic equation***.

Example

We can cancel out everything that is exactly the same on both sides:

$$\cancel{Na^+_{(aq)}} + Cl^-_{(aq)} + Ag^+_{(aq)} + \cancel{NO_3^-{}_{(aq)}}$$

$$\rightarrow \cancel{Na^+_{(aq)}} + \cancel{NO_3^-{}_{(aq)}} + AgCl_{(s)}$$

The stuff that remains is called the ***net ionic equation***.

$$Cl^-_{(aq)} + Ag^+_{(aq)} \rightarrow AgCl_{(s)}$$

The things that get canceled out (Na^+ and NO_3^-) are called ***spectator ions***.

To Summarize

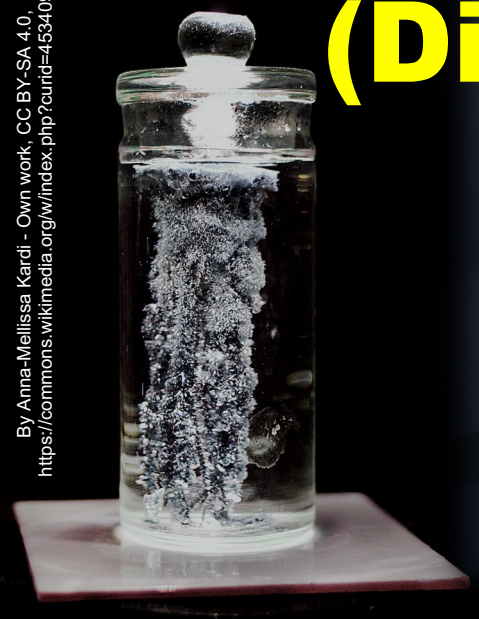
To generate a ***net ionic equation***:

1. Balance the chemical equation.
2. Use your ***solubility-rule*** knowledge to label everything as **(s)**, **(l)**, **(g)** or **(aq)**.
3. Cut all the **(aq)**'s "in half," separating their cations from their anions.
4. Cancel out all the species that are exactly the same on both sides. canceled-out species are called ***spectator ions***.

Question

Please write down the ***complete ionic equation***, ***total ionic equation***, and ***net ionic equation*** for the reaction of ammonium iodide with lead (II) nitrate. What are the ***spectator ions*** in this reaction?

Solutions
(Dissolving Solids and Gases)



Solids' Solubility

Solids, of course, can often dissolve in liquids. Generally speaking, solids get more soluble (they dissolve more) as you raise the temperature:

↑ **Temperature** ≈ ↑ **Solids' solubility**

You've likely experienced the results of this: it's much easier to dissolve sugar in hot drinks than in cold drinks.

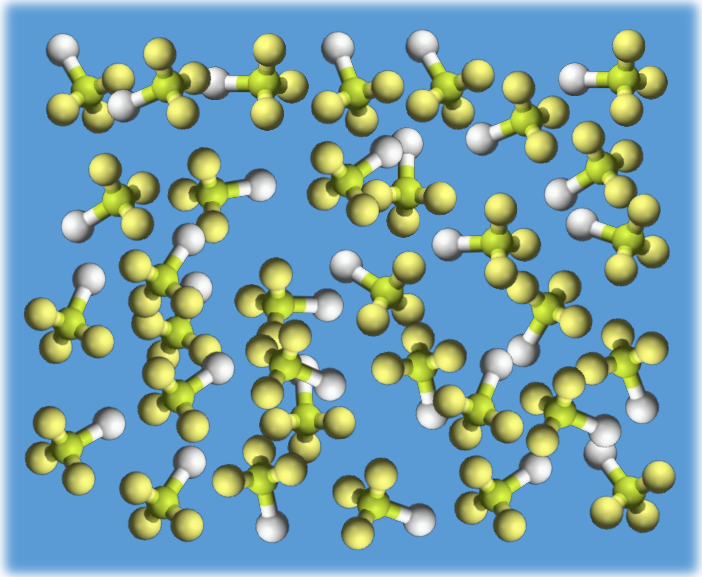
By TheCulinaryGeek from Chicago, USA - Coffee with Cream and SugarUploaded by the wub, CC BY 2.0, https://commons.wikimedia.org/w/index.php?curid=27753493

Gases' Solubility

Gases are the opposite. Unlike solids, gases get less soluble (they dissolve less) as you raise temperature:

↑ **Temperature** ≈ ↓ **Gases' solubility**

Why? The reason is because gas molecules move faster as temperature goes up. When a gas is dissolved in a liquid, raising the temperature will increase the gas molecules' speeds, which causes more and more of those gas molecules to escape and leave the liquid.



Gases' Solubility

It hopefully makes sense, then, that:

↑ **Temperature** ≈ ↓ **Gases' solubility**

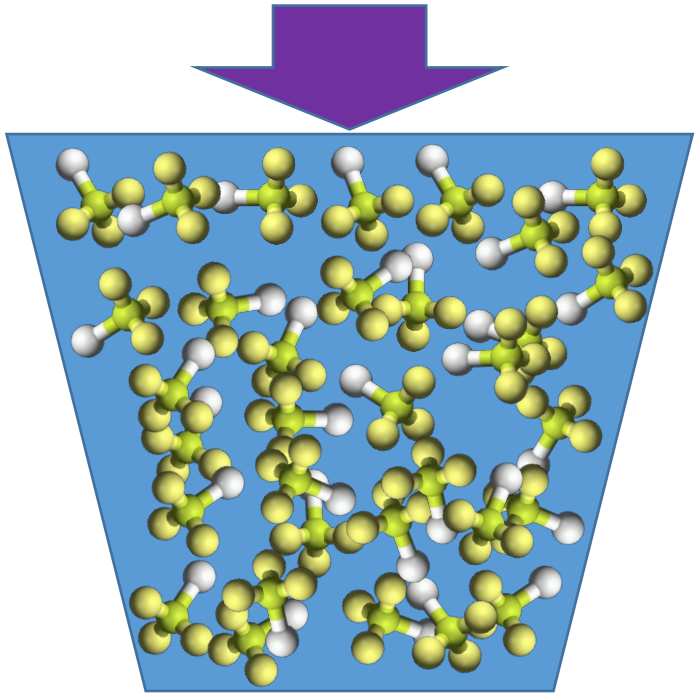
You can dissolve more gas into a liquid at *low* temperatures.

This is why soda is served cold. Cold liquid soda allows more carbon dioxide (CO_2) to dissolve in your drink, which gives it more fizz.

Gases' Solubility

Gases are also more soluble at high pressure:

↑ Pressure ≈ ↑ Gases' solubility



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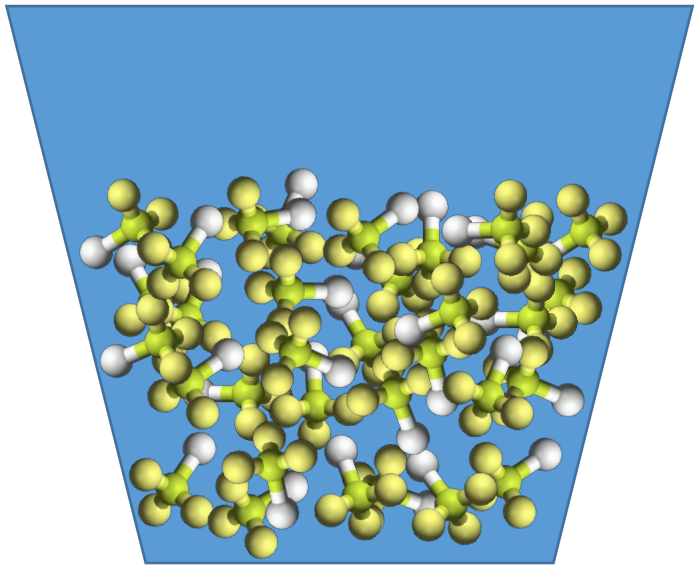

This is why a bottle of soda is under pressure before you open it. More pressure in the bottle means more CO_2 is dissolved in your drink.

