Ch. 11 Liquids, Solids, & Materials
-Binding Forces & Properties

A. Intro: Phases and Attractive/Binding Forces

① T-84 Brown
② T75 Gillespie

1. 3 Phases

disorder

a) Crystalline Solid: complete order
- strong binding forces, no "molecular flow"

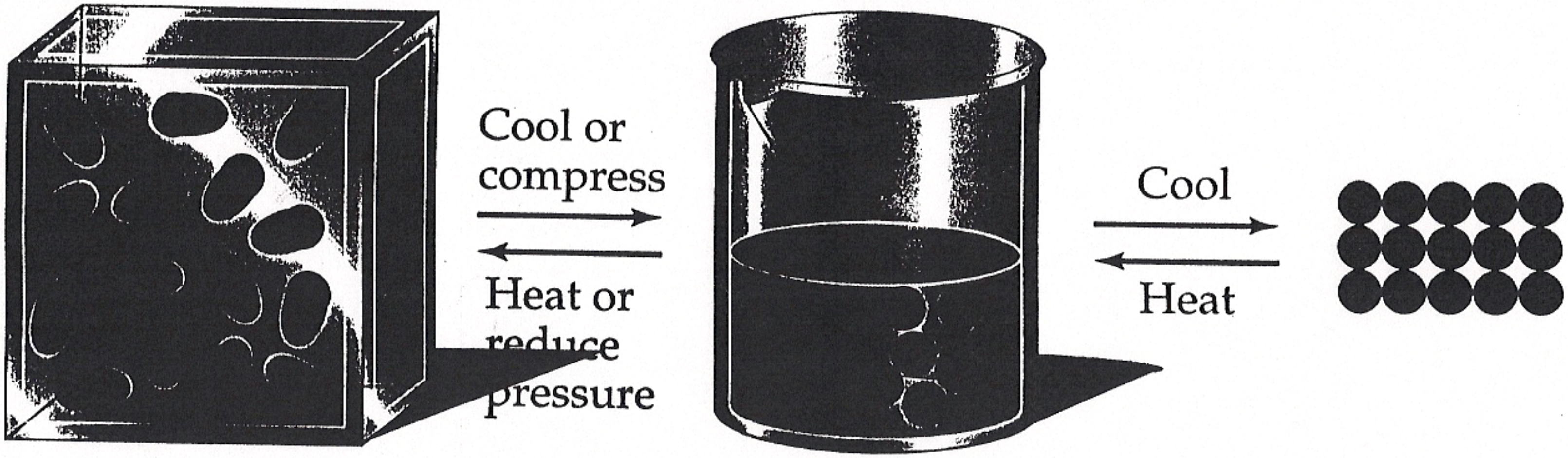
b) Liquid: limited disorder
- held together, but not fixed;
- fluid, free "molecular flow"

c) Gas: complete disorder
- far apart, no binding between molecules
- gases can be compressed; solids/liquids are "condensed" states, can't be compressed (much)

2. The phase of a substance depends on balance between binding force (holds together) vs. kinetic energy (fly apart)

- ~~energy~~

Fig. 11.1 Molecular Comparison of Gases, Liquids, and Solids



Gas

Total disorder; much empty space; particles have complete freedom of motion; particles far apart.

Liquid

Disorder; particles or clusters of particles are free to move relative to each other; particles close together.

Crystalline solid

Ordered arrangement; particles are essentially in fixed positions; particles close together.

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① Temp/heat

- raising temp increases KE ⇒ increased disorder

solid $\xrightarrow{heat}$ liquid $\xrightarrow{heat}$ gas

② At ~~[illegible]~~ given temp, strength of binding force determines phase

solid liquid gas

strongest binding ←———————— weakest binding

⇒ identifying nature and strength of attractive forces is key to predicting (mp, bp, etc...)

B. (9.5) Types of Binding Forces

1. Covalent Bonds
2. Ionic Bonds
3. Metallic Bonds

} Full Bonds, Strong!

4. Noncovalent Interactions (weak!)
   - do not involve full bonds (see Brown T2 Fig 1.5)
   - <10% strength of full
   - easy to overcome ⇒ low mp, bp ...

IMF { "Intermolecular Force" between molecules

{ "Intramolecular Force" within parts of molecule

↳ key to biology

(H bonding/ hydrophobic/ hydrophilic ...)

Chapter 11 + 9: Polarity, Binding Forces, Noncovalent Interactions, and Dependent Physical Properties

Types of Binding Forces

1) Covalent bonds (between nonmetals)
2) Ionic bonds (between metal and nonmetal)
3) Metallic bonds (within pure metals)
4) Noncovalent Interactions (in decreasing strength)
    - Hydrogen Bonds (O-H, N-H, or F-H bonds)
    - Dipole-dipole Attractions (for polar molecules, but lacking any O-H, N-H, or F-H bonds)
    - London Forces (only binding force between nonpolar molecules, but also applies in polar and H-bonding molecules. Increases with increasing molecular weight.)

Network vs Molecular Substances

- Network (usually identifiable by presence of a metal!)
    - Ionic
    - Metallic
    - Network Covalent (diamond...)
- Molecular (usually identifiable by absence of a metal)

Network binding forces are strong, molecular binding forces are relatively weak

Recognizing Polarity

1. AB_m	Nonpolar	If no lone pairs on central, and all outside atoms same.
2. AB_mL_N	Polar.	If lone pairs on central atom.
3. AB_mL_n	Weakly polar	No lone pairs on central atom, but attached atoms not same.
4. Hydrocarbons	Nonpolar	
5. Halocarbons	Very weakly polar	

Ranking substances in terms of relative binding forces:

1. Identify network substances versus molecular substances
2. For molecular substances
    a. H-bonding?
    b. polar?
3. For molecular substances, what is the molecular weight? Greater molecular weight gives greater binding force.

Note: when molecular weight effects counteract H-bonding or polar effects (such as methanol versus acetone versus hexane!) you are not expected to be able to predict which has stronger binding forces, except for special cases like H_2O.

Predictable Properties that Depend on Binding Forces

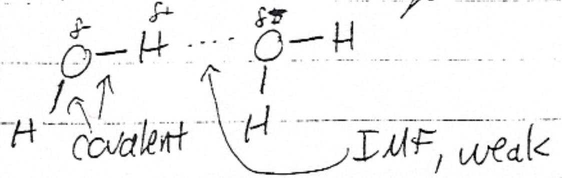
1. mp	Higher binding force –> higher mp
2. ΔH_f	Higher binding force –> higher ΔH_f
3. bp	Higher binding force –> higher bp
4. ΔH_v	Higher binding force –> higher ΔH_v
5. evaporation rate/volatility	Higher binding force –> lower evaporation rate
6. vapor pressure	Higher binding force –> lower vapor pressure
7. viscosity	Higher binding force –> lower viscosity
8. surface tension	Higher binding force –> lower surface tension
9. solubility	higher solute/solvent binding force –> higher solubility

C. Types of Substances

1. Molecular

- held together by IMF

IMF vs. Covalent
(weak) (strong)



- a given single molecule has atoms linked by strong, full covalent bonds; but 2 molecules attracted only by weak IMF
- molecular ⇒ weak IMF ⇒ easy to disorganize
- one molecule can be removed from others without losing full bond

2. "Network type"

sually
etal
nvolved

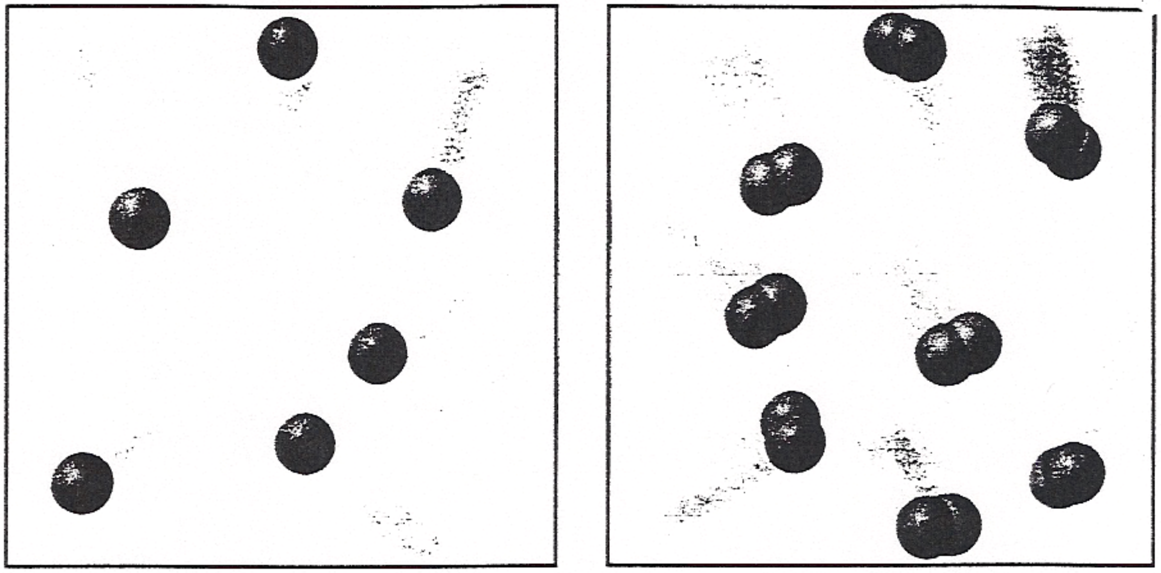
a) ionic NaCl
rare → b) network covalent C (diamond), SiO_2 (quartz)
c) metallic Fe

- reshow (T75 Gillespie 11.5)
- full lattice
- strong binding (full bonds)
- can't remove one without losing full bonds
- always solids

Key: network $\underbrace{\text{substance} \gg}_{\text{binding}}$ molecular ← nonmetals only

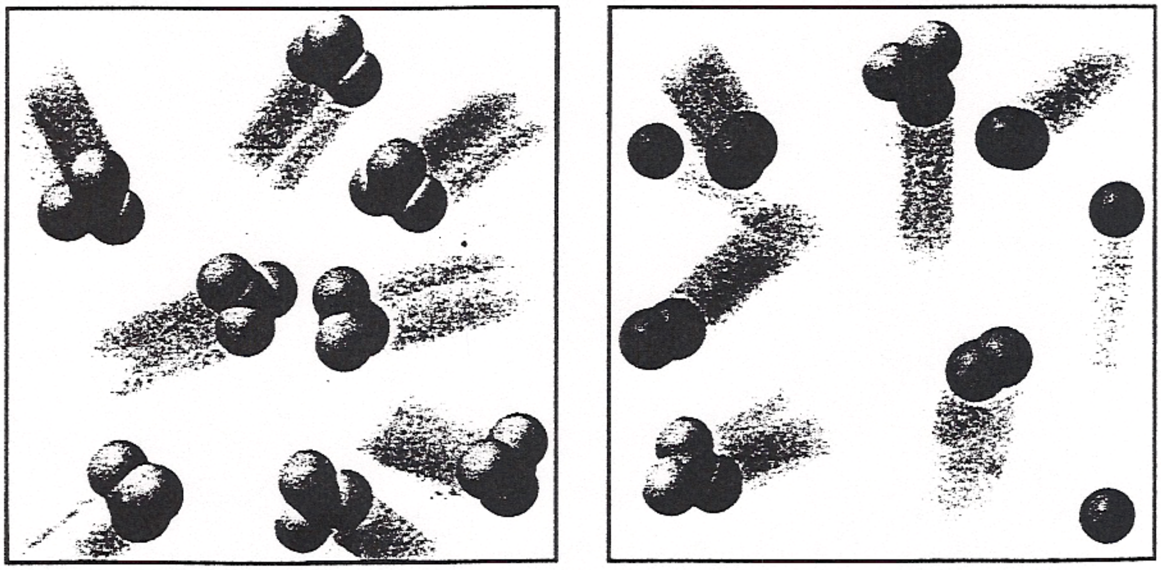
⇒ molecular have much lower mp, bp, etc.

