

Chemistry Unit 6

Primary reference: **CHEMISTRY**, Addison-Wesley

B3
Spr 17

Topic	Essential Knowledge	Study Support
Atomic Structure 2.6 SOL 2f	<p>Electronegativity is the measure of an atom's attraction for electrons in a bond. Electronegativity increases across a period toward the halogens and decreases down a group. The most electronegative atom is fluorine. The least electronegative element (excluding noble gases) is Francium, Fr.</p>	<p>Ch 14: Read pp. 405</p>
Nomenclature, Formulas, and Reactions 3.6 SOL 3d, 3e, 3f	<p>Exothermic reactions release heat whereas endothermic reactions absorb heat. Heat of reaction is the amount of energy absorbed or released during a chemical change. Exothermic reactions have a negative ΔH_{rxn}, whereas endothermic reactions have a positive ΔH_{rxn}. Examples of writing an exothermic reaction equation are:</p> $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + 890 \text{ kJ}$ <p style="text-align: center;">or</p> $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}, \Delta H_{rxn} = -890 \text{ kJ/mol}$ <p>Polar covalent bonds form between elements with very different electronegativities. The more electronegative atom will attract the electrons more strongly and this will result in it having a slight negative charge. The less electronegative atom then takes on a slight positive charge. A non-polar covalent bonds form between atoms of similar electronegativities.</p> <p>A polar molecule has unequally distributed electrons around the central atom. This is caused by unsymmetrical polar bonds or a lone pair on the central atom. The positive end of the molecule has a positive dipole and the negative end has a negative dipole. Polar molecules have dipole-dipole intermolecular attractions as well as London dispersion intermolecular attractions. Non-polar molecules only have London dispersion intermolecular attractions.</p> <p>with O-H, N-H or F-H bonds have hydrogen bonding attractions.</p> <p>Kinetics is the study of reaction rates. increase with increased temperature, reactant concentration, increased and the use of a catalyst. Activation minimum energy needed to initiate a High activation energies correspond to rates. Catalysts speed up reactions decreasing the activation energy. Potential energy diagrams are used to analyze reaction energy changes.</p> <div style="text-align: center;"> <p>EXOTHERMIC</p> </div>	<p>Ch 11: Read pp. 303-304</p> <p>Ch 16: Read pp. 460-466</p> <p>Ch 19: Read pp. 533-538.</p>
Molar Relationships 4.6	<p>Stoichiometry can be combined with heat of reaction, ΔH_{rxn}, to calculate the amount of heat produced from a known amount of reactant.</p>	
Phases of Matter and Kinetic Molecular Theory 5.6 SOL 5b, 5c, 5d	<p>Forces of attraction (intermolecular forces) between molecules determine their state of matter at a given temperature. Forces of attraction include hydrogen bonding, dipole-dipole attraction, and London dispersion (van der Waals) forces.</p> <p>Vapor pressure is the pressure of the vapor found directly above a liquid in a closed container. When the vapor pressure equals the atmospheric pressure, a liquid boils. Volatile liquids have high vapor pressures, weak intermolecular forces, and low boiling points. Nonvolatile liquids have low vapor pressures, strong intermolecular forces and high boiling points. Sublimation is the phase change from solid to gas without passing through the liquid phase. A substance's triple point, is the pressure and temperature conditions where all three phases coexist in dynamic equilibrium.</p> <p>The following mathematical relationship between the pressure, volume and temperature of a gas is used to describe the behavior of gases:</p> $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$ <p>An Ideal Gas does not exist, but this concept is used to model gas behavior. A Real Gas exists, has intermolecular forces and particle volume, and can change states. The Ideal Gas Law states that</p> $PV = nRT$ <p>R is the ideal gas law constant and has two values depending on the pressure units. They are R = 8.314 L·kPa/mol·K and R = 0.0821 L·atm/mol·K</p> <p>Dalton's Law of Partial Pressures says the sum of the partial pressures of all the components in a gas mixture equals the total pressure of the gas mixture.</p> $P_{\text{total}} = P_A + P_B + P_C \quad \text{and} \quad n_A/n_{\text{total}} = P_A/P_{\text{total}} = V_A/V_{\text{total}}$ <p>Graham's Law says gas molecules with the lightest mass travel fastest. $K.E = 0.5 mv^2$</p>	<p>Ch 10: Read pp. 269-280 and pp. 284-286.</p> <p>Ch 12: Read pp. 350-353.</p>

Unit 6 Objectives
Chemistry, Addison-Wesley, 2002

- I) Endothermic and Exothermic Reactions
 - A) Classifying Reactions
 - B) Stoichiometry and Calculating Heats of Reaction
- II) Intermolecular Forces (IMFs)
 - A) Polar bonds
 - B) Polar molecules
 - C) Intermolecular Attractions and Physical Properties
 - 1) Intermolecular forces
 - (a) London Dispersion forces
 - (b) Dipole-Dipole attractions
 - (c) Intermolecular Hydrogen bonding
 - 2) Effect of Intermolecular Forces on Physical properties
 - 3) Comparing molecular and ionic compounds
- III) Phase Changes and Intermolecular Forces (IMFs)
 - A) Kinetic Energy, Particle Velocity, and Kelvins
 - B) Kinetic Energy and Liquids
 - 1) Vapor pressure
 - 2) Boiling points and atmospheric pressure
 - C) Kinetic Energy and Solids
 - D) Phase Changes and Phase Diagrams
- IV) Gas Laws: Combined, Ideal, Dalton's Law and Graham's Law
- V) Reaction Rates
 - A) Collision Theory
 - B) Potential Energy Diagrams
 - 1) Activation Energy
 - 2) Catalysts

(SOL) Learning Objective

1. (3e) Identify a reaction as endothermic or exothermic based on its thermochemical equation and/or sign of ΔH .
2. (4b) Calculate the heat change of a reaction using stoichiometry and heats of reaction.
3. (2f) Compare the electronegativity of two elements based on their position on the periodic table.
4. (3d) Compare relative bond polarity based on the two elements position on the periodic table.
5. (3d) Use VSEPR theory and electronegativity to identify polar and non-polar molecules.
6. (5d) Identify and compare the three types of intermolecular forces (dipole interaction, hydrogen bonding, London dispersion (van der Waals) forces)
7. (5d) Predict the relative melting and boiling points of molecular and ionic substances based on intermolecular forces.
8. (5d) Explain the relationship between kinetic energy and temperature
9. (5d) Interpret a graph of percent molecules vs kinetic energy
10. (5b) Explain why real gases condense whereas ideal gases do not condense using IMFs and kinetic energy.
11. (5b) interpret vapor pressure graphs.
12. (5d) Explain what happens as a solid melts using IMFs and kinetic energy.
13. (5d) Explain the relationship between a substance's vapor pressure and boiling point and the strength of the substance's IMFs.
14. (5d) Interpret the effect of temperature and pressure on states of matter using a phase diagram.
15. (5d) Identify the triple point on a phase diagram and identify which states of matter exist at the triple point.
16. (5d) Identify phase changes on a phase diagram of water including fusion, solidification, vaporization, condensation and sublimation.
17. (5b) Solve gas law problems using the Combined Gas Law and the Ideal Gas Law.
18. (5b) Explain the difference between a real gas and an ideal gas.
19. (5b) Predict when a gas will behave most ideally.
20. (5b) Use Dalton's Law to calculate partial pressures
21. (5b) Use Graham's Law to compare rates of effusion and diffusion of two gases
22. (3f) Draw a reaction's potential energy diagram with axes labeled, and ΔH , activation energy, product energy, reactant energy, transition state, and catalyst shift clearly identified for exothermic and endothermic reactions.
23. (3f) Explain how a catalyst increases reaction rate.
24. (3f) Identify and explain the effect the following factors have on the rate of a chemical reaction: (catalyst, temperature, concentration, and reactant particle size).

Chapter 11 Part 2: Endothermic and Exothermic Reactions

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Classifying Reactions as Endothermic or Exothermic

Exothermic Reactions

✓ we know

Endothermic Reactions

Heat of Reaction, $\Delta H_{\text{reaction}}$ is the heat absorbed or released by a reaction.

$+\Delta H_{\text{rxn}}$ endo \uparrow $-\Delta H_{\text{rxn}}$ exo \downarrow

Thermochemical Equations

like reg. eqn... but w/ heat.
 $A + 30 \text{ KJ} \rightarrow 2D + C$

There are two general ways they're written. Both are acceptable:

Heat as a Product (exo) or Reactant (endo):



Heat shown as a change in Enthalpy (ΔH)

$-\Delta H$ means heat is lost (exothermic)

$+\Delta H$ means heat was absorbed (endothermic)



$\Delta H = -20 \text{ KJ}$

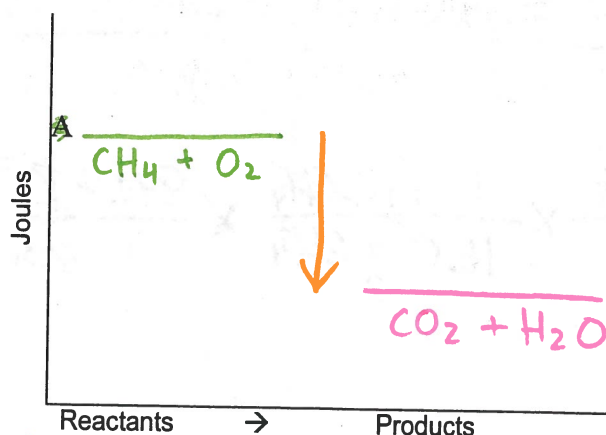
Example one: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + 890.4 \text{ kJ}$

so $\Delta H = -890.4 \text{ KJ}$

(heat of combustion)

always
exothermic

Potential Energy Diagram



$\Delta H = -890.4 \text{ kJ}$

The system released/produced heat so ΔH is negative/positive.

The reaction is endothermic/exothermic?

Law of Conservation of Energy:

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$E=mc^2$



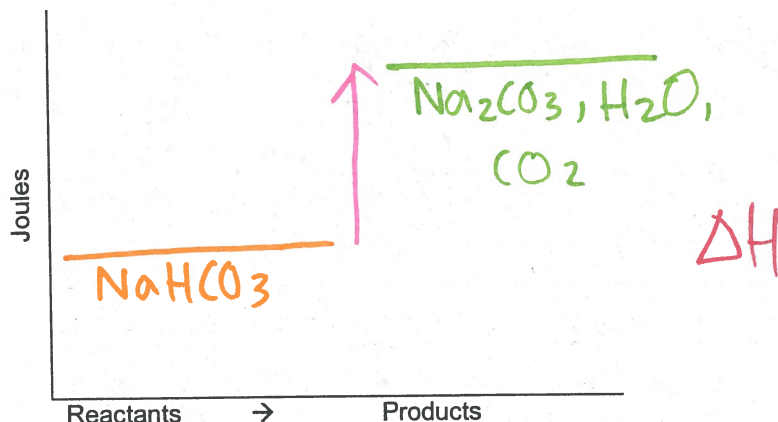
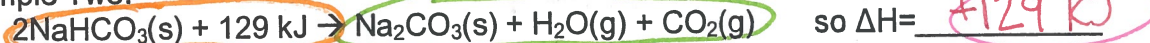
$$W = F \cdot d$$

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Energy = WORK + HEAT + Potential Energy

Gasoline = Distance + Heat + CO₂, H₂O

Example Two:



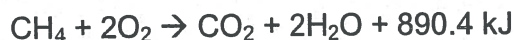
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5/4

The system absorbed heat so ΔH_{rxn} is negative/positive.

C. Heat of Reaction Calculations (Thermal Stoichiometry)

Example 1:

How many kilojoules of energy are produced by burning 821 grams of methane with excess oxygen? (Ans = 45600 J)



$$\frac{821 \text{ g CH}_4}{1} \times \frac{1 \text{ mol CH}_4}{16.0 \text{ g CH}_4} \times \frac{890.4 \text{ kJ}}{1 \text{ mol CH}_4} = \boxed{} \text{ kJ}$$

0/0

How many liters of oxygen would be consumed at STP to produce 122 kJ of heat in the below reaction? (Ans = 6.14 L)



$$\frac{122 \text{ kJ}}{1} \times \frac{2 \text{ mol O}_2}{890.4 \text{ kJ}} \times \frac{22.4 \text{ L O}_2}{1 \text{ mol O}_2} = \boxed{} \text{ L O}_2$$

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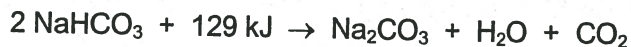
Solving Heat of Reaction Problems

Treat heat (in J or kJ) the same as any reactant or product in a chemical equation

1. How much heat is produced by the reaction of 25.7 g of CaO in the equation below? (Ans = 29.9 kJ)

$$\text{CaO (s)} + \text{H}_2\text{O (l)} \rightarrow \text{Ca(OH)}_2 \text{ (s)} + 65.2 \text{ kJ}$$

2. How many grams of NaHCO_3 are needed to react completely when 980 kJ of heat are used in the equation below? (Ans = 1300 g)



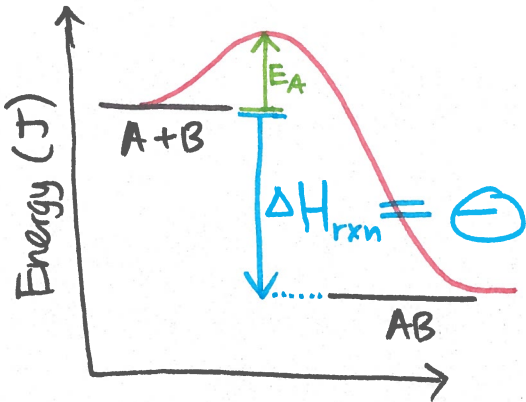
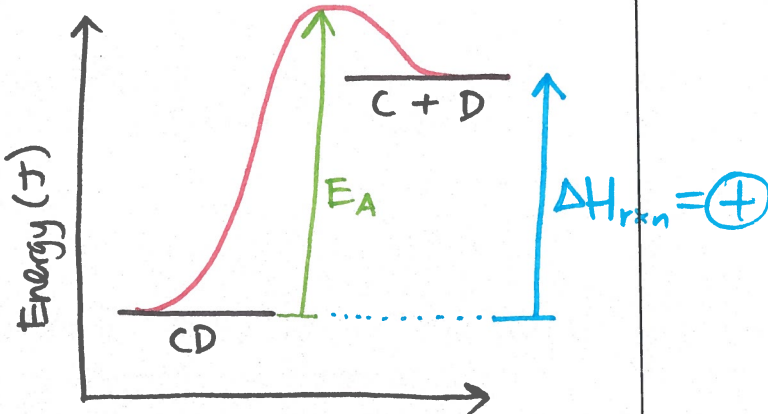
3. Using the same equation, how many kJ of heat must be used to produce 55.7 liters of CO_2 at STP? (Ans = 321 kJ)



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Comparing Endothermic and Exothermic Reactions

Exothermic Reactions	Endothermic Reactions
The reaction <u>loses</u> heat	The reaction <u>gains</u> heat
Example of exothermic reaction $A + B \rightarrow AB + \text{heat}$	Example of endothermic reaction $CD + \text{heat} \rightarrow C + D$
Potential Energy vs. Reaction Path for Exothermic Reactions 	Potential Energy vs. Reaction Path for Endothermic Reactions 
An exothermic reaction has a net <u>⊖</u> ΔH	An endothermic reaction has a net <u>⊕</u> ΔH
In exothermic reactions, the product's energy is <u>lower</u> than the reactant's energy	In endothermic reactions, the product's energy is <u>higher</u> than the reactant's energy

Word bank: positive, releases, negative, absorbs, lower, higher

E_A = Activation Energy
 ΔH = amt of energy lost/gained



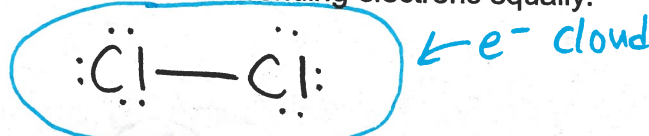
Chapter 16: Polar Bonds and Polar Molecules and Intermolecular Forces

Electronegativity: Tendency of an atom to "hog" or attract e^-

A. Non-Polar and Polar Covalent Bonds

1) non-polar covalent bonds: atoms share bonding electrons equally.

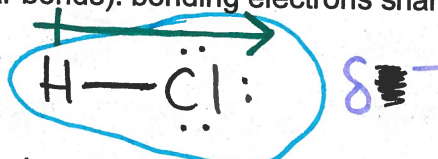
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Example: Cl_2
 $14e^-$



2) polar covalent bonds (polar bonds): bonding electrons shared unequally.



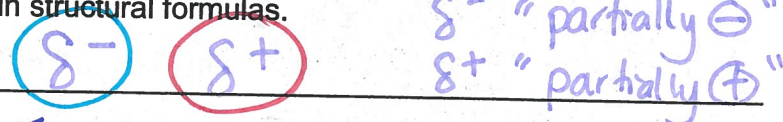
Example: HCl
 $8e^-$



3) Electron sharing based on electronegativity differences.

- more electronegative atom attracts the electrons more closely and acquires a slight negative charge.
- less electronegative atom then acquires a slight positive charge.
- unequal sharing creates "polarized" bonds with opposite charges.
- Two ways to show polarity in structural formulas.

lower case greek deltas:



slashed arrows:



The type of bond depends on electronegativity differences between the atoms

Electronegativity Difference	Guideline: Type of Bond	Example (electronegativity difference)
0.0 - 0.4	Non-polar Covalent	C-H in CH_4 $2.5 - 2.1 = 0.4$
0.4 - 2.0	Polar Covalent	HF $4.0 - 2.1 = 1.9$
>2.0	Ionic	NaCl $3.0 - 0.9 = 2.1$

Selected Electronegativity Values

H	Li	Be	B	C	N	O	F
2.1	1.0	1.5	2.0	2.5	3.0	3.5	4.0
	Na	Mg	Al	Si	P	S	Cl
	0.9	1.2	1.5	1.8	2.1	2.5	3.0
	K	Ca	Ga	Ge	As	Se	Br
	0.8	1.0	1.6	1.8	2.0	2.4	2.8

Metals:
Low EN

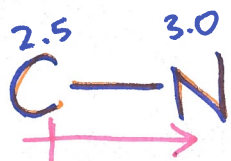
Non-metals... High EN

Is the bond polar, non-polar or ionic?

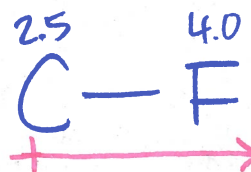
C=O in CO_2 ?	$ 2.5 - 3.5 = 1.0$	PC
Si-H in SiH_4	$ 1.8 - 2.1 = 0.3$	NPC
C-F in CF_4	$ 2.5 - 4.0 = 1.5$	PC
N-Cl in NCl_3	$ 3.0 - 3.0 = 0$	NPC
KCl	$ 3.0 - 0.8 = 2.2$	I

Answer this question without looking at the table on the previous page.

Which bond is most polar; C-N or C-F?



VS



B: POLAR MOLECULES

Polar Molecules: One end of the molecule is slightly negative, and one end is slightly positive.

δ^-

δ^+


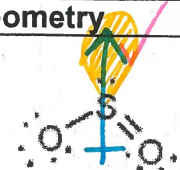
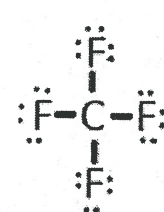
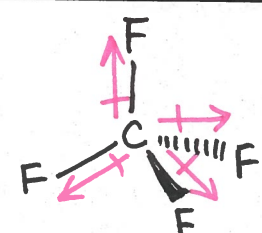
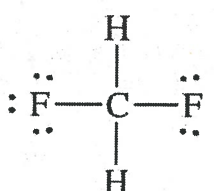
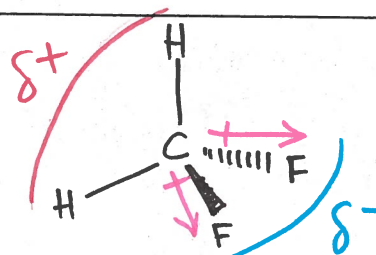
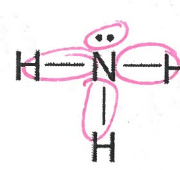
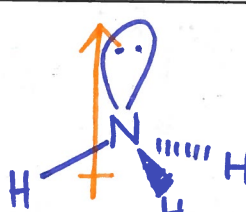
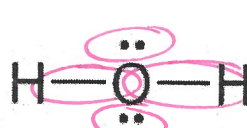
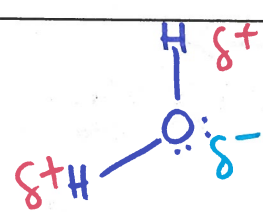
Dipole: a molecule with two poles (one negative, one positive or δ^- , δ^+)

What makes a molecule polar?	A molecule is polar if the electrons are pulled to one side of the molecule. The molecule is lopsided (assymetrical).
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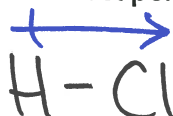
Determining if a molecule is polar.

1. Draw the Lewis structure
2. Determine the molecular geometry
3. Look for lone pairs on central atom (automatically polar)
4. Are there polar bonds?
5. If yes, are the polar bonds unsymmetrical in 3-D around the molecule's center?

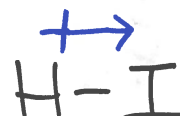
Molecule	Lewis Structure and Geometry	Polar or Non Polar
CO		Polar
CO ₂	<p>★ Has 2 polar bonds... BUT they're opposite / equal</p>	Non-polar
SO ₃		Non-polar

Molecule	Lewis Structure and Geometry	Polar or Non Polar
SO ₂	 	Polar
CF ₄	 	Non-polar
CH ₂ F ₂	 	polar
NH ₃	 	polar
H ₂ O	 	polar

Which molecule is most polar: HCl or HI?



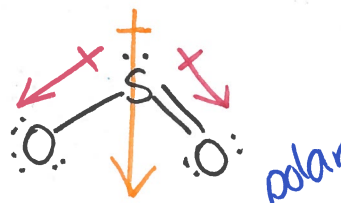
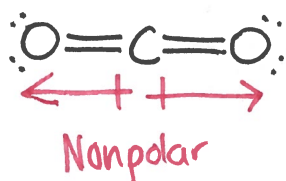
vs



Which molecule is most polar; CO₂ or SO₂?

CO₂ = 16ve

SO₂ = 18ve



between

C: INTERMOLECULAR FORCES

Intermolecular Forces are attractions between molecules due to three forces

1. **London Dispersion Forces** (weakest) temporary attractions between molecules due to temporary dipoles caused by shifting electron clouds. Dispersion forces are greater in more massive molecules with larger electron "clouds". **All molecules have LDFs, and larger atoms/molecules have stronger LDF because they have more electrons.**

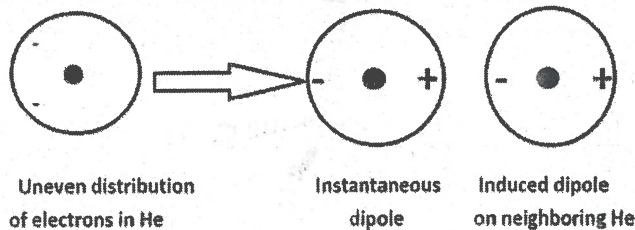
for all the NP stuff

Non-polar

CO₂ (44 g/mol)

vs

CS₂ (76 g/mol)



more e⁻ in their clouds

★ All things can do LDF

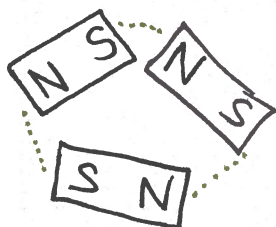
★★ BIGGER = STRONGER

2. **Dipole-dipole attraction:** polar molecules are attracted to each other (like magnets).