Henry's Law

In a sealed vessel containing dissolved gas, the gas' pressure and concentration are proportional, according to ***Henry's Law:***

$$P_A = k_H [A]$$

- P_A is the gas' pressure in the solution.
- **[A]** is the gas' concentration in moles per liter (***molarity***).
- k_H is a constant that differs for each gas, solvent, and temperature.

Henry's Law

The point is: a gas' pressure in solution is proportional to the amount of dissolved gas in the liquid.

If you double the amount of pressure, you double the amount of gas that can dissolve in solution, and vice versa.

Henry's Law

Nitrogen gas' solubility at ambient temperature and 1 atm of pressure is 6.8×10^{-4} mol/L. If the pressure of nitrogen gas in the air is reduced to 0.5 atm, then what is its new concentration?

a. 6.8×10^{-4} M

b. 3.4×10^{-4} M

c. 1.4×10^{-3} M

d. 4.5×10^{-4} M

e. 1.7×10^{-4} M

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÷2 → $P_A = k_H [A]$ ← ÷2

$6.8 \times 10^{-4} \div 2 = 3.4 \times 10^{-4}$

Henry's Law

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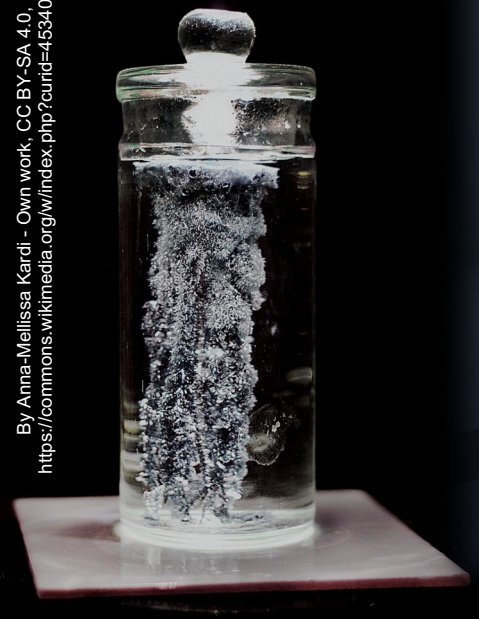
Solutions
(Colligative Properties)



Freezing Point

Colligative properties are properties of a solution that change as you add more ***solute***.

For example, if you dissolve a substance in a liquid, the liquid's ***freezing point*** goes down.

You may have experienced this if you ever threw salt on an icy sidewalk to get the ice to melt.

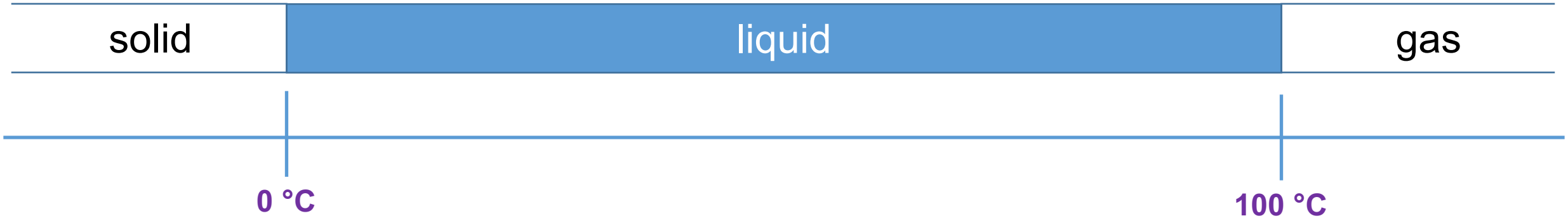
↑ **solute** ≈ ↓ **freezing point**

Boiling Point

In contrast, dissolving a substance in a liquid causes the liquid's ***boiling point*** to go up:

↑**solute** ≈ ↑**boiling point**

Pure water at ***STP:***

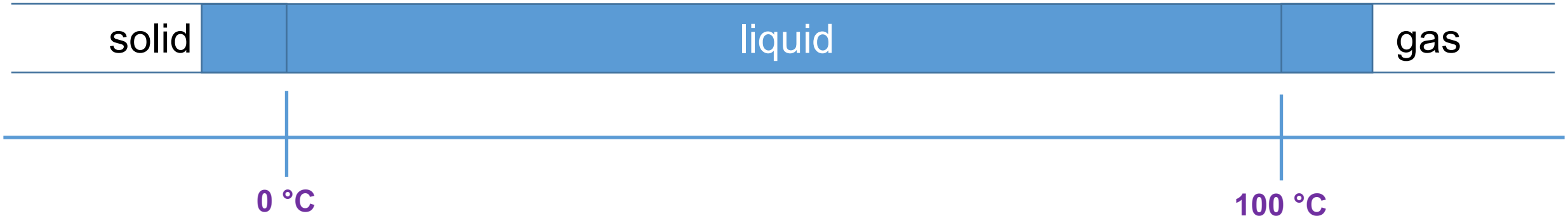


Boiling Point

In contrast, dissolving a substance in a liquid causes the liquid's ***boiling point*** to go up:

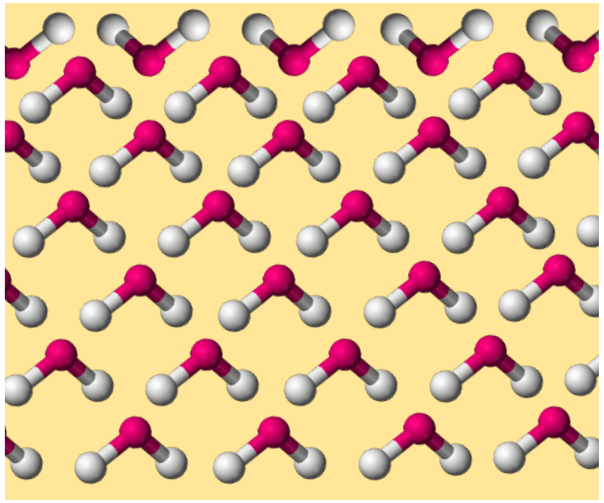
↑**solute** ≈ ↑**boiling point**

Water at ***STP*** with a ***solute*** dissolved in it:

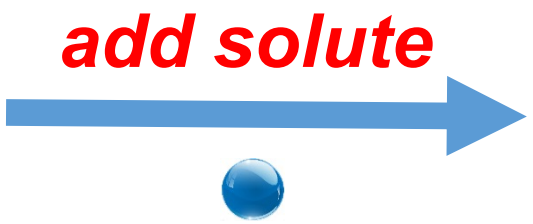


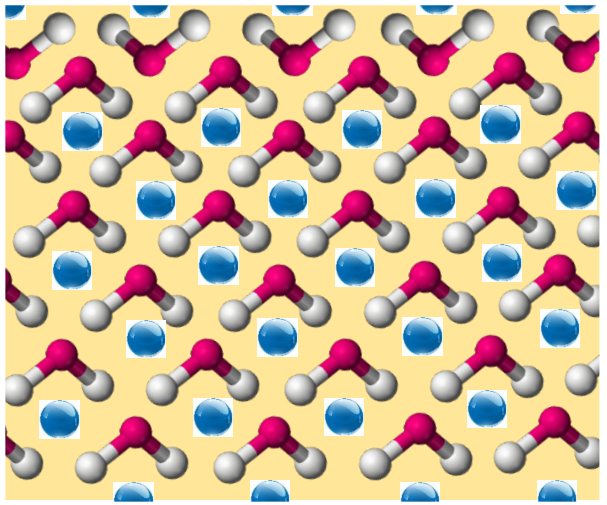
Ethylene glycol (***antifreeze***) expands the temperature at which water stays liquid.