T–2 Fig. 1.5 Separation of Matter (mixtures, elements, compounds)



(a) Atoms of an element

(b) Molecules
of an element



(c) Molecules
of a compound

(d) Mixture of elements
and a compound

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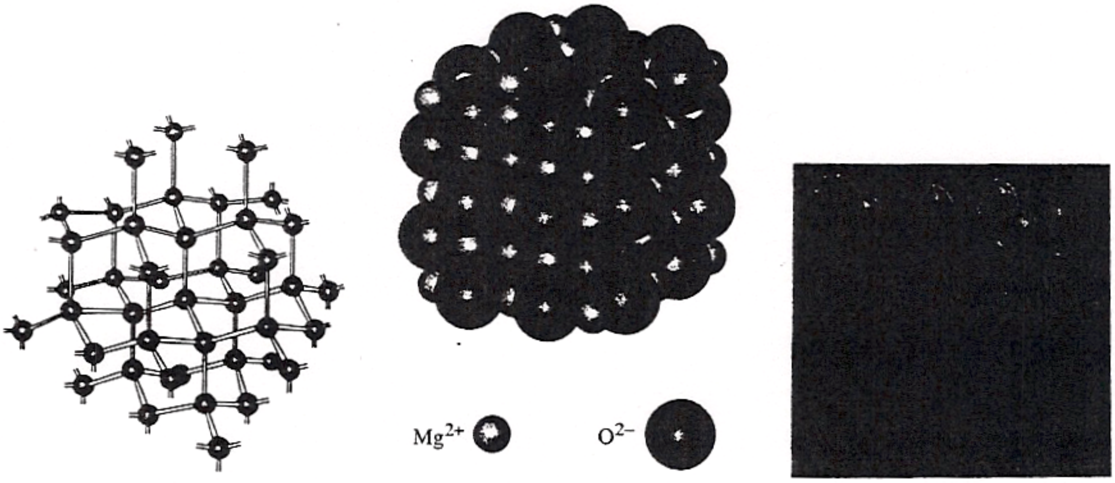
Fig. 11.5 Three-Dimensional Network Solids



Diamond

(a)

Magnesium oxide

(b)

Copper

(c)

ATOMS MOLECULES
AND REACTIONS
Gillespie/Eaton/
Humphreys/Robinson

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Chapter 11 + 9: Polarity, Binding Forces, Noncovalent Interactions, and Dependent Physical Properties

Types of Binding Forces

Full bonds, strong

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    - Hydrogen Bonds (O-H, N-H, or F-H bonds)
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Recognizing Polarity for polyatomic molecules

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9. solubility	higher solute/solvent binding force –> higher solubility

11-4

(IMF)

9.5 Noncovalent Forces in Molecular Substances

- if not for IMF, all moleculars ⇒ gases!!
- IMF < 10% full
- strength of IMF determines properties
- nature based on +/− attraction

1. Dipole-dipole
   - polar molecules

$$\begin{array}{ccccc} \overset{\delta+}{H}-\overset{\delta-}{Br} & \cdots & \overset{\delta+}{H}-\overset{\delta-}{Br} \\ \vdots\quad\vdots & & \vdots\quad\vdots \\ Br-H & \cdots & Br-H \end{array}$$

   - $\boxed{\text{more polar} \Rightarrow \text{stronger}}$ (weight equal)

	Mass	Dipole	BP (K)
CH_3OCH_3	46	1,3	248
CH_3CHO	44	2,7	294
CH_3CN	41	3,9	355

Recognizing Polarity: Handout $\boxed{\text{See 9,4}}$

Br−Br   CO NCl_3 CH_4 $CHCl_3$

- diatomics: same or diff?
- polyatomics
  1) lone pair on central? ⇒ Polar
  2) no lone pair ⇒ nonpolar = outside all same
     ⇒ weakly polar = outside not same

2. Hydrogen-Bonds

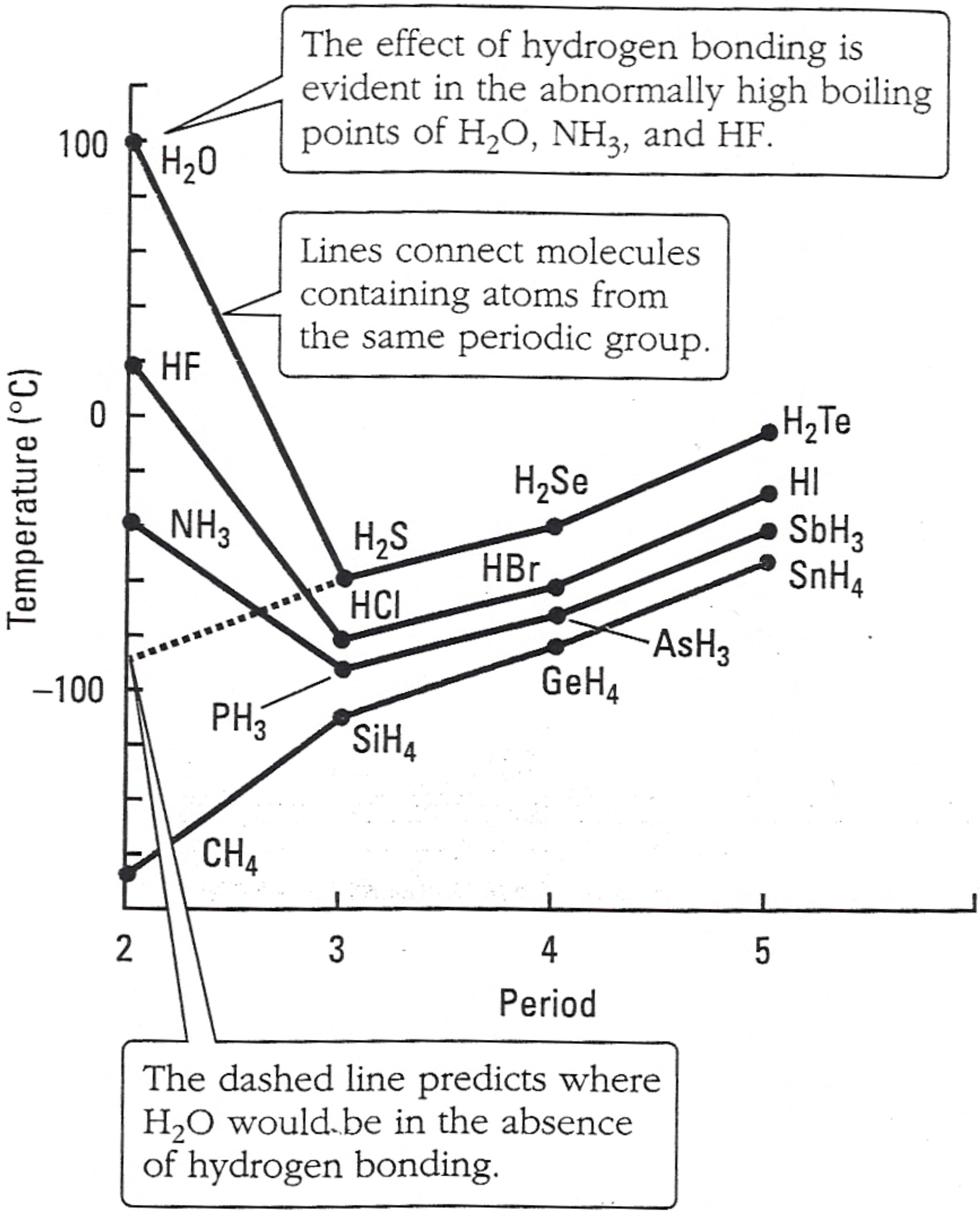
requires O–H, N–H, or F–H bond

- ultra-strong dipole-dipole between $H^{\delta+}$ and a lone pair on N, O, or F (still <10% of full bond)

ex $\overset{\delta-}{HO}-\overset{\delta+}{H} - :\ddot{O}H_2$ $H_2N-\overset{\delta+}{H}\cdots\overset{\delta-}{N}H_3$

~~[illegible]~~ Fig 9.21

- $H^{\delta+}$ especially $\delta+$ OH > NH NOF
- each $H^{\delta+}$ and each lone pair likes an H-bond
- causes huge jump in IMF, bp, etc
- why water so special
  (bp, density, dissolving power, air conditioning ...)
- H-bonding (⇒ "hydrophilic") dominating force in biological organization



The effect of hydrogen bonding is evident in the abnormally high boiling points of H_2O, NH_3, and HF.

Lines connect molecules containing atoms from the same periodic group.

The dashed line predicts where H_2O would be in the absence of hydrogen bonding.

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Figure 9.21

3. London Force (LF)

$Cl_2(g)$ $Br_2(l)$ $I_2(s)$

- all nonpolar, so what binds
- nonpolar lack permanent dipole, but electron flow ⇒ temporary dipole

$\boxed{\text{Brown 11.5}}$

"Polarizability" - ease of electron distortion

★ More e's ≡ more mw ⇒ more LF

⇒ $\boxed{\text{IMF increases with mw}}$

$\boxed{\text{Table 9.5}}$

mw ↓

	BP(K)
F_2	85
Cl_2	239
Br_2	332
I_2	458

IMF ↓

- LF only force in nonpolars, but also very important in polar / H-bonders

$\boxed{\text{Table 9.6}}$

bp: CH_3OH << CH_3CH_2OH

Effect of Numbers of Electrons on Boiling Points of Nonpolar Molecular Substances

Noble gases			Halogens			Hydrocarbons		
	No. e's	bp (°C)		No. e's	bp (°C)		No. e's	bp (°C)
He	2	−269	F_2	18	−188	CH_4	10	−161
Ne	10	−246	Cl_2	34	−34	C_2H_6	18	−88
Ar	18	−186	Br_2	70	59	C_3H_8	26	−42
Kr	36	−152	I_2	106	184	C_4H_{10}*	34	0

* Butane.