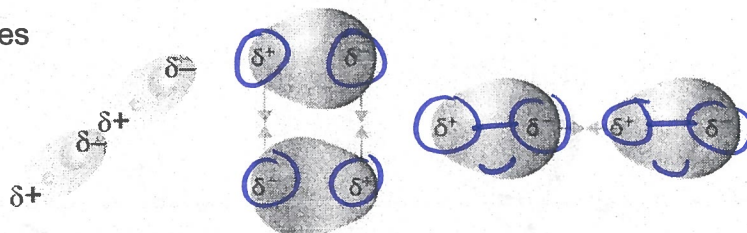
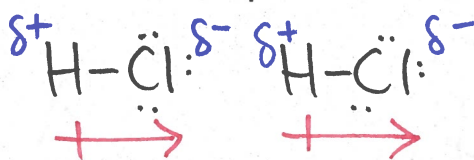
DP-DP

The positive dipole of one molecule is attracted to the negative dipole of another.

These occur when molecules have uneven distribution of electrons due to electronegative elements (like N, O, Cl, F, Br, etc.).



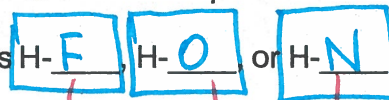
Example: HCl molecules



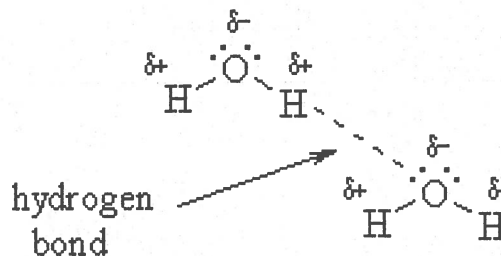
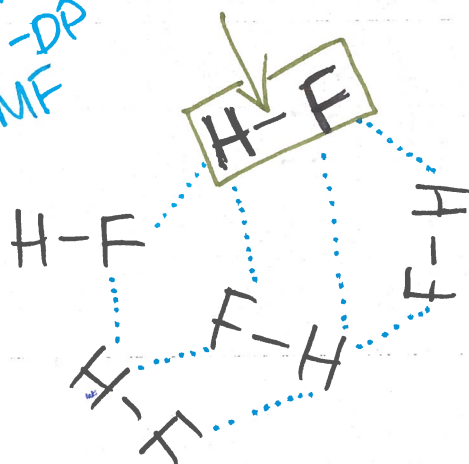
★ must have polar molecules!!!

3. **Intermolecular Hydrogen bonding:** hydrogen that is covalently bonded to a very electronegative atom is also weakly bonded to the unshared pair of another electronegative atom.

H-Bonding only occurs when an molecule has H-F, H-O, or H-N bonds. Why these three atoms specifically?



H-bonds are just super-strong DP-DP IMF



★ occurs between "H" in ONE molecule, and N,O,F in ANOTHER!

more/stronger IMF \rightarrow harder to separate.
 5 of 6
 melting/boiling

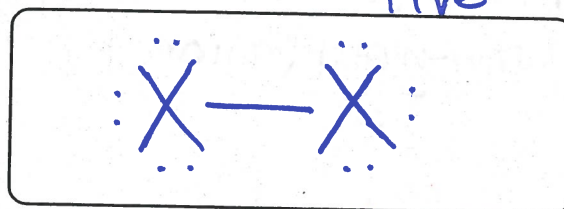
4. Intermolecular Forces and Molecular Physical Properties

As intermolecular forces (IMF) increases (meaning gets stronger), the melting and boiling points increase because more kinetic energy is needed to overcome the IMFs between molecules.

An Analysis of the Halogens and their Physical States:

The Halogens are Diatomic:
 (Lewis diagram \rightarrow)

"X" is any generic halogen \rightarrow X_2



all pure X_2 (halogens) are NON-POLAR

Are pure diatomic halogen elements polar? Explain & draw dipole vectors & partial charges.

NO!!!

What is the only IMF that all non-polar things can do?

\leftarrow all non-polar

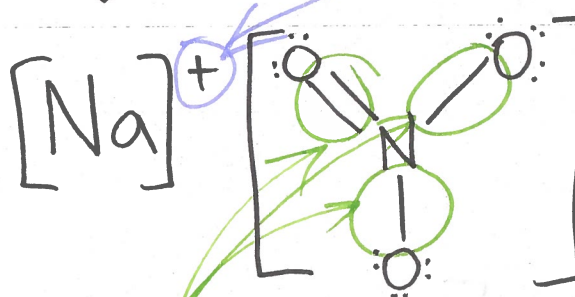
	Halogen (X)	<u>Molar Mass (X_2)</u>	State at Room Temp.
BIGGER \downarrow	F	$F_2 = 38.0 \text{ g/mol}$	<u>gas</u>
	Cl	$Cl_2 = 70.9 \text{ g/mol}$	<u>gas</u>
	Br	$Br_2 = 159.8 \text{ g/mol}$	<u>liquid</u>
	I	$I_2 = 253.8 \text{ g/mol}$	<u>solid</u>

★ LDF increases as size increases

Comparing Ionic and Covalent Compounds

Characteristic	Ionic Compound	Molecular Compound
bond formation	TRANSFER of e^- $\text{Na} \cdot \curvearrowright \cdot \ddot{\text{Cl}}: \rightarrow [\text{Na}^+][\ddot{\text{Cl}}:]^-$	SHARE e^- $:\text{C} \equiv \text{O}:$
Types of elements in compound	<ul style="list-style-type: none"> • Metal (cation) • Non-metal (anion) 	Non-metals
Physical state at <u>25°C</u>	SOLID (powders)	<u>ANY!</u>
Melting point	HIGH AF	relatively low (but variable)
Electrical Conductivity in aqueous solution	\rightarrow Conduct electricity (usually) <u>electrolytes</u>	<u>Do NOT</u> conduct electricity (insulation)

Sodium nitrate



covalent.

Sometimes, something can be both.

ionic!

motion of molecules

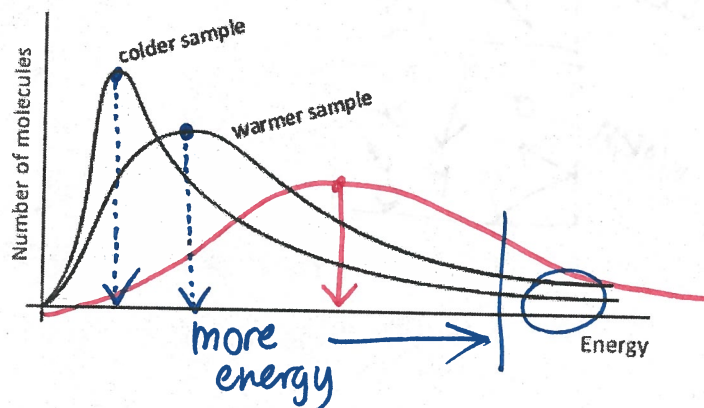
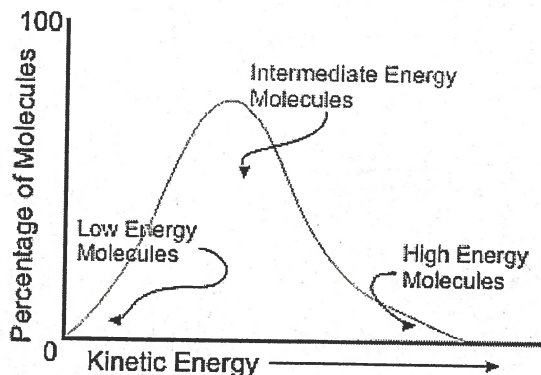
Chapter 10 Kinetic Theory, IMFs, and Phase Changes

Kinetic Molecular Theory: The tiny particles in all forms of matter are in constant motion.

S → L → G
fastest

I Kinetic Energy and Kelvin temperature scale

A) Temperature measures average kinetic energy



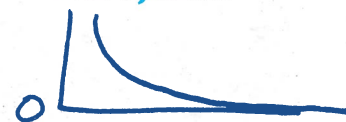
$$KE = \frac{1}{2}mv^2$$

B) Gas particle's kinetic energy increases as velocity increases

C) Kelvin Temperature scale is absolute (0 → UP)

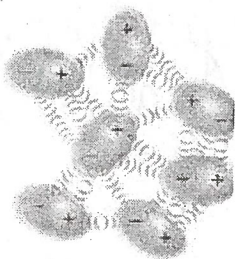
0/10 273 K = 0 °C

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$



II. Kinetic Energy and Liquids

Intermolecular forces (between molecules) hold particles together in solid or liquid phases.



Kinetic energy keeps the molecules moving but not with enough energy to overcome the IMFs.

Evaporation, Vapor Pressure and Temperature

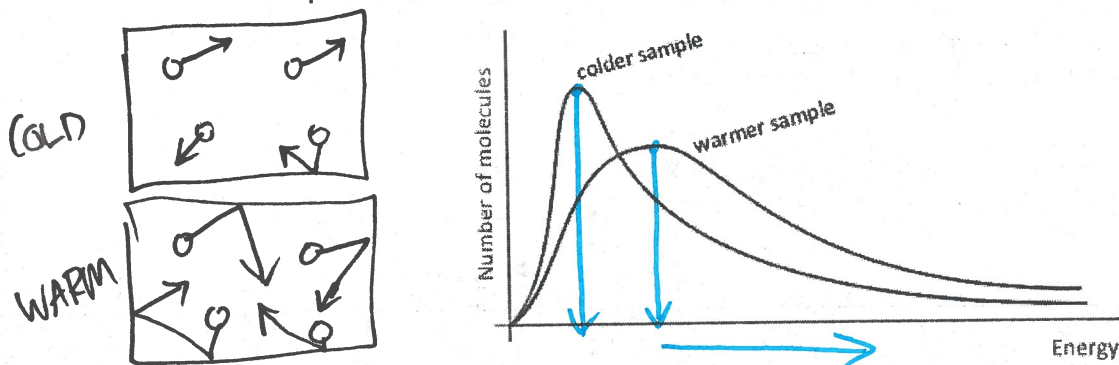
Evaporation:

Phase change from L → G
at the non-boiling point.



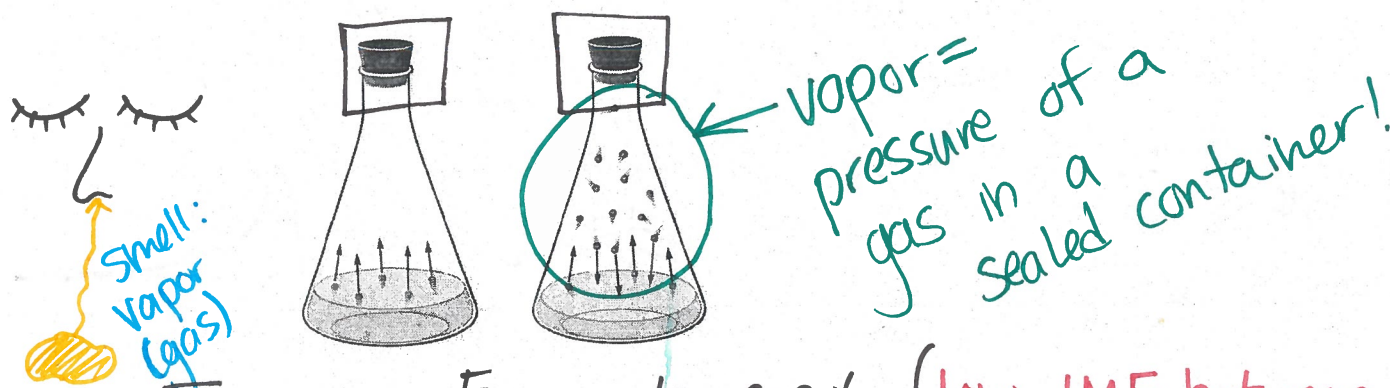
endothermic

- Particles with enough kinetic energy to overcome intermolecular forces escape into gas phase
- Evaporation rate increases as temperature increases



Evaporation in a closed container produces vapor pressure.

Increasing temperature increases vapor pressure over a liquid until a dynamic equilibrium is reached.



Volatile liquids

Evaporate easily (low IMF between particles)

Nonvolatile liquids

Doesn't easily evaporate! (high IMF)

have higher VP →

Volatile
(Low IMF)

Ⓛ →

Non-volatile
(high IMF)

Ⓛ →

Less vapor!

IMF = "self-attractiveness"

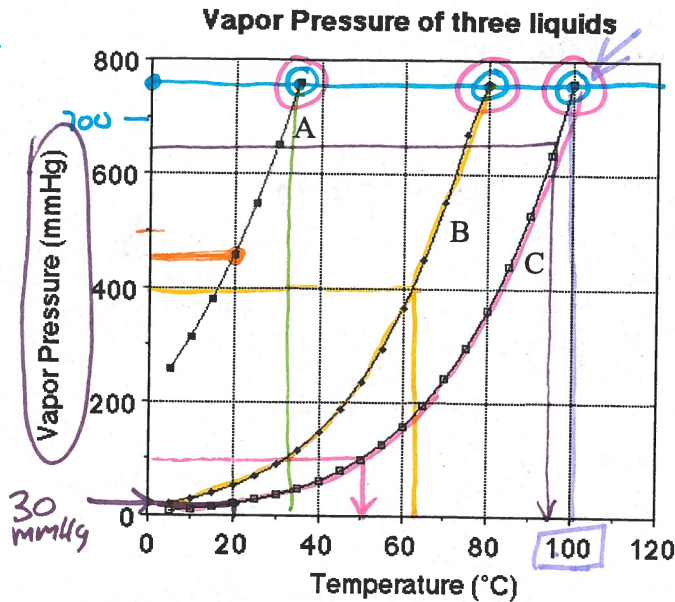
Boiling Point of a Liquid (open container)

Boiling Point: Temp @ which ALL particles have sufficient kinetic energy to leave Ⓛ → Ⓜ

Sea Level = 1 atm = 760 mmHg (Normal)

A) Boiling point changes as external pressure changes/

⊙ = "normal" pressure (Standard) (= Boiling Point)



What is the vapor pressure of liquid A at 20°C?_

450 mmHg

Which liquid represents water?

"C"

What is the boiling point of B when the external pressure is 400 mmHg?

63°C

Which liquid is most volatile?

A

lowest BP

Which liquid has the strongest intermolecular forces?

C

least volatile (highest B)

How hot does water need to be to boil at 100 mmHg?

50°C

What will be the boiling point of water on Pike's Peak (elevation = 14,000 ft, atmospheric pressure = 640 mmHg)?

~93°C

↑ P decrease

III. Kinetic Energy and Solids

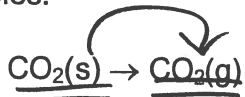
melting point:

Temp @ which ALL particles can easily go from (S) → (L)

sublimation:

the (S) → (G) phase (skips Liq.)

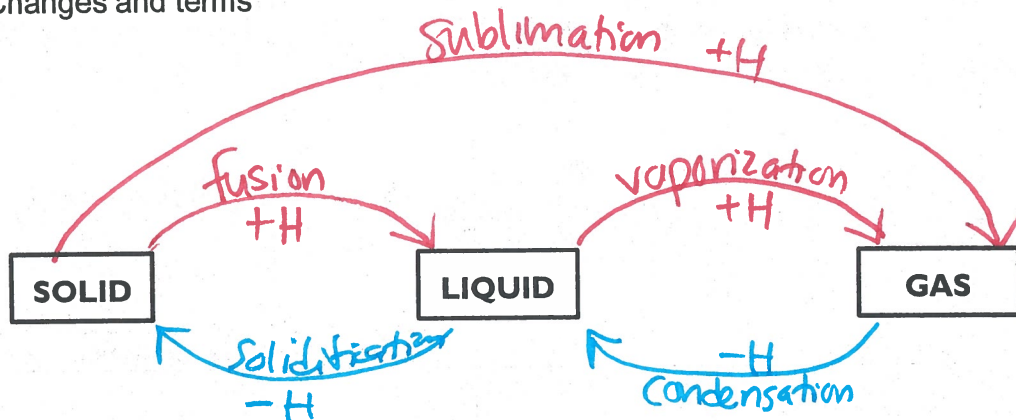
Examples:



} Sublimation

IV. Phase Changes and Phase Diagrams

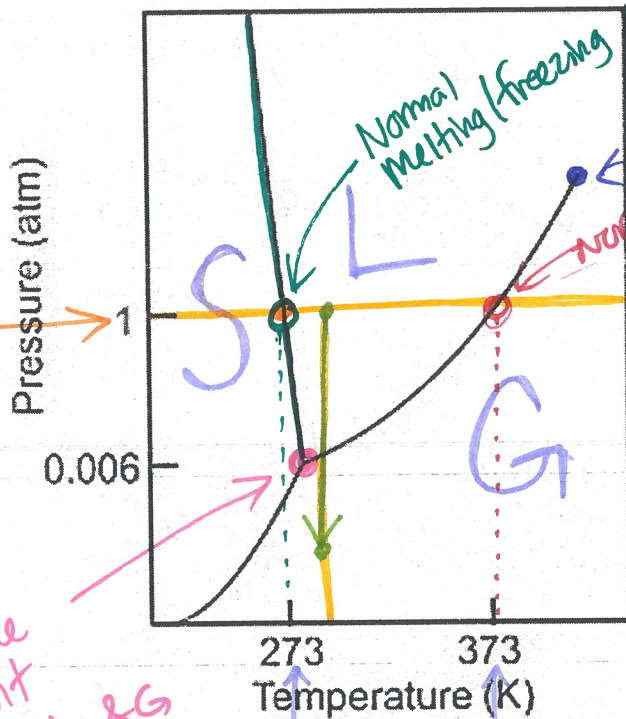
Phase Changes and terms



Phase Diagram and Triple Point

WATER

Normal/Standard



tells you the state (phase) of ANY substance @ ANY P & T

Critical point: L & G are indistinguishable

triple point (where S, L & G are all together)

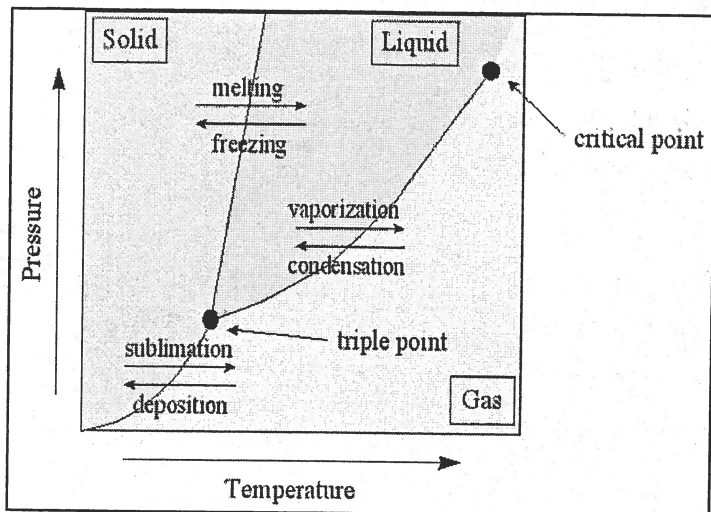
273 0°C
373 100°C

constant temp = isothermic

Phase Diagrams

A **phase diagram** is a graphical way to depict the effects of **pressure** and **temperature** on the phase of a substance:

The **CURVES** indicate the conditions of **temperature** and **pressure** under which "equilibrium" between different phases of a substance can exist. **BOTH** phases exist on these lines:



Melting/Freezing: Any point on this line (pressure & temperature) the substance is both **solid** and **liquid**

Sublimation/Deposition: Any point on this line (pressure & temperature) the substance is both **solid** and **gas**

Vaporization/Condensation: Any point on this line (pressure & temperature) the substance is both **liquid** and **gas**

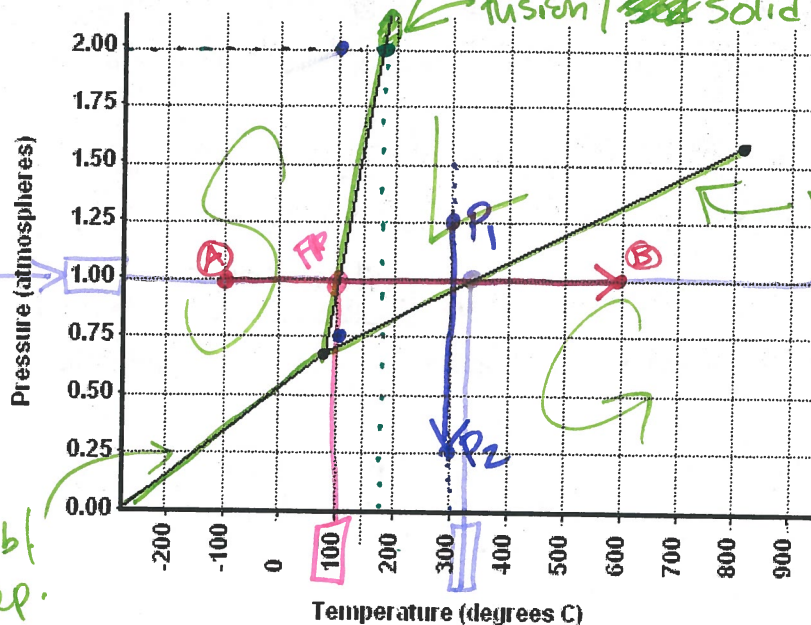
NOTE: the vapor pressure curve ends at the **critical point**, the temperature above which the gas cannot be liquefied no matter how much pressure is applied (the kinetic energy simply is too great for attractive forces to overcome). Any substance beyond this critical point is called a **supercritical fluid** – indistinguishable between gas or liquid (neither one)

The **TRIPLE POINT** is the condition of temperature and pressure where ALL THREE phases exist in equilibrium (solid, liquid, gas)

Remember that pressure can be expressed in many units where: **1 atm = 101.3 kpa = 760 mmHg = 760 torr = 14.7 psi**

Refer to the phase diagram below when answering the questions.

NOTE: "Normal" refers to STP – Standard Temperature and Pressure.



A → B
• Fus, then vap.
B → A
• cond, then solid.

Not water.