Why (Freezing Point)

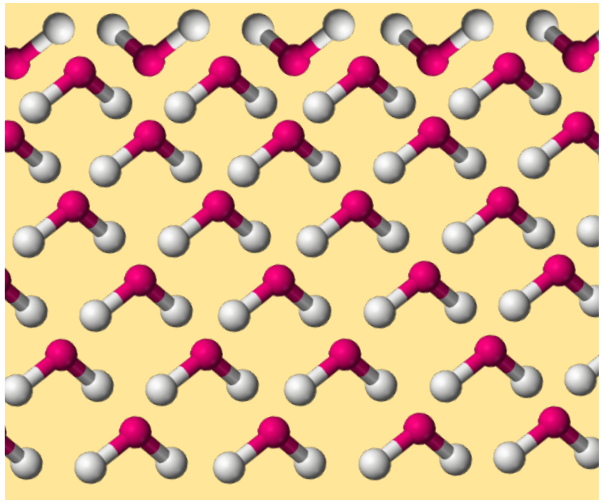


ice



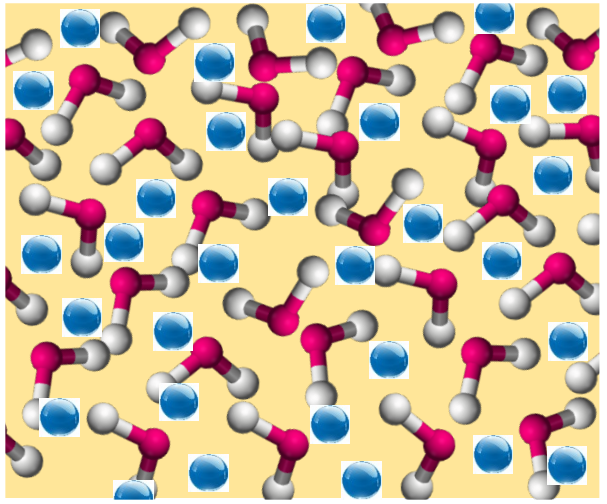


Why (Freezing Point)



ice

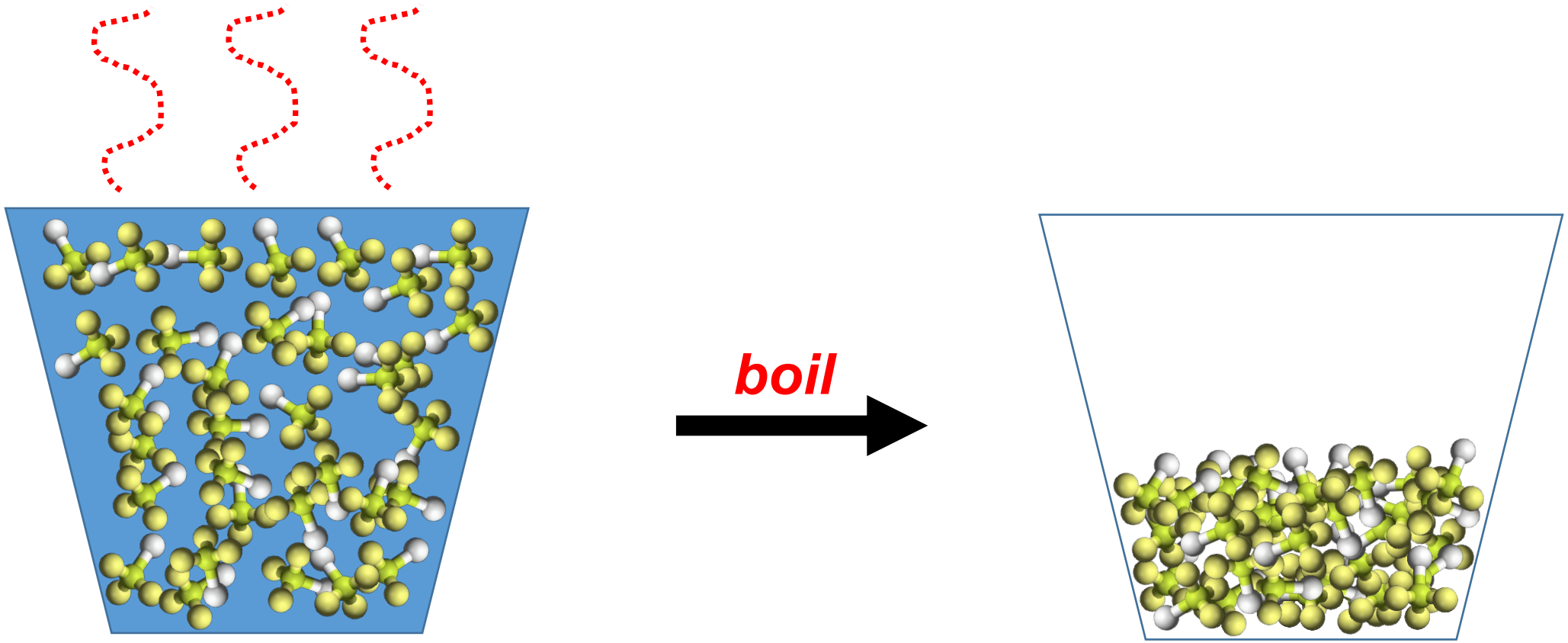
add solute



liquid

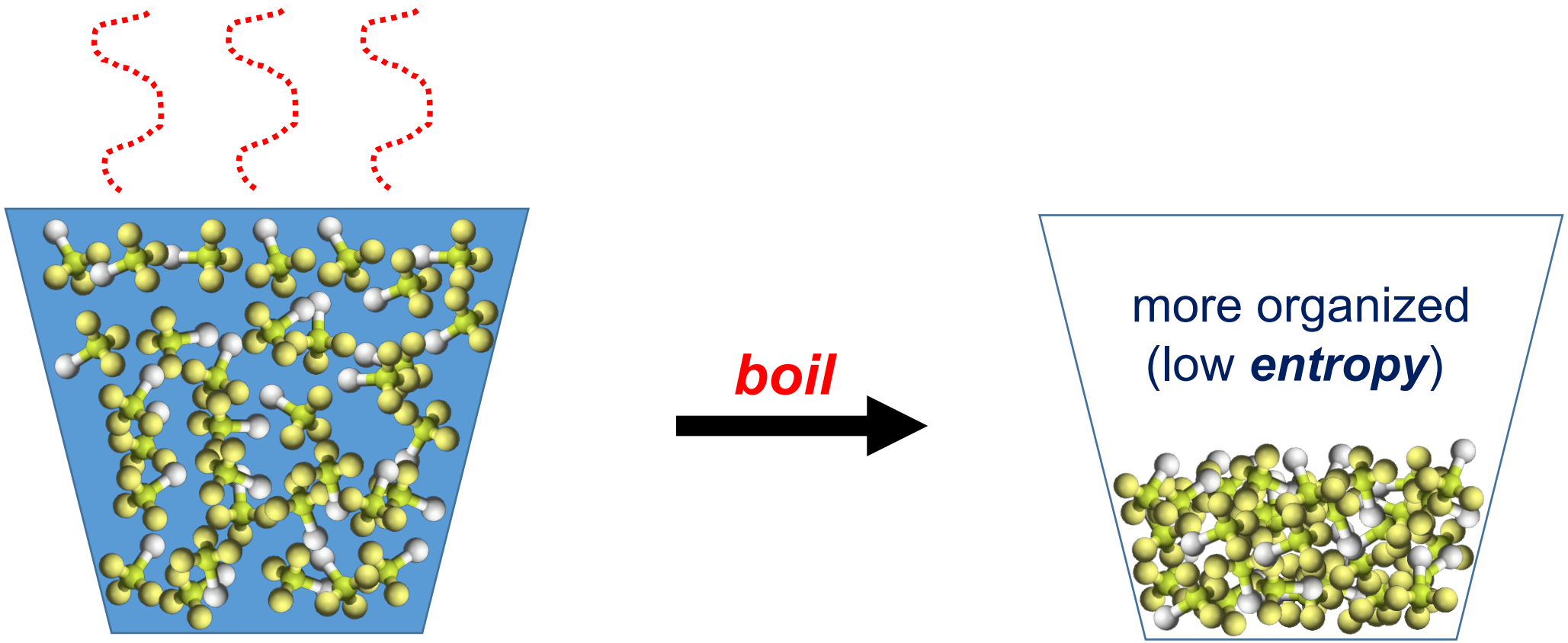
Why (Boiling Point)

When a solute is dissolved in a liquid, if you boil the liquid off, then you're left with solute molecules encrusted together at the bottom of your container.