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Table 9.5

Numbers of Electrons and Boiling Points of Nonpolar and Polar Substances

Nonpolar molecules			Polar molecules		
	No. e's	bp (°C)		No. e's	bp (°C)
N_2	14	−196	CO	14	−192
SiH_4	18	−112	PH_3	18	−88
GeH_4	36	−90	AsH_3	36	−62
Br_2	70	59	ICl	70	97

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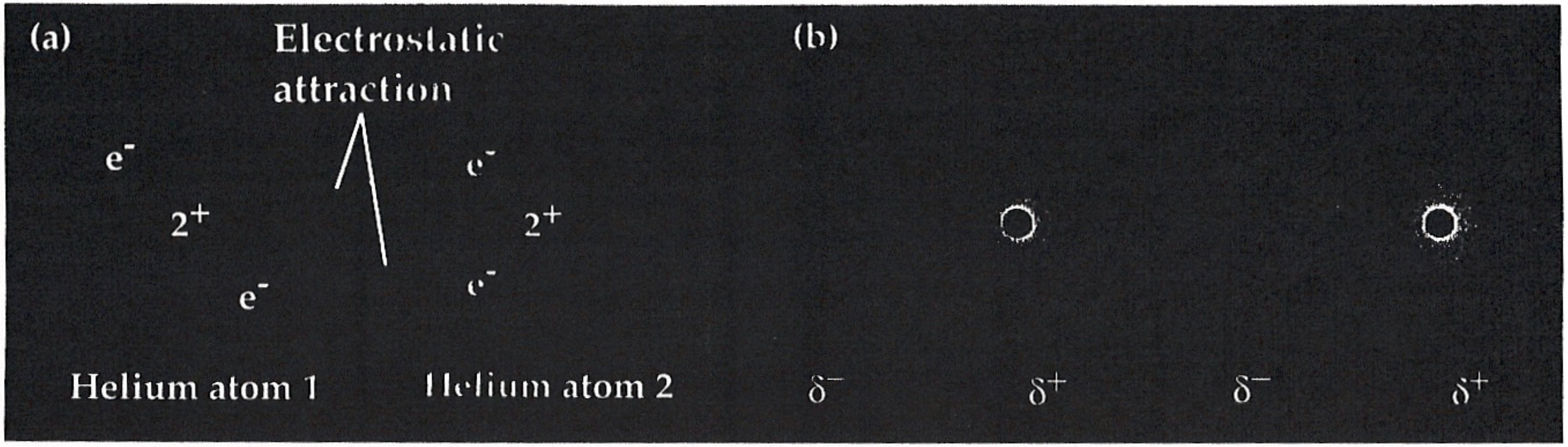
Table 9.

Fig. 11.5 London Dispersion Forces



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Comparing/Identifying/Ranking binding Forces

Importance: Network (full bonds) > H-bonding > London > Dipole

⇑

mw factor

Fig 11.12

Q's: ① Network or molecular?
(metals, ionic?)
② H-bonding? (OH, NH)
③ Polar?
★④ What is mw?

Tiebreakers

a) if mw = equal, H-bonding or polarity breaks tie

b) if H-bonding or polarity equal, use mw to break tie

Fig. 11.12 Flowchart of Intermolecular Forces

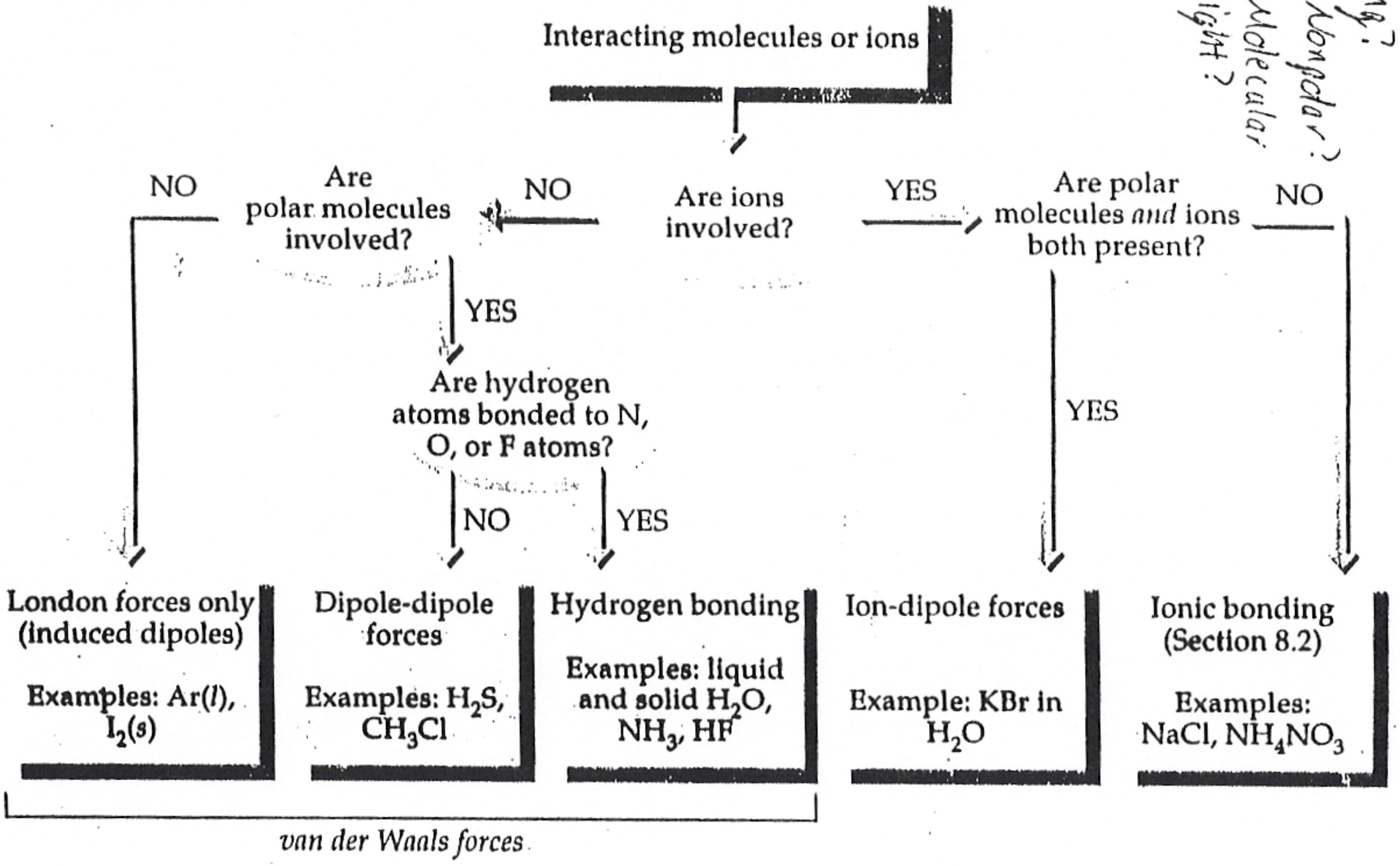
Questions

① Ionic or Molecular? (Network or Molecular)

② H-Bonding?

③ Polar or Nonpolar?

④ What is Molecular Weight?



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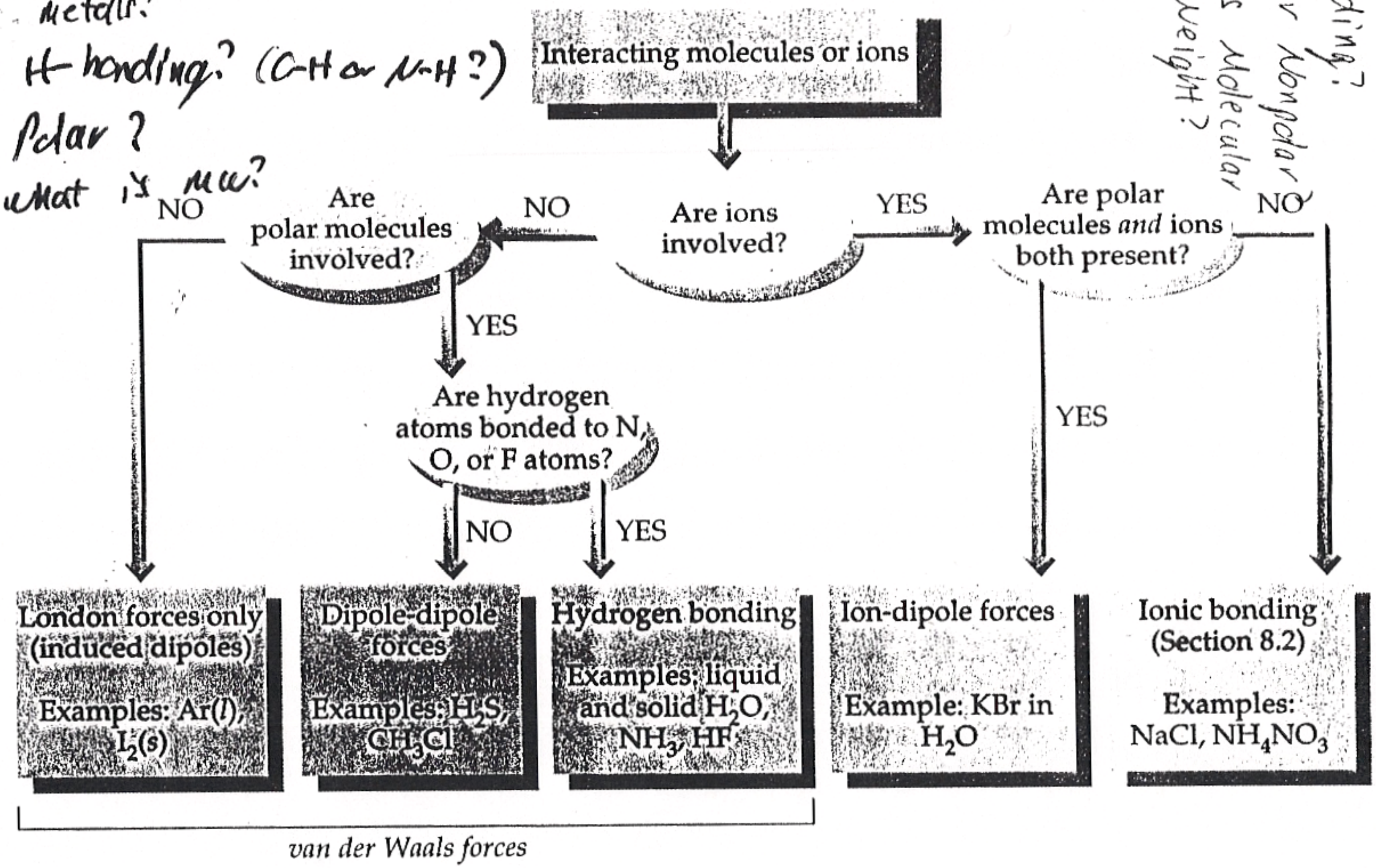
T-87

Fig. 11.12 Flowchart of Intermolecular Forces

Questions
① Network or Molecular?
metals?
② H-bonding? (O-H or N-H?)
③ Polar?
★ ④ what is MW?



Questions
① Ionic or Molecular? (Network or Molecular)
② H-Bonding?
③ Polar or Nonpolar
④ What is Molecular weight?

