

COLLIGATIVE PROPERTIES

Complete Teaching Guide

*Understanding how dissolved particles change the
physical properties of solutions*

CHM-152L

General Chemistry 2

Workshop 1 Preparation

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INTRODUCTION

What Are Colligative Properties?

Imagine you're making pasta. You add salt to the water, and suddenly, it takes longer to boil. Or think about why we put salt on icy roads in winter. These everyday phenomena are examples of **colligative properties** in action.

THE BIG IDEA: Colligative properties depend only on the *number* of dissolved particles, not on what those particles are. It doesn't matter if you dissolve sugar, salt, or antifreeze—what matters is *how many particles* you add to the solution.

The Four Colligative Properties

Property	What Happens	Real-World Example
Vapor Pressure Lowering	Solution has lower vapor pressure than pure solvent	Why ocean water evaporates more slowly than pure water
Boiling Point Elevation	Solution boils at a higher temperature	Salted pasta water takes longer to boil
Freezing Point Depression	Solution freezes at a lower temperature	Salt on icy roads; antifreeze in cars
Osmotic Pressure	Pressure needed to stop water flow across membrane	IV solutions must match blood osmotic pressure

Before we can work with these properties, we need to master **concentration units**—the language we use to describe how much solute is in a solution.

Concentration Units: The Language of Solutions

Think of concentration units like different ways to give directions. You could say "2 miles north" or "a 5-minute drive"—both describe the same destination, but one might be more useful depending on the situation. In chemistry, we have several ways to express concentration, and choosing the right one matters.

The Two Most Important Units: Molarity vs. Molality

	Molarity (M)	Molality (m)
Formula	$M = \text{mol solute} / \text{L of solution}$	$m = \text{mol solute} / \text{kg of solvent}$
Denominator	Volume of SOLUTION (solute + solvent)	Mass of SOLVENT only
Units	mol/L or M	mol/kg or m
Temperature	Changes with temperature (volume expands)	Doesn't change (mass is constant)
Best For	Lab work; measuring volumes	Colligative property calculations

THE #1 MISTAKE STUDENTS MAKE: Confusing *solution* (solute + solvent) with *solvent* (just the dissolving liquid). Molarity uses solution volume. Molality uses solvent mass. Get this wrong, and every calculation falls apart!

MEMORY HOOK

"M is for Measuring volume, m is for Mass of solvent"

Molarity has an "r" like "liter" (volume). Molality has an "a" like "mass."

Other Important Concentration Units

Mole Fraction (χ)

MOLE FRACTION

$$\chi_A = \text{moles of component A} / \text{total moles of all components}$$

Key properties:

- Dimensionless (no units)—it's just a ratio
- All mole fractions in a mixture must add up to exactly 1
- Used extensively in Raoult's Law for vapor pressure calculations

THINK OF IT LIKE A PIZZA PARTY: If you have 8 slices of pepperoni and 2 slices of cheese, the mole fraction of pepperoni is $8/10 = 0.80$. Mole fraction tells you what *fraction* of all the particles belong to one component.

Mass Percent (Mass %)

MASS PERCENT

$$\text{Mass \%} = (\text{mass of solute} / \text{mass of solution}) \times 100\%$$

Key properties:

- Expressed as a percentage
- Uses mass of total solution (not just solvent)
- Common on product labels (e.g., "70% isopropyl alcohol")

THE 100-GRAM TRICK: When a problem gives you mass percent, assume you have *100 grams of solution*. Then the mass percent directly equals grams of solute! A 5% solution? That's 5 g solute in 100 g solution. Easy.

Converting Between Concentration Units

Converting between M, m, χ , and mass % is like translating between languages. You need a common reference point and a systematic approach.

THE SECRET WEAPON—ASSUME A CONVENIENT AMOUNT:

- Given mass %? Assume **100 g of solution**
 - Given molarity? Assume **1 L of solution**
- Then calculate everything else from there!