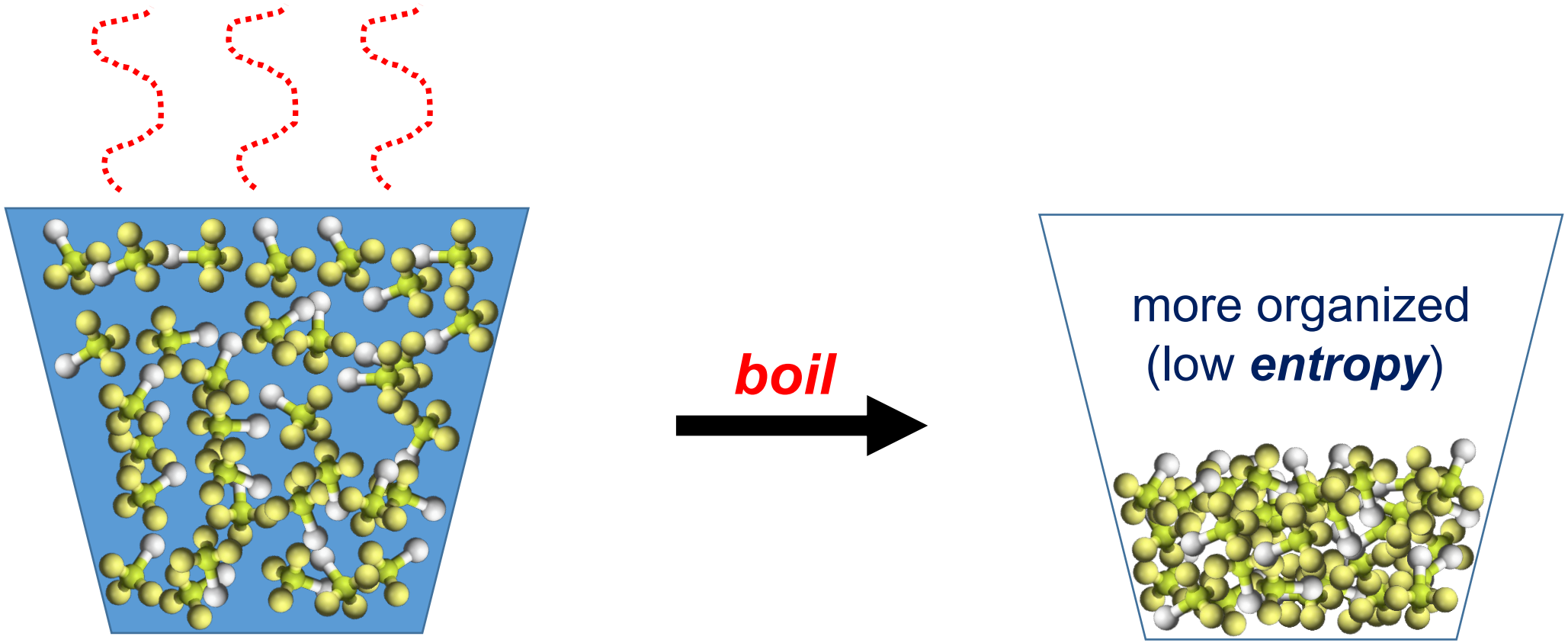
Why (Boiling Point)

For the solute molecules, their crystallized state is much more organized (lower in ***entropy***) than they were before the liquid evaporated. To avoid this more-organized (lower ***entropy***) state, the solution will avoid vaporizing (evaporating).



Why (Boiling Point)

↑solute ≈ ↑boiling point



Math Stuff (Freezing Point)

We can calculate the change in ***freezing point*** that a liquid will experience when a ***solute*** is dissolved in it, by using the following equation:

$$\Delta T_F = -iK_F m$$

- ΔT_F is the change to the liquid's freezing temperature.
- i is the ***Van't Hoff factor***.
- K_F is the freezing point constant.
- m is the solution's ***molality*** (moles of solute / kg solvent).

Math Stuff (Freezing Point)

The ***Van't Hoff factor*** increases as the number of molecules or ions in the solution goes up.

$$\Delta T_F = -iK_F m$$

For example, **NaCl** dissolves into **Na^+** and **Cl^-**, which is two things. **$Ca(NO_3)_2$** dissolves to form **Ca^{2+}** and two **NO_3^-** ions, which is three things. Thus, dissolving **$Ca(NO_3)_2$** would decrease water's freezing point more than dissolving the same amount of **NaCl**. **Glucose** dissolves water, but doesn't produce any ions, so you only get one thing (one mole of dissolved **glucose**) per mole of **glucose** that you add.

Math Stuff (Boiling Point)

We can calculate the change in ***boiling point*** that a liquid will experience when a ***solute*** is dissolved in it, by using the following equation:

$$\Delta T_B = iK_B m$$

- ΔT_B is the change to the liquid's boiling temperature.
- i is the ***Van't Hoff factor***.
- K_B is the boiling point constant.
- m is the solution's ***molality*** (moles of solute / kg solvent).

Questions

Which of the following solutes would decrease water's freezing point by the greatest amount, per mole of solute added?

A. $NaCl$ → **$Na^+ + Cl^-$ (two things)**

B. NH_4NO_3 → **$NH_4^+ + NO_3^-$ (two things)**

C. NH_4Cl → **$NH_4^+ + Cl^-$ (two things)**

D. Na_2SO_4 → **$2Na^+ + SO_4^{2-}$ (three things)**

E. sucrose → **sucrose (one thing)**

Questions

Which of the following solutes would decrease water's freezing point by the greatest amount, per mole of solute added?

A. $NaCl$

B. NH_4NO_3

C. NH_4Cl

D. Na_2SO_4

E. sucrose

Questions

Which of the following solutes would change water's boiling point by the greatest amount?

A. 1.2 *m* $C_6H_{12}O_6$ → **$C_6H_{12}O_6$ (one thing)**

B. 0.8 *m* $Ca(NO_3)_2$

C. 1.0 *m* KCl

Questions

Which of the following solutes would change water's boiling point by the greatest amount, per mole of solute added?

A. 1.2 *m* $C_6H_{12}O_6$

B. 0.8 *m* $Ca(NO_3)_2$

C. 1.0 *m* KCl

Questions

Which of the following solutes would change water's boiling point by the greatest amount, per mole of solute added?

A. 1.2 *m* $C_6H_{12}O_6$ → **$C_6H_{12}O_6$ (one mole) × 1.2 *m* = 1.2 *m***

B. 0.8 *m* $Ca(NO_3)_2$ → **Ca^{2+} + 2 NO_3^- (3 moles) × 0.8 *m* = 2.4 *m***

C. 1.0 *m* KCl → **K^+ + Cl^- (2 moles) × 1.0 *m* = 2.0 *m***

Questions

If you have a 1.0 *m* aqueous solution of NaCl, by how much will it increase the water's boiling point, if K_B = 0.512 C/m? In other words, what is the boiling point elevation (increase)?

Solutions
(Raoult's Law)

A Colligative Reminder

As I mentioned in a prior video, ***colligative properties*** are properties of a solution that change as you add more ***solute***.

For example, if you dissolve a substance in a liquid, the liquid's ***freezing point*** goes down, and its ***boiling point*** goes up.

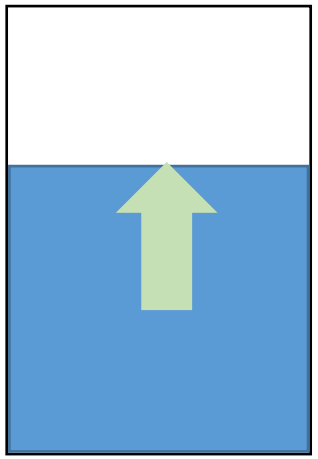
↑ **solute** ≈ ↓ **freezing point** ≈ ↑ **boiling point**

We also learned that it is a drive to avoid decreased ***entropy*** that causes a solution's ***boiling point*** to go up when you add a ***solute***.

Boiling Point and Vapor Pressure

You should also remember from previous videos that ***boiling point*** and ***vapor pressure*** are connected.