- 1) What are the values for temperature and pressure at STP? T= _____, P= _____
- 2) What is the **normal** freezing point of this substance? 100°C
- 3) What is the **normal** boiling point of this substance? 330°C

- 4) What is the normal melting point of this substance? _____
- 5) What is the phase (s, l, g) of a substance at **2.0 atm** and **100 °C**? S
- 6) What is the phase (s, l, g) of a substance at **0.75 atm** and **100 °C**? L
- 7) What is the phase (s, l, g) of a substance at **0.5 atm** and **100 °C**? _____
- 8) What is the phase (s, l, g) of a substance at **1.5 atm** and **50 °C**? _____
- 9) What is the phase (s, l, g) of a substance at **1.5 atm** and **200 °C**? _____
- 10) What is the phase (s, l, g) of a substance at **1.5 atm** and **800 °C**? _____
- 11) What is the condition of the **triple point** of this substance? $T =$ _____, $P =$ _____

12) If a quantity of this substance was at an initial pressure of 1.25 atm and a temperature of 300° C was lowered to a pressure of 0.25 atm, what phase transition(s) would occur? Vap

13) If a quantity of this substance was at an initial pressure of 1.25 atm and a temperature of 0° C was lowered to a pressure of 0.25 atm, what phase transition(s) would occur? _____

14) If a quantity of this substance was at an initial pressure of 1.0 atm and a temperature of 200° C was lowered to a temperature of -200° C, what phase transition(s) would occur? _____

15) If a quantity of this substance was at an initial pressure of 0.5 atm and a temperature of 200° C was lowered to a temperature of -200° C, what phase transition(s) would occur? _____

16) If this substance was at a pressure of 2.0 atm, at what temperature would it **melt**? _____

17) If this substance was at a pressure of 2.0 atm, at what temperature would it **boil**? _____

18) If this substance was at a pressure of 0.75 atm, at what temperature would it **melt**? _____

19) If this substance was at a pressure of 0.75 atm, at what temperature would it **boil**? _____

20) At what temperature do the gas and liquid phases become indistinguishable from each other? _____

21) At what pressure would it be possible to find this substance in the gas, liquid, and solid phase? _____

22) If I had a quantity of this substance at a pressure of 1.00 atm and a temperature of -100° C, what phase change(s) would occur if I **increased the temperature to 600° C**? At what temperature(s) would they occur? (**NOTE: multiple answers needed for this question**)

23) If I had a quantity of this substance at a pressure of 2.00 atm and a temperature of -150° C, what phase change(s) would occur if I **decreased the pressure to 0.25 atm**? At what pressure(s) would they occur? (**NOTE: multiple answers needed for this question**)

★ vocab
exert

Chapter 12 Dalton's Law and Graham's Law

Mixtures of Gases: Dalton's Law of Partial Pressure

Partial Pressure: The pressure due to a single gas in a mixture of gases.

Σ

Verbally: At constant pressure and temperature, the total pressure exerted by a mixture of gases is equal to the sum of the partial pressures of the component gases.

Math Equation: $P_T = P_A + P_B + P_C \dots$

Example 1: What is the partial pressure of oxygen in air at STP (101.3 kPa) if $P_{N_2} = 79.1$ kPa, $P_{CO_2} = 0.040$ kPa and $P_{others} = 0.94$ kPa?

$$P_T = P_{N_2} + P_{O_2} + P_{CO_2} + P_{others}$$

$$101.3 = 79.1 + P_{O_2} + 0.040 + 0.94$$

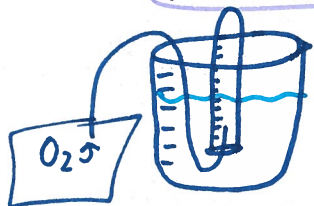
$$\therefore P_{O_2} = 21.22 \text{ kPa}$$

Example 2: A sample of oxygen gas is collected over water at 20.°C. The vapor pressure of water at 20.°C is 15 mm Hg. If the total pressure is 420 mm Hg, what is the partial pressure of the oxygen?

$$P_T = P_{O_2} + P_{H_2O(g)}$$

$$420 \text{ mmHg} = P_{O_2} + 15 \text{ mmHg}$$

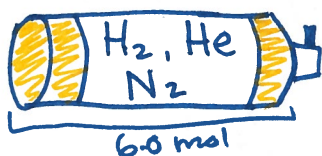
$$\therefore P_{O_2} = 405 \text{ mmHg}$$



Variants of Dalton's Law

$$\frac{\text{mole}_A}{\text{mole}_{\text{Total}}} = \frac{P_A}{P_{\text{Total}}} = \frac{\%V_A}{100 \%V_{\text{Total}}}$$

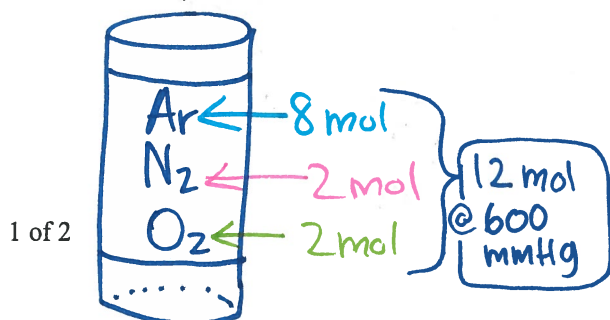
Example 1: A tank contains 6.0 moles of a mixture of hydrogen, helium, and nitrogen at 102 kPa. If there are 2.0 moles of hydrogen in the tank, what is the partial pressure of hydrogen?



$$\frac{\text{mol } A}{\text{mol Total}} = \frac{P_A}{P_T} \rightarrow \frac{2.0 \text{ mol}}{6.0 \text{ mol}} = \frac{P_{H_2}}{102 \text{ kPa}}$$

$$\therefore P_{H_2} = 34 \text{ kPa}$$

Example 2: A gas cylinder contains 8.0 moles of argon, 2.0 moles of nitrogen, and 2.0 moles of oxygen at 600. mmHg. What is the partial pressure of nitrogen in the cylinder? (Ans = 1.0×10^2 mmHg)



$$\frac{\text{mol } N_2}{\text{mol, total}} = \frac{P_{N_2}}{P_{\text{total}}}$$

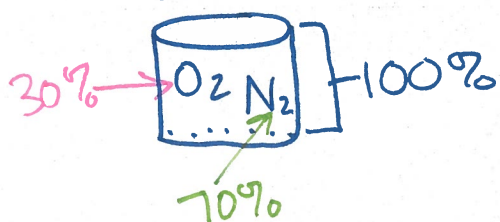
Example 3: A mixture of gases with a pressure of 950 mm Hg contains 20% hydrogen and 80% neon by volume. What is the partial pressure of neon gas in the mixture? (Ans = 760 mmHg)



$$\frac{80\% \text{ Ne}}{100\%} = \frac{P_{\text{Ne}}}{950 \text{ mmHg}}$$

$$P_{\text{Ne}} = 760 \text{ mmHg}$$

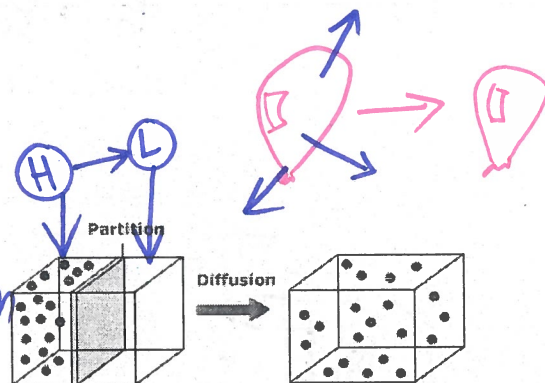
Example 4: In a mixture of oxygen and nitrogen gas, 70.0 percent of the total gas pressure is exerted by the nitrogen. If the total pressure is 150 kPa, what pressure does oxygen exert? (Ans = 45 kPa)



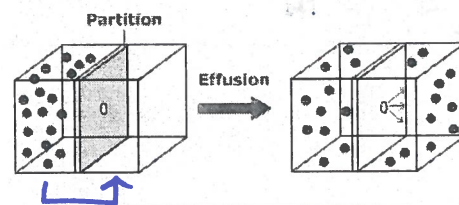
$$(0.30)(150) = 45 \text{ kPa}$$

Graham's Law of Effusion

Diffusion Gas particles move from HIGH to LOW concentration



Effusion Gas particles in motion thru a hole ★ LIGHT gases move FASTEST



★ ★ Graham's Law Verbally: Gas molecules with the lightest mass travel fastest.

Graham's Law Equation:

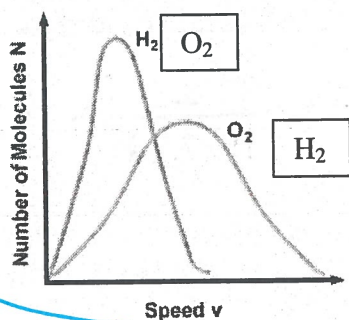
$$KE = \frac{1}{2} m v^2$$

← molar mass

← velocity (rate)

← mass

← inverse



N_2 (FAST) vs F_2 (SLOW)

H_2S (FAST) vs Cl_2 (SLOW)

Which gas will escape slowest from a tiny hole in a balloon; He, C_3H_8 , or Xe?

Ar vs O_2 (FAST)
SLOW

propane C_3H_8 44 g/mol vs Xe 131 g/mol

Partial Pressures Practice (Lee)

NDB→

- 1) Argon, oxygen, and nitrogen are mixed together and pressurized in a tank. P_{O_2} is 155 kPa, P_{N_2} is 415 kPa, and P_{Ar} is 285 kPa.

a. What is the total pressure of the gas mixture?

$$P_T = P_{O_2} + P_{N_2} + P_{Ar}$$

$$\therefore P_T = 855 \text{ kPa}$$

b. What is the % O_2 ?

$$\frac{155 \text{ kPa}}{855 \text{ kPa}} = \frac{18\%}{100\%}$$

$$18\% \text{ } O_2$$

c. What is the % N_2 ?

64%



- 2) The total pressure in a tank is 1200.0 mmHg. Krypton has a partial pressure (P_{Kr}) of 680.0 mmHg, and methane has a partial pressure that is half that of krypton. The third gas in the container is chlorine.

a. Calculate the partial pressures of each gas with units:

$$1200 = P_{Kr} + \frac{1}{2}P_{Kr} + P_{Cl}$$

$$P_{CH_4} = 340 \text{ mmHg}$$

$$P_{Kr} = 680 \text{ mmHg}$$

$$P_{Cl_2} = 180 \text{ mmHg}$$

b. Determine the % volume of each gas:

$$\% Cl_2 = 15\%$$

$$\rightarrow \frac{180 \text{ mmHg}}{1200 \text{ mmHg}} = \frac{15\%}{100\%}$$

$$\% Kr = \underline{\hspace{2cm}}$$

$$\% CH_4 = \underline{\hspace{2cm}}$$

- 3) A highly pressurized (46.2 atm) mixture of gases contains a total of 330. moles of gases. The technologist who created the mixture added fluorine, chlorine, helium, and hydrogen. Twice as many moles of helium were present than fluorine. Three times as many moles of chlorine were added than fluorine, and the half as many moles of hydrogen were added as fluorine.

a. Set up an algebraic equation using variables that expresses the situation above.

$$330 = \underbrace{2x}_{He} + \underbrace{3x}_{Cl_2} + \underbrace{0.5x}_{H_2} + \underbrace{x}_F$$

$$330 = 6.5x$$

b. Determine the number of moles of each gas.

$$\bullet x = \text{mol F} = 50.77 \text{ mol}$$

$$\bullet 101.54 = \text{mol He}$$

$$\bullet 152.31 \text{ mol} \Rightarrow Cl_2$$

$$\bullet 25.39 \text{ mol } H_2$$

c. Determine the partial pressures of each gas. Use proper notation. (Example: To express the partial pressure of X, write P_X)

$$\frac{\text{mol } F_2}{\text{total mol}} = \frac{P_{F_2}}{P_T} \rightarrow \frac{50.77}{330} = \frac{P_{F_2}}{46.2} \therefore P_{F_2} = 7.1 \text{ atm}$$

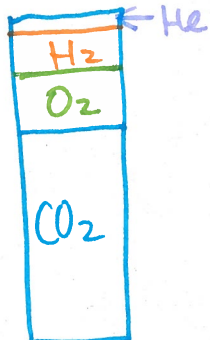
101.3 kPa

Earth

202.6 kPa

Planet

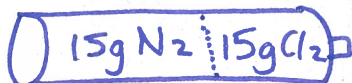
- 4) A certain planet was discovered whose atmospheric composition was 70% CO₂, 20% O₂, and 7% H₂, and 3% He. The atmospheric air pressure on the planet was determined to be twice that of normal atmospheric pressure on Earth. Calculate the partial pressure of each gas in kPa.



$$\frac{70\% \text{ CO}_2}{100\%} = \frac{P_{\text{CO}_2}}{202.6 \text{ kPa}}$$

- 5) 15.0 g of nitrogen gas, and 15.0 grams of chlorine gas were added to a container that exists at STP.

$$1.0 \text{ atm} = 101.3 \text{ kPa}$$



- a. What is the number of moles of nitrogen in the container?

$$\frac{15 \text{ g N}_2}{1} \times \frac{1 \text{ mol}}{28.0 \text{ g N}_2}$$

- b. What is the number of moles of chlorine in the container?

$$\frac{15 \text{ g Cl}_2}{1} \times \frac{1 \text{ mol}}{70.9 \text{ g Cl}_2}$$

- c. How many total moles of gas are in the container?

$$0.536 \text{ mol N}_2$$

+

$$0.2116 \text{ mol Cl}_2$$

=

$$0.747 \text{ mol, total}$$

- d. What is the P_{N_2} in kPa?

$$\frac{\text{mol N}_2}{\text{total mol}} = \frac{P_{\text{N}_2}}{\text{total P}} \rightarrow \frac{0.536}{0.747} = \frac{P_{\text{N}_2}}{101.3}$$

$$P_{\text{N}_2} = 72.7 \text{ kPa}$$

- e. What is the P_{Cl_2} in kPa?

$$\frac{\text{mol Cl}_2}{\text{total}} = \frac{P_{\text{Cl}_2}}{\text{total P}}$$

$$P_{\text{Cl}_2} = 29.6 \text{ kPa}$$

Temp \Rightarrow energy

Gas A vs. Gas B

Same temp... Same ^{Kinetic} energy

Graham's Law of Effusion (Applied Quantitative Practice)

Graham's Law can be derived from the equation for kinetic energy

$$(K = \frac{1}{2}mv^2)$$

Derivation: ① $KE_A = KE_B$ ^(because) \therefore they're the same temp!!!

$$\textcircled{2} KE_A = \frac{1}{2}m_A v_A^2 \quad \text{and} \quad KE_B = \frac{1}{2}m_B v_B^2$$

equal!!! (same temp.)

$$\textcircled{3} \cancel{\frac{1}{2}}m_A v_A^2 = \cancel{\frac{1}{2}}m_B v_B^2 \quad \text{"1/2" cancels b/c it's a constant.} \quad \textcircled{4} m_A v_A^2 = m_B v_B^2$$

$$\textcircled{5} \frac{v_A^2}{v_B^2} = \frac{m_B}{m_A}$$

$$\textcircled{6} \frac{v_A}{v_B} = \sqrt{\frac{m_B}{m_A}}$$

⑦
FINAL

$$\frac{\text{rate A}}{\text{rate B}} = \sqrt{\frac{MM_B}{MM_A}}$$

Graham's Law (final equation) starts that lighter gases move more quickly than heavier gases in an inverse square proportion. (According to kinetic theory, when the same amount of energy is available to different bodies with different masses, they will move at inverse square velocities relative to each other.) A fundamental assumption when using Graham's Law of Effusion is that the gases have the same amount of energy... so they're at the same temperature on the Kelvin scale. Their difference in average velocity (aka: rate, speed...) is due to mass difference.

Analogy: A 60-kg girl eats 2 eggs and 3 slices of bacon. A 150 kg sumo wrestler eats the same thing. The girl will run faster because she's smaller, even though they had the same breakfast (energy).

Graham's Law of Effusion

Inverse-Square relationship

$$\frac{r_A}{r_B} = \sqrt{\frac{MM_B}{MM_A}}$$

Tips for Use:

Heavier gas as "B" (or "2")

Worked Example 1: The average velocity of oxygen (O_2) molecules will be faster than the average velocity of chlorine (Cl_2) molecules because oxygen has a smaller molar mass. What is the relative rate (speed) of oxygen molecules to chlorine molecules if they are at the same temperature? (i.e., how many times faster will molecules of oxygen move?)

$$\frac{\text{rate A}}{\text{rate B}} = \sqrt{\frac{MM_B}{MM_A}} = \sqrt{\frac{70.9}{32}}$$

O_2 is 1.49 times faster

Part 1: Determine which gas will effuse or move fastest, and determine how many times faster it moves.

- 1) H_2 vs. He $\leftarrow 2.02 \text{ g/mol}$ $\leftarrow 4.00 \text{ g/mol}$
- 2) O_2 vs. Ne $\leftarrow 32.0 \text{ g/mol}$ $\leftarrow 20.0 \text{ g/mol}$

3) CH_4 vs. NCl_3

4) Ammonia vs. Hydrogen Sulfide (H_2S)

5) Xenon vs. Argon
 131.3 39.9

Part 2: Determine which gas effuse or move slowest, and determine how many times slower it moves.

6) Oxygen vs. chlorine

7) Sulfur dioxide vs. methane

8) Laughing gas (dinitrogen monoxide) vs. carbon monoxide

9) Sulfur Hexafluoride vs. carbon tetrafluoride

10) Silane (SiH_4) vs. hydrogen

Part 3 (Advanced): Determining the molar mass of an unknown gas; or identifying the gas by calculating the molar mass from relative rates.

- 11) A sample of hydrogen gas effuse through a porous container 9 times faster than an unknown gas. Estimate the molar mass of the unknown gas. Would it be reasonable to assume this gas is silicon tetrafluoride? Explain and justify your answer using mathematics and complete sentences.

$$\frac{9}{1} = \sqrt{\frac{MM_B}{2.0}}$$

$$81 = \frac{MM_B}{2.0}$$

$$\therefore MM_B = 162 \text{ g/mol}$$

NO, it's not SiF_4 !

- 12) At a certain temperature, hydrogen molecules move at an average velocity of 1.84×10^3 m/s. Estimate the molar mass of a gas whose molecules have an average velocity of 311 m/s.

1840 m/s rate H_2

$$\frac{\text{rate } A_{H_2}}{\text{rate } B_{?}} = \sqrt{\frac{MM_B}{MM_A}} \rightarrow \frac{1840 \text{ m/s}}{311 \text{ m/s}} = \sqrt{\frac{MM_B}{2.0}}$$

$$35 = \sqrt{\frac{MM_B}{2}}$$

$$MM_B = 70.0 \text{ g/mol}$$

- 13) Nitrogen gas (N_2) effuses at a rate 2.17 faster than an unknown noble gas. Identify the noble gas

N_2 is faster... so it's lighter.

$$\frac{2.17}{1} = \sqrt{\frac{MM_B}{28.0}}$$

Xe

- 14) A sample of $Br_2(g)$ takes 10.0 min to effuse from one side of a room that is 86 feet long. How long would it take the same amount of $Ar(g)$ to effuse the same distance?

Adv

BONUS

- 15) Explain why carbon monoxide and nitrogen effuse at nearly the same rate. Use complete sentences and justify your answer.

Chapter 19 Part 1: Reaction Rates

http://www.wwnorton.com/college/chemistry/gilbert/index/site_map.htm

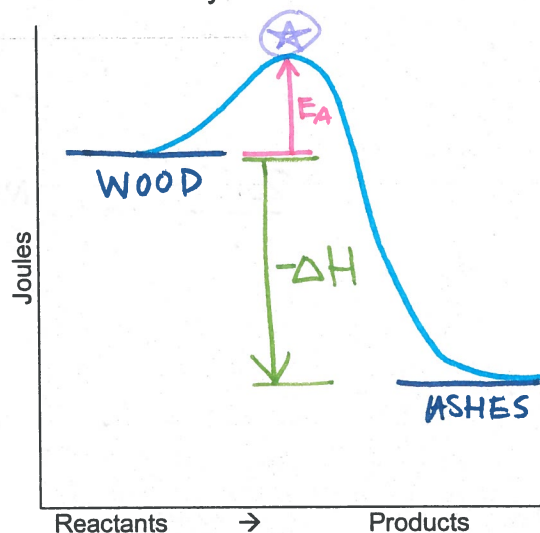
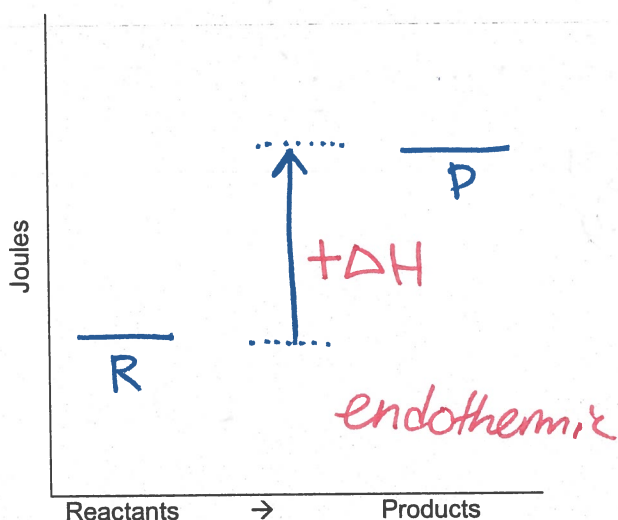
Reaction rates are measured as mol/time units.

A. Collision Theory: Atoms & molecules will react only when they have enough **energy** & correct **orientation**

Reaction Coordinate Diagram for Exothermic Process (forward direction)

Chapter 11: Thermochemistry

Collision Theory



Activation Energy: E_A = Activation energy... The amt of **(E)** needed to **START rxn.**

(★) Activation Complex: The "top of the hill" (unstable, temporary)

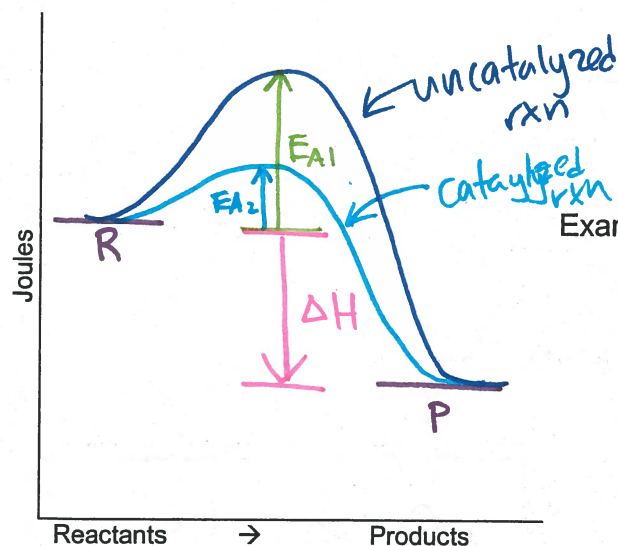
B. Factors Influencing Reaction Rate

1. Temperature \uparrow increase Temp..., speeds up rxn. (faster molecules)
2. Concentration \uparrow increase concentration (either MOLARITY or PRESSURE)
3. Particle Size
Smaller particles react faster.
4. Catalysts
speeds up rxn... by LOWERING act.en. (E_A)



Catalyst Effect on Reaction Path

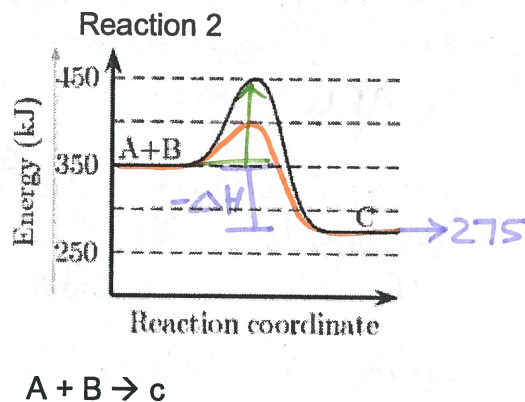
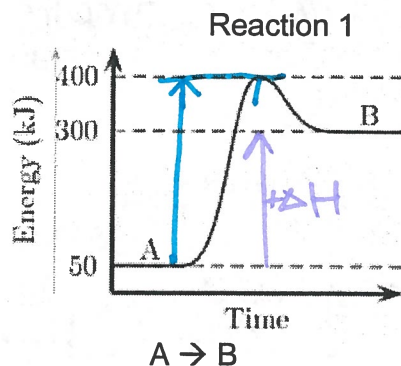
Why do catalysts increase reaction rate?