The Conversion Roadmap

- 1 **Start with what you know.** Use the given concentration to establish your reference amount (100 g or 1 L).
- 2 **Calculate moles of solute.** Use molar mass: $\text{mol} = \text{grams} / \text{MM}$
- 3 **Find mass of solvent.** Remember: $\text{mass of solvent} = \text{mass of solution} - \text{mass of solute}$
- 4 **Use density as a bridge.** Density connects mass and volume ($d = m/V$), allowing conversion between volume-based and mass-based units.
- 5 **Calculate the target unit.** Plug your values into the appropriate formula.

CRITICAL REMINDER:

Mass of solvent = Mass of solution – Mass of solute

The solution includes everything. The solvent is just the liquid doing the dissolving. Always subtract the solute mass!

THINK ABOUT IT:

- If you have a very dilute aqueous solution (density $\approx 1.00 \text{ g/mL}$), why are molarity and molality approximately equal?
- What happens to molarity when you heat a solution? What about molality? Why?

PART 2

Solution Preparation & Dilutions

You've made lemonade before, right? When it's too strong, you add water. The amount of lemon flavor stays the same—you're just spreading it through more liquid. This is exactly how dilutions work in chemistry.

The Dilution Equation

THE DILUTION EQUATION

$$M_1V_1 = M_2V_2$$

WHY THIS WORKS: $M \times V$ = moles of solute. When you dilute, the moles of solute don't change—you're only adding solvent. So moles before = moles after, which means $M_1V_1 = M_2V_2$.

Understanding the Subscripts

Symbol	Meaning	Also Called
M_1	Initial (stock) concentration	The concentrated solution you start with
V_1	Volume of stock solution used	What you pipette out
M_2	Final (diluted) concentration	The concentration you want to achieve
V_2	Final total volume	Total volume after adding solvent

QUICK SANITY CHECK: When diluting, $M_2 < M_1$ (concentration goes DOWN) and $V_2 > V_1$ (volume goes UP). If your answer shows concentration increasing, you made an error!

Serial Dilutions in the Lab

When creating a calibration curve (like for Beer's Law), you need multiple standards at different concentrations—all prepared from the same stock solution.

The Standard Procedure

- 1 **Calculate V_1 for each standard.** Rearrange: $V_1 = (M_2 \times V_2) / M_1$
- 2 **Pipette V_1 into a volumetric flask.** Use an appropriately sized flask for your final volume.
- 3 **Dilute to the mark.** Add solvent until you reach V_2 (the calibration line on the flask).

THE DILUTION FACTOR: If you go from 0.50 M to 0.10 M, that's a 5× dilution ($0.50 \div 0.10 = 5$). Quick check: the stock volume should be 1/5 of the final volume. If $V_2 = 10$ mL, then $V_1 = 2$ mL. Does that match $M_1V_1 = M_2V_2$? Always verify!

When Concentrations Are in Different Units

Sometimes you need a solution in g/L but your stock is in M (mol/L). You must convert first!

CONVERTING g/L TO MOLARITY

$$M \text{ (mol/L)} = (\text{g/L}) \div (\text{g/mol})$$

Dimensional analysis proves it: $(\text{g/L}) \div (\text{g/mol}) = (\text{g/L}) \times (\text{mol/g}) = \text{mol/L} = M$

THINK ABOUT IT:

- If you calculate $V_1 = 0.05$ mL for a very dilute solution, is that practical to measure? What would you do instead?
- Why do we use volumetric flasks instead of beakers for accurate dilutions?

PART 3

Vapor Pressure & Raoult's Law

What is Vapor Pressure?

Even when water isn't boiling, some molecules at the surface have enough energy to escape into the air. This creates **vapor pressure**—the pressure exerted by the gaseous molecules above the liquid. At equilibrium, the rate of evaporation equals the rate of condensation.