↓ **boiling point** ≈ ↑ **vapor pressure**



This means:

↑ **solute** ≈ ↑ **boiling point** ≈ ↓ **vapor pressure**

So, by what amount will a solution's ***vapor pressure*** decrease, when a ***solute*** is added? To calculate this, we use something called ***Raoult's law:***

$$P_A = X_A P_{pure}$$

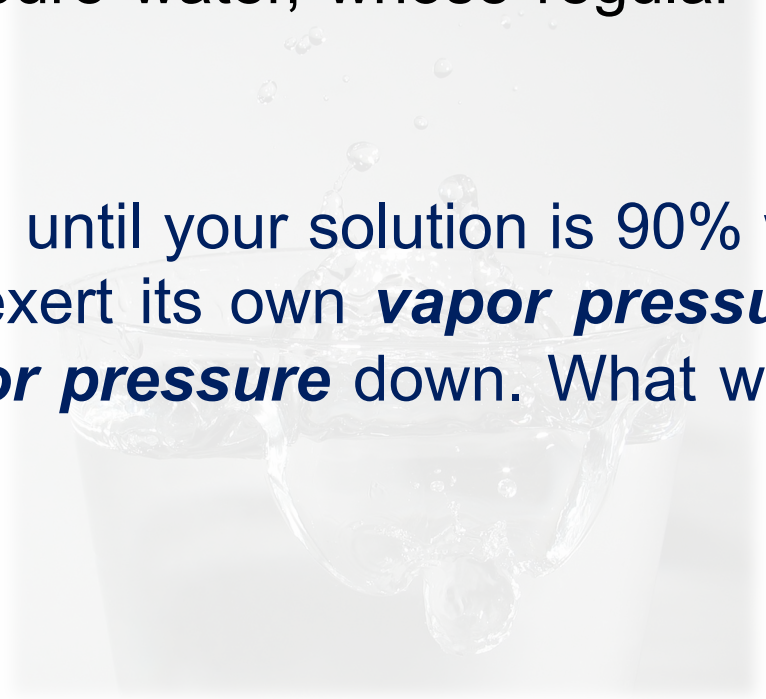
Where:

- P_{pure} is the vapor pressure of the original solvent, when it's pure.
- X_A is the percentage of solvent in the new solution.
- P_A is the solution's *new* vapor pressure.

Raoult's Law

For example, let's pretend that you have a glass of pure water, whose regular vapor pressure, P_{pure}, is 100 Torr.

Now, let's imagine that you dilute it by adding ethanol, until your solution is 90% water and 10% EtOH. When this happens, the EtOH will exert its own ***vapor pressure*** in competition with the water, which drives water's ***vapor pressure*** down. What will the new vapor pressure, P_A, of the water be?



To answer this, of course, we go to ***Raoult's Law!***

$$P_A = X_A P_{pure}$$

In this example, P_{pure} is 100 Torr. X_A, the percentage of solvent in the new solution, is 0.90. Thus, the new pressure, P_A, will equal (0.90) × (100 Torr) = 90 Torr.

Question

What is the vapor pressure of a solution that contains hexane dissolved in dichloromethane, where the mole fraction of the dichloromethane is 0.8? (The vapor pressure of pure hexane is 130 mmHg, while that of dichloromethane is 540 mmHg.)

A. 108 mmHg

B. 350 mmHg

C. 26 mmHg

D. 458 mmHg

E. 432 mmHg

Solutions
(Osmosis)

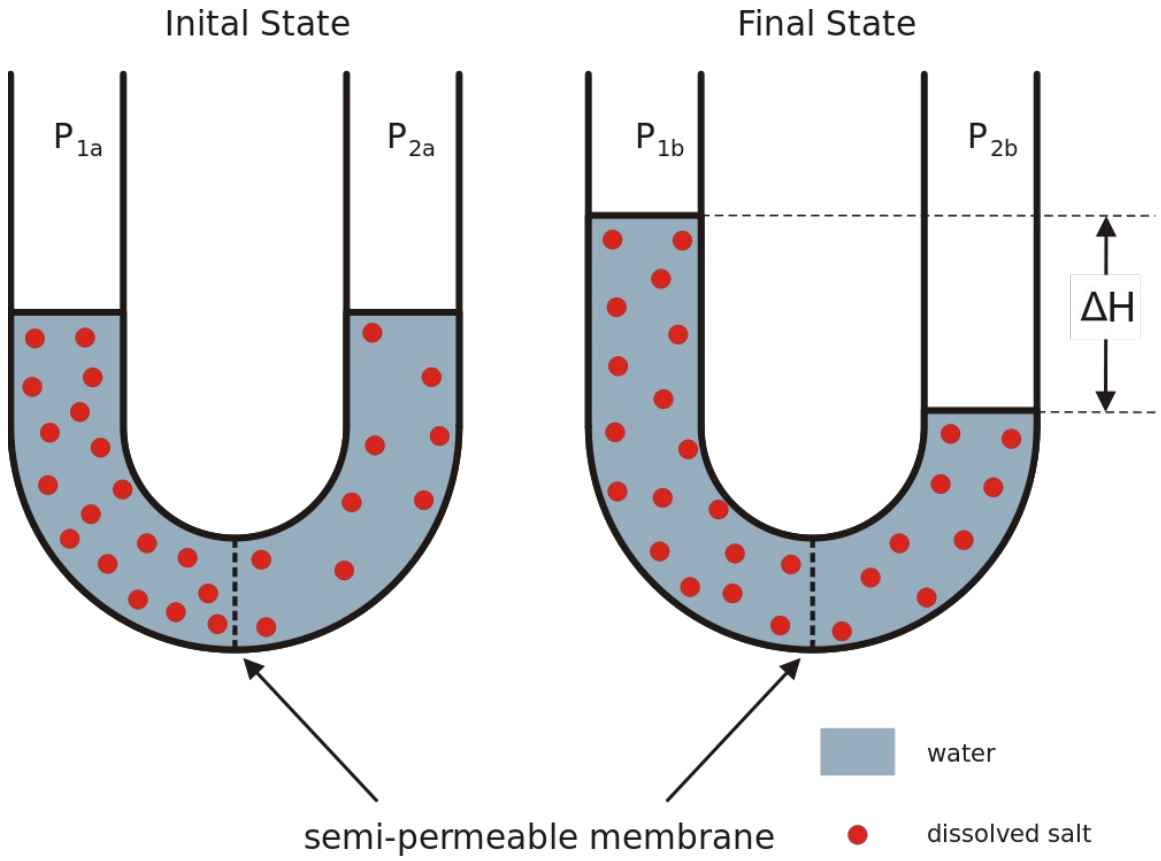
Osmosis

Osmosis is the flow of water from an area of high concentration *of water* (low concentration *of solute*) to an area of low concentration *of water* (high concentration *of solute*).

Osmosis

Osmosis is the flow of water from one area to another, in order to balance those areas' *solute* concentrations.

For example, if you take a container that has two chambers separated by a semi-permeable membrane, where one chamber contains more salt than the other, then water will spontaneously pass through the membrane until the concentrations of both sides become equal.



