

Show all work with units and correct significant figures. In explanations, name the specific force for **both** substances, not just "stronger" or "weaker."

Constants	$R = 0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$	$R = 62.36 \text{ L}\cdot\text{torr}/(\text{mol}\cdot\text{K})$	$1 \text{ atm} = 760 \text{ torr} = 760 \text{ mmHg}$	STP: 273 K and 1.00 atm 1 mol gas = 22.4 L
	$h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$	$c = 3.00 \times 10^8 \text{ m/s}$	$N_A = 6.022 \times 10^{23} \text{ mol}^{-1}$	$K = ^\circ\text{C} + 273$
Equations	$PV = nRT$	$P_A = P_{\text{total}} \times X_A$	$P_{\text{total}} = P_A + P_B + \dots$	$KE = \frac{1}{2}mv^2$
	$M = \text{mol solute} \div \text{L solution}$	$M_1V_1 = M_2V_2$	$c = \lambda\nu \quad E = h\nu$	$A = \epsilon bc$

Part 1. Intermolecular Forces (Topic 3.1)

1. Complete the table. IMF choices: London dispersion forces (LDF), dipole-dipole, hydrogen bonding.

Substance	Polar or nonpolar?	All IMFs present between its molecules	H-bonds with itself? (Y/N)
Kr	<i>Nonpolar (single atom)</i>	<i>LDF only</i>	<i>N</i>
H ₂ S	<i>Polar (bent)</i>	<i>LDF, dipole-dipole</i>	<i>N (H is bonded to S, not N, O, or F)</i>
CH ₃ NH ₂	<i>Polar</i>	<i>LDF, dipole-dipole, H-bonding</i>	<i>Y (N–H bond)</i>
CCl ₄	<i>Nonpolar (symmetric tetrahedral)</i>	<i>LDF only</i>	<i>N</i>
H ₂ CO (formaldehyde)	<i>Polar (C=O)</i>	<i>LDF, dipole-dipole</i>	<i>N (H is bonded to C; but its O can accept an H-bond from water)</i>

2. Account for each difference using the types and strengths of IMFs. Discuss **both** substances.

(a) Propane, C₃H₈ (44 g/mol), boils at 231 K. Acetaldehyde, CH₃CHO (44 g/mol), boils at 294 K.

Both molecules have about the same number of electrons, so their LDFs are similar in strength. Acetaldehyde is polar (polar C=O bond), so it has dipole-dipole forces *in addition to* LDF. Propane is nonpolar and has only LDF. More energy is needed to overcome the IMFs in acetaldehyde, so its boiling point is higher.

(b) Nonpolar CCl₄ (bp 350 K) boils at a higher temperature than polar CH₂Cl₂ (bp 313 K).

CCl₄ has 74 electrons and CH₂Cl₂ has 42. The larger electron cloud of CCl₄ is more polarizable, so its LDFs are much stronger. CH₂Cl₂ has dipole-dipole forces plus weaker LDFs. Here the greater LDF in CCl₄ outweighs the dipole-dipole forces in CH₂Cl₂. Polar does not automatically mean a higher boiling point.

3. A student claims, "When liquid bromine, Br₂, boils, the Br–Br bonds break and Br atoms form." Do you agree or disagree? Justify your answer.

Disagree. Boiling is a physical change. Only the LDFs *between* Br₂ molecules (intermolecular) are overcome. The covalent Br–Br bond *within* each molecule (intramolecular) stays intact, so the gas is still made of Br₂ molecules. Breaking covalent bonds takes far more energy and is a chemical change.

Part 2. Properties of Solids (Topic 3.2)

4. Classify each solid and describe what holds it together.

Solid	Type of solid	Particles that make up the solid	Forces holding particles together
MgO	<i>Ionic</i>	<i>Mg²⁺ and O²⁻ ions</i>	<i>Ionic bonds (Coulombic attraction between ions)</i>
SiC	<i>Covalent network</i>	<i>Si and C atoms</i>	<i>Covalent bonds in a 3-D network</i>
Cu	<i>Metallic</i>	<i>Cu cations in a sea of mobile valence electrons</i>	<i>Metallic bonding</i>
C ₁₂ H ₂₂ O ₁₁ (sucrose)	<i>Molecular</i>	<i>Individual sucrose molecules</i>	<i>IMFs: LDF, dipole-dipole, H-bonding</i>

5. Rank KBr, NaF, and MgO from **lowest** to **highest** melting point. Justify using Coulomb's law.

KBr < NaF < MgO. Coulombic attraction increases with larger ion charges and smaller distance between ions. MgO has 2+ and 2- ions, so its attraction is strongest. NaF and KBr both have 1+ and 1- ions, but Na⁺ and F⁻ are smaller than K⁺ and Br⁻, so the ions are closer together and the attraction in NaF is stronger.