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(11-8)

① Identify any Forces (Other than London) for following:

br_2

CH_3OH

CH_3OCH_3

Fe_2O_3

SiH_4

$H-\ddot{N}=O$

CO

NH_3

② which has London as <u>only</u> binding force?

H_2S Cl_2 CH_2O H_2O

(11-8)

① Identify any Forces (Other than London) for following:

Br_2 mol → LF

CH_3OH mol → H-bonding

← lone pairs

$CH_3\ddot{O}CH_3$ mol → dipole-dipole

Fe_2O_3 ionic! Network

SiH_4 mol → London

$H-\ddot{N}=O$ H-bonding

CO dipole-dipole

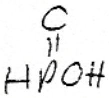
NH_3 H-bonding

② Which has London as <u>only</u> binding force?

H_2S [✓ Cl_2] CH_2O H_2O

11-9

① which will have highest bp? mp?

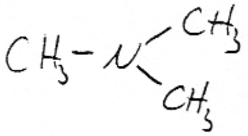
	SO_2	Cl_2	 $H\overset{O}{\overset{\|}{C}}OH$	$CH_3CH_2CH_2CH_3$
mw =	64	70	65	72

② which has highest mp? $NaCl$ $HOCH_2CH_3$ CH_3OCH_3
lowest bp?

Highest bp:

③ OBr_2 OF_2

④ $CH_3CH_2CH_2NH_2$



⑤ CH_3OH CH_3CH_2OH

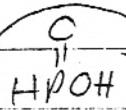
⑥

Br_2	F_2	SiH_4	CO_2
160	38	32	44

⑦ Rank, from highest to lowest bp?

CH_3NO_2	LiF	Cl_2	CH_3OH
70	26	70	32

① which will have highest bp? mp?



	SO_2	Cl_2	$H\overset{O}{\overset{\|}{P}}OH$	$CH_3CH_2CH_2CH_3$
mw=	64	70	65	72
			H-bonding	

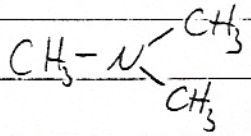
② which has highest mp? lowest bp?

$NaCl$	$HOCH_2CH_3$	CH_3OCH_3
ionic, network	↑ H-bonding	lowest, Dipole-dipole

Highest bp:

③ OBr_2 OF_2 mw ≡ London Force

④ $CH_3CH_2CH_2NH_2$



$CH_3-N(CH_3)_2$ H-bonding

⑤ CH_3OH CH_3CH_2OH H-bonding

⑥

Br_2	F_2	SiH_4	CO_2	mw=London Force
160	38	32	44	

⑦ Rank, from highest to lowest bp?

CH_3NO_2	LiF	Cl_2	CH_3OH
70	26	70	32
③	①	④	②
polar, dipole-dipole	network, ionic	nonpolar, London only	H-bonding

11.1 Liquid State

— molecular flow

— binding: too weak for solid
too strong to be gas

Ketchup, syrup, motor oil

A. "Viscosity": resistance to flow, "thickness" of liquid

— depends on binding vs. KE
(anti-flow) (pro flow)

Greater IMF ⇒ more viscous
Greater temp ⇒ less "

"molasses in January"

(discuss motor oil)

b. Surface Tension

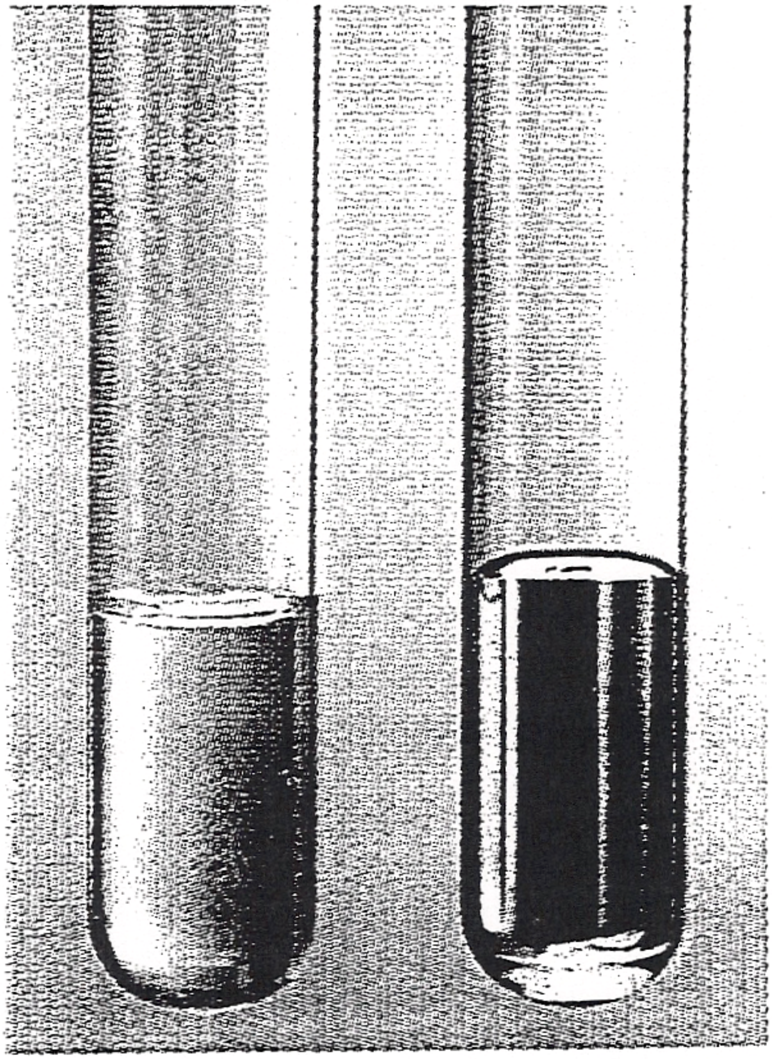
— water "beads up" on smooth surface or overfilled glass

— mercury more, rubbing alcohol, turpentine less

Max IMF ⇒ more S.T.
temp ⇒ less

IMF pulls molecules together, ball up, pack liquid as if it had skin

— water bugs "walk" on water



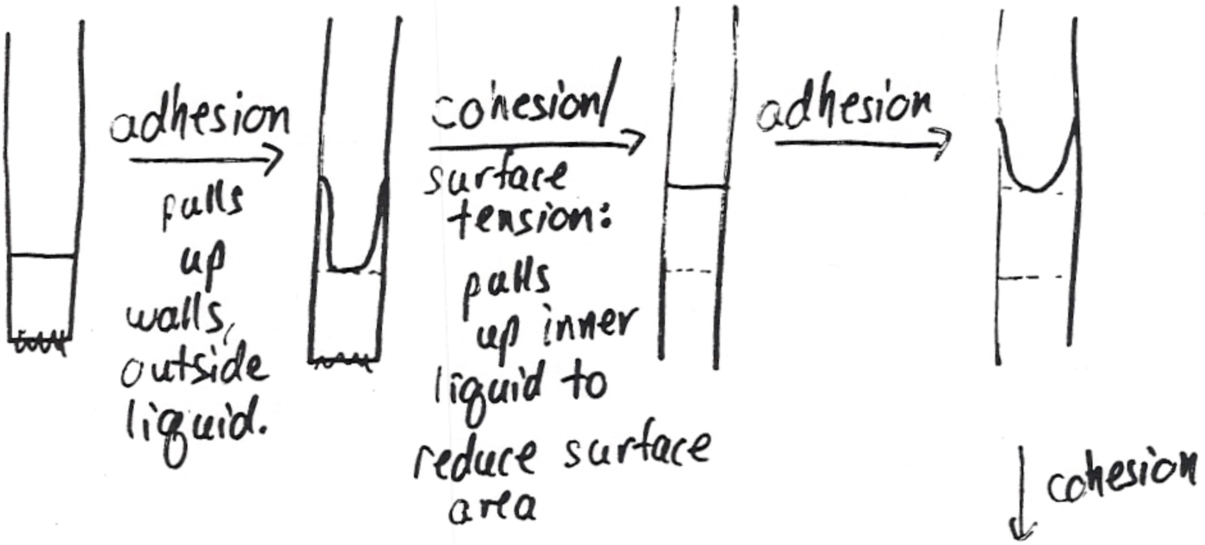
▲ FIGURE 11.16 The water meniscus in a glass tube compared with the mercury meniscus.

Meniscus Chemistry

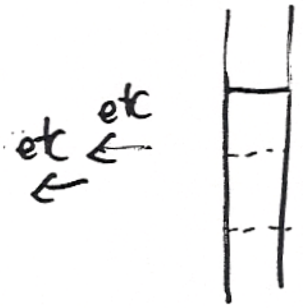
H_2O in glass: adhesion > cohesion

Hg in glass: cohesion > adhesion

Capillary Action: Liquids climb up narrow tubes! (against gravity)



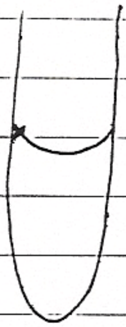
Combined adhesion/ cohesion "steps" liquid up a narrow tube (sometimes!) Key to plant biology!



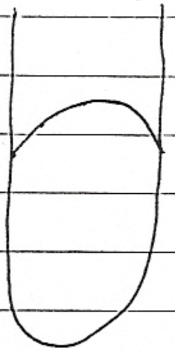
C. Meniscus + Capillary Chemistry

"cohesive force" - attraction of liquid to itself
"adhesive " " " surface

<u>meniscus</u> shape depends on cohesive vs. adhesive:
(IMF to self or surface)



adhesive > cohesive
(water)



cohesive > adhesive (mercury)

<u>capillary action</u>: liquid climbs narrow tube,
or absorbant material
(demo: paper towel in water)
- key to plants, getting water (+ nutrients) up to leaves

- how: adhesive/cohesive pull!!
⇓ outside ⇓ inside

Motor oil: want right viscosity,
but changes between -30° F → 90° F
(why change oil between summer + winter)

- higher SAE Number (Soc Auto Engineers), higher viscosity

5W/30 (winter)
at -18° F (winter) ≈ to SAE 5 (thin)
at 99° F (summer) ≈ SAE 30 (thick)
10W/30 or 40 good summer, thick winter

Future: synthetic, less sensitive to temp, better able to act thin enough in winter & thick enough in summer

- ideally not break down (sludge "thickens")
- don't let engine clears get destroyed as fast

11.2 Vapor Pressure

"Volatility" - tendency to evaporate
- why do some things evaporate faster? later
- why does Pepsi go flat? vs. oil?
- why does anything evaporate below bp?

A. Theory

① temp reflects average KE

② some above-average molecules have enough KE to break away

⇒ molecules always escaping

⇒ evaporation below bp!!