Examples: Catalytic Convertors

LATER
☺

Interpret the following potential energy diagrams



1. Which reaction is endothermic? ①

→ 2. What is the activation energy of Reaction 1? +350 KJ

3. What is the ΔH_{rxn} of reaction 1? +250 KJ

4. What is the activation energy of reaction 2? +100 KJ

5. What is the ΔH_{rxn} of reaction 2? -75 KJ

6. Sketch the effect of a catalyst on both reactions

7. Does a catalyst effect the ΔH_{rxn} ?

by a decrease of 50 KJ

Biologically (in terms of **biochemistry**), enzymes are catalysts.

How are they usually used? help BREAK stuff down or BUILD them up!

_____ is the process by which life is _____.

metabolism

anabolism

(meaning "to build up")

catabolism

(meaning "to break down")

Anabolic ("constructive") processes are when larger, more complex biomolecules are created from simpler ones.

Example #1: When your body is building muscle proteins using amino acids.

Example #2: **PHOTOSYNTHESIS**

Plants use H₂O & CO₂ to create glucose - C₆H₁₂O₆ which is a larger molecule.

Anabolism = Energy REQUIRED

Catabolic ("destructive") processes are when large, complicated molecules are broken down into smaller ones biomolecules.

Example #1: **RESPIRATION & GLYCOLYSIS**

Your body breaks down carbohydrates/ sugars to produce energy, CO₂ & H₂O.

Example #2: When you exercise, fats are broken down.

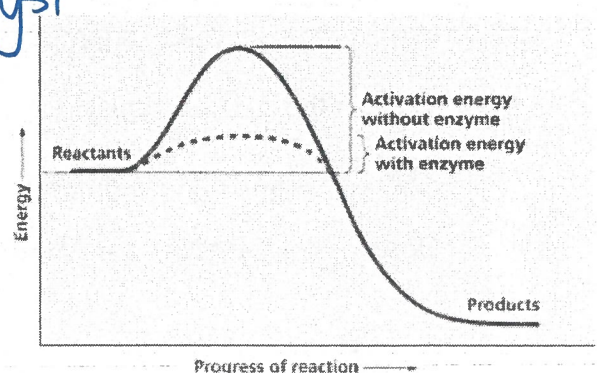
Catabolism = Energy RELEASED

Enzymes (which are biological catalyst) speed up reactions, but they are NOT consumed in the reaction.

They are proteins, which means they're made of amino acids.

Enzymes usually end in the suffix -ase.

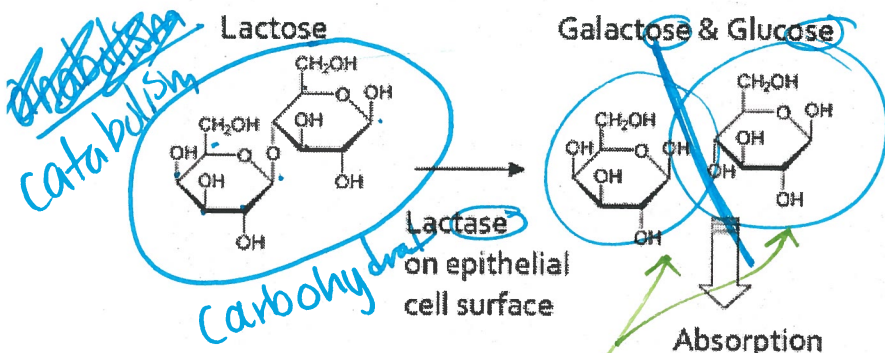
Carbohydrates often end in -ose.



Lactose Intolerance:

Lactose is a sugar that is found naturally in dairy products.

People who are lactose intolerant have a difficult time breaking down the lactose molecule.



What suffix do carbohydrates often have?

-ose

If you see dextrose, maltose, or sucrose on a food label, would you call them fat? Sugar? Or protein?

Sugar

Lactose is a 2-ring sugar. It must be broken by the enzyme lactase.

This will turn it into 2 simpler sugar, which the body can then use.

People who are *lactose intolerant* don't have enough **lactase**

enzymes in their gut to break down the **lactose** sugar.

THINK: What kinds of problems result from consuming lactose with insufficient or non-existent gut lactase?

Other Enzymes - Research their bodily functions on your own:

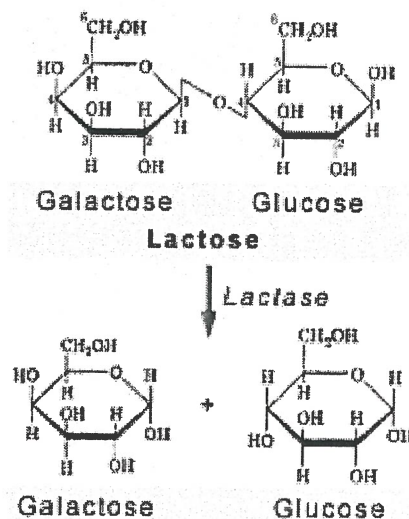
DNA Helicase _____

DNA Polymerase _____

Amylase _____

Protease ("PRO-tee-ase") _____

Lipase _____



Block 2
F'16

Chemistry Unit 6

Primary reference: **CHEMISTRY**, Addison-Wesley

Topic	Essential Knowledge	Study Support
Atomic Structure 2.6 SOL 2f	Electronegativity is the measure of an atom's attraction for electrons in a bond. Electronegativity increases across a period toward the halogens and decreases down a group. The most electronegative atom is fluorine. The least electronegative element (excluding noble gases) is Francium, Fr.	Ch 14: Read p. 405
Nomenclature, Formulas, and Reactions 3.6 SOL 3d, 3e, 3f	<p>Exothermic reactions release heat whereas endothermic reactions absorb heat. Heat of reaction is the amount of energy absorbed or released during a chemical change. Exothermic reactions have a negative ΔH_{rxn}, whereas endothermic reactions have a positive ΔH_{rxn}. Examples of writing an exothermic reaction equation are:</p> $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + 890 \text{ kJ}$ <p>or</p> $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}, \Delta H_{rxn} = -890 \text{ kJ/mol}$ <p>Polar covalent bonds form between elements with very different electronegativities. The more electronegative atom will attract the electrons more strongly and this will result in it having a slight negative charge. The less electronegative atom then takes on a slight positive charge. A non-polar covalent bond forms between atoms of similar electronegativities. A polar molecule has unequally distributed electrons around the central atom. This is caused by unsymmetrical polar bonds or a lone pair on the central atom. The positive end of the molecule has a positive dipole and the negative end has a negative dipole. Polar molecules have dipole-dipole intermolecular attractions as well as London dispersion intermolecular attractions. Non-polar molecules only have London dispersion intermolecular attractions. with O-H, N-H or F-H bonds have hydrogen bonding attractions.</p> <p>Kinetics is the study of reaction rates. increase with increased temperature, reactant concentration, increased and the use of a catalyst. Activation minimum energy needed to initiate a High activation energies correspond to rates. Catalysts speed up reactions decreasing the activation energy. Potential energy diagrams are used to analyze reaction energy changes.</p> <p>EXOTHERMIC</p>	Ch 11: Read pp. 303-304 Ch 16: Read pp. 460-466 Ch 19: Read pp. 533-538.
Molar Relationships 4.6	<p>Stoichiometry can be combined with heat of reaction, ΔH_{rxn}, to calculate the amount of heat produced from a known amount of reactant.</p>	
Phases of Matter and Kinetic Molecular Theory 5.6 SOL 5b, 5c, 5d	<p>Forces of attraction (intermolecular forces) between molecules determine their state of matter at a given temperature. Forces of attraction include hydrogen bonding, dipole-dipole attraction, and London dispersion (van der Waals) forces. Vapor pressure is the pressure of the vapor found directly above a liquid in a closed container. When the vapor pressure equals the atmospheric pressure, a liquid boils. Volatile liquids have high vapor pressures, weak intermolecular forces, and low boiling points. Nonvolatile liquids have low vapor pressures, strong intermolecular forces and high boiling points. Sublimation is the phase change from solid to gas without passing through the liquid phase. A substance's triple point, is the pressure and temperature conditions where all three phases coexist in dynamic equilibrium. The following mathematical relationship between the pressure, volume and temperature of a gas is used to describe the behavior of gases:</p> $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$ <p>An Ideal Gas does not exist, but this concept is used to model gas behavior. A Real Gas exists, has intermolecular forces and particle volume, and can change states. The Ideal Gas Law states that</p> $PV = nRT$ <p>R is the ideal gas law constant and has two values depending on the pressure units. They are $R = 8.314 \text{ L}\cdot\text{kPa}/\text{mol}\cdot\text{K}$ and $R = 0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$</p> <p>Dalton's Law of Partial Pressures says the sum of the partial pressures of all the components in a gas mixture equals the total pressure of the gas mixture.</p> $P_{\text{total}} = P_A + P_B + P_C \quad \text{and} \quad n_A/n_{\text{total}} = P_A/P_{\text{total}} = V_A/V_{\text{total}}$ <p>Graham's Law says gas molecules with the lightest mass travel fastest. $K.E = 0.5 mv^2$</p>	Ch 10: Read pp. 269-280 and pp. 284-286. Ch 12: Read pp. 350-353.

Unit 6 Objectives
Chemistry, Addison-Wesley, 2002

- I) Endothermic and Exothermic Reactions
 - A) Classifying Reactions
 - B) Stoichiometry and Calculating Heats of Reaction
- II) Intermolecular Forces (IMFs)
 - A) Polar bonds
 - B) Polar molecules
 - C) Intermolecular Attractions and Physical Properties
 - 1) Intermolecular forces
 - (a) London Dispersion forces
 - (b) Dipole-Dipole attractions
 - (c) Intermolecular Hydrogen bonding
 - 2) Effect of Intermolecular Forces on Physical properties
 - 3) Comparing molecular and ionic compounds
- III) Phase Changes and Intermolecular Forces (IMFs)
 - A) Kinetic Energy, Particle Velocity, and Kelvins
 - B) Kinetic Energy and Liquids
 - 1) Vapor pressure
 - 2) Boiling points and atmospheric pressure
 - C) Kinetic Energy and Solids
 - D) Phase Changes and Phase Diagrams
- IV) Gas Laws: Combined, Ideal, Dalton's Law and Graham's Law
- V) Reaction Rates
 - A) Collision Theory
 - B) Potential Energy Diagrams
 - 1) Activation Energy
 - 2) Catalysts

(SOL) Learning Objective

1. (3e) Identify a reaction as endothermic or exothermic based on its thermochemical equation and/or sign of ΔH .
2. (4b) Calculate the heat change of a reaction using stoichiometry and heats of reaction.
3. (2f) Compare the electronegativity of two elements based on their position on the periodic table.
4. (3d) Compare relative bond polarity based on the two elements position on the periodic table.
5. (3d) Use VSEPR theory and electronegativity to identify polar and non-polar molecules.
6. (5d) Identify and compare the three types of intermolecular forces (dipole interaction, hydrogen bonding, London dispersion (van der Waals) forces)
7. (5d) Predict the relative melting and boiling points of molecular and ionic substances based on intermolecular forces.
8. (5d) Explain the relationship between kinetic energy and temperature
9. (5d) Interpret a graph of percent molecules vs kinetic energy
10. (5b) Explain why real gases condense whereas ideal gases do not condense using IMFs and kinetic energy.
11. (5b) interpret vapor pressure graphs.
12. (5d) Explain what happens as a solid melts using IMFs and kinetic energy.
13. (5d) Explain the relationship between a substance's vapor pressure and boiling point and the strength of the substance's IMFs.
14. (5d) Interpret the effect of temperature and pressure on states of matter using a phase diagram.
15. (5d) Identify the triple point on a phase diagram and identify which states of matter exist at the triple point.
16. (5d) Identify phase changes on a phase diagram of water including fusion, solidification, vaporization, condensation and sublimation.
17. (5b) Solve gas law problems using the Combined Gas Law and the Ideal Gas Law.
18. (5b) Explain the difference between a real gas and an ideal gas.
19. (5b) Predict when a gas will behave most ideally.
20. (5b) Use Dalton's Law to calculate partial pressures
21. (5b) Use Graham's Law to compare rates of effusion and diffusion of two gases
22. (3f) Draw a reaction's potential energy diagram with axes labeled, and ΔH , activation energy, product energy, reactant energy, transition state, and catalyst shift clearly identified for exothermic and endothermic reactions.
23. (3f) Explain how a catalyst increases reaction rate.
24. (3f) Identify and explain the effect the following factors have on the rate of a chemical reaction: (catalyst, temperature, concentration, and reactant particle size).

Chapter 11 Part 2: Endothermic and Exothermic Reactions

1 of 4

Classifying Reactions as Endothermic or Exothermic

Exothermic Reactions _____

Endothermic Reactions _____

Heat of Reaction, $\Delta H_{\text{reaction}}$ is the heat absorbed or released by a reaction.

$+\Delta H_{\text{rxn}}$ endo \uparrow $-\Delta H_{\text{rxn}}$ exo \downarrow

Thermochemical Equations like regular eqns, but include heat flow (+ ΔH or - ΔH)

There are two general ways they're written. Both are acceptable:

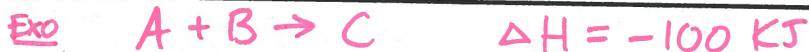
Heat as a Product (exo) or Reactant (endo):



Heat shown as a change in Enthalpy (ΔH)

$-\Delta H$ means heat is lost (exothermic)

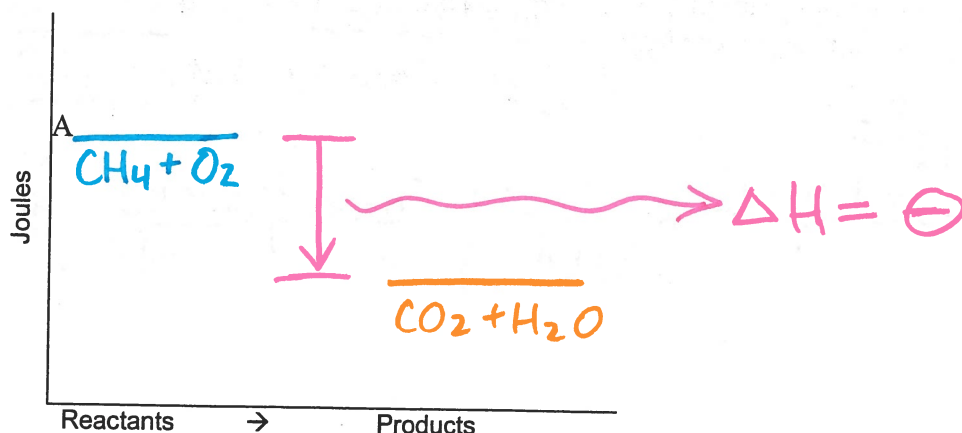
$+\Delta H$ means heat was absorbed (endothermic)



Example one: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + 890.4 \text{ kJ}$

so $\Delta H = -890.4 \text{ KJ}$

Potential Energy Diagram



The system released / produced heat so ΔH is negative/positive.

The reaction is endothermic/exothermic?

Law of Conservation of Energy:

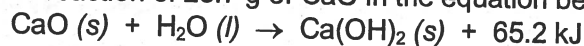
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Solving Heat of Reaction Problems

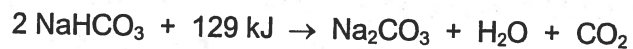
3 of 4

Treat heat (in J or kJ) the same as any reactant or product in a chemical equation

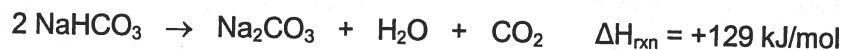
1. How much heat is produced by the reaction of 25.7 g of CaO in the equation below? (Ans = 29.9 kJ)



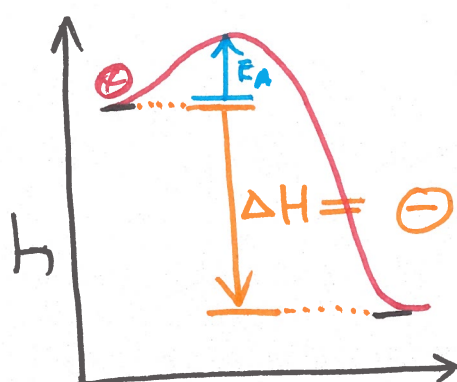
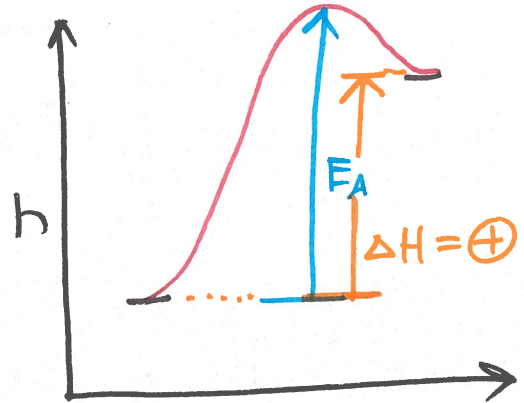
2. How many grams of NaHCO₃ are needed to react completely when 980 kJ of heat are used in the equation below? (Ans = 1300 g)



3. Using the same equation, how many kJ of heat must be used to produce 55.7 liters of CO₂ at STP? (Ans = 321 kJ)



Comparing Endothermic and Exothermic Reactions

Exothermic Reactions	Endothermic Reactions
The reaction <u>produces</u> heat	The reaction <u>absorbs</u> heat
Example of exothermic reaction	Example of endothermic reaction
<p>Potential Energy vs. Reaction Path for Exothermic Reactions</p>  <p>E_A = Activation Energy</p>	<p>Potential Energy vs. Reaction Path for Endothermic Reactions</p> 
An exothermic reaction has a net <u>⊖</u> ΔH	An endothermic reaction has a net <u>⊕</u> ΔH
In exothermic reactions, the product's energy is <u>lower</u> than the reactant's energy	In endothermic reactions, the product's energy is <u>higher</u> than the reactant's energy

Word bank: positive, releases, negative, absorbs, lower, higher

Chapter 16: Polar Bonds and Polar Molecules and Intermolecular Forces

Electronegativity: Tendency of an atom to "hog" or attract e^- (F^o)

A. Non-Polar and Polar Covalent Bonds

- 1) non-polar covalent bonds: atoms share bonding electrons equally.

Non-Polar

Draw
Example: Cl_2

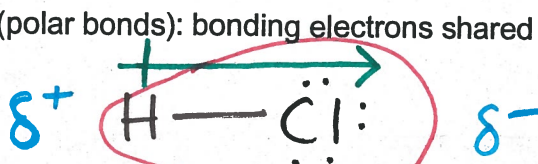


↑
most EN

- 2) polar covalent bonds (polar bonds): bonding electrons shared unequally.

N S

Example: HCl



← less EN → more EN

- 3) Electron sharing based on electronegativity differences.
- more electronegative atom attracts the electrons more closely and acquires a slight negative charge.
 - less electronegative atom then acquires a slight positive charge.
 - unequal sharing creates "polarized" bonds with opposite charges.
 - Two ways to show polarity in structural formulas.

lower case greek deltas: δ^- , δ^+

Partially \ominus negative
partially \oplus positive

slashed arrows: \rightarrow

The type of bond depends on electronegativity differences between the atoms

Electronegativity Difference	Guideline: Type of Bond	Example (electronegativity difference)
0.0 – 0.4	Non-polar Covalent	C-H in CH_4
0.4 – 2.0	Polar Covalent	HF
>2.0	Ionic	NaCl

Selected Electronegativity Values

H 2.1						
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8

metals:
low EN

Non-metals:
high EN

Is the bond polar, non-polar or ionic?

C=O in CO_2 ?	$ 3.5 - 2.5 = 1.0$
Si-H in SiH_4	$ 1.8 - 2.1 = 0.3$
C-F in CF_4	$ 2.5 - 4.0 = 1.5$
N-Cl in NCl_3	$ 3.0 - 3.0 = 0$
KCl	$ 3.0 - 0.8 = 2.2$

Answer this question without looking at the table on the previous page.

Which bond is most polar; C-N or C-F?

B: POLAR MOLECULES

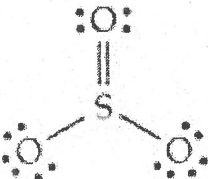
Polar Molecules: One end of the molecule is slightly negative, and one end is slightly positive.

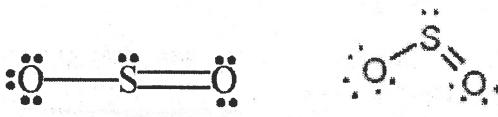
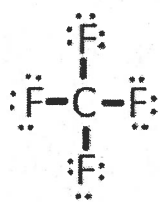
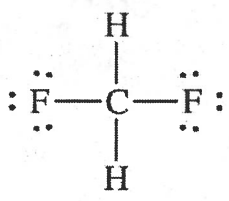
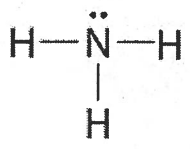
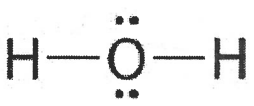
Dipole: a molecule with two poles (one negative, one positive or δ^- , δ^+)

What makes a molecule polar?	A molecule is polar if the electrons are pulled to one side of the molecule. The molecule is lopsided (assymetrical).
------------------------------	---

Determining if a molecule is polar.

1. Draw the Lewis structure
2. Determine the molecular geometry
3. Look for lone pairs on central atom (automatically polar)
4. Are there polar bonds?
5. If yes, are the polar bonds unsymmetrical in 3-D around the molecule's center?

Molecule	Lewis Structure and Geometry	Polar or Non Polar
CO	$:\text{C}\equiv\text{O}:$	
CO ₂	$:\ddot{\text{O}}=\text{C}=\ddot{\text{O}}:$	
SO ₃		

Molecule	Lewis Structure and Geometry	Polar or Non Polar
SO ₂		
CF ₄		
CH ₂ F ₂		
NH ₃		
H ₂ O		

Which molecule is most polar: HCl or HI?

Which molecule is most polar; CO₂ or SO₂?

C: INTERMOLECULAR FORCES

4 of 6

Intermolecular Forces are attractions between molecules due to three forces

1. **London Dispersion Forces** (weakest) temporary attractions between molecules due to temporary dipoles caused by shifting electron clouds. Dispersion forces are greater in more massive molecules with larger electron "clouds". **All molecules have LDFs, and larger atoms/molecules have stronger LDF because they have more electrons.**

✱ Important in non-polar

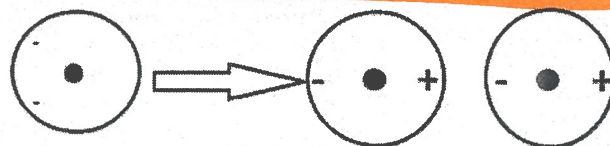
CO₂

VS

CS₂

stronger LDF

larger



Uneven distribution of electrons in He

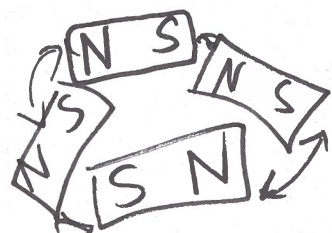
Instantaneous dipole

Induced dipole on neighboring He

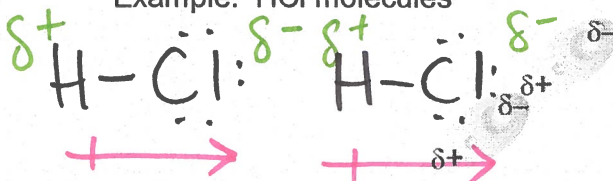
✱ In everything
✱ Stronger in bigger stuff

2. **Dipole-dipole attraction** **polar** molecules are attracted to each other (like magnets). The positive dipole of one molecule is attracted to the negative dipole of another. These occur when molecules have uneven distribution of electrons due to electronegative elements (like N, O, Cl, F, Br, etc.).