6. Explain each observation at the particle level.

(a) An ionic crystal shatters when struck, but a piece of copper bends.

Ionic: a shift brings like-charged ions next to each other, the repulsion splits the crystal (brittle). Metal: cations sit in a sea of mobile electrons, so layers slide without repulsion (malleable).

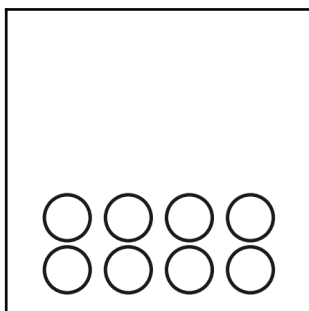
(b) Graphite is soft, but diamond is very hard. Steel (C atoms in Fe) is less malleable than pure Fe.

Diamond: rigid 3-D covalent network with fixed bond angles. Graphite: 2-D layers that slide past each other easily. Steel is an interstitial alloy: small C atoms fill gaps in the Fe lattice, making it more rigid.

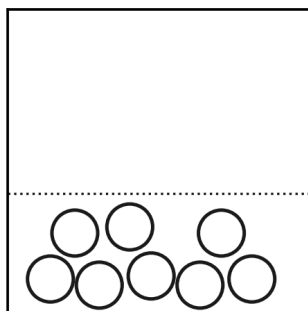
Part 3. Solids, Liquids, and Gases and Properties of Liquids (Topics 3.2 and 3.3)

7. Draw about 8 molecules (as small circles) in each box to show the same substance as a solid, liquid, and gas. Then explain why the solid and liquid have similar molar volumes but the gas does not.

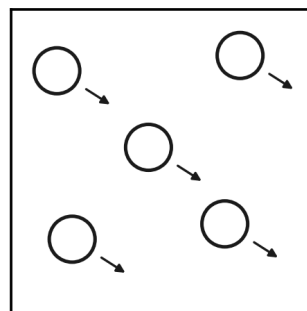
Solid



Liquid



Gas



Solid: particles touching in a regular, ordered pattern, vibrating in place. Liquid: particles still touching but disordered and able to slide past each other. Gas: particles far apart in random, constant motion and filling the container. In both solids and liquids the particles are in close contact, so the molar volumes are similar. Gas particles are separated by mostly empty space, so a mole of gas takes up far more volume.

8. At 25 °C, the vapor pressure of diethyl ether is 530 torr and the vapor pressure of water is 24 torr. Circle the liquid that fits each description, then answer the question below the table.

Property	Circle one	Property	Circle one
Stronger IMFs	Water	Higher viscosity	Water
Higher boiling point	Water	Larger heat of vaporization	Water
Evaporates faster	Ether	Feels colder on skin	Ether

Explain why the liquid you chose for "feels colder on skin" cools the skin more.

Evaporation is endothermic: molecules need energy to overcome IMFs and escape, and they take that energy from the skin. Ether has weaker IMFs (only LDF and weak dipole-dipole) than water (H-bonding), so it has a higher vapor pressure, evaporates faster, and removes heat from the skin faster. Water also has higher viscosity because H-bonding makes molecules resist flowing past one another.

9. In a narrow glass tube, water rises and forms a concave (U-shaped) meniscus. Mercury drops and forms a convex meniscus. Explain using cohesive and adhesive forces.

Water: adhesion (H-bonding and dipole attractions to the polar O atoms on the glass surface) is strong, and cohesion (H-bonds between water molecules) pulls more water up behind it, so water climbs. Mercury: cohesion (metallic bonding) is much stronger than its weak adhesion to glass, so mercury pulls in on itself.