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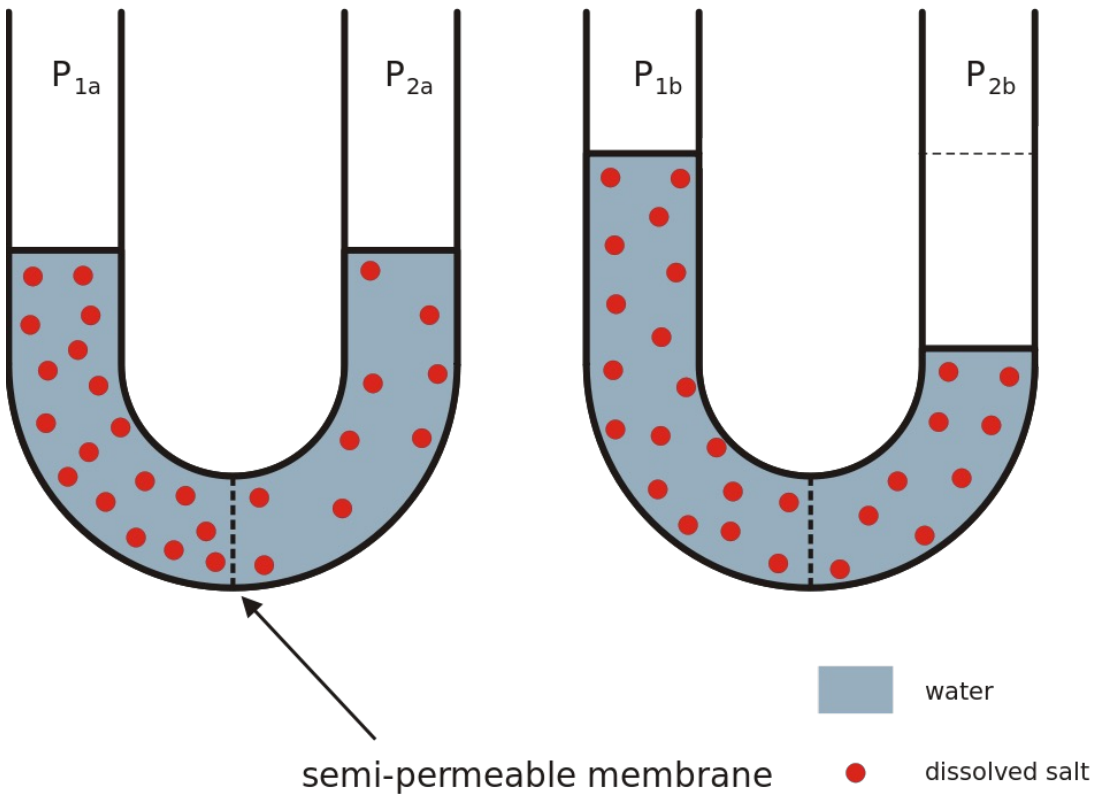

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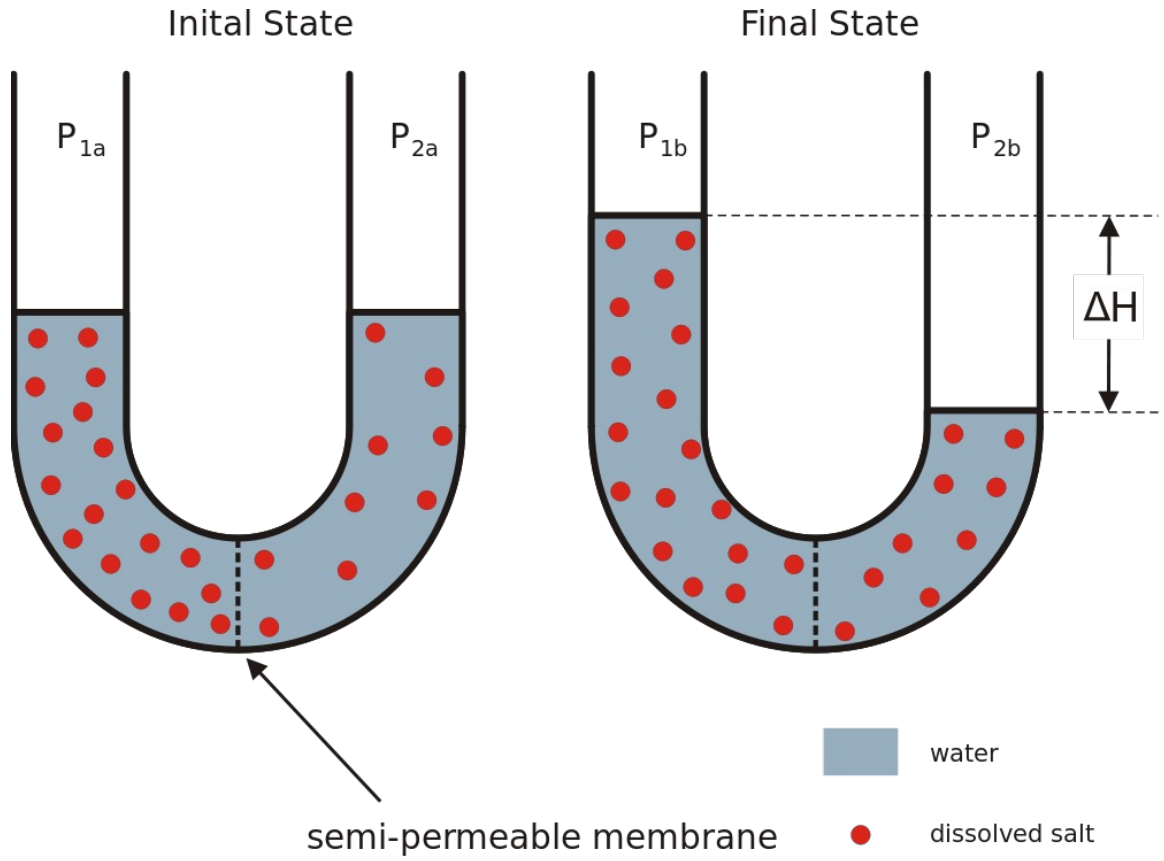


Osmosis

It's easy to mess this up if you don't understand what is causing the concentration to change on both sides.

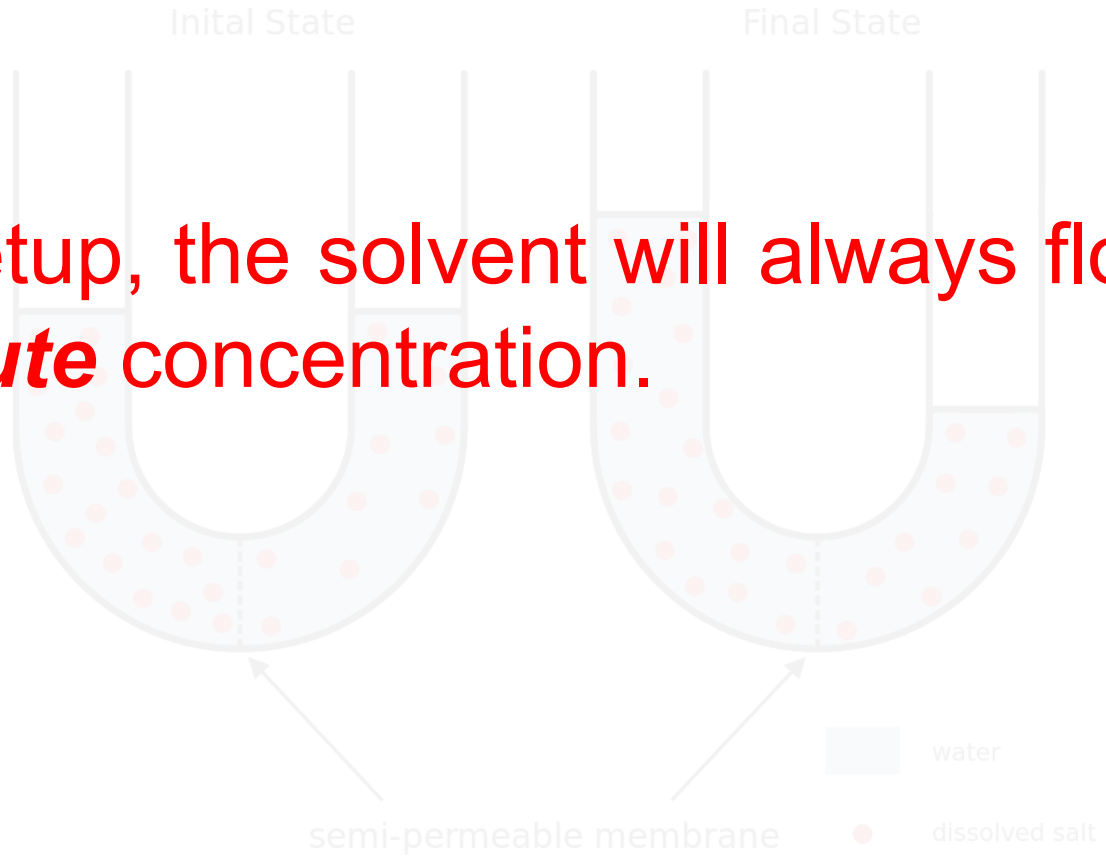
Note: only the *water* can pass through the membrane, so it is the flow of water we have to consider, not the flow of salt!