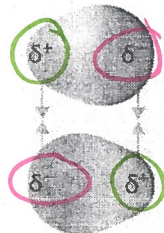
other polar



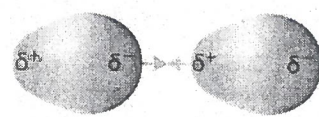
Example: HCl molecules



✱ molecules must be polar!



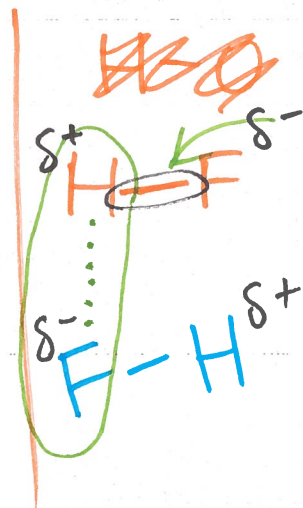
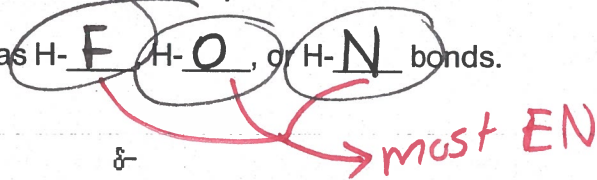
(a) Attraction



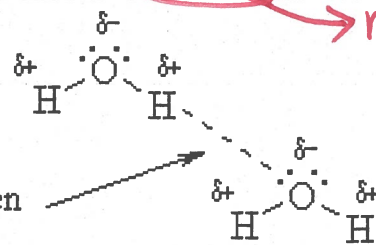
(b) Attraction

3. **Intermolecular Hydrogen bonding**: hydrogen that is covalently bonded to a **very** electronegative atom is also weakly bonded to the unshared pair of another electronegative atom. **H-Bonding** only occurs when a molecule has H-**F**, H-**O**, or H-**N** bonds. Why these three atoms specifically?

✱ H-bonding is just a very strong dipole-dipole DP-DP



hydrogen bond



more/stronger IMF \rightarrow harder to sep.

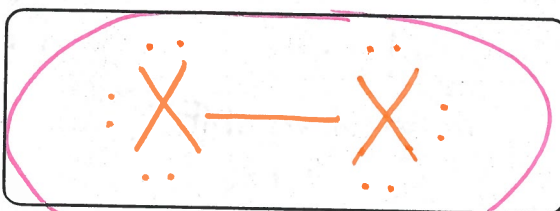
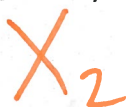
5 of 6
melting \rightarrow boiling

4. Intermolecular Forces and Molecular Physical Properties

As intermolecular forces (IMF) increases (meaning gets stronger), the melting and boiling points increase because more kinetic energy is needed to overcome the IMFs between molecules.

An Analysis of the Halogens and their Physical States:

The Halogens are Diatomic:
(Lewis diagram \rightarrow)



"X" represents halogens

Non-polar

Are pure diatomic halogen elements polar? Explain & draw dipole vectors & partial charges.

No.




What is the only IMF that all non-polar things can do?

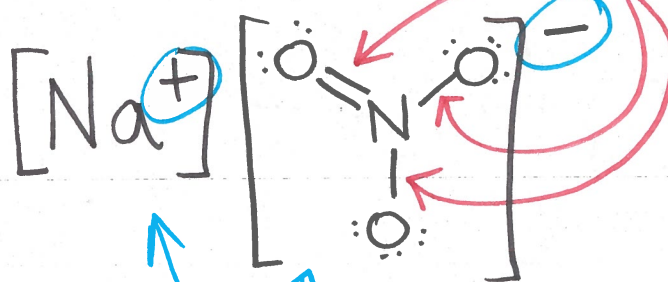
	Halogen (X)	Molar Mass (X_2)	State at Room Temp.	
<div>All NP</div> <div>Bigger</div>	F	$F_2 = 38.0 \text{ g/mol}$	<u>gas</u>	
	Cl	$Cl_2 = 70.9 \text{ g/mol}$	<u>gas</u>	
	Br	$Br_2 = 159.8 \text{ g/mol}$	<u>liquid</u>	
	I	$I_2 = 253.8 \text{ g/mol}$	<u>solid</u>	

LDF:
bigger = stronger

Comparing Ionic and Covalent Compounds

Characteristic	Ionic Compound	Molecular Compound
bond formation	TRANSFER of e^- $\text{Na} \cdot \curvearrowright \cdot \ddot{\text{Cl}}: \rightarrow \text{Na}^+ : \ddot{\text{Cl}}:^-$	SHARE of e^- $\ddot{\text{O}}=\text{C}=\ddot{\text{O}}:$
Types of elements in compound	METALS NON-METAL	Non-metals
Physical state at 25°C	ion = strong IMF!!! SOLID 	Any!
Melting point	HIGH AF	relatively low
Electrical Conductivity in aqueous solution	conducts e^- (usually)	doesn't.

~~Sodium~~ nitrate



electrostatic attraction (ionic)

↙ motion

$$KE = \frac{1}{2} m v^2$$

velocity

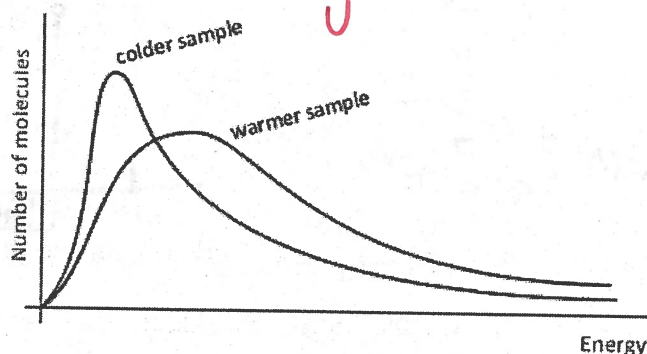
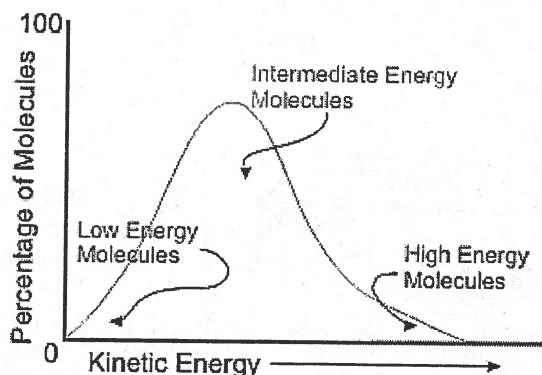


Chapter 10 Kinetic Theory, IMFs, and Phase Changes

Kinetic Molecular Theory: The tiny particles in all forms of matter are in constant motion.

I Kinetic Energy and Kelvin temperature scale

A) Temperature measures average kinetic energy



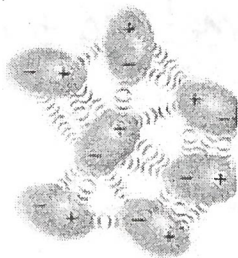
B) Gas particle's kinetic energy increases as velocity increases

C) Kelvin Temperature scale is absolute 0 → UP

0 K 0 K = 0 °C

II. Kinetic Energy and Liquids

Intermolecular forces (between molecules) hold particles together in solid or liquid phases.



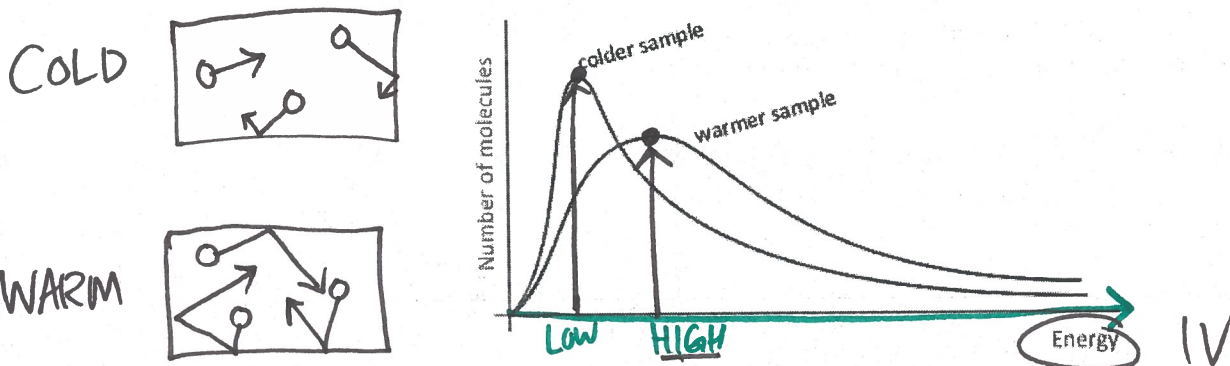
Kinetic energy keeps the molecules moving but not with enough energy to overcome the IMFs.

Evaporation, Vapor Pressure and Temperature

Evaporation:

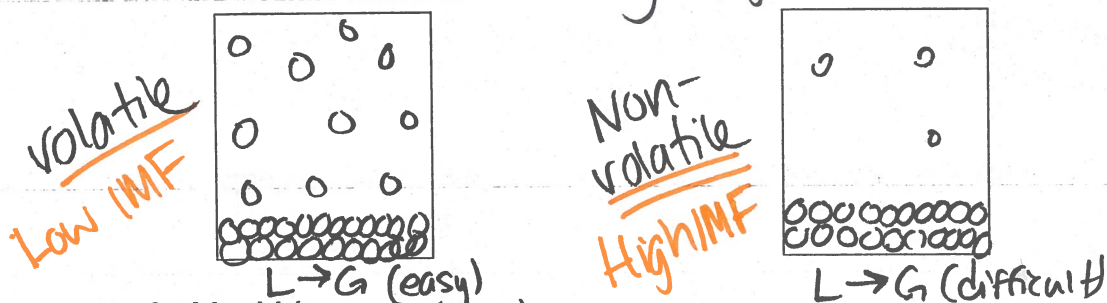
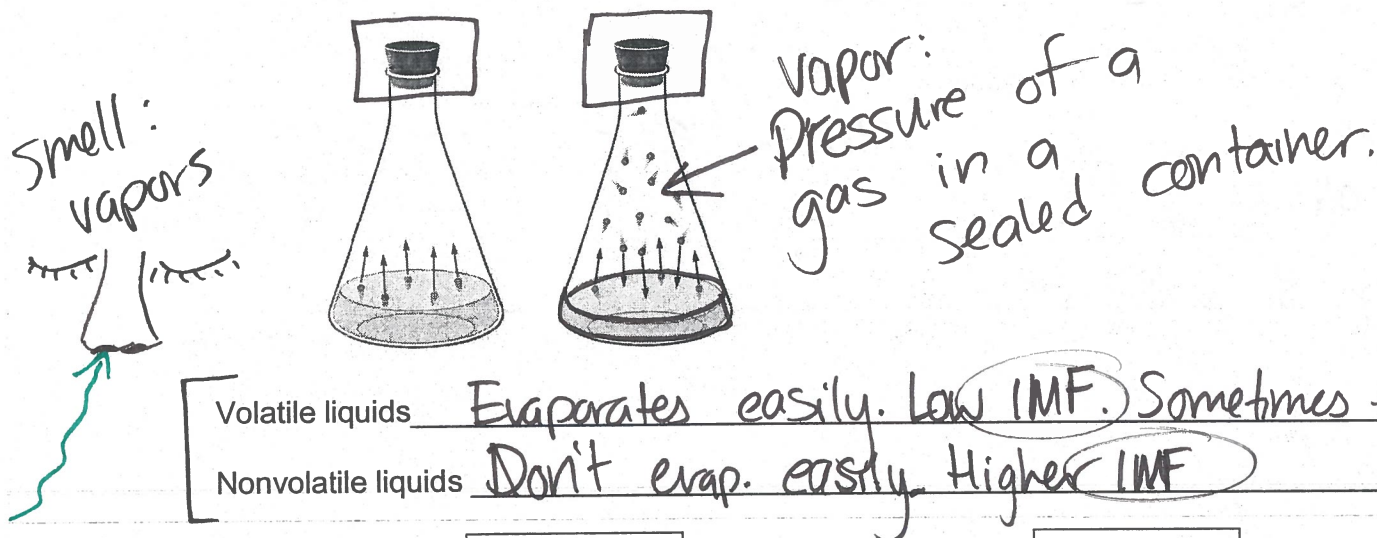
Phase change from (L) → (G)
@ the non-boiling temperature

- Particles with enough kinetic energy to overcome intermolecular forces escape into gas phase
- Evaporation rate increases as temperature increases



Evaporation in a closed container produces vapor pressure.

Increasing temperature increases vapor pressure over a liquid until a dynamic equilibrium is reached.

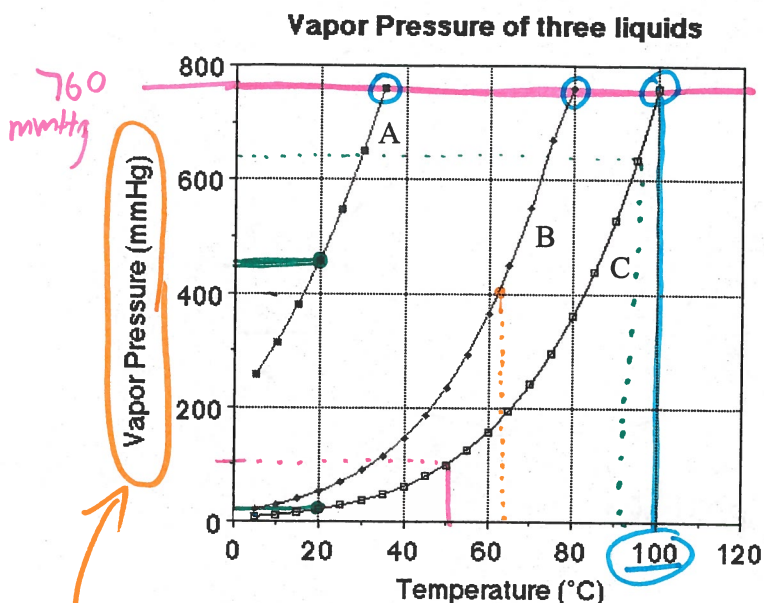


Boiling Point of a Liquid (open container)

Boiling Point: Temp. @ which all molecules
in a liquid have enough energy
from L to G.

⊙ = boiling point (@ sea level)

A) Boiling point changes as external pressure changes/



What is the vapor pressure of liquid A at 20°C?_

450 mmHg

Which liquid represents water?

"C"

What is the boiling point of B when the external pressure is 400 mmHg?

63°C

→ Which liquid is most volatile? L → G

A

Which liquid has the strongest intermolecular forces?

C

How hot does water need to be to boil at 100 mmHg?

50°C

What will be the boiling point of water on Pike's Peak (elevation = 14,000 ft, atmospheric pressure = 640 mmHg)?

~92°C

A

"Normal" (Standard)

= 1 atm = 760 mmHg (y-axis)

III. Kinetic Energy and Solids

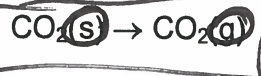
melting point:

Temp @ which S → L
(Same as Freezing Point)

sublimation:

S → G (skips liquid)

Examples:

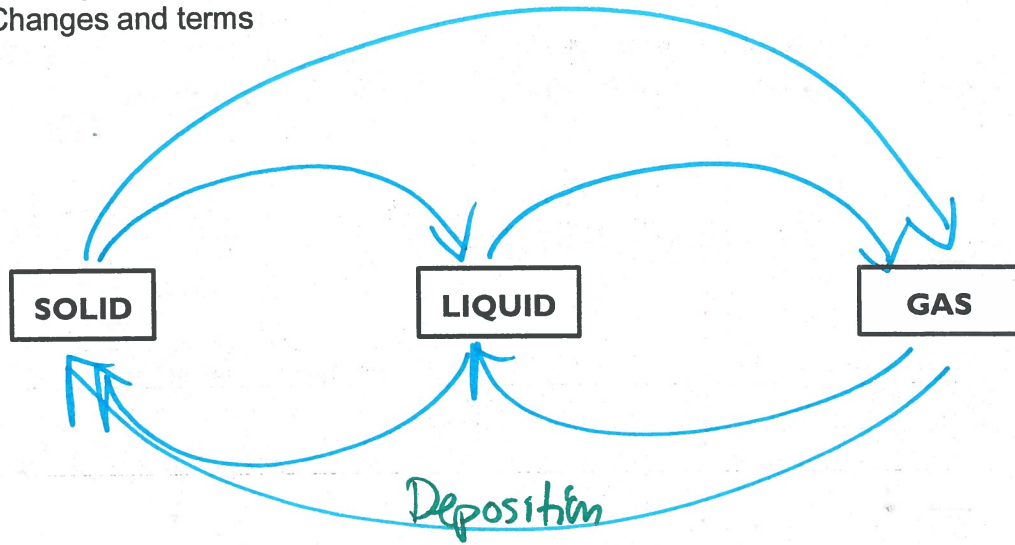


Dry Ice



IV. Phase Changes and Phase Diagrams

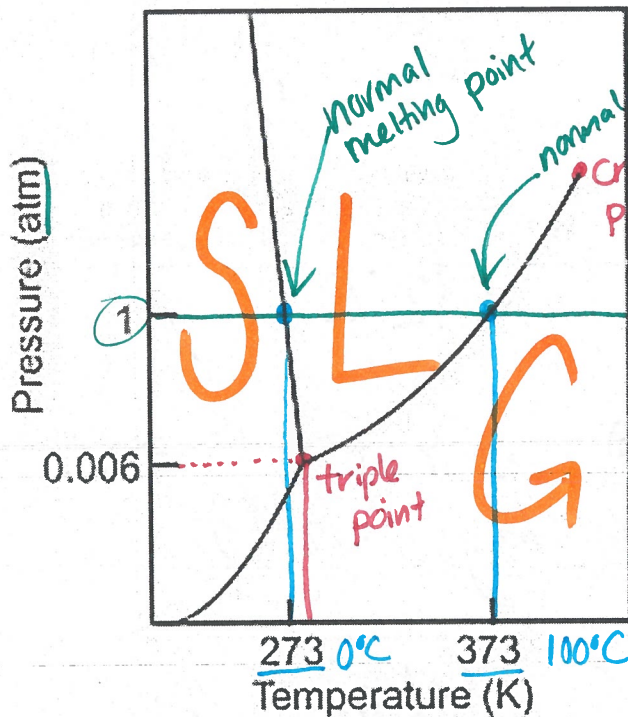
Phase Changes and terms



12/16

12/19

Phase Diagram and Triple Point



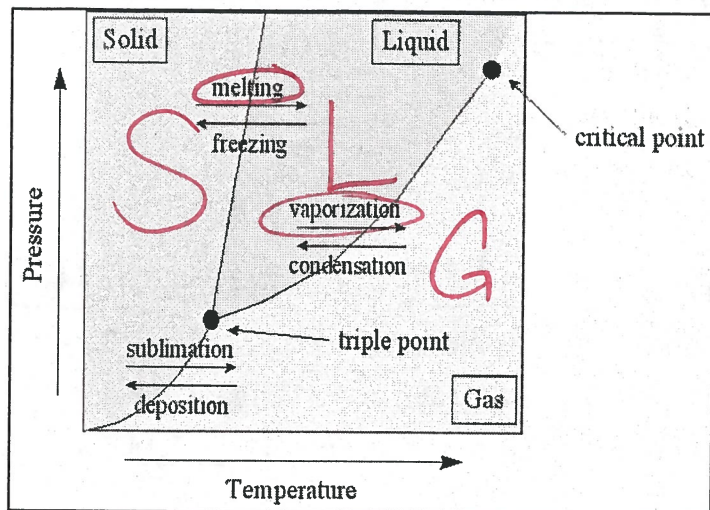
Phase Diagrams:

Tell ~~you~~ you the state/phase (s, l, g) of a substance @ any P or T.

Phase Diagrams

A **phase diagram** is a graphical way to depict the effects of **pressure** and **temperature** on the phase of a substance:

The **CURVES** indicate the conditions of **temperature** and **pressure** under which "equilibrium" between different phases of a substance can exist. **BOTH** phases exist on these lines:



Melting/Freezing: Any point on this line (pressure & temperature) the substance is both **solid** and **liquid**

Sublimation/Deposition: Any point on this line (pressure & temperature) the substance is both **solid** and **gas**

Vaporization/Condensation: Any point on this line (pressure & temperature) the substance is both **liquid** and **gas**

NOTE: the vapor pressure curve ends at the **critical point**, the temperature above which the gas cannot be liquefied no matter how much pressure is applied (the kinetic energy simply is too great for attractive forces to overcome). Any substance beyond this critical point is called a **supercritical fluid** – indistinguishable between gas or liquid (neither one)

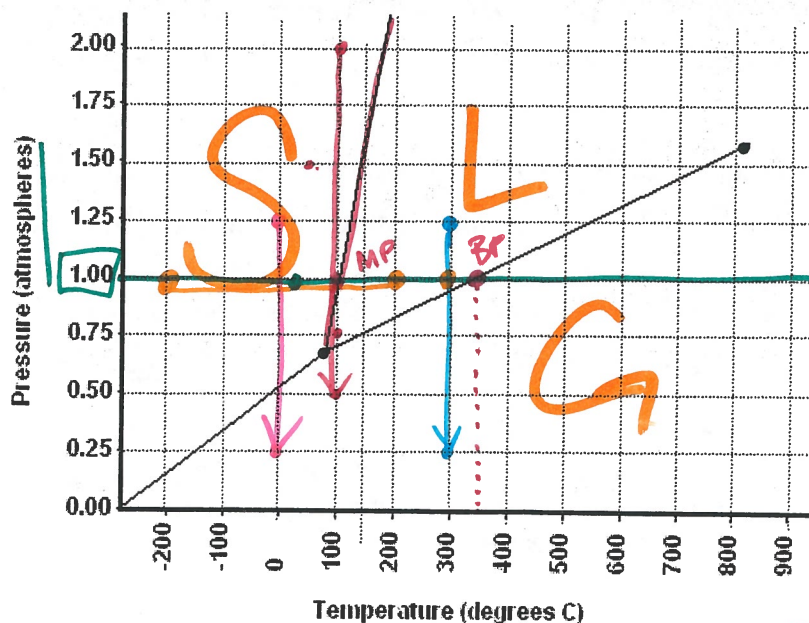
The **TRIPLE POINT** is the condition of temperature and pressure where ALL THREE phases exist in equilibrium (solid, liquid, gas)

Remember that pressure can be expressed in many units where: **1 atm = 101.3 kpa = 760 mmHg = 760 torr = 14.7 psi**

'Normal'

Refer to the phase diagram below when answering the questions.

NOTE: "Normal" refers to STP – Standard Temperature and Pressure.