10. A sealed flask holds a liquid and its vapor at constant temperature. Write T or F. Correct any false statement.

(a) Once the pressure is constant, evaporation and condensation stop.

F. It is a dynamic equilibrium: evaporation and condensation continue at equal rates.

(b) Raising the temperature increases the vapor pressure.

T. More molecules have enough KE to overcome IMFs and escape.

(c) Doubling the flask volume (liquid still present) permanently lowers the vapor pressure.

F. More liquid evaporates until the vapor pressure returns to the same value. Vapor pressure depends only on temperature and IMFs.

Part 4. Gases: Ideal Gas Law, Kinetic Molecular Theory, Real Gases (Topics 3.4 to 3.6)

11. Zinc reacts with acid: $\text{Zn(s)} + 2 \text{HCl(aq)} \rightarrow \text{ZnCl}_2\text{(aq)} + \text{H}_2\text{(g)}$. A student collects 0.325 L of H_2 over water at 25 °C and a total pressure of 748 torr. The vapor pressure of water at 25 °C is 23.8 torr.

(a) Calculate the partial pressure of the dry H_2 .

$P_{\text{H}_2} = P_{\text{total}} - P_{\text{H}_2\text{O}} = 748 - 23.8 = \mathbf{724 \text{ torr}}$. The gas in the bottle is a mixture of H_2 and water vapor (Dalton's law).

(c) Calculate the mass of Zn that reacted.

1 mol Zn : 1 mol H_2 , so $0.0127 \text{ mol Zn} \times 65.38 \text{ g/mol} = \mathbf{0.828 \text{ g Zn}}$.

(b) Calculate the moles of H_2 collected.

$n = PV/RT = (724.2 \text{ torr})(0.325 \text{ L}) / [(62.36 \text{ L}\cdot\text{torr/mol}\cdot\text{K})(298 \text{ K})] = \mathbf{0.0127 \text{ mol H}_2}$. Use the R value whose units match torr.

(d) Calculate the volume of the dry H_2 at STP.

$V = 0.0127 \text{ mol} \times 22.4 \text{ L/mol} = \mathbf{0.284 \text{ L}}$ (same result from $V = nRT/P$ with 273 K and 1.00 atm).

12. A rigid container holds 2.0 mol N_2 , 3.0 mol O_2 , and 1.0 mol He at a total pressure of 3.0 atm.

(a) Calculate the partial pressure of O_2 .

$X_{\text{O}_2} = 3.0/6.0 = 0.50$. $P_{\text{O}_2} = 0.50 \times 3.0 \text{ atm} = \mathbf{1.5 \text{ atm}}$.

(b) 1.0 mol Ar is added at constant T. Does P_{He} increase, decrease, or stay the same? Explain.

Stays the same (0.50 atm). Each gas exerts its partial pressure independently; n_{He} , V, and T did not change. Only the total pressure (to 3.5 atm) and X_{He} change.

13. For He, N_2 , CO_2 , and SO_2 , all at the same temperature and pressure, write the gas that fits.

Description	Gas
Highest average kinetic energy	All equal (same T)
Highest average speed	He (lowest mass)
Slowest rate of effusion	SO_2 (highest mass)
Weakest IMFs	He (fewest e^-)
Greatest deviation from ideal behavior	SO_2 (polar, most e^-)

14. Real gases deviate from ideal behavior. Explain each observation at the particle level.

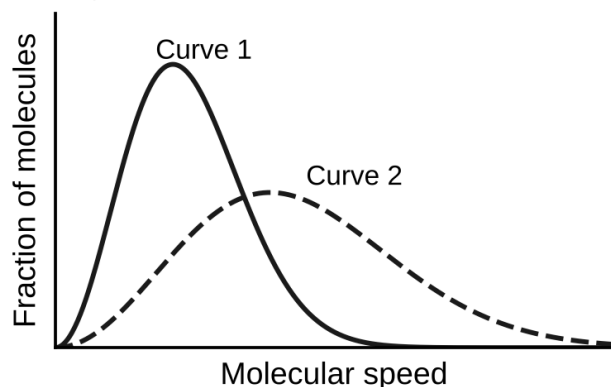
(a) At very high pressure, the measured volume is *greater* than the ideal volume.