Moore Fig 11.4

③ higher temp ⇒ higher fraction/volatility (why water evaporates faster in hot day than cold)

At given temp, a molecule with stronger binding is less volatile

B. Vapor Pressure

- in closed container, gas can't escape, & builds up
- gas exerts pressure; also can condense
- achieves "dynamic equilibrium":
  rate of vaporization = condensation
- constant supply of gas
(steady state)

illustrate: roller coaster at Valley Fair always the same people getting off and on, but steady state #

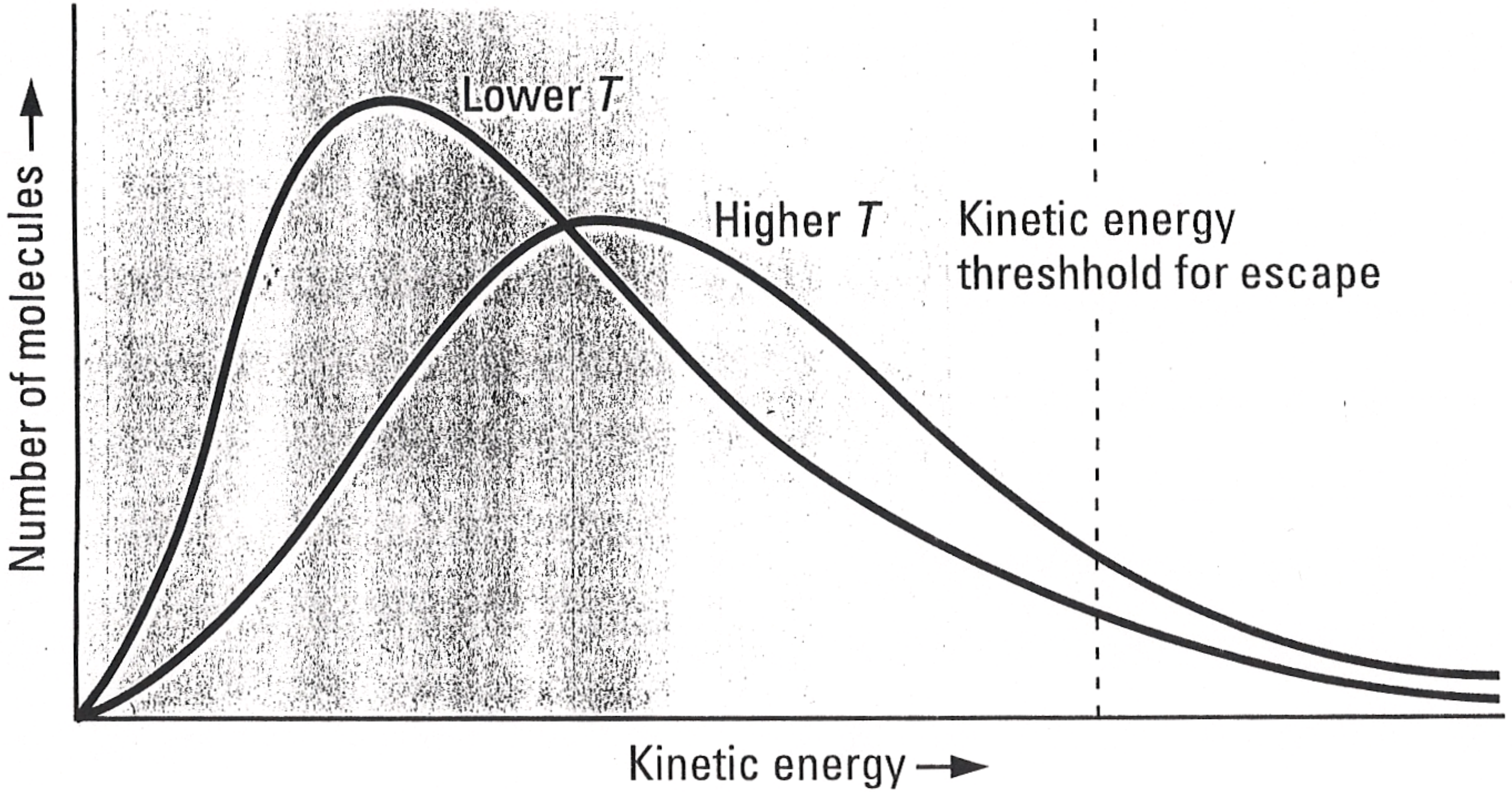


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Figure 11.4

11-13

Vapor Pressure: when liquid/gas in equilibrium (or gas pressure)

– more stuff in gas phase (escaping) ≡ higher vapor pressure

2 Factors Fig 11.5

① Temp ⇒ Higher VP

② IMF (higher IMF ⇒ lower VP)

– relative VP ordering never changes

C. VP and BP and Pressure

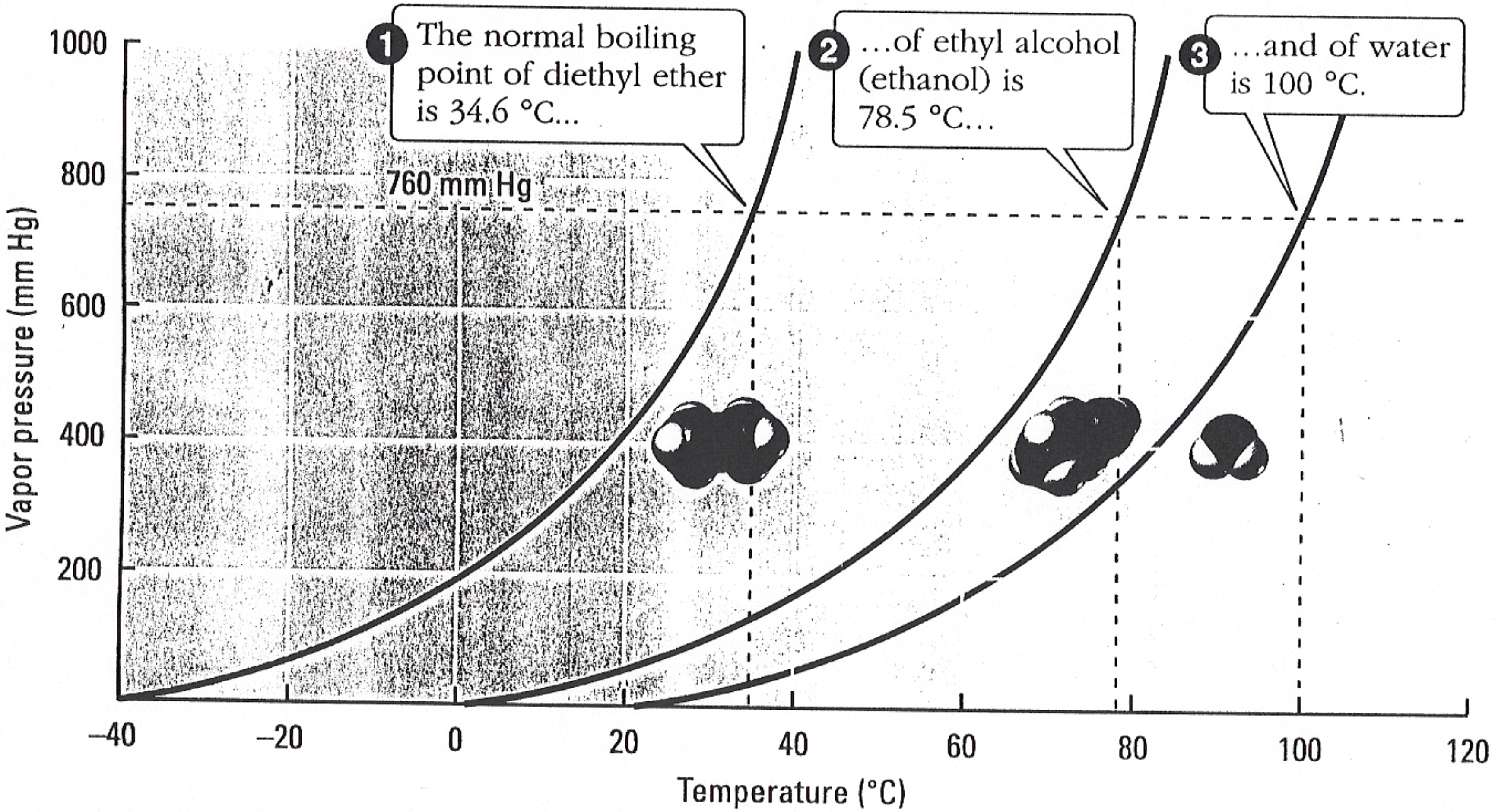
– liquid boils when VP (escaping gas) equals external pressure

"normal bp" = bp at 1atm pressure = 760 mm Hg = 760 torr

higher IMF ⟺ higher bp

higher external pressure ⇒ higher bp (pressure cooker ⇓ higher bp ⇓ hotter, faster cook)

lower ⇒ lower bp (longer to boil in egg in Denver)



① The normal boiling point of diethyl ether is 34.6 °C...

② ...of ethyl alcohol (ethanol) is 78.5 °C...

③ ...and of water is 100 °C.

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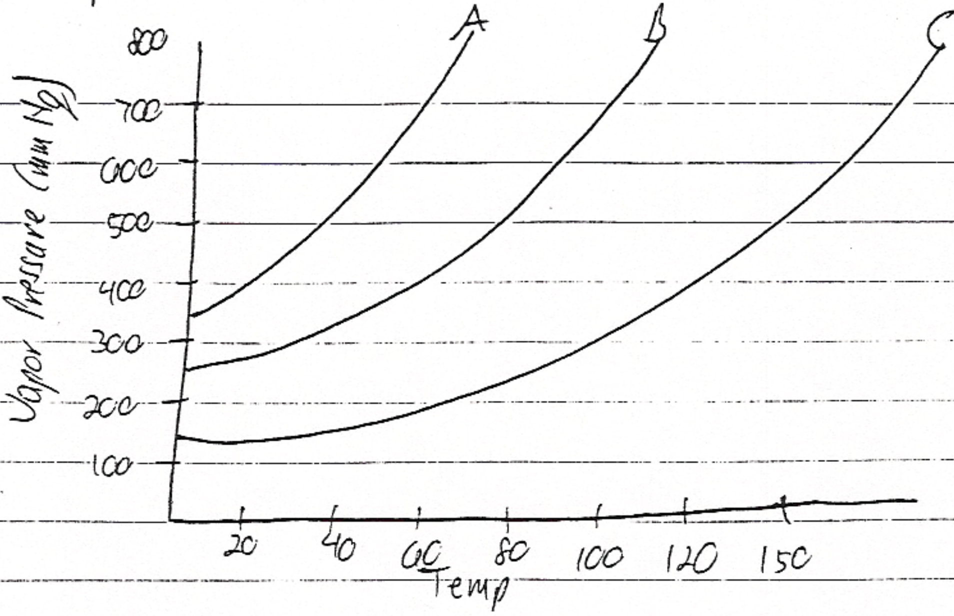
Figure 11.5

(11-14)

Vapor Pressure Curve

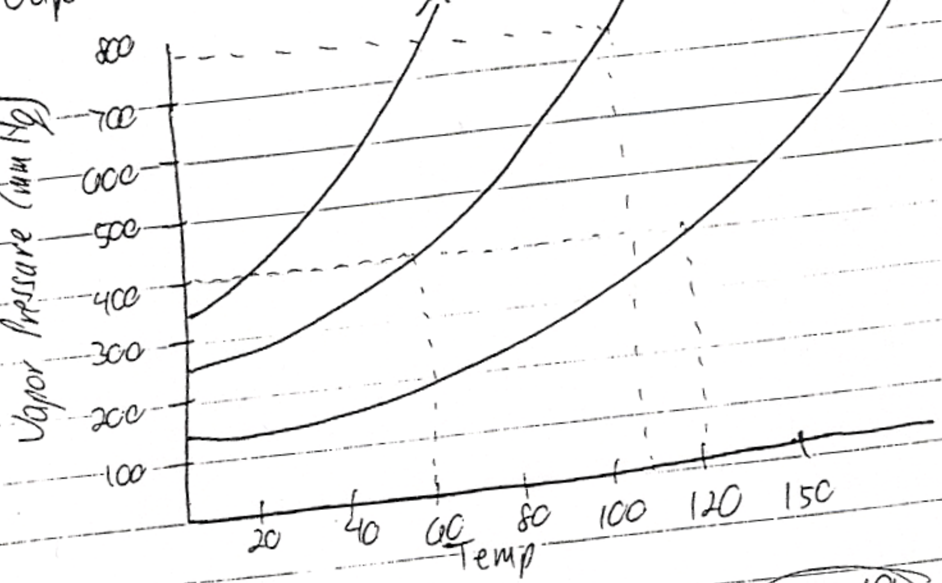


① whats the "normal" bp for B?