- 1) What are the values for temperature and pressure at STP? T= 0°C, P= 1 atm
- 2) What is the normal freezing point of this substance? 100°C
- 3) What is the normal boiling point of this substance? 350°C

- 4) What is the normal melting point of this substance? 100°C
- 5) What is the phase (s, l, g) of a substance at **2.0 atm** and 100 °C? S
- 6) What is the phase (s, l, g) of a substance at **0.75 atm** and 100 °C? _____
- 7) What is the phase (s, l, g) of a substance at **0.5 atm** and 100 °C? _____
- 8) What is the phase (s, l, g) of a substance at 1.5 atm and 50 °C? _____
- 9) What is the phase (s, l, g) of a substance at **1.5 atm** and **200 °C**? _____
- 10) What is the phase (s, l, g) of a substance at **1.5 atm** and **800 °C**? _____
- 11) What is the condition of the triple point of this substance? T= _____, P= _____
- 12) If a quantity of this substance was at an initial pressure of 1.25 atm and a temperature of 300° C was lowered to a pressure of 0.25 atm, what phase transition(s) would occur? Vaporization
- 13) If a quantity of this substance was at an initial pressure of 1.25 atm and a temperature of 0° C was lowered to a pressure of 0.25 atm, what phase transition(s) would occur? sub.
- 14) If a quantity of this substance was at an initial pressure of **1.0 atm** and a temperature of **200° C** was lowered to a temperature of -200° C, what phase transition(s) would occur? Solid.
- 15) If a quantity of this substance was at an initial pressure of **0.5 atm** and a temperature of **200° C** was lowered to a temperature of -200° C, what phase transition(s) would occur? _____
- 16) If this substance was at a pressure of **2.0 atm**, at what temperature would it **melt**? _____
- 17) If this substance was at a pressure of **2.0 atm**, at what temperature would it **boil**? _____
- 18) If this substance was at a pressure of **0.75 atm**, at what temperature would it **melt**? _____
- 19) If this substance was at a pressure of **0.75 atm**, at what temperature would it **boil**? _____
- 20) At what temperature do the gas and liquid phases become indistinguishable from each other? _____
- 21) At what pressure would it be possible to find this substance in the gas, liquid, **and** solid phase? _____
- 22) If I had a quantity of this substance at a pressure of **1.00 atm** and a temperature of **-100° C**, what phase change(s) would occur if I **increased the temperature to 600° C**? At what temperature(s) would they occur? (**NOTE: multiple answers needed for this question**)
- 22) If I had a quantity of this substance at a pressure of **2.00 atm** and a temperature of **-150° C**, what phase change(s) would occur if I **decreased the pressure to 0.25 atm**? At what pressure(s) would they occur? (**NOTE: multiple answers needed for this question**)

Chapter 12 Dalton's Law and Graham's Law

Mixtures of Gases: Dalton's Law of Partial Pressure

Partial Pressure: Pressure due to a single gas in a mixture of gases

$$\sum_{n=1}^x P_n$$

Verbally: At constant pressure and temperature, the total pressure exerted by a mixture of gases is equal to the sum of the partial pressures of the component gases.

Math Equation: $P_T = P_A + P_B + P_C \dots$

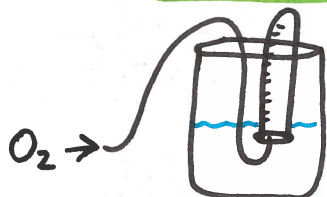
Example 1: What is the partial pressure of oxygen in air at STP (101.3 kPa) if $P_{N_2} = 79.1$ kPa, $P_{CO_2} = 0.040$ kPa and $P_{others} = 0.94$ kPa?

$$101.3 \text{ kPa} = 79.1 \text{ kPa} + 0.040 \text{ kPa} + 0.94 \text{ kPa} +$$

$$P_{O_2}$$

$$P_{O_2} = 21.2 \text{ kPa}$$

Example 2: A sample of oxygen gas is collected over water at 20°C . The vapor pressure of water at 20°C is 15 mm Hg. If the total pressure is 420 mm Hg, what is the partial pressure of the oxygen?



$$P_T = P_{O_2} + P_{H_2O(g)}$$

$$420 = P_{O_2} + 15$$

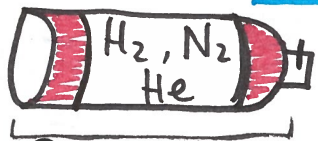
$$\therefore P_{O_2} = 405 \text{ mmHg}$$

Variants of Dalton's Law

$$\frac{20}{50} = \frac{2}{5} \leftarrow \text{mole fraction}$$

$$\frac{\text{mole}_A}{\text{mole}_{\text{Total}}} = \frac{P_A}{P_{\text{Total}}} = \frac{\%V_A}{100 \%V_{\text{Total}}}$$

Example 1: A tank contains 6.0 moles of a mixture of hydrogen, helium, and nitrogen at 102 kPa. If there are 2.0 moles of hydrogen in the tank, what is the partial pressure of hydrogen?

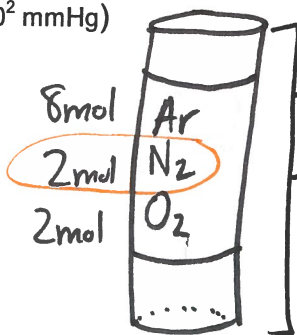


$$P_T = 102 \text{ kPa}$$

$$\frac{\text{mol } H_2}{\text{total mol}} = \frac{P_{H_2}}{P_{\text{Total}}} \rightarrow \frac{2 \text{ mol}}{6 \text{ mol}} = \frac{P_{H_2}}{102 \text{ kPa}}$$

$$= 34 \text{ kPa}$$

Example 2: A gas cylinder contains 8.0 moles of argon, 2.0 moles of nitrogen, and 2.0 moles of oxygen at 600. mmHg. What is the partial pressure of nitrogen in the cylinder? (Ans = 1.0×10^2 mmHg)



$$600 \text{ mmHg} = P_T$$

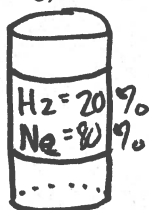
$$12 \text{ mol, total}$$

$$\frac{\text{mol } N_2}{\text{total mol}} = \frac{P_{N_2}}{P_T}$$

$$\frac{2 \text{ mol } N_2}{12 \text{ mol } N_2} = \frac{P_{N_2}}{600 \text{ mmHg}}$$

$$P_{N_2} = 100 \text{ mmHg}$$

Example 3: A mixture of gases with a pressure of 950 mm Hg contains 20% hydrogen and 80% neon by volume. What is the partial pressure of neon gas in the mixture? (Ans = 760 mmHg)

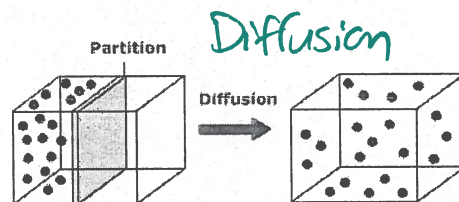


$$\frac{80\% \text{ Ne}}{100\% \text{ total}} = \frac{P_{\text{Ne}}}{950 \text{ mmHg}} \therefore P_{\text{Ne}} = 760 \text{ mmHg}$$

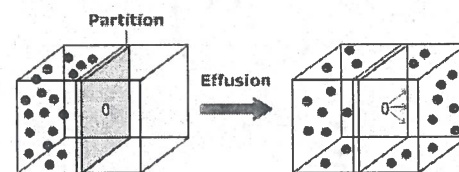
Example 4: In a mixture of oxygen and nitrogen gas, 70.0 percent of the total gas pressure is exerted by the nitrogen. If the total pressure is 150 kPa, what pressure does oxygen exert? (Ans = 45 kPa)

Graham's Law of Effusion

Diffusion Gas particles move from HIGH to LOW Concentration



Effusion Gas particles moving thru a hole lighter gases move/travel fastest

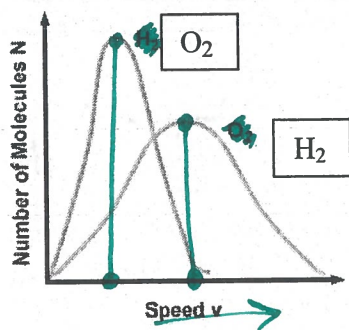


Graham's Law Verbally: Gas molecules with the lightest mass travel fastest.

Graham's Law Equation:

$$KE = \frac{1}{2} m v^2$$

Annotations:
 - m is labeled 'mass (molar mass)'
 - v is labeled 'velocity ("speed")'
 - v is also labeled 'rate' with a box around it.



N2 vs F2
 faster

Ar vs O2
 faster

H2S vs Cl2
 slower

Which gas will escape slowest from a tiny hole in a balloon; He, C3H8, or Xe?

slowest.

Partial Pressures Practice (Lee)

NDB→

- 1) Argon, oxygen, and nitrogen are mixed together and pressurized in a tank. P_{O_2} is 155 kPa, P_{N_2} is 415 kPa, and P_{Ar} is 285 kPa.

a. What is the total pressure of the gas mixture?

b. What is the % O_2 ?

c. What is the % N_2 ?

- 2) The total pressure in a tank is 1200.0 mmHg. Krypton has a partial pressure (P_{Kr}) of 680.0 mmHg, and methane has a partial pressure that is half that of krypton. The third gas in the container is chlorine.

a. Calculate the partial pressures of each gas with units:

$$P_{CH_4} = 340 \text{ mmHg}$$

$$P_{Kr} = 680 \text{ mmHg}$$

$$P_{Cl} = 180 \text{ mmHg}$$

$$1200 = 680 + 340 + \underline{180} \text{ Cl}$$

b. Determine the % volume of each gas:

$$\% Cl = \frac{180}{1200} \times 100 = 15\%$$

$$\% Kr = \frac{680}{1200} \times 100 = 56\%$$

$$\% CH_4 = \frac{340}{1200} \times 100 = 29\%$$

- 3) A highly pressurized (46.2 atm) mixture of gases contains a total of 330. moles of gases. The technologist who created the mixture added fluorine, chlorine, helium, and hydrogen. Twice as many moles of helium were present than fluorine. Three times as many moles of chlorine were added than fluorine, and the half as many moles of hydrogen were added as fluorine.

a. Set up an algebraic equation using variables that expresses the situation above.

$$330 \text{ mol} = \underset{(F_2)}{x} + \underset{(He)}{2x} + \underset{(Cl_2)}{3x} + \underset{(H_2)}{0.5x}$$

b. Determine the number of moles of each gas.

$$330 \text{ mol} = 6.5x$$

$$x = \# \text{ of } F_2 \text{ moles} = 50.77 \text{ mol}$$

$$Cl_2 = 152.31 \text{ mol}$$

$$He = 101.53 \text{ mol}$$

$$H_2 = 25.39 \text{ mol}$$

c. Determine the partial pressures of each gas. Use proper notation. (Example: To express the partial pressure of X, write P_X)

$$\frac{50.77 \text{ mol } F_2}{330 \text{ mol, total}} = \frac{P_{F_2}}{46.2 \text{ atm}} \therefore P_{F_2} = 7.11 \text{ atm}$$

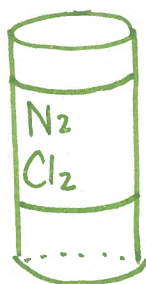
Earth's pressure = 101.3 kPa

Planet X press = 202.6 kPa

- 4) A certain planet was discovered whose atmospheric composition was 70% CO₂, 20% O₂, and 7% H₂, and 3% He. The atmospheric air pressure on the planet was determined to be twice that of normal atmospheric pressure on Earth. Calculate the partial pressure of each gas in kPa.

$$P_{\text{CO}_2} \rightarrow \frac{70\% \text{ CO}_2}{100\% \text{ total}} = \frac{P_{\text{CO}_2}}{202.6 \text{ kPa}} \therefore P_{\text{CO}_2} = 141.8 \text{ kPa}$$

- 5) 15.0 g of nitrogen gas, and 15.0 grams of chlorine gas were added to a container that exists at STP



→ 101.3 kPa

- a. What is the number of moles of nitrogen in the container?

$$\frac{15 \text{ g N}_2}{1} \times \frac{1 \text{ mol N}_2}{28 \text{ g N}_2}$$

0.5357 mol N₂

- b. What is the number of moles of chlorine in the container?

$$\frac{15 \text{ g Cl}_2}{1} \times \frac{1 \text{ mol Cl}_2}{70.9 \text{ g Cl}_2}$$

0.2116 mol ~~N₂~~ Cl₂

- c. How many total moles of gas are in the container?

0.7473 moles, total.

- d. What is the P_{N₂} in kPa?

72.6 kPa

$$\frac{0.5357 \text{ mol N}_2}{0.7473 \text{ mol, total}} = \frac{P_{\text{N}_2}}{101.3 \text{ kPa}}$$

- e. What is the P_{Cl₂} in kPa?

0.40 kPa

Gas A vs Gas B

Graham's Law of Effusion (Applied Quantitative Practice)

Graham's Law can be derived from the equation for kinetic energy

$$(K = \frac{1}{2}mv^2)$$

Derivation: ① $KE_A = KE_B$ because they're the same temperature.

$$\textcircled{2} KE_A = \frac{1}{2}m_A v_A^2 \quad \text{and} \quad KE_B = \frac{1}{2}m_B v_B^2$$

equal energy b/c same temp.

$$\textcircled{3} \cancel{\frac{1}{2}}m_A v_A^2 = \cancel{\frac{1}{2}}m_B v_B^2 \quad (\text{constants cancel}) \quad \textcircled{4} \text{ get velocity (rate) ratio: } \frac{m_A v_A^2}{m_B v_B^2} = m_B$$

$$\textcircled{5} \frac{v_A^2}{v_B^2} = \frac{m_B}{m_A}$$

$$\textcircled{6} \frac{v_A}{v_B} = \sqrt{\frac{m_B}{m_A}}$$

$$\frac{\text{rate}_A}{\text{rate}_B} = \sqrt{\frac{MM_B}{MM_A}}$$

Graham's Law (final equation) starts that lighter gases move more quickly than heavier gases in an inverse square proportion. (According to kinetic theory, when the same amount of energy is available to different bodies with different masses, they will move at inverse square velocities relative to each other.) A fundamental assumption when using Graham's Law of Effusion is that the gases have the same amount of energy... so they're at the same temperature on the Kelvin scale. Their difference in average velocity (aka: rate, speed...) is due to mass difference.

Analogy: A 60-kg girl eats 2 eggs and 3 slices of bacon. A 150 kg sumo wrestler eats the same thing. The girl will run faster because she's smaller, even though they had the same breakfast (energy).

Graham's Law of Effusion

$$\frac{r_A}{r_B} = \sqrt{\frac{MM_B}{MM_A}}$$

Tips for Use:

Heavier gas as "B" (lighter = "A")

Worked Example 1: The average velocity of oxygen (O_2) molecules will be faster than the average velocity of chlorine (Cl_2) molecules because oxygen has a smaller molar mass. What is the relative rate (speed) of oxygen molecules to chlorine molecules if they are at the same temperature? (i.e., how many times faster will molecules of oxygen move?)

$$\frac{\text{rate}_A}{\text{rate}_B} = \sqrt{\frac{MM_B}{MM_A}} \rightarrow \frac{\text{rate}_{O_2}}{\text{rate}_{Cl_2}} = \sqrt{\frac{70.9 \text{ g/mol}}{32.0 \text{ g/mol}}}$$

Cl_2 has a velocity of 310. m/s

$$\frac{1.49}{1}$$

O_2 moves 1.49 times faster.

Part 1: Determine which gas will effuse or move fastest, and determine how many times faster it moves.

1) H_2 vs. He

2) O_2 vs. Ne

3) CH_4 vs. NCl_3

4) Ammonia vs. Hydrogen Sulfide (H_2S)

5) Xenon vs. Argon

Part 2: ~~Determine which gas effuses or move slowest, and determine how many times slower it moves~~

6) Oxygen vs. chlorine

7) SO_2 vs. CH_4 Sulfur dioxide vs. methane

8) Laughing gas (dinitrogen monoxide) vs. carbon monoxide

9) Sulfur Hexafluoride vs. carbon tetrafluoride

10) Silane (SiH_4) vs. hydrogen

Part 3 (Advanced): Determining the molar mass of an unknown gas; or identifying the gas by calculating the molar mass from relative rates.

11) A sample of hydrogen gas effuses through a porous container 9 times faster than an unknown gas. Estimate the molar mass of the unknown gas. Would it be reasonable to assume this gas is silicon tetrafluoride? Explain and justify your answer using mathematics and complete sentences.

- 12) At a certain temperature, hydrogen molecules move at an average velocity of 1.84×10^3 m/s. Estimate the molar mass of a gas whose molecules have an average velocity of 311 m/s
- 13) Nitrogen gas (N_2) effuses at a rate 2.17 faster than an unknown noble gas. Identify the noble gas
- 14) A sample of $\text{Br}_2(\text{g})$ takes 10.0 min to effuse from one side of a room that is 86 feet long. How long would it take the same amount of $\text{Ar}(\text{g})$ to effuse the same distance?
- 15) Explain why carbon monoxide and nitrogen effuse at nearly the same rate. Use complete sentences and justify your answer.

Chapter 19 Part 1: Reaction Rates

http://www.wwnorton.com/college/chemistry/gilbert/index/site_map.htm

Reaction rates are measured as mol/time units.

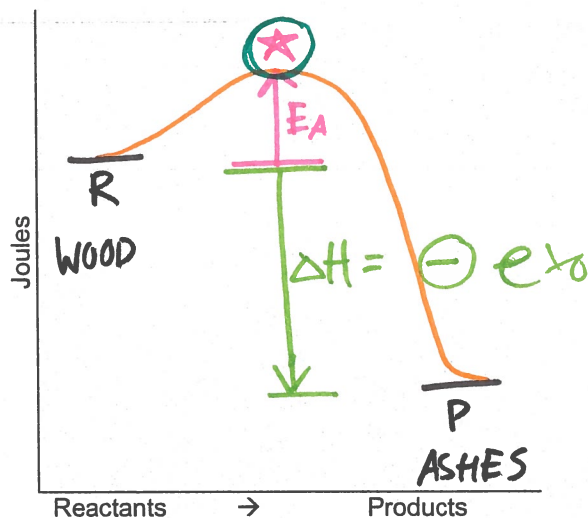
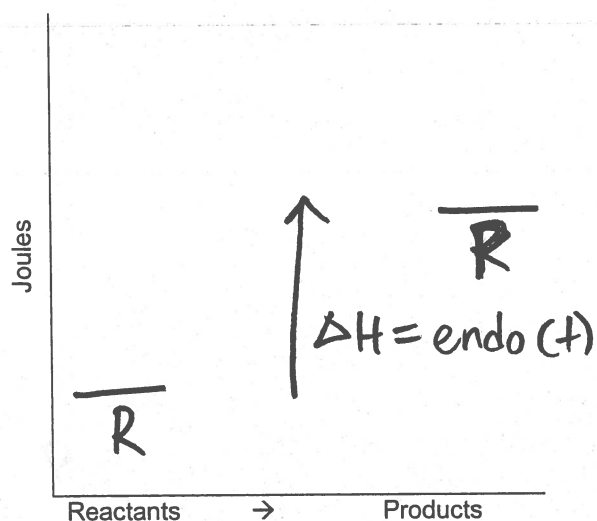
A. Collision Theory: Atoms/molecules will participate in rxns only when they have enough energy

Reaction Coordinate Diagram for Exothermic Process (forward direction)

A correct orientation

Chapter 11: Thermochemistry

Collision Theory



Activation Energy: Minimum energy needed to start reaction

Activation Complex: The high-energy "hill top" (unstable... temporary)

B. Factors Influencing Reaction Rate

1. Temperature

Increasing Temp ($T \uparrow$) speeds up rxn.

2. Concentration

Increase concentration (Molarity \uparrow), speeds up rxn.

3. Particle Size

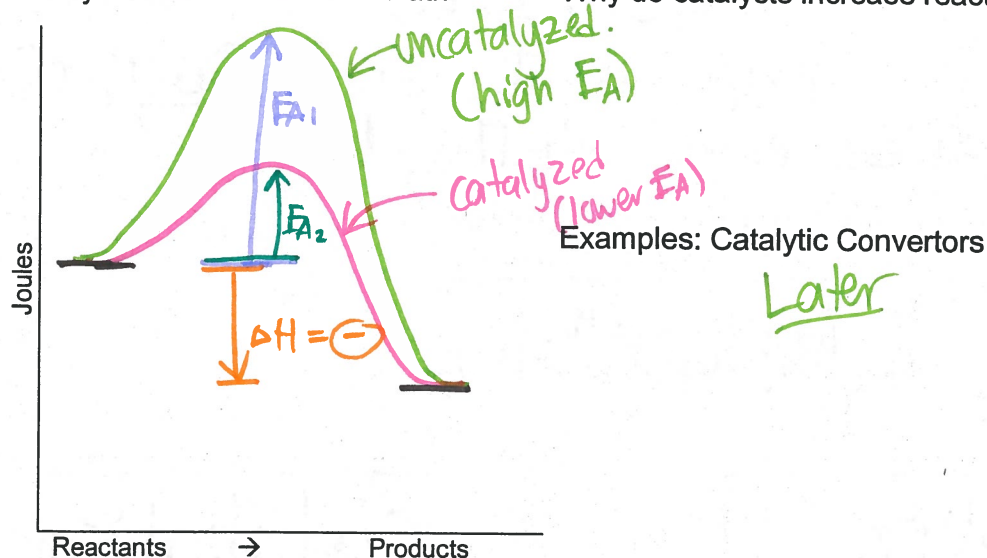
Smaller particles react more rapidly!!

4. Catalysts

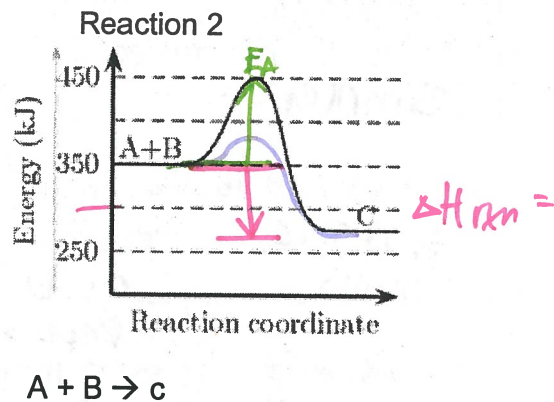
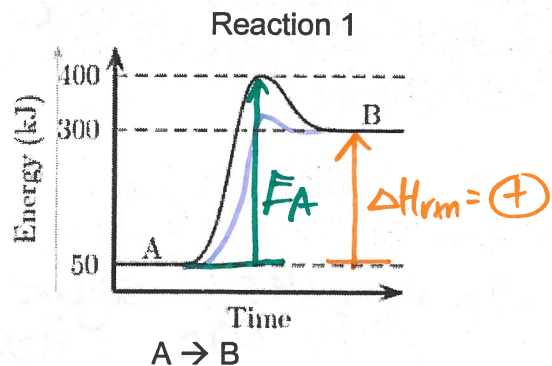
lowering the activation energy... (hill top)

Catalyst Effect on Reaction Path

Why do catalysts increase reaction rate?



Interpret the following potential energy diagrams



1. Which reaction is endothermic? ①

2. What is the activation energy of Reaction 1? ~~350~~ $E_a = 350$ kJ

3. What is the ΔH_{rxn} of reaction 1? \oplus 250 kJ

4. What is the activation energy of reaction 2? 100 kJ

5. What is the ΔH_{rxn} of reaction 2? \ominus 100 kJ

6. Sketch the effect of a catalyst on both reactions

7. Does a catalyst effect the ΔH_{rxn} ?

NO!!!

Biologically (in terms of biochemistry), enzymes are catalysts.

How are they usually used? help break stuff DOWN or
help build new stuff UP

_____ is the process by which life is _____.

Metabolism

Anabolism

(meaning "to build up")

Anabolic ("constructive") processes are when larger, more complex biomolecules are created from simpler ones.

Example #1: When your body is building muscle proteins using amino acid.

↳ peptides

Example #2: **PHOTOSYNTHESIS**
Plants use H₂O & CO₂ to create glucose - C₆H₁₂O₆, which is a larger molecule.