The gas particles themselves take up space. At high pressure, particle volume is no longer negligible compared to the container, so the actual volume is larger than predicted.

(b) At low temperature (near condensation), the measured pressure is *less* than ideal.

Particles move slowly, so IMFs between them matter. Attractions pull particles toward each other, so they hit the walls less often and with less force. Gases behave most ideally at high T and low P.

15. The graph shows the speed distribution of N_2 molecules at two temperatures.



- (a) Which curve is at the higher temperature? Justify.

Curve 2. Higher T means higher average KE, so the distribution shifts to higher speeds and spreads out (flatter, broader peak).

- (b) Why is the area under both curves the same?

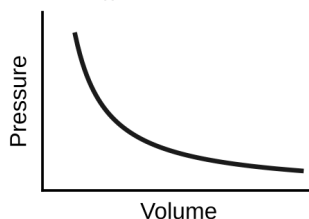
Both represent the same total number of molecules.

- (c) If the curves instead showed He and Ar at the same T , which curve is He? Why?

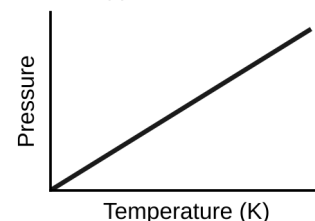
Curve 2. Same T gives the same average $KE = \frac{1}{2}mv^2$, so the lighter He must move faster on average.

16. Sketch each relationship for an ideal gas.

(i) constant n , T



(ii) constant n , V



(i) Inverse curve: P is proportional to $1/V$ (doubling V halves P). (ii) Straight line through the origin: P is directly proportional to T in kelvins. Particles hit the walls harder and more often at higher T .

Part 5. Solutions and Representations of Solutions (Topics 3.7 and 3.8)

17. 2.84 g of Na_2SO_4 (142.04 g/mol) is dissolved in water to make 250.0 mL of solution.

- (a) Calculate the molarity of Na_2SO_4 .

$2.84 \text{ g} \div 142.04 \text{ g/mol} = 0.0200 \text{ mol}$. $0.0200 \text{ mol} \div 0.2500 \text{ L} = \mathbf{0.0800 \text{ M}}$.

- (b) Calculate $[Na^+]$ and the number of Na^+ ions.

$Na_2SO_4 \rightarrow 2 Na^+ + SO_4^{2-}$, so $[Na^+] = 2 \times 0.0800 = \mathbf{0.160 \text{ M}}$. Ions: $0.0400 \text{ mol} \times 6.022 \times 10^{23} = \mathbf{2.41 \times 10^{22} Na^+}$.

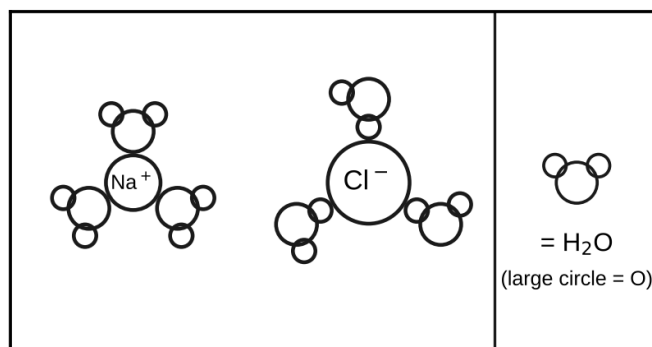
18. Describe how to prepare 500.0 mL of 0.150 M $CuSO_4$ (159.61 g/mol) from the solid. Why is "dissolve the solid in 500.0 mL of water" incorrect?

$0.5000 \text{ L} \times 0.150 \text{ mol/L} = 0.0750 \text{ mol} \times 159.61 \text{ g/mol} = \mathbf{12.0 \text{ g}}$. Dissolve 12.0 g in some water in a 500-mL volumetric flask, then add water up to the mark. Adding 500.0 mL of water makes the *total* volume more than 500.0 mL (the solid adds volume), so the concentration would be too low.

19. How many mL of 12.0 M HCl are needed to make 250. mL of 1.50 M HCl? How much water is added (assume volumes are additive)?

$V_1 = (1.50 \text{ M})(250. \text{ mL}) / 12.0 \text{ M} = \mathbf{31.3 \text{ mL}}$ of 12.0 M HCl.
Water added = $250. - 31.3 = \mathbf{219 \text{ mL}}$.