Every pure liquid has a characteristic vapor pressure at a given temperature. For example, at 25°C:

Substance	Vapor Pressure at 25°C	Relative Volatility
Water	23.8 mmHg	Low (evaporates slowly)
Methanol	94 mmHg	High (evaporates quickly)
Glucose	≈ 0 mmHg	Non-volatile (doesn't evaporate)

THE KEY INSIGHT: When you dissolve a solute in water, the solute molecules take up space at the surface. Fewer water molecules can escape → lower vapor pressure. This is called **vapor pressure lowering**, and it's a colligative property.

THE NIGHTCLUB ANALOGY: Imagine water molecules trying to leave through a door (evaporate). Now add some big, non-moving solute molecules standing near the exit. They block some water molecules from leaving! Fewer molecules escape → lower vapor pressure.

Raoult's Law

$$\text{RAOULT'S LAW}$$
$$P_{\text{solution}} = \chi_{\text{solvent}} \times P^{\circ}_{\text{solvent}}$$

Symbol	Meaning	Units
P_{solution}	Vapor pressure of the solution	mmHg (or atm, torr)
χ_{solvent}	Mole fraction of the solvent	No units (dimensionless)
$P^{\circ}_{\text{solvent}}$	Vapor pressure of PURE solvent	mmHg (same as P_{solution})

CRITICAL POINT: Raoult's Law uses the mole fraction of the *solvent* (the component that evaporates), not the solute! Since $\chi_{\text{solvent}} < 1$ when solute is present, $P_{\text{solution}} < P^{\circ}_{\text{solvent}}$. The vapor pressure always drops.

Calculating Vapor Pressure: Step by Step

- 1 **Find moles of each component.** Convert grams to moles using molar mass ($\text{mol} = \text{g} / \text{MM}$).
- 2 **Calculate mole fraction of solvent.** $\chi_{\text{solvent}} = \text{mol solvent} / (\text{mol solvent} + \text{mol solute})$
- 3 **Apply Raoult's Law.** Multiply χ_{solvent} by $P^{\circ}_{\text{solvent}}$ to get P_{solution} .

Two Volatile Components

What if BOTH substances in your mixture can evaporate? Glucose dissolved in water involves only one volatile component (water). But methanol mixed with water? Both are volatile—both contribute to the total vapor pressure!

FOR TWO VOLATILE COMPONENTS: Apply Raoult's Law to EACH component separately, then add the partial pressures together (Dalton's Law).

$$\text{TOTAL VAPOR PRESSURE (TWO VOLATILE COMPONENTS)}$$

$$P_{\text{total}} = P_A + P_B = (\chi_A \times P^{\circ}_A) + (\chi_B \times P^{\circ}_B)$$

Non-Volatile Solute (e.g., glucose in water)	Two Volatile Components (e.g., methanol + water)
Formula: $P_{\text{total}} = \chi_{\text{water}} \times P^{\circ}_{\text{water}}$ Only the solvent contributes to vapor pressure. The solute has negligible vapor pressure.	Formula: $P_{\text{total}} = P_{\text{methanol}} + P_{\text{water}}$ Both components contribute. Calculate each partial pressure separately, then add them.

VOLATILITY MATTERS! A component with high P° (very volatile) will contribute more to the total vapor pressure even if its mole fraction is small. Methanol ($P^{\circ} = 94 \text{ mmHg}$) contributes way more per mole than water ($P^{\circ} = 23.8 \text{ mmHg}$).

THINK ABOUT IT:

- If you mix equal moles of methanol and water, which contributes more to the total vapor pressure? Why?
- How can you tell from a problem whether to use the one-volatile or two-volatile approach?

PART 4

Osmotic Pressure

What is Osmosis?

Place a cell in pure water, and water rushes in. Place it in concentrated salt water, and water rushes out. This movement of water across membranes is called **osmosis**, and understanding it is crucial for biology and medicine.

OSMOSIS DEFINED: Osmosis is the movement of solvent (water) across a semipermeable membrane from an area of LOW solute concentration to HIGH solute concentration. The membrane lets water through but blocks solute particles.