Anabolism = Energy required

Catabolism

(meaning "to break down")

Catabolic ("destructive") processes are when large, complicated molecules are broken down into smaller ones biomolecules.

Example #1: **RESPIRATION & GLYCOLYSIS**: Your body breaks down carbohydrates / sugars to produce energy, CO₂ & H₂O.

Example #2: When you exercise, fats are broken down.

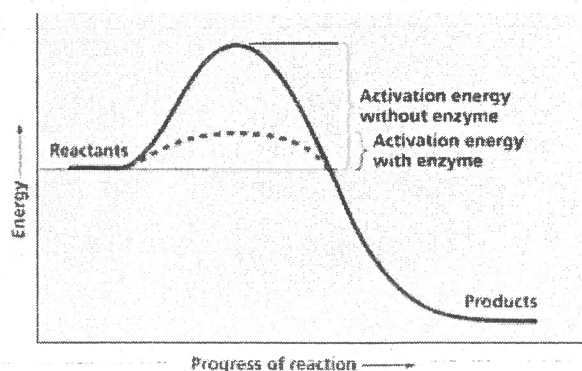
Catabolism = Energy released

Enzymes (which are biological catalysts) speed up reactions, but they are NOT consumed in the reaction.

They are proteins, which means they're made of amino acids.

Enzymes usually end in the suffix -ase.

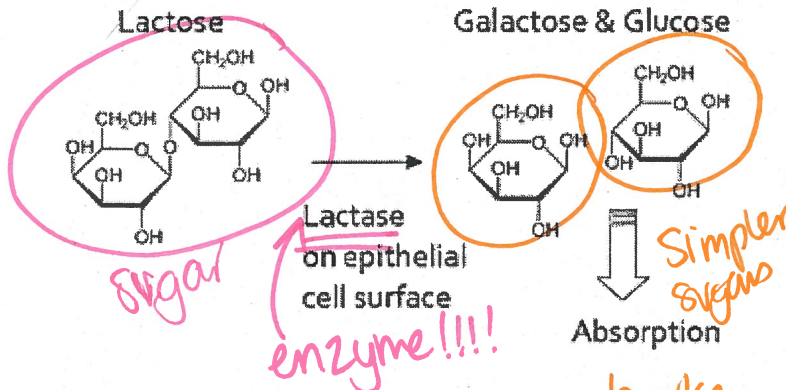
Carbohydrates often end in -ose.



Lactose Intolerance:

Lactose is a sugar that is found naturally in dairy products.

People are who lactose intolerant have a difficult time breaking down the lactose molecule.



What suffix do carbohydrates often have?

-ose

If you see dextrose, maltose, or sucrose on a food label, would you call them fat? Sugar? Or protein?

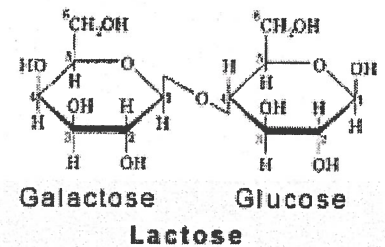
Sugars

1/3 Lactose is a 2-ring sugar. It must be broken by the enzyme lactase.

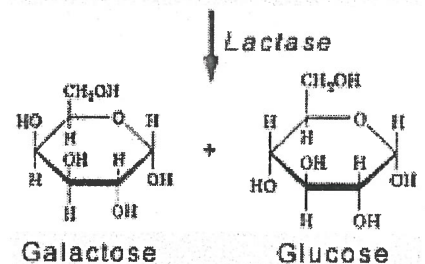
1/4 This will turn it into 2 simpler sugars, which the body can then use.

People who are *lactose intolerant* don't have enough lactase

enzymes in their gut to break down the lactose sugar.



0/0 **THINK:** What kinds of problems result from consuming lactose with insufficient or non-existent gut lactase?



Other Enzymes - Research their bodily functions on your own:

DNA Helicase _____

DNA Polymerase _____

Amylase _____

Protease ("PRO-tee-ase") _____

Lipase _____

Chemistry Unit 6

Primary reference: **CHEMISTRY, Addison-Wesley**

Topic	Essential Knowledge	Study Support
Atomic Structure 2.6 SOL 2f	Electronegativity is the measure of an atom's attraction for electrons in a bond. Electronegativity increases across a period toward the halogens and decreases down a group. The most electronegative atom is fluorine. The least electronegative element (excluding noble gases) is Francium, Fr.	Ch 14: Read p. 405
Nomenclature, Formulas, and Reactions 3.6 SOL 3d, 3e, 3f	<p>Exothermic reactions release heat whereas endothermic reactions absorb heat. Heat of reaction is the amount of energy absorbed or released during a chemical change. Exothermic reactions have a negative ΔH_{rxn}, whereas endothermic reactions have a positive ΔH_{rxn}. Examples of writing an exothermic reaction equation are:</p> $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + 890 \text{ kJ}$ <p style="text-align: center;">or</p> $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}, \Delta H_{rxn} = -890 \text{ kJ/mol}$ <p>Polar covalent bonds form between elements with very different electronegativities. The more electronegative atom will attract the electrons more strongly and this will result in it having a slight negative charge. The less electronegative atom then takes on a slight positive charge. A non-polar covalent bond forms between atoms of similar electronegativities. A polar molecule has unequally distributed electrons around the central atom. This is caused by unsymmetrical polar bonds or a lone pair on the central atom. The positive end of the molecule has a positive dipole and the negative end has a negative dipole. Polar molecules have dipole-dipole intermolecular attractions as well as London dispersion intermolecular attractions. Non-polar molecules only have London dispersion intermolecular attractions. with O-H, N-H or F-H bonds have hydrogen bonding attractions.</p> <p>Kinetics is the study of reaction rates. increase with increased temperature, reactant concentration, increased and the use of a catalyst. Activation minimum energy needed to initiate a High activation energies correspond to rates. Catalysts speed up reactions decreasing the activation energy. Potential energy diagrams are used to analyze reaction energy changes.</p> <div style="text-align: center;"> <p>EXOTHERMIC</p> </div>	<p>Ch 11: Read pp. 303-304</p> <p>Ch 16: Read pp. 460-466</p> <p>Ch 19: Read pp. 533-538.</p>
Molar Relationships 4.6	Stoichiometry can be combined with heat of reaction, ΔH_{rxn} , to calculate the amount of heat produced from a known amount of reactant.	
Phases of Matter and Kinetic Molecular Theory 5.6 SOL 5b, 5c, 5d	<p>Forces of attraction (intermolecular forces) between molecules determine their state of matter at a given temperature. Forces of attraction include hydrogen bonding, dipole-dipole attraction, and London dispersion (van der Waals) forces. Vapor pressure is the pressure of the vapor found directly above a liquid in a closed container. When the vapor pressure equals the atmospheric pressure, a liquid boils. Volatile liquids have high vapor pressures, weak intermolecular forces, and low boiling points. Nonvolatile liquids have low vapor pressures, strong intermolecular forces and high boiling points. Sublimation is the phase change from solid to gas without passing through the liquid phase. A substance's triple point, is the pressure and temperature conditions where all three phases coexist in dynamic equilibrium. The following mathematical relationship between the pressure, volume and temperature of a gas is used to describe the behavior of gases:</p> $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$ <p>An Ideal Gas does not exist, but this concept is used to model gas behavior. A Real Gas exists, has intermolecular forces and particle volume, and can change states. The Ideal Gas Law states that</p> $PV = nRT$ <p>R is the ideal gas law constant and has two values depending on the pressure units. They are $R = 8.314 \text{ L}\cdot\text{kPa}/\text{mol}\cdot\text{K}$ and $R = 0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$</p> <p>Dalton's Law of Partial Pressures says the sum of the partial pressures of all the components in a gas mixture equals the total pressure of the gas mixture.</p> $P_{\text{total}} = P_A + P_B + P_C \text{ and } \text{moles}_A/\text{moles}_{\text{total}} = P_A/P_{\text{total}} = V_A/V_{\text{total}}$ <p>Graham's Law says gas molecules with the lightest mass travel fastest.</p>	<p>Ch 10: Read pp. 269-280 and pp. 284-286.</p> <p>Ch 12: Read pp. 350-353.</p>

Unit 6 Objectives
Chemistry, Addison-Wesley, 2002

- I) Endothermic and Exothermic Reactions
 - A) Classifying Reactions
 - B) Stoichiometry and Calculating Heats of Reaction
- II) Intermolecular Forces (IMFs)
 - A) Polar bonds
 - B) Polar molecules
 - C) Intermolecular Attractions and Physical Properties
 - 1) Intermolecular forces
 - (a) London Dispersion forces
 - (b) Dipole-Dipole attractions
 - (c) Intermolecular Hydrogen bonding
 - 2) Effect of Intermolecular Forces on Physical properties
 - 3) Comparing molecular and ionic compounds
- III) Phase Changes and Intermolecular Forces (IMFs)
 - A) Kinetic Energy, Particle Velocity, and Kelvins
 - B) Kinetic Energy and Liquids
 - 1) Vapor pressure
 - 2) Boiling points and atmospheric pressure
 - C) Kinetic Energy and Solids
 - D) Phase Changes and Phase Diagrams
- IV) Gas Laws: Combined, Ideal, Dalton's Law and Graham's Law
- V) Reaction Rates
 - A) Collision Theory
 - B) Potential Energy Diagrams
 - 1) Activation Energy
 - 2) Catalysts

(SOL) Learning Objective

- 1. (3e) Identify a reaction as endothermic or exothermic based on its thermochemical equation and/or sign of ΔH .
- 2. (4b) Calculate the heat change of a reaction using stoichiometry and heats of reaction.
- 3. (2f) Compare the electronegativity of two elements based on their position on the periodic table.
- 4. (3d) Compare relative bond polarity based on the two elements position on the periodic table.
- 5. (3d) Use VSEPR theory and electronegativity to identify polar and non-polar molecules.
- 6. (5d) Identify and compare the three types of intermolecular forces (dipole interaction, hydrogen bonding, London dispersion (van der Waals) forces)
- 7. (5d) Predict the relative melting and boiling points of molecular and ionic substances based on intermolecular forces.
- 8. (5d) Explain the relationship between kinetic energy and temperature
- 9. (5b) Explain why real gases condense whereas ideal gases do not condense using IMFs and kinetic energy.
- 10. (5b) interpret vapor pressure graphs.
- 11. (5d) Explain what happens as a solid melts using IMFs and kinetic energy.
- 12. (5d) Explain the relationship between a substance's vapor pressure and boiling point and the strength of the substance's IMFs.
- 13. (5d) Interpret the effect of temperature and pressure on states of matter using a phase diagram.
- 14. (5d) Identify the triple point on a phase diagram and identify which states of matter exist at the triple point.
- 15. (5d) Identify phase changes on a phase diagram of water including fusion, solidification, vaporization, condensation and sublimation.
- 16. (5b) Use Dalton's Law to calculate partial pressures
- 17. (5b) Use Graham's Law to compare rates of effusion and diffusion of two gases
- 18. (3f) Draw a reaction's potential energy diagram with axes labeled, and ΔH , activation energy, product energy, reactant energy, transition state, and catalyst shift clearly identified for exothermic and endothermic reactions.
- 19. (3f) Explain how a catalyst increases reaction rate.
- 20. (3f) Identify and explain the effect the following factors have on the rate of a chemical reaction: (catalyst, temperature, concentration, and reactant particle size).

Chapter 11 Part 2: Endothermic and Exothermic Reactions

1 of 4

Classifying Reactions as Endothermic or Exothermic

Exothermic Reactions produce (releases) heat ($-\Delta H$)

Endothermic Reactions absorbs heat ($+\Delta H$)

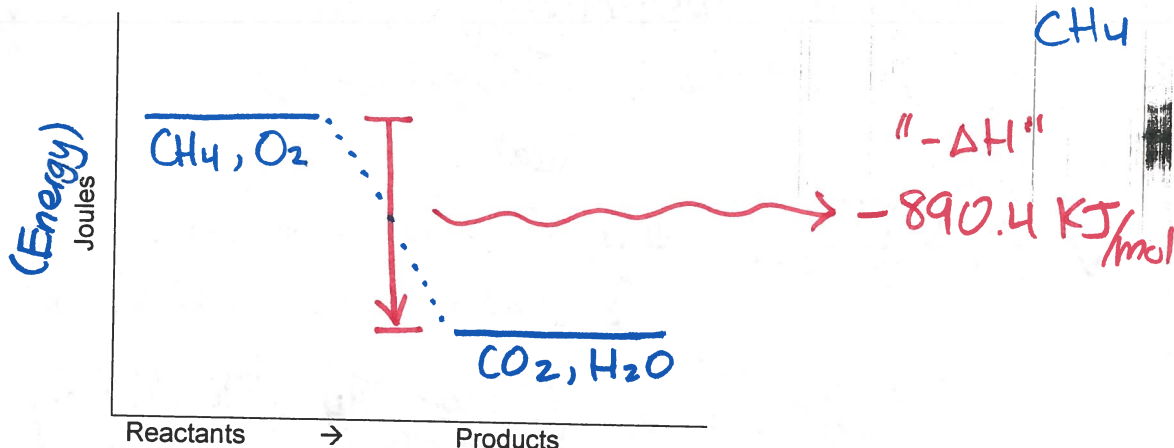
Heat of Reaction, $\Delta H_{\text{reaction}}$ is the heat absorbed or released by a reaction.

$+\Delta H_{\text{rxn}}$ endo $-\Delta H_{\text{rxn}}$ exo

Thermochemical Equations are balanced stoichiometric equations showing the heat change.

Example one: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + 890.4 \text{ kJ}$ exo so $\Delta H =$ -890.4 kJ/mol

Potential Energy Diagram



The system releasing heat so ΔH is negative/positive.

The reaction is endothermic/exothermic?

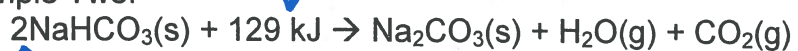
Law of Conservation of Energy: (E) is never created/destroyed
... only transformed

Energy = work + heat + potential energy

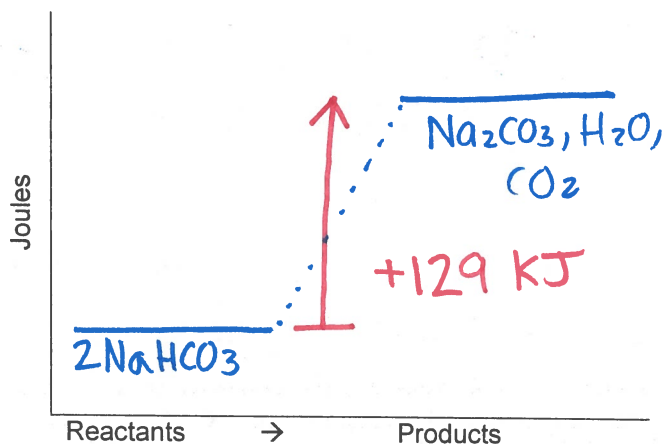
Gasoline = C_8H_{18} miles + heat + $\text{CO}_2 + \text{H}_2\text{O}$ ←
(work) CO, SO_x, NO_x

catalytic converters

Example Two:



so $\Delta H = +129 \text{ KJ}$

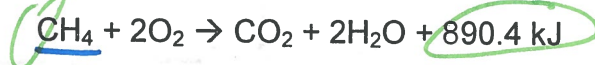


The system is absorbing heat so ΔH_{rxn} is negative/positive.

C. Heat of Reaction Calculations (Thermal Stoichiometry)

Example 1:

✓ How many kilojoules of energy are produced by burning 821 grams of methane with excess oxygen? (Ans = 45600 J)



$$\frac{821 \text{ g CH}_4}{1} \times \frac{1 \text{ mol CH}_4}{16 \text{ g CH}_4} \times \frac{890.4 \text{ KJ heat}}{1 \text{ mol CH}_4}$$

✓ How many liters of oxygen would be consumed at STP to produce 122 kJ of heat in the below reaction? (Ans = 6.14 L)



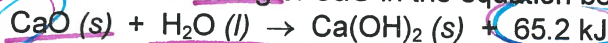
Solving Heat of Reaction Problems

3 of 4

Treat heat (in J or kJ) the same as any reactant or product in a chemical equation

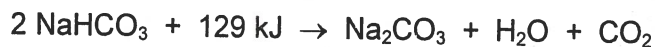
calcium oxide
M NM

1. How much heat is produced by the reaction of 25.7 g of CaO in the equation below? (Ans = 29.9 kJ)



$$\frac{25.7 \text{ g CaO}}{1} \times \frac{1 \text{ mol CaO}}{56.0 \text{ g}} \times \frac{65.2 \text{ kJ}}{1 \text{ mol CaO}} =$$

2. How many grams of NaHCO_3 are needed to react completely when 980 kJ of heat are used in the equation below? (Ans = 1300 g)



3. Using the same equation, how many kJ of heat must be used to produce 55.7 liters of CO_2 at STP? (Ans = 321 kJ)



Comparing Endothermic and Exothermic Reactions

Exothermic Reactions	Endothermic Reactions
<p>The reaction _____ heat</p> <p>Example of exothermic reaction</p>	<p>The reaction _____ heat</p> <p>Example of endothermic reaction</p>
<p>Potential Energy vs. Reaction Path for Exothermic Reactions</p>	<p>Potential Energy vs. Reaction Path for Endothermic Reactions</p>
<p>An exothermic reaction has a net _____ ΔH</p>	<p>An endothermic reaction has a net _____ ΔH</p>
<p>In exothermic reactions, the product's energy is _____ than the reactant's energy</p>	<p>In endothermic reactions, the product's energy is _____ than the reactant's energy</p>

Word bank: positive, releases, negative, absorbs, lower, higher

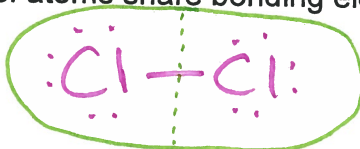
Chapter 16: Polar Bonds and Polar Molecules and Intermolecular Forces

Electronegativity: tendency of an atom to attract electrons

A. Non-Polar and Polar Covalent Bonds

1) non-polar covalent bonds: atoms share bonding electrons equally.

Example: Cl_2



2) polar covalent bonds (polar bonds): bonding electrons shared unequally.

Example: HCl



3) Electron sharing based on electronegativity differences.

- more electronegative atom attracts the electrons more closely and acquires a slight negative charge.
- less electronegative atom then acquires a slight positive charge.
- unequal sharing creates "polarized" bonds with opposite charges.
- Two ways to show polarity in structural formulas.

lower case greek deltas: δ^- & δ^+ (partial charges)

slashed arrows: (points to more electroneg. atom)

The type of bond depends on electronegativity differences between the atoms

Electronegativity Difference	Guideline: Type of Bond	Example (electronegativity difference)
<u>0.0 - 0.4</u>	Non-polar Covalent	C-H in CH_4 ($2.5 - 2.1 = 0.4$)
<u>0.4 - 2.0</u>	Polar Covalent	HF ($4.0 - 2.1 = 1.9$)
<u>>2.0</u>	Ionic	NaCl ($3.0 - 0.9 = 2.1$)

Selected Electronegativity Values

H 2.1						
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8

Is the bond polar, non-polar or ionic?

C=O in CO_2 ?	$3.5 - 2.5 = 1$	PC
Si-H in SiH_4	$1.8 - 2.1 = 0.3$	NP
C-F in CF_4	1.5	PC
N-Cl in NCl_3	$3.0 - 3.0 = 0$	NP
KCl	$3 - 0.8 = 2.2$	I

Answer this question without looking at the table on the previous page.

Which bond is most polar; C-N or C-F?



B: POLAR MOLECULES

→ Polar Molecules: One end of the molecule is slightly negative, and one end is slightly positive.

Dipole: a molecule with two poles (one negative, one positive or δ^- , δ^+)

What makes a molecule polar?

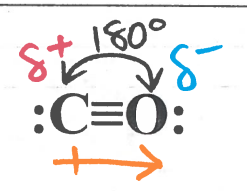
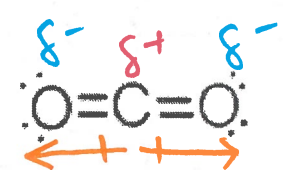
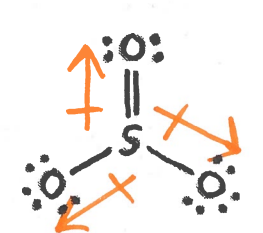
A molecule is polar if the electrons are pulled to one side of the molecule. The molecule is lopsided (assymetrical).

Determining if a molecule is polar.

1. Draw the Lewis structure
2. Determine the molecular geometry
3. Look for lone pairs on central atom (automatically polar)
4. Are there polar bonds?
5. If yes, are the polar bonds unsymmetrical in 3-D around the molecule's center?

NaCl with H₂O

← automatically BENT & automatically polar

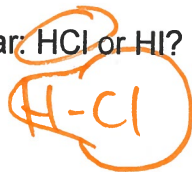
Molecule	Lewis Structure and Geometry	Polar or Non Polar
CO carbon monoxide	 <p>END = 1.0 (PC)</p>	polar
CO ₂ linear		Non-polar
SO ₃ trigonal planar	 <p>all vectors (dipoles) cancel.</p>	Non-polar

★ Will be polar if theres LP on central atom 6

Polar things dissolve in H₂O
"LIKE dissolves LIKE"

Molecule	Lewis Structure and Geometry	Polar or Non Polar
<u>SO₂</u> <i>Bent</i>		<i>polar</i>
CF ₄		<i>Non-polar</i>
CH ₂ F ₂		<i>polar</i>
NH ₃ <i>trig. pyr.</i>		<i>polar</i>
H ₂ O		<i>polar</i>

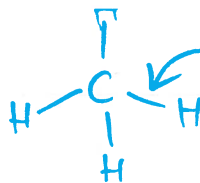
Which molecule is most polar: HCl or HI?



HW due 5/3 :
PS Unit 6,
p.1 & 2

→ Which molecule is most polar; CO₂ or SO₂?

C: INTERMOLECULAR FORCES

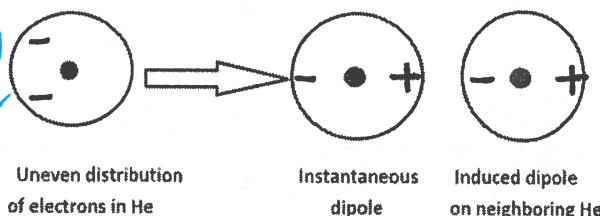


Intermolecular Forces are attractions between molecules due to three forces

LDF

1. London Dispersion Forces (weakest) temporary attractions between molecules due to temporary dipoles caused by shifting electron clouds. Dispersion forces are greater in more massive molecules with larger electron "clouds". All molecules have LDFs.

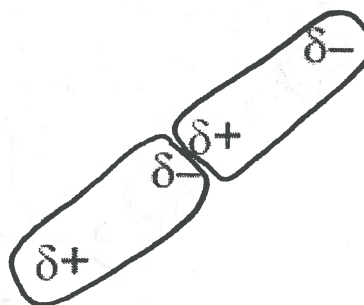
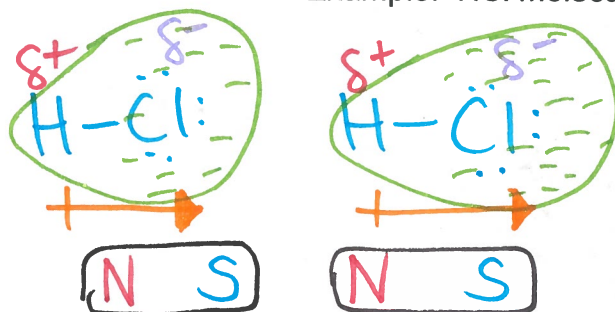
- Present in everything
- Bigger things = stronger LDF



more electrons
= stronger LDF
(Bigger = ↑ LDF)

2. Dipole-dipole attraction: polar molecules are attracted to each other. The positive dipole of one molecule is attracted to the negative dipole of another.