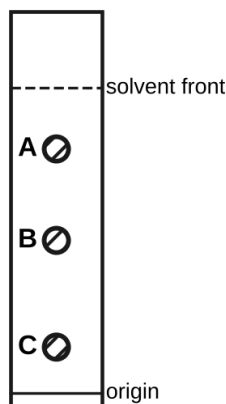
20. Draw 3 water molecules around each ion to show how water orients around the ions in $NaCl(aq)$.



Around Na^+ , the partially negative O ends point toward the cation. Around Cl^- , a partially positive H end points toward the anion. This is an ion-dipole attraction. Ions are separated (not paired), and Cl^- is drawn larger than Na^+ .

Part 6. Separation of Solutions and Mixtures (Topic 3.9)

21. A mixture is separated by thin-layer chromatography on silica (polar stationary phase) using 20% ethyl acetate in hexane (mostly nonpolar mobile phase). Distances from the origin: front 8.0 cm, A 6.4 cm, B 4.0 cm, C 1.2 cm.



(a) Which component is most polar? Explain.

C. It is attracted most strongly to the polar silica, so it moves least with the nonpolar solvent.

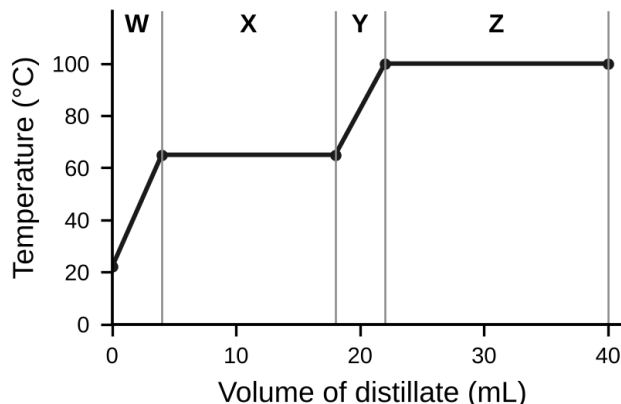
(b) Calculate R_f for B.

$$R_f = 4.0 \div 8.0 = 0.50$$

(c) To make all spots travel *farther*, should the percentage of ethyl acetate be increased or decreased? Explain.

Increased. Ethyl acetate is more polar than hexane, so a more polar mobile phase competes better with the silica for the components and carries them farther. (Decreasing it would make the spots travel less.)

22. A mixture of methanol, CH_3OH , and water is separated by fractional distillation.



(a) Which substance is collected in range X? Explain using IMFs.

Methanol (plateau at 65 °C). Both H-bond, but water has two O–H bonds and two lone pairs per molecule and forms more H-bonds, so its IMFs are stronger and it boils at the higher plateau (range Z, 100 °C).

(b) Why can't this mixture be separated by filtration?

It is a solution. The molecules are mixed at the particle level and pass right through filter paper. Distillation works because the components have different IMF strengths (different boiling points).

Part 7. Solubility (Topic 3.10)

23. Circle the solvent each substance dissolves in better and name the key solute-solvent attraction.

Solute	Water or hexane?	Key attraction
KBr	<u>Water</u>	<i>Ion-dipole</i>
C_8H_{18}	<u>Hexane</u>	<i>LDF</i>
I_2	<u>Hexane</u>	<i>LDF</i>
CH_3COOH	<u>Water</u>	<i>H-bonding</i>

24. Acetone, CH_3COCH_3 , cannot hydrogen bond with itself, yet it mixes with water in all proportions. Explain.

Acetone's O atom has lone pairs and a partial negative charge, so it can *accept* H-bonds from water's H atoms. Acetone and water also have dipole-dipole attractions. Similar IMFs means they are miscible.

25. Nonpolar N_2 dissolves slightly in water. What attraction allows this?

Dipole-induced dipole: polar water distorts N_2 's electron cloud, creating a temporary dipole (plus LDF). This attraction is weak, so solubility is low.