Osmotic pressure (π) is the pressure required to STOP osmosis—to prevent water from moving across the membrane. The higher the solute concentration, the higher the osmotic pressure.

THE CROWD ANALOGY: Imagine two rooms connected by a door that only lets small people through. One room is packed with big people (solute). Small people (water) from the less crowded room squeeze through to equalize the "crowding." Osmotic pressure is like a bouncer at the door, pushing back to stop this flow.

The Osmotic Pressure Equation

OSMOTIC PRESSURE

$$\pi = iMRT$$

Notice this equation is structured like the ideal gas law ($PV = nRT$)!

Symbol	Meaning	Units
π	Osmotic pressure	atm
i	Van't Hoff factor (number of particles)	No units
M	Molarity of solution	mol/L
R	Gas constant	0.08206 L·atm/(mol·K)
T	Temperature	Kelvin ($K = ^\circ C + 273$)

TEMPERATURE MUST BE IN KELVIN!
 $K = ^\circ C + 273$. Using Celsius will give you completely wrong answers. This is a non-negotiable step—always convert temperature first!

The Van't Hoff Factor (i)

This is where colligative properties get interesting—and where students make the most mistakes!

WHAT IS i? The van't Hoff factor tells you how many particles one formula unit creates when dissolved. Colligative properties depend on particle count, not formula units!

Determining the Van't Hoff Factor

Type of Compound	Example	Dissociation	i Value
Nonelectrolyte	Glucose, Urea, Sucrose	Stays intact—no dissociation	1
Strong electrolyte (1:1)	NaCl, KBr, HCl	$NaCl \rightarrow Na^+ + Cl^-$	2
Strong electrolyte (1:2)	$MgCl_2$, $CaCl_2$	$MgCl_2 \rightarrow Mg^{2+} + 2Cl^-$	3
Strong electrolyte (2:1)	Na_2SO_4 , K_2CO_3	$Na_2SO_4 \rightarrow 2Na^+ + SO_4^{2-}$	3
Strong electrolyte (1:3)	$FeCl_3$, $AlBr_3$	$FeCl_3 \rightarrow Fe^{3+} + 3Cl^-$	4

MEMORY HOOK
"i = ions in solution"

Count EVERY particle that forms when the compound dissolves.
 $MgCl_2 \rightarrow Mg^{2+} + Cl^- + Cl^- = 3$ particles, so $i = 3$

ALWAYS WRITE THE DISSOCIATION EQUATION:

$MgCl_2 \rightarrow Mg^{2+} + 2Cl^-$
Count: $1 + 2 = 3$ particles. Therefore $i = 3$.

This systematic approach prevents errors!

Why i Matters So Much

Consider two solutions, both at 1.0 M concentration:

Solution	i Value	Effective Particle Concentration
1.0 M glucose	1	1.0 M particles
1.0 M NaCl	2	2.0 M particles
1.0 M $MgCl_2$	3	3.0 M particles

The $MgCl_2$ solution has **3× the osmotic pressure** of the glucose solution because it has 3× the particles—even though both are "1.0 M"!

Clinical Connections: Why This Matters

Blood has an osmotic pressure of about **7.4 atm**. IV solutions must match this to avoid damaging cells:

Solution Type	Osmotic Pressure	Effect on Cells
Hypotonic	Lower than blood	Water enters cells → Cells BURST (hemolysis)
Isotonic	Equal to blood	No net water movement → Cells stay healthy
Hypertonic	Higher than blood	Water leaves cells → Cells SHRINK (crenation)

This is why normal saline (0.9% NaCl) and D5W (5% dextrose) are carefully formulated to be **isotonic** with blood plasma.