Initially, there is more water on the right side of the tube than the left side. So the water flows to the left to equalize the salt concentration. Thus, we are *not* looking at the concentration of the salts. We are looking at the concentration of the *water*. This spontaneous flow of water to equalize concentration is called ***osmosis***.

Osmosis

Osmosis is the flow of water from an area of high concentration to an area of low concentration.

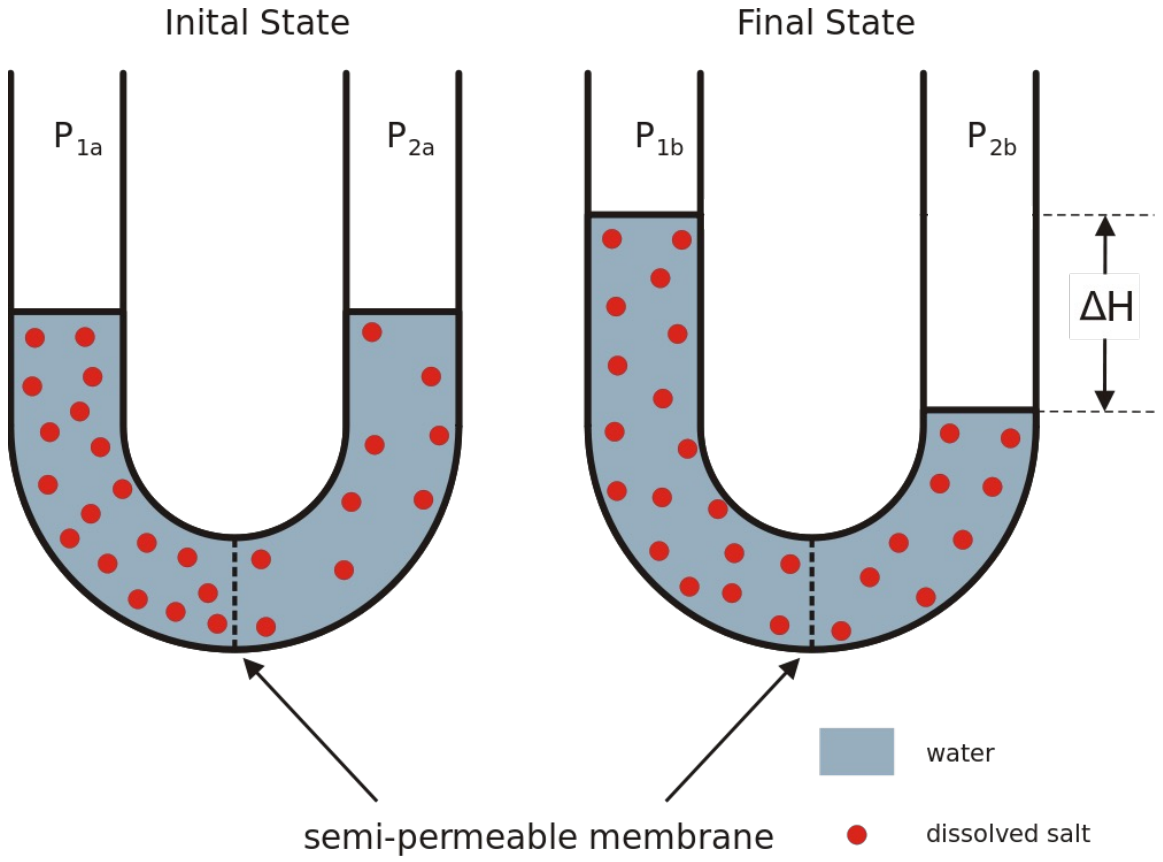
For example, if you take a container that has two chambers separated by a semi-permeable membrane, where one chamber contains more salt than the other, then water will spontaneously pass through the membrane until the concentration of salt is equal on both sides.



We can say that in this type of setup, the solvent will always flow toward the side with a higher ***solute*** concentration.

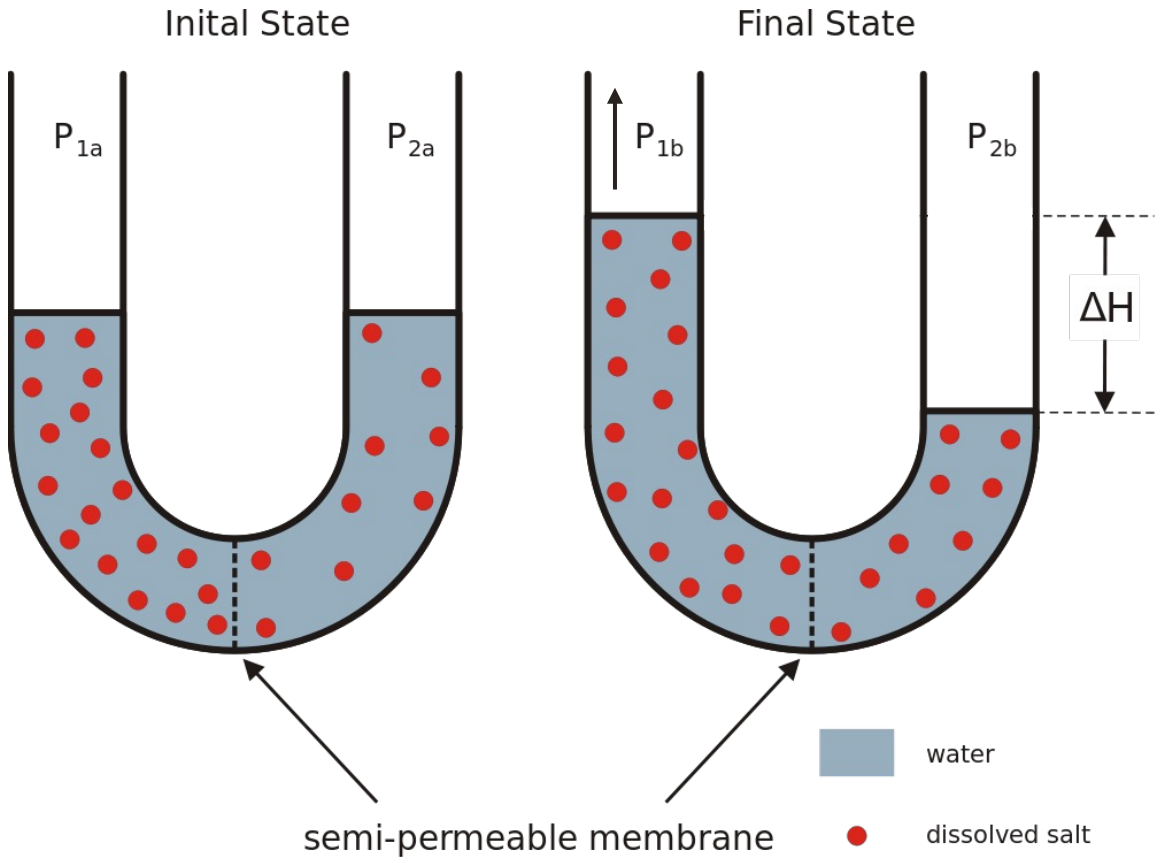
Osmotic Pressure

When the water shifts to equalize the concentrations of both chambers, it changes the pressure above the liquid on both sides.