Part 8. Spectroscopy, Photons, and the Beer-Lambert Law (Topics 3.11 to 3.13)

26. Complete the table.

Radiation absorbed	Type of transition
Microwave	<i>Molecular rotation (rotational levels)</i>
Infrared (IR)	<i>Molecular vibration (bond stretching/bending)</i>
Ultraviolet/visible	<i>Electronic (electrons change energy levels)</i>

IR reveals the types and strengths of bonds because each bond vibrates at characteristic energies. When a photon is absorbed, the molecule's energy increases by exactly the photon's energy.

27. Blue light has a wavelength of 450. nm.

(a) Calculate its frequency and the energy of one photon.

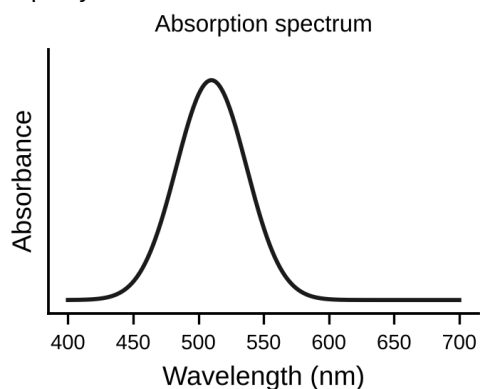
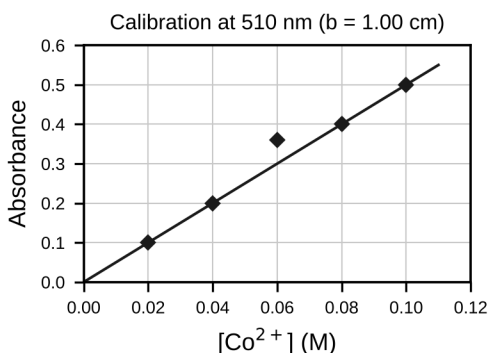
$$\nu = c/\lambda = (3.00 \times 10^8 \text{ m/s}) \div (450. \times 10^{-9} \text{ m}) = 6.67 \times 10^{14} \text{ s}^{-1}. E = h\nu = (6.626 \times 10^{-34}) (6.67 \times 10^{14}) = 4.42 \times 10^{-19} \text{ J}.$$

(b) *No calculator:* estimate E for a photon with $\nu = 2 \times 10^{15} \text{ s}^{-1}$.

$$E \approx (6.6 \times 10^{-34}) (2 \times 10^{15}) \approx 1 \times 10^{-18} \text{ J (about } 1.3 \times 10^{-18} \text{ J).}$$

Check the power of ten: $-34 + 15 = -19$, and $6.6 \times 2 \approx 13$ adds one more.

28. A student uses a spectrophotometer to find the mass percent of $\text{Co}(\text{NO}_3)_2$ (182.94 g/mol) in an impure 1.50 g sample. The impurity is colorless.



- (a) Which wavelength should be used? Why?

510 nm, the wavelength of maximum absorbance. It gives the greatest sensitivity (largest change in A per change in c).

- (b) Determine ϵ from the line ($b = 1.00 \text{ cm}$).

Slope = $\Delta A / \Delta c = 0.500 \div 0.100 \text{ M} = 5.0 \text{ M}^{-1}$. $\epsilon = \text{slope}/b = 5.0 \text{ M}^{-1} \text{ cm}^{-1}$.

- (c) The 0.060 M point is above the line. Give one likely cause.

The cuvette had fingerprints or smudges (was not wiped clean) or was scratched, so extra light was blocked and the absorbance read too high.

- (d) The sample is dissolved to make 100.0 mL of solution, and its absorbance is 0.320. Calculate the mass percent of $\text{Co}(\text{NO}_3)_2$.

$c = A/\epsilon b = 0.320 \div 5.0 = 0.0640 \text{ M}$. $\text{mol} = 0.0640 \text{ M} \times 0.1000 \text{ L} = 0.00640 \text{ mol}$.
 $\text{mass} = 0.00640 \times 182.94 = 1.17 \text{ g}$. $\text{Mass \%} = 1.17 \div 1.50 \times 100 = 78.1\%$.

- (e) Water droplets were left in the cuvette before adding the sample. Is the calculated mass percent too high or too low? Explain.

Too low. The droplets dilute the sample, so the measured absorbance and calculated concentration are lower than the true values. (Always rinse the cuvette with the solution and zero the instrument with a blank of pure solvent.)