Example: HCl molecules



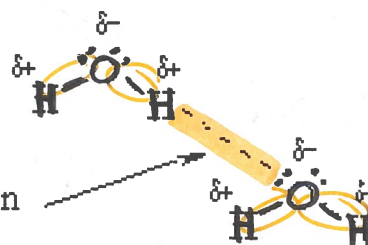
3. Intermolecular Hydrogen bonding: hydrogen that is covalently bonded to a very electronegative atom is also weakly bonded to the unshared pair of another electronegative atom.

STRONG!!!

H-N, H-O, H-F

★ The H in one molecule is attracted to the N, O, or F in another molecule(s)

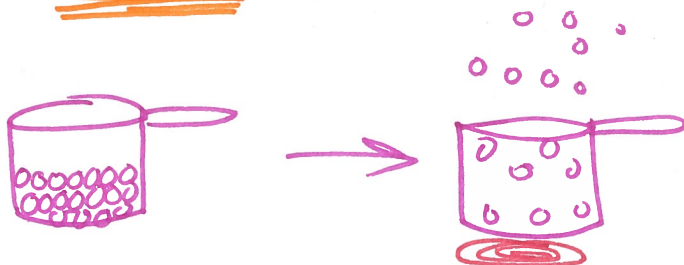
hydrogen bond



N O F
✓ ✓ ✓

4. Intermolecular Forces and Molecular Physical Properties

As intermolecular forces increases, the melting and boiling points increase because more kinetic energy is needed to overcome the IMFs between molecules.



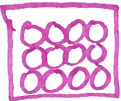
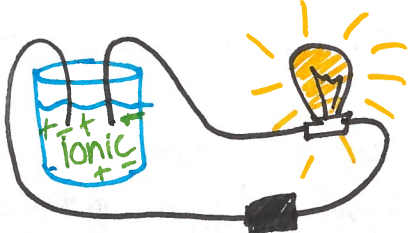
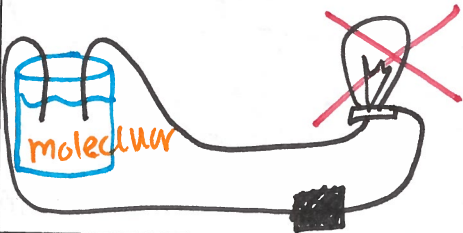
IMF broken

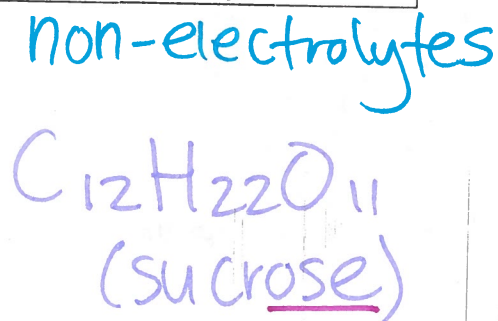
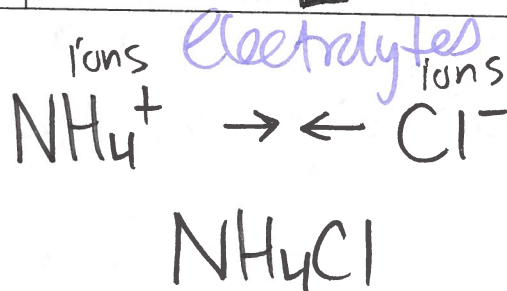
Stronger IMF
= ↑ MP & ↑ BP

Comparing Ionic and Covalent Compounds

← covalent



Characteristic	Ionic Compound	Molecular Compound
bond formation	electrostatic attraction btw ions $\oplus \ominus$	Sharing of valence e^-
Types of elements in compound	$M + NM$	$NM + NM$
Physical state at 25°C ↑ room temp	SOLID 	most: gases — liquids — solids
Melting point	high af	low
Electrical Conductivity in aqueous solution →		

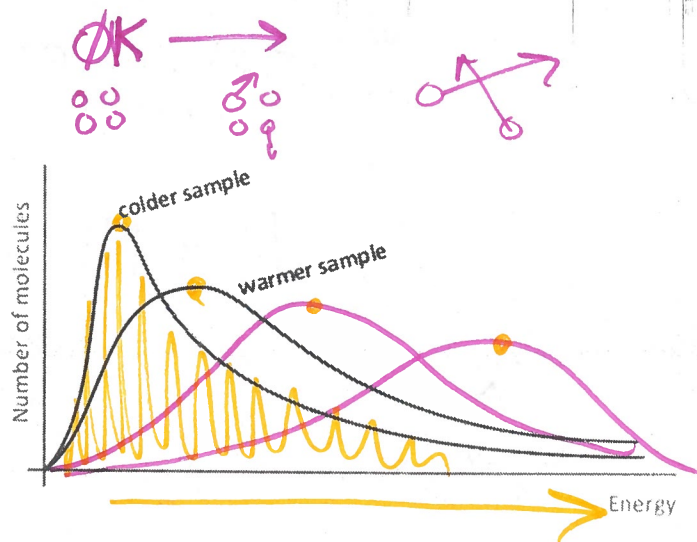
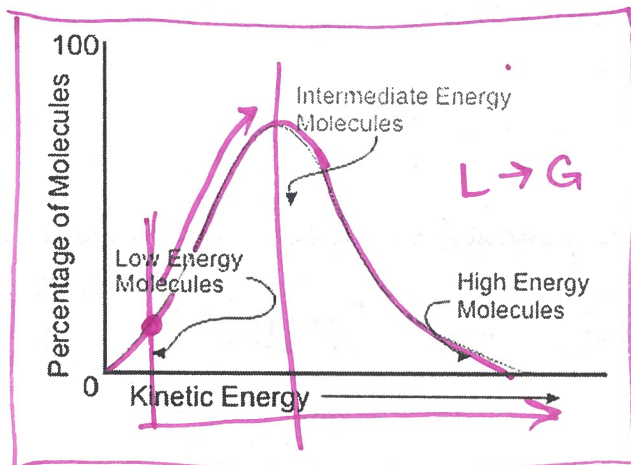


Chapter 10 Kinetic Theory, IMFs, and Phase Changes

Kinetic Molecular Theory: The tiny particles in all forms of matter are in constant motion.

I Kinetic Energy and Kelvin temperature scale

A) Temperature measures average kinetic energy

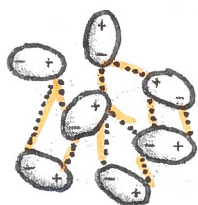


B) Gas particle's kinetic energy increases as temperature (K) rises

C) Kelvin Temperature scale is absolute & directly proportional
273 K = 0 °C

II. Kinetic Energy and Liquids

Intermolecular forces (between molecules) hold particles together in solid or liquid phases.



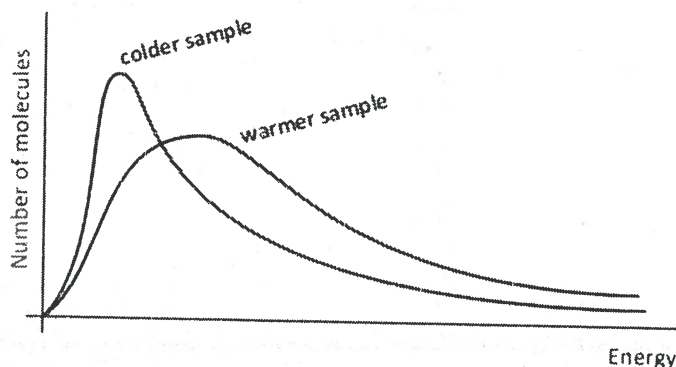
Stronger IMF take more HEAT energy. (MP & BP are higher for things w/ stronger IMF)

Kinetic energy keeps the molecules moving but not with enough energy to overcome the IMFs.

Evaporation, Vapor Pressure and Temperature

Evaporation: The L → G transition that happens BELOW the BP.

- Particles with enough kinetic energy to overcome intermolecular forces escape into gas phase
- Evaporation rate increases as temperature increases



Evaporation in a closed container produces vapor pressure.

Increasing temperature increases vapor pressure over a liquid until a dynamic equilibrium is reached.

high vapor pressure
(\therefore low IMF)

low VP
(\therefore high IMF)

Volatile liquids

evaporate quickly!

Nonvolatile liquids

evaporate slowly or not at all.

volatile

low IMF
high VP

non-volatile

high IMF
low VP

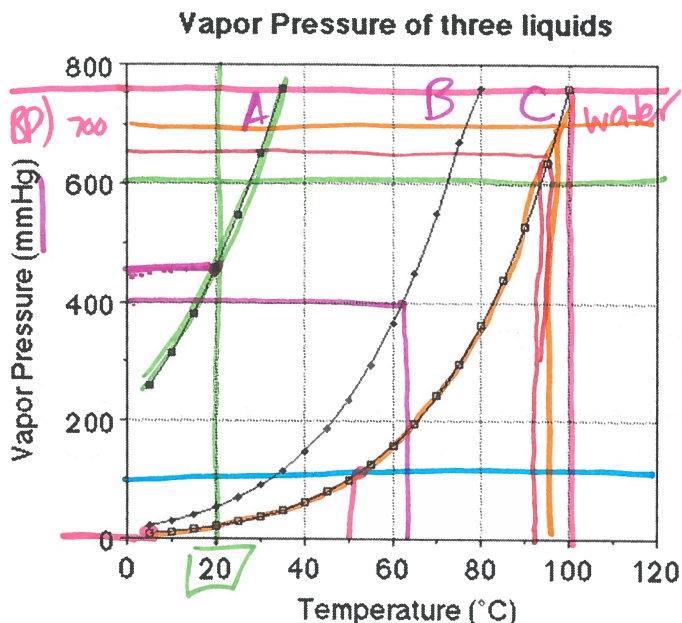
Boiling Point of a Liquid (open container)

Boiling Point: T Temp @ which enough heat is present to break ALL IMFs

$L \rightarrow G$

high IMF \rightarrow low VP \rightarrow high BP

A) Boiling point changes as external pressure changes/



What is the vapor pressure of liquid A at 20°C? 450 mmHg

What is the boiling point of B when the external pressure is 400 mmHg? 63°C

Which liquid is most volatile? A

Which liquid has the strongest intermolecular forces? C

Strong/high IMF = low VP

How hot does water need to be to boil at 100 mmHg? 50°C

What will be the boiling point of water on Pike's Peak (elevation = 14,000 ft, atmospheric pressure = 640 mmHg)? ~97°C

BP goes \downarrow when pressure goes \downarrow

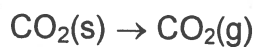
III. Kinetic Energy and Solids

melting point: temp of S \rightarrow L transition @ Normal atmosphere.

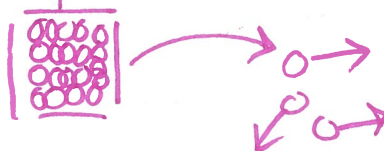
S \rightarrow L : Break most IMF that hold in place.

sublimation: S \rightarrow G (skips liquid)

Examples:

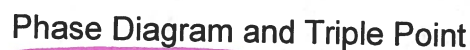


DRY ICE



F ₂	(g)
Cl ₂	(g)
Br ₂	(l)
I ₂	(s)
At	(s)

Phase Changes and terms



Borling = bubbles

★ = triple point S, L, & G
as coexisting

\dagger = critical point: L & G are indistinguishable

SKIP for now

Chapter 19 Part 1: Reaction Rates

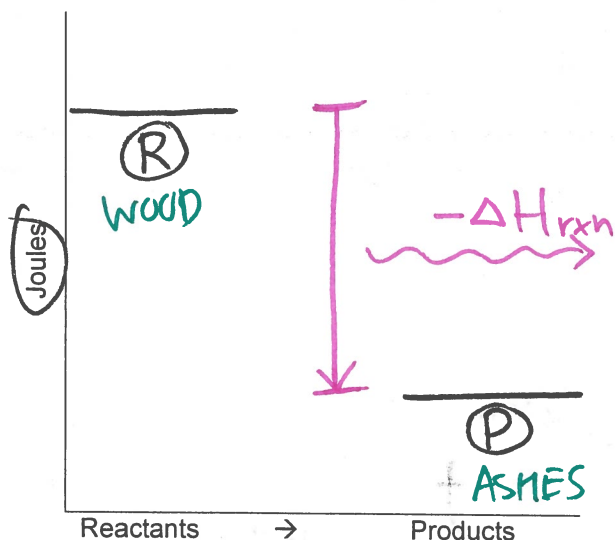
http://www.wwnorton.com/college/chemistry/gilbert/index/site_map.htm

Reaction rates are measured as mol/time units.

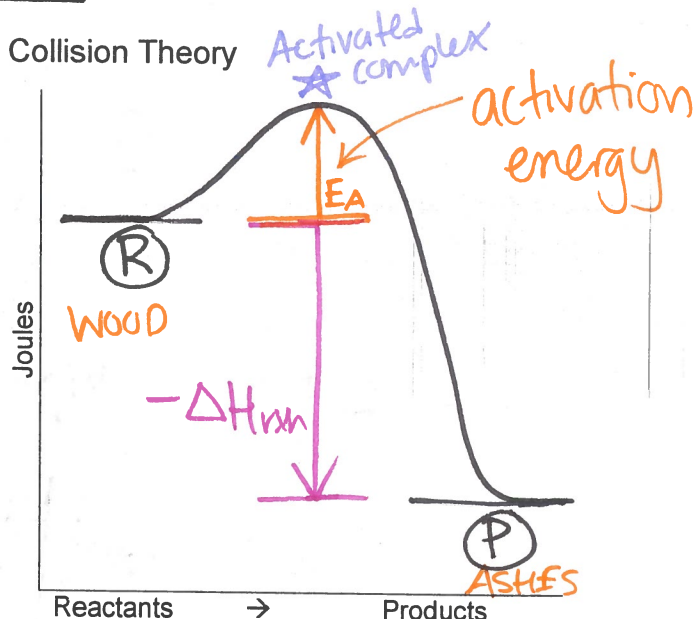
A. Collision Theory: Atoms/Molecules react w/ each other only if there's enough (E) & correct orientation

Reaction Coordinate Diagram for Exothermic Process (forward direction)

Chapter 11: Thermochemistry



Collision Theory



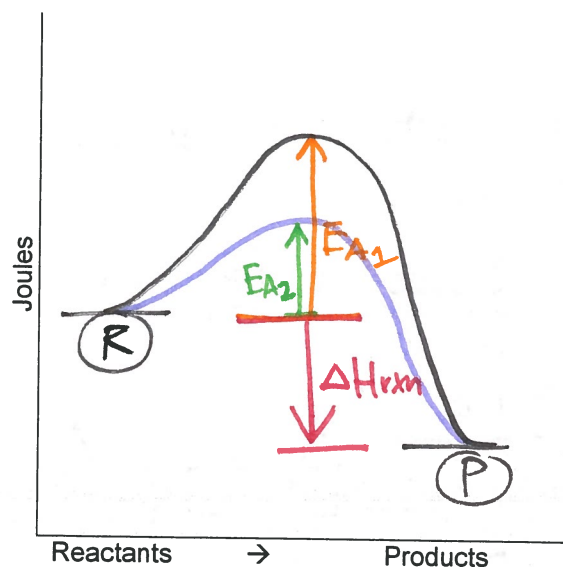
Activation Energy: (E_a) The amt of (E) needed to start a reaction. (amt of energy to complete)

Activated Activation Complex: The arrangement of atoms & molecules at highest energy stage ★

B. Factors Influencing Reaction Rate

1. Temperature \uparrow = faster molecules = more collisions = more product
faster music
2. Concentration \uparrow = increase # of molecules = more collisions
add more kids
3. Particle Size \downarrow = increase surface area = more collisions
(smaller particles collisions)
4. Catalysts \uparrow speeds up reaction (but are not reactants)

Catalyst Effect on Reaction Path



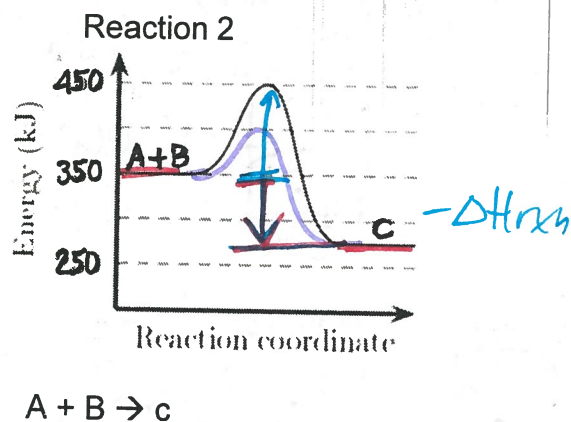
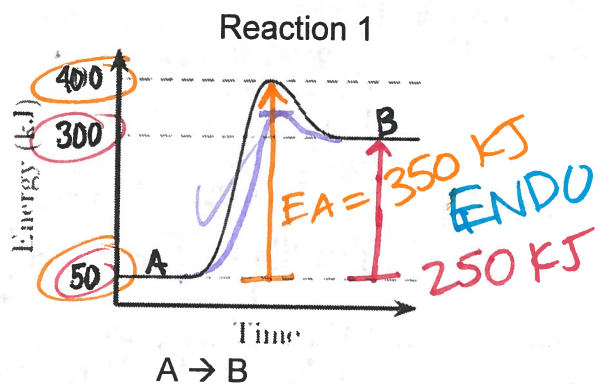
Why do catalysts increase reaction rate?

lower hill/
(lower E_A)

Examples: Catalytic Convertors

LATER

Interpret the following potential energy diagrams



- Which reaction is endothermic? 1
- What is the activation energy of Reaction 1? 350 kJ
- What is the ΔH_{rxn} of reaction 1? 250 kJ
- What is the activation energy of reaction 2? 100 kJ
- What is the ΔH_{rxn} of reaction 2? ~~100~~ -75 kJ
- Sketch the effect of a catalyst on both reactions ✓ ✓
- Does a catalyst effect the ΔH_{rxn} ? NO!

5/5/16

Chapter 12 Dalton's Law and Graham's Law

Mixtures of Gases: Dalton's Law of Partial Pressure

Partial Pressure: Pressure due to ONE gas in a confined mixture of gases (in a container)

Verbally: At constant pressure and temperature, the total pressure exerted by a mixture of gases is equal to the sum of the partial pressures of the component gases.

Math Equation: $P_{\text{Total}} = P_A + P_B + P_C \dots$

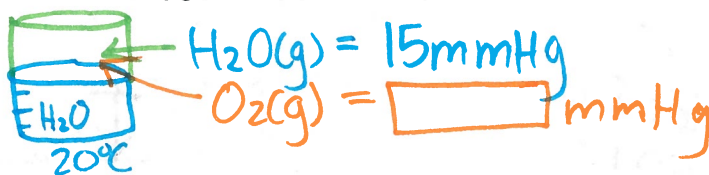
Example 1: What is the partial pressure of oxygen in air at STP (101.3 kPa) if $P_{N_2} = 79.1$ kPa, $P_{CO_2} = 0.040$ kPa and $P_{\text{others}} = 0.94$ kPa?

$$P_T = P_A + P_B + P_C + P_D \dots$$

$$101.3 \text{ kPa} = 79.1 \text{ kPa} + 0.040 \text{ kPa} + 0.94 \text{ kPa} + P_{O_2}$$

$$P_{O_2} = 21.22 \text{ kPa}$$

Example 2: A sample of oxygen gas is collected over water at 20.°C. The vapor pressure of water at 20.°C is 15 mm Hg. If the total pressure is 420 mm Hg, what is the partial pressure of the oxygen?



$$P_T = P_A + P_B$$

$$P_T = P_{H_2O} + P_{O_2}$$

$$420 \text{ mmHg} = 15 \text{ mmHg} + P_{O_2}$$

$$P_{O_2} = 405 \text{ mmHg}$$

Variants of Dalton's Law

$$\frac{\text{mole}_A}{\text{mole}_{\text{Total}}} = \frac{P_A}{P_{\text{Total}}} = \frac{\%V_A}{100 \%V_{\text{Total}}}$$

Example 1: A tank contains 6.0 moles of a mixture of hydrogen, helium, and nitrogen at 102 kPa. If there are 2.0 moles of hydrogen in the tank, what is the partial pressure of hydrogen?

$$\frac{\text{mol } H_2}{\text{mol}_{\text{Total}}} = \frac{P_{H_2}}{P_T} \rightarrow \frac{2 \text{ mol } H_2}{6 \text{ mol total}} = \frac{P_{H_2}}{102 \text{ kPa}}$$

$$P_{H_2} = 33.97 \text{ kPa}$$

Example 2: A gas cylinder contains 8.0 moles of argon, 2.0 moles of nitrogen, and 2.0 moles of oxygen at 600. mmHg. What is the partial pressure of nitrogen in the cylinder? (Ans = $1.0 \times 10^2 \text{ mmHg}$)

$$\frac{2 \text{ mol } N_2}{12 \text{ mol total}} = \frac{P_{N_2}}{600 \text{ mmHg}}$$

$$P_{N_2} = 100 \text{ mmHg}$$

Example 3: A mixture of gases with a pressure of 950 mm Hg contains 20% hydrogen and 80% neon by volume. What is the partial pressure of neon gas in the mixture? (Ans = 760 mmHg)

$$\frac{P_{H_2}}{P_{Ne}} = \frac{20\% H_2}{100\% \text{ total}} = \frac{P_{H_2}}{950 \text{ mmHg}} \rightarrow 190 \text{ mmHg } H_2$$

$$P_{Ne} = 760 \text{ mmHg} \checkmark$$

Example 4: In a mixture of oxygen and nitrogen gas, 70.0 percent of the total gas pressure is exerted by the nitrogen. If the total pressure is 150 kPa, what pressure does oxygen exert? (Ans = 45 kPa)

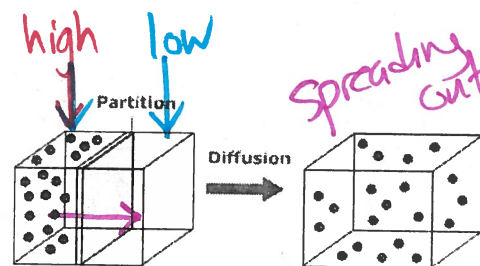
$$\frac{70\% N_2}{100\% \text{ total}} = \frac{P_{N_2}}{150 \text{ kPa total}}$$

$$P_{N_2} = 105 \text{ kPa}$$

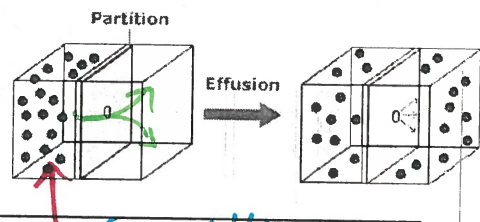
$$P_{O_2} = 45 \text{ kPa}$$

Graham's Law of Effusion

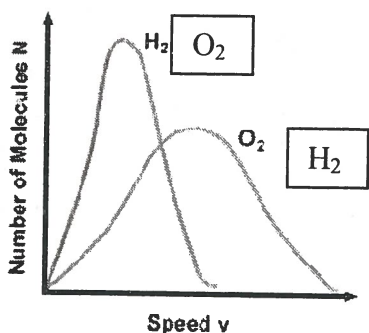
Diffusion