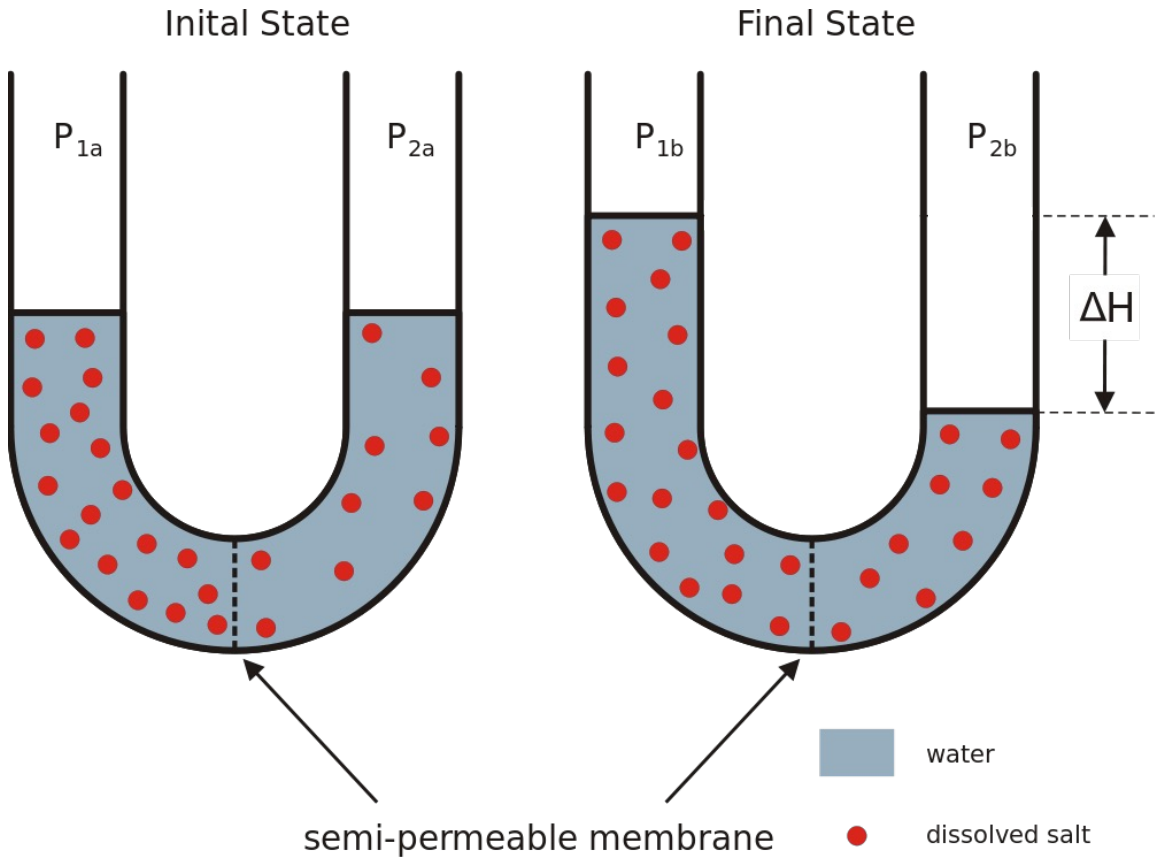
Osmotic Pressure

For example, let's assume that in its initial state, $\mathbf{P_{1a}} = \mathbf{P_{2a}}$. As ***osmosis*** occurs, the growing amount of water on the left side exerts an upward pressure, $\mathbf{P_{1b}}$.



Osmotic Pressure

Eventually, the water stops flowing. At this point, if we exert a pressure equal to $\mathbf{P_{1b}}$ down upon the left chamber, we can get the water levels to become equal again.



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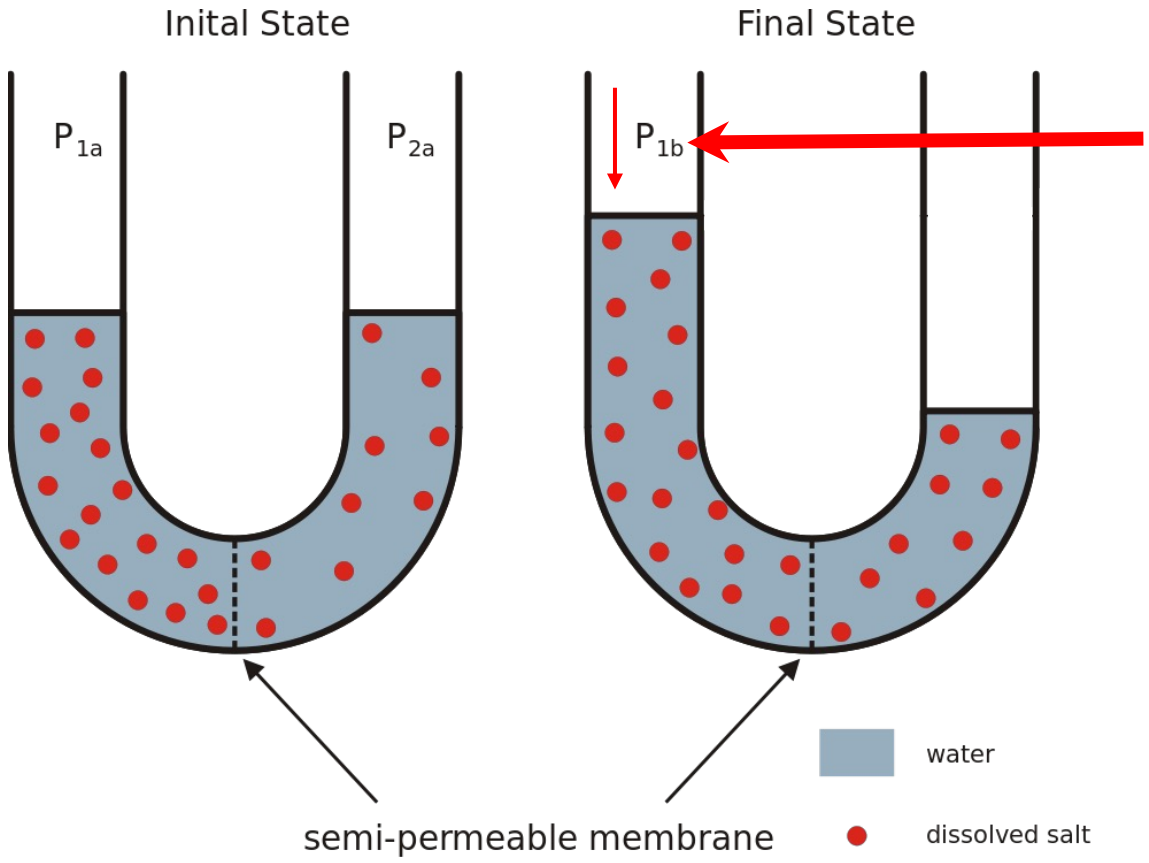

This downward pressure, which equals $\mathbf{P_{1b}}$ (just applied in reverse) is called ***osmotic pressure***.

Osmotic pressure, then, is the amount of pressure you have to apply to stop ***osmosis***.

Osmotic Pressure

We can perform ***osmotic pressure*** calculations by using an equation that is very similar to the ***ideal gas law:***

$$\Pi V = nRT$$

Where:

- Π is ***osmotic pressure***.
- V is the solution's ***volume***.
- n is the number of moles of ***solute*** (not ***solvent***).
- R is the ***ideal constant constant*** (0.0821 L-atm/mol-K), and T is the temp in K.

Osmotic Pressure

If you rearrange this equation a little, you can get the following:

$$\Pi = (n/V)\,RT$$

You might notice that ***(n/V)*** has units of moles/liter, which is the same thing as ***molarity***, or concentration.

Because ***n*** is the number of moles of dissolved ***solute***, ***n*** actually has to be multiplied by the number of ions your solute produces, if your solute is an ionic compound.

For example, **NaCl** dissolves to form two ions, Na^+ and Cl^-. Thus, if **NaCl** were my solute, I would have to multiply it by two to get ***n***, which is really the total number of dissolved ions or molecules in the solution.

Osmotic Pressure

We can modify the ***osmotic pressure*** equation, then, to be the following:

$$\Pi = i\,(n/V)\,RT$$

Where ***i*** is the number of dissolved solvent molecules or ions that your solute produces. You might recognize ***i*** as the ***Van't Hoff factor***, which we discussed in an earlier video.

Osmosis (EXAM Tip)

Osmotic pressure is a preferred colligative property to use for measuring high-molecular-weight molecules (1000 g/mole or higher).

The reason is because it's much easier to accurately measure a small change in ***osmotic pressure*** than to measure a tiny difference in freezing/boiling point.

So, if we use ***boiling point elevation*** or ***vapor pressure*** or ***freezing point depression*** to calculate MW, we'll get best results for molecules with molecular weights of 1000 g/mole or less.

Question

Arrange the following aqueous solutions in order of decreasing osmotic pressure.

I. 0.20 M KCl
II. 0.30 M urea
III. 0.50 M sucrose

A. **III > I > II**
B. **II > III > I**
C. **I > II > III**
D. **III > II > I**
E. **II > I > III**