② whats the bp for B at 400 mm Hg?

③ What temp needed to get 400 mm Hg for C?

④ Rank the IMF for A, B + C.



① whats the "normal" bp for B? ~100°

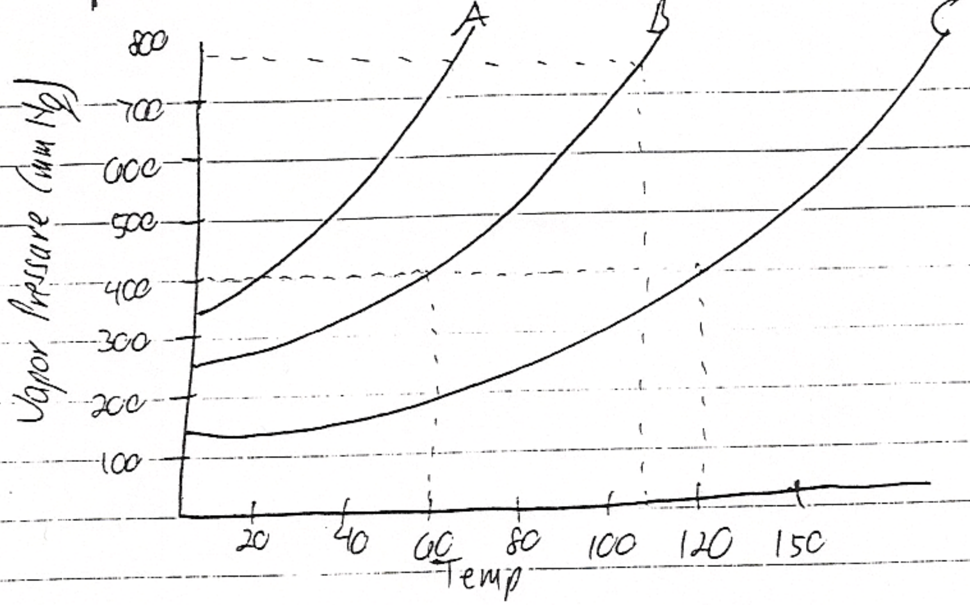
② whats the bp for B at 400 mm Hg? 60°

③ what temp needed to get 400 mm Hg for C? 120°C

④ Rank the IMF for A, B + C. $\boxed{C > B > A}$

(11-14)

Vapor Pressure curve



① whats the "normal" bp for B? ~100°

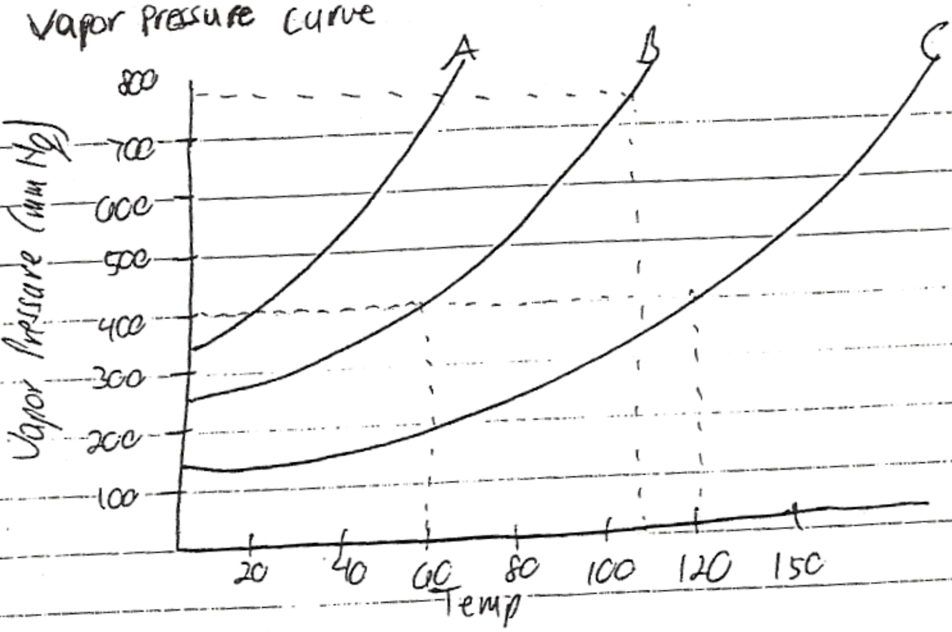
② whats the bp for B at 400 mm Hg? 60°

③ What temp needed to get 400 mm Hg for C? 120°C

④ Rank the IMF for A, B + C. C > B > A

(11-14)

Vapor Pressure curve



① whats the "normal" bp for B? ~100°

② whats the bp for B at 400 mm Hg? 60°

③ What temp needed to get 400 mm Hg for C? 120°C

④ Rank the IMF for A, B + C. $\boxed{C > B > A}$

(11-15)

11.3 Phase Changes

A. Terms [Fig 11.7] (discuss)

"fusion" (melting)
"vaporization"
"deposition" (car frost, CD coating)
"sublimation"
"crystallization" ≡ freezing

B. Energetics

- energy used to change to more disordered state (need to up KE)
- " released " ordered
- ~~at phase change, all~~
- at bp or mp, all energy goes into phase change, not temp change

melting, vaporization, sublimation: endothermic
freezing, condensation, deposition: exo

- water as air conditioner -- (without freezing water temp 30° cooler!) evaporation is cooling!
  - → fridge: liquid evaporates (endo)
    electrical compressor resqueezes gas to liquid

11-15

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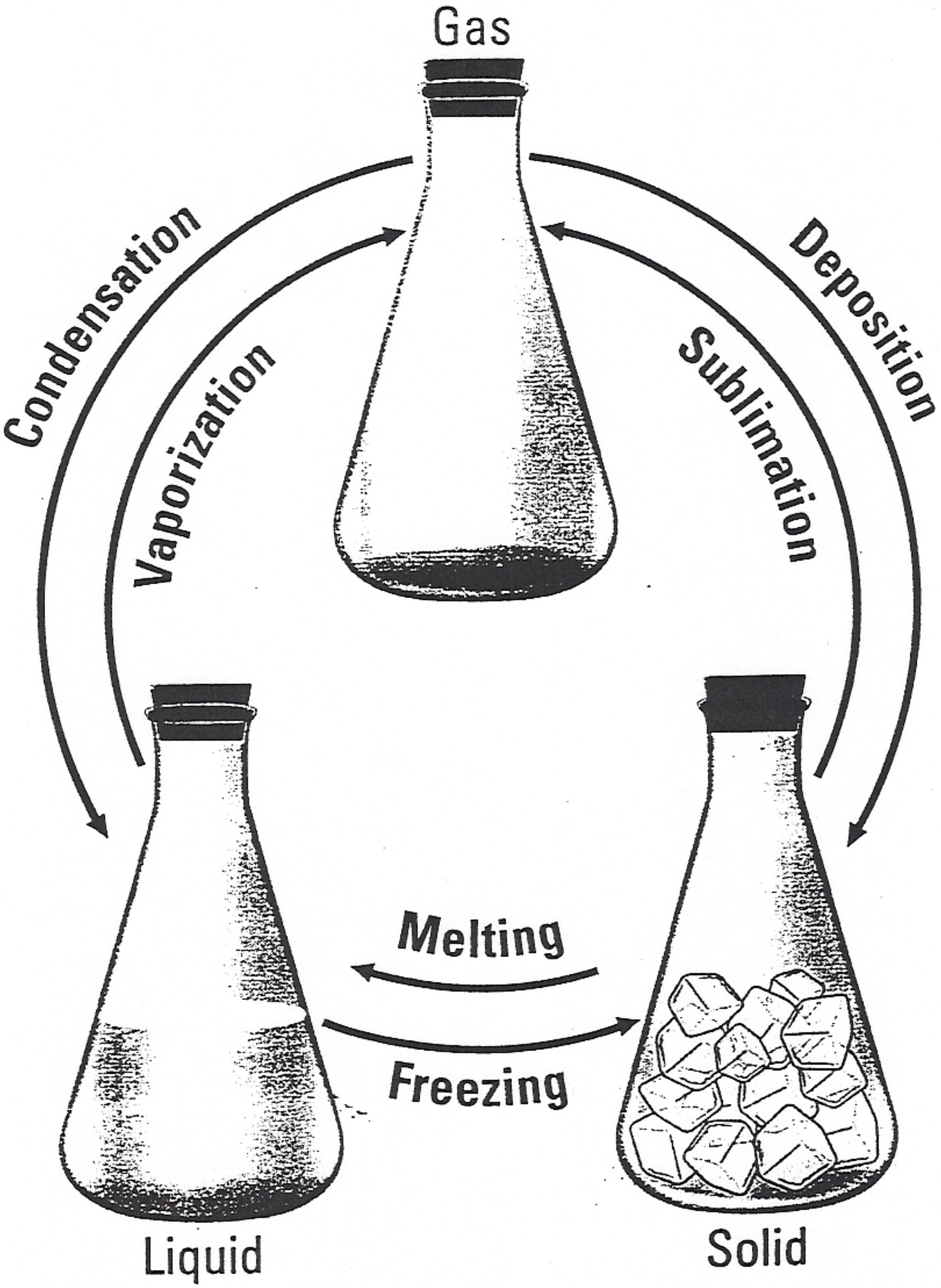
Figure 11.7



11-16

$\text{heat} + \text{liquid} \rightleftharpoons \text{gas}$

$\Delta H_{vaporize} = -\Delta H_{condense}$

ΔH_{vap} = energy to vaporize 1 mol (or 1g)
ΔH_{fus} = " melt
$-\Delta H_{vap}$ = condense
$-\Delta H_{fus}$ = freeze

larger IMF ⇒ larger ΔH_v, ΔH_f
$\Delta H_v \gg \Delta H_{fus}$

C. Heating Curves Fig 11.12, handout
"horizontals" = phase change
- 2 phases present
- heat goes into phase change, not temp