Step-by-Step: Calculating Osmotic Pressure

- 1 **Identify the solute and determine i .** Is it an electrolyte? Write the dissociation equation and count particles.
- 2 **Calculate moles of solute.** $n = \text{mass (g)} / \text{molar mass (g/mol)}$
- 3 **Calculate molarity.** $M = \text{moles} / \text{volume (L)}$. Make sure volume is in liters!
- 4 **Convert temperature to Kelvin.** $K = ^\circ\text{C} + 273$. Don't skip this!
- 5 **Plug into $\pi = iMRT$.** Use $R = 0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$. Units should cancel to give atm.
- 6 **Convert units if needed.** $1 \text{ atm} = 760 \text{ mmHg} = 760 \text{ torr}$

COMMON MISTAKES TO AVOID:

- **Forgetting i :** This is the #1 error! Never leave out the van't Hoff factor for electrolytes.
- **Using $^\circ\text{C}$:** Temperature must be in Kelvin for this equation.
- **Wrong i value:** Count ALL particles from dissociation ($\text{MgCl}_2 = 3$, not 2).
- **Using molality:** Osmotic pressure uses molarity (M), not molality (m).

THINK ABOUT IT:

- If osmotic pressure depends on particle concentration, why would 0.9% NaCl ($i=2$) be isotonic but 0.9% glucose ($i=1$) would not be?
- A solution has $\pi = 60 \text{ atm}$. Is this safe for IV use? How does it compare to blood ($\pi \approx 7.4 \text{ atm}$)?

SUMMARY

Master Formula Reference

All Key Formulas in One Place

Concept	Formula	Key Point
Molarity	$M = \text{mol} / \text{L solution}$	Uses total solution volume
Molality	$m = \text{mol} / \text{kg solvent}$	Uses solvent mass only
Mole Fraction	$\chi = \text{mol}_A / \text{total mol}$	All χ values must sum to 1
Mass Percent	$\text{Mass \%} = (\text{g solute} / \text{g solution}) \times 100\%$	Assume 100g for easy calculation
Dilution	$M_1V_1 = M_2V_2$	Moles are conserved
g/L to Molarity	$M = (\text{g/L}) / (\text{g/mol})$	Divide by molar mass
Raoult's Law	$P = \chi \times P^\circ$	Use χ of volatile component
Dalton's Law	$P_{\text{total}} = P_A + P_B$	Add partial pressures
Osmotic Pressure	$\pi = iMRT$	Don't forget i ! T in Kelvin!

Van't Hoff Factor Quick Reference

i Value	Type	Examples
1	Nonelectrolytes	Glucose, Urea, Sucrose, Ethanol
2	1:1 Electrolytes	NaCl, KBr, HCl, NaOH, KNO ₃
3	1:2 or 2:1 Electrolytes	MgCl ₂ , CaCl ₂ , Na ₂ SO ₄ , K ₂ CO ₃ , Ba(NO ₃) ₂
4	1:3 Electrolytes	FeCl ₃ , AlBr ₃ , Al(NO ₃) ₃

Constants You Need

Constant	Value	When to Use
R (gas constant)	0.08206 L·atm/(mol·K)	Osmotic pressure calculations
P° _{water} at 25°C	23.8 mmHg	Vapor pressure problems
P° _{methanol} at 25°C	94 mmHg	Two-volatile vapor pressure
Density of water	1.00 g/mL	Converting mL water to grams
1 atm	760 mmHg	Pressure unit conversions

Success Strategies for Your Workshop

- 1 **Read the entire problem first.** Identify what you're given and what you need to find. Write down the knowns and unknowns.
- 2 **Choose your starting point wisely.** Mass %? Assume 100g. Molarity? Assume 1L. This makes calculations cleaner.
- 3 **Show ALL unit conversions.** Partial credit comes from clear methodology. Units tell a story and help you catch errors!
- 4 **Box your final answers.** Make it easy for graders to find your answers. Always include units!
- 5 **Sanity check your answers.** Does a vapor pressure > P° make sense? (No!) Does M = m for dilute aqueous solutions? (Yes!)

FINAL REMINDER

Colligative properties count PARTICLES, not molecules!

Whether it's vapor pressure, freezing point, boiling point, or osmotic pressure—it's all about how many particles are in solution. Electrolytes multiply your particle count!