Skills: ID phase(s) at any point

energy calculations
- each portion calculate independantly
  Use ΔH's for phase changes
  use "specific heats" for temp changes

11-16

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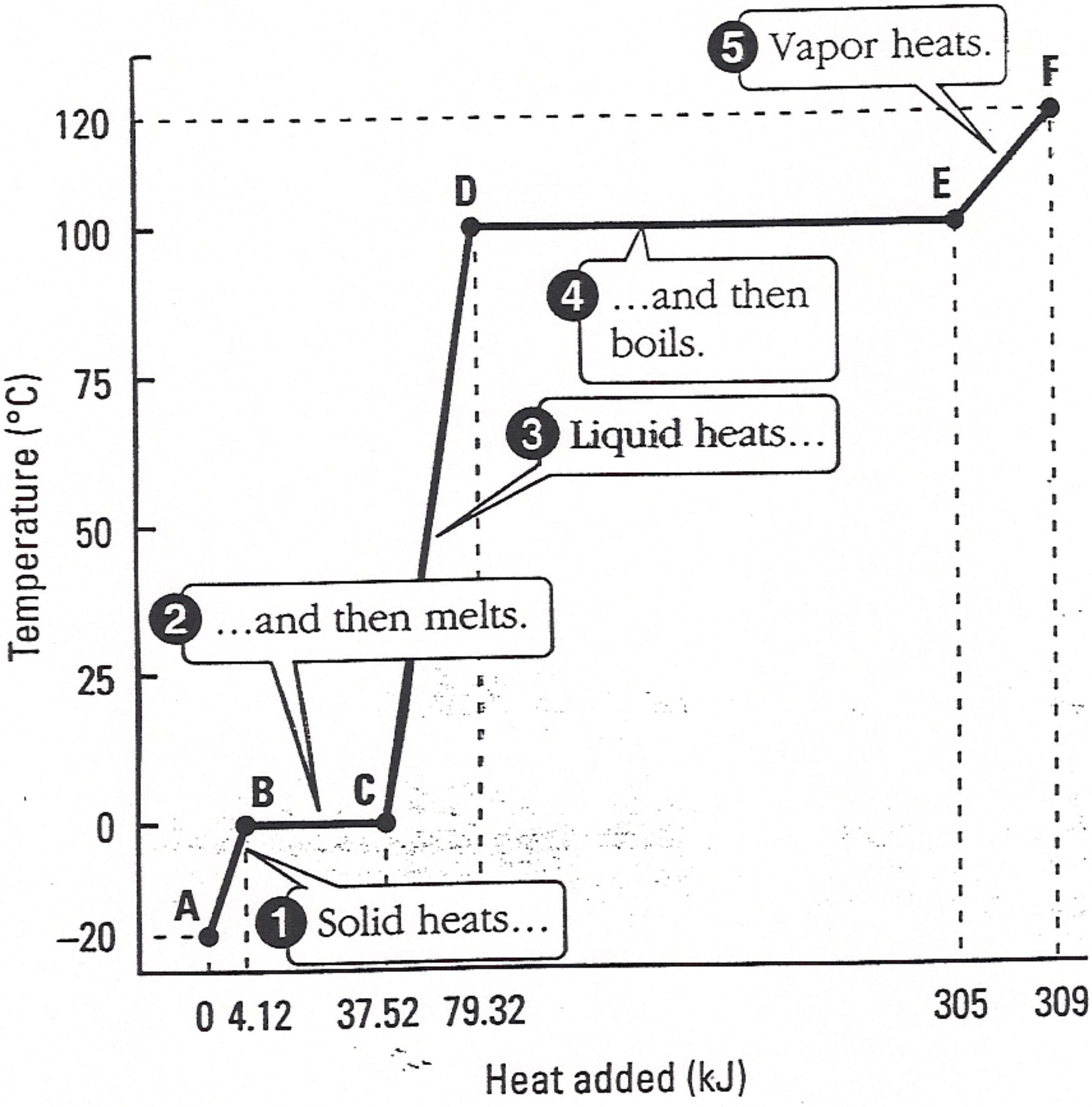
Figure 11.12



11-17

Heating Curve Sample Problem

How much energy (kJ) to heat 36 g ice (18g/mol) from −50°C → +50°C?

Given: ΔH_{fusion} = 6.01 kJ/mol

specific heat (ice) = 2.09 J/g·K

specific heat (water) = 4.18 J/g·K

11-17

Heating Curve Sample Problem

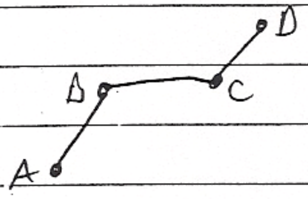
How much energy (kJ) to heat 36 g ice (18g/mol) from −50°C → +50°C?

→ 2 mol

Given: ΔH_{fusion} = 6.01 kJ/mol

specific heat (ice) = 2.09 J/g·K

specific heat (water) = 4.18 J/g·K



3 steps:

A → B warm ice

B → C melt ice

C → D warm water

A → B −50°C → 0°C

$$x\,J = \frac{2.09\,J}{g\cdot K}\Big|\,36g\,\Big|\,50K = 3.762\,kJ$$

B → C

$$\frac{6.01\,kJ}{mol}\Big|\,2\,mol = 12.02\,kJ$$

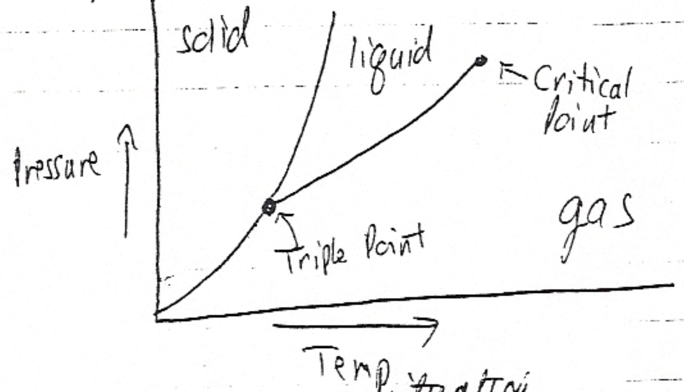
C → D

$$\frac{4.18\,J}{g\cdot K}\Big|\,36g\,\Big|\,50K = 7.524\,kJ$$

Sum total $= \boxed{23.3\ kJ}$

D. Phase Diagram

Fig 11.13



- on line: phase change ~~point~~ situation, 2 phase at once
- triple point: unique when all 3 phases present

Phases

max temp, min pressure ⟹ gas (lower right)
min temp, max pressure ⟹ solid (upper left)
- liquid between!

Increasing temp: solid → liquid → gas or solid → gas
Increasing pressure: gas → liquid → solid or gas → solid

"normal" pressure ≡ 1 atm ≡ 760 mm Hg

Critical point: above this, liquid/gas merge
"supercritical fluid"
- liquid density, gas-like flow

11-18

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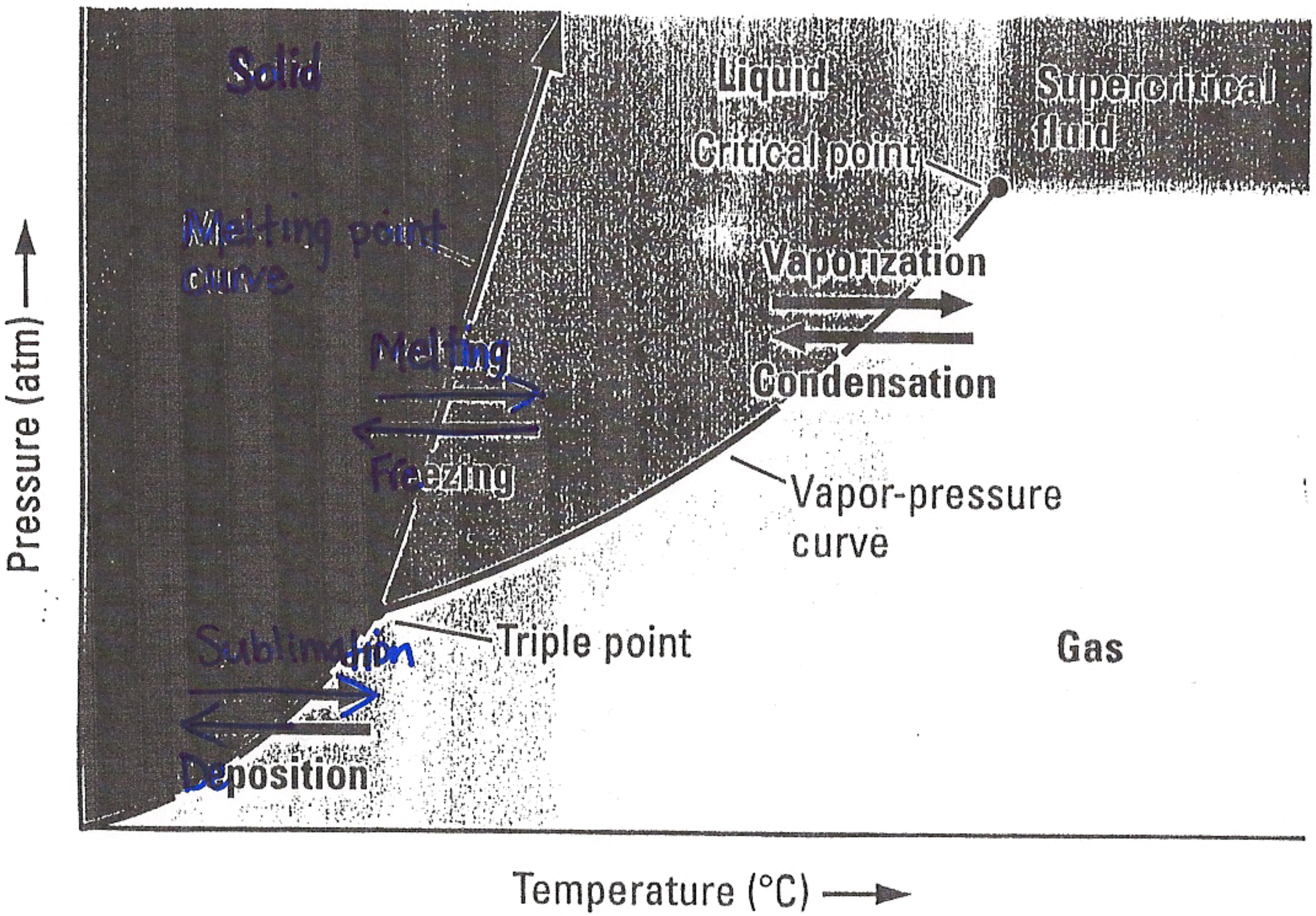
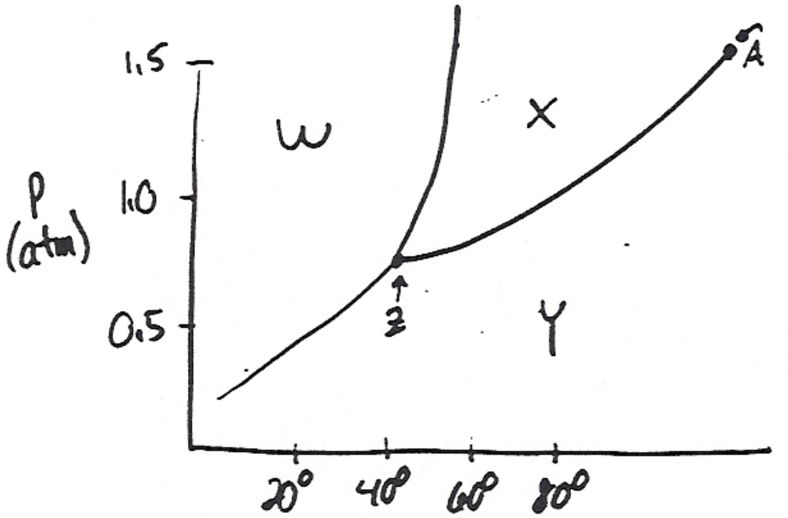
Figure 11.13





which letter represents:

① Gas Phase

② Liquid

③ Solid

④ Triple Point

⑤ Critical Point

⑤ what is the normal bp?

a) 20° b) 40° c) 60° d) 80°

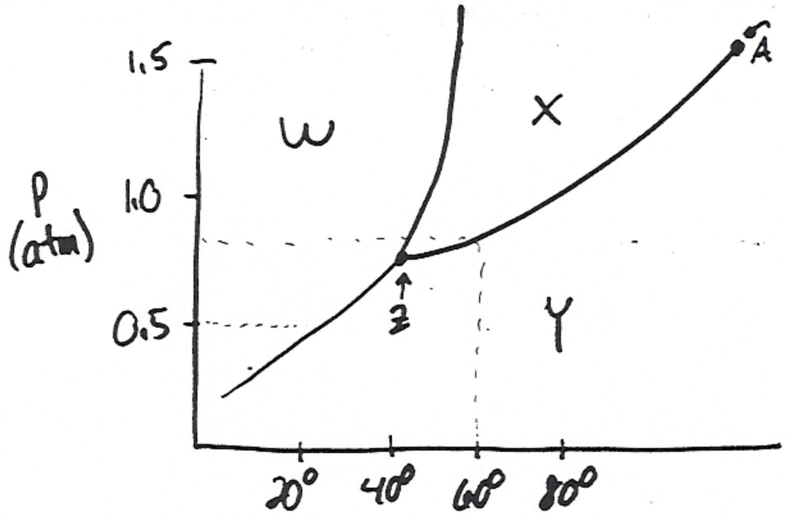
⑥ what is the normal mp?

a) 20° b) 50° c) 80°

⑦ when a liquid at 40° has pressure adjusted, at what pressure does it vaporize?

⑧ when solid at 0.5 atm is warmed, does it

a) melt or b) sublime



Which letter represents:

① Gas Phase Y

② Liquid X

③ Solid W

④ Triple Point Z

⑤ Critical Point A

⑤ What is the normal bp?

a) 20° b) 40° c) 60° [✓ d) 80°]

⑥ What is the normal mp?

a) 20° [✓ b) 50°] c) 80°

⑦ When a liquid at 60° has pressure adjusted, at what pressure does it vaporize? .8 atm

⑧ When solid at 0.5 atm is warmed, does it

a) melt or [✓ b) sublime]

11-20

11.4 Wonderful Water: Special Properties

① Given small size/mass, amazing that its liquid!

	Mass	bp		Mass	bp	
N_2	28	−196	HF	20	−164	
O_2	32	−183	H_2S	34	−60	
CO_2	44	−78	NH_3	17	−33	
CH_4	18	−161	H_2O	18	+100	liqu

② "Universal Solvent"

- far more substances dissolve in water than any other liquid ⇒ both ionic + molecula

③ Density (unique phase diagram)

- other solids more dense than liquids ⇒ sink!
- solid water (ice) unique in that it floats

⇒ impact on water life

- "turnover" (4° water sinks ⇒ O_2, nutrient exchan

④ Earth's "Air Conditioner" (+ Body's)

- exceptional heat capacity
- coolant in summer, heater in winter
- sweat/evaporation keeps us from overheating

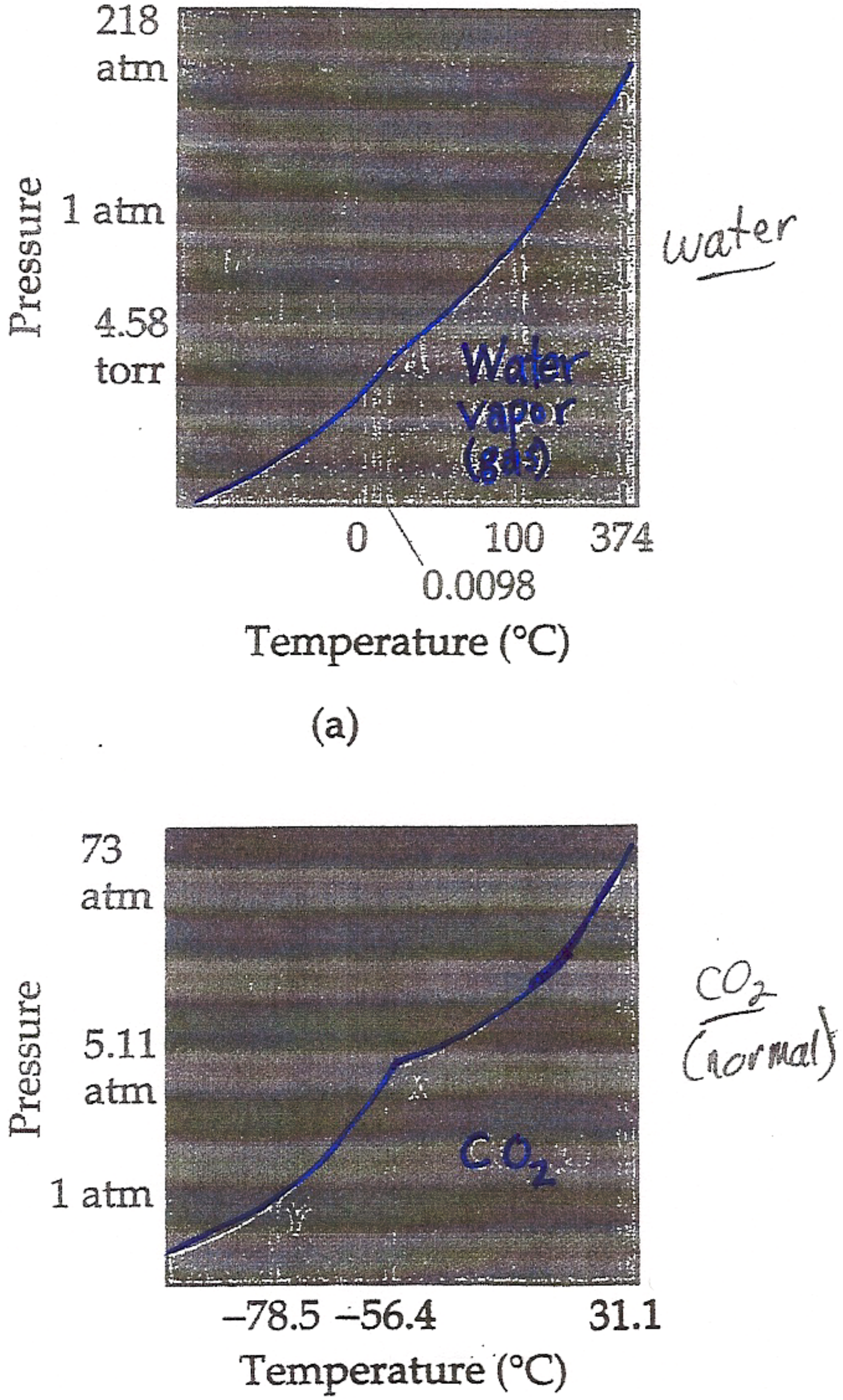
⑤ High surface tension

- bugs, [illegible] ships

H-Bonding Special Reason

T-91B

Fig. 11.25 Phase Diagram of H_2O and CO_2



(a)

(b)

CHEMISTRY: The Central Science 7/e
by Brown/Lemay/Bursten

11.8 Metals

- view bonding as metal cations in sea of valence electrons
- e's mobile ⇒ electrical conductivity

Properties

☆ electrical + thermal conductivity
- very high bp: can't vaporize
- "malleable" + "ductile" (not brittle)
- ≡ not brittle; can pound, bend, shape, draw into wires...
- variable hardness, mp
- shiny
- insoluble

Skip 11.6, 7, 9, 11

11.10 Covalent-Network Solids Brown T494

diamond: 3-d lattice ⇒ strong!

graphite: 2-d lattice ⇒ less strong, slip

11.5 Overview of Types of Solids
+ Properties

Key: Handout (11-23) Table 11-5

Amorphous solids: like network covalent in having covalent bonds
- solid, high melting
- nonuniform organization ⇒ a given solid melts over wide range

Key properties that discriminate:

① electrical conductivity
molecular poor
ionic poor in solid, good as liquid
metallic good

② hardness
molecular soft
ionic: hard, brittle
metallic: never brittle, malleable/ductile

③ solubility in water
molecular: variable
ionic: often good
metallic: no

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Table 11.5

Structures and Properties of Various Types of Solid Substances

Type	Examples	Structural units	Forces holding units together	Typical properties
Ionic (⊖ *p. 86*)	NaCl, K_2SO_4, $CaCl_2$, $(NH_4)_3PO_4$	Positive and negative ions (some polyatomic); no discrete molecules	Ionic bonding; attractions among charges on positive and negative ions	Hard; brittle; high melting point; poor electrical conductivity as solid, good as liquid; often water-soluble
Metallic (Section 11.8)	Iron, silver, copper, other metals and alloys	Metal atoms (or positive metal ions surrounded by an electron sea)	Metallic bonding; electrostatic attraction among metal ions and electrons	Malleability; ductility; good electrical conductivity in solid and liquid; good heat conductivity; wide range of hardness and melting points
Molecular (⊖ *p. 75*)	H_2, O_2, I_2, H_2O, CO_2, CH_4, CH_3OH, CH_3COOH	Molecules with covalent bonds	London forces, dipole-dipole forces, hydrogen bonds	Low to moderate melting points and boiling points; soft; poor electrical conductivity in solid and liquid
Network (Section 11.10)	Graphite, diamond, quartz, feldspars, mica	Atoms held in an infinite one-, two-, or three-dimensional network	Covalent bonds; directional electron-pair bonds	Wide range of hardnesses and melting points (three-dimensional bonding > two-dimensional bonding > one-dimensional bonding); poor electrical conductivity, with some exceptions
Amorphous (glassy)	Glass, polyethylene, nylon	Covalently bonded networks of atoms or collections of large molecules with no long-range regularity in their arrangement	Covalent; directional electron-pair bonds	Noncrystalline; wide temperature range for melting; poor electrical conductivity, with some exceptions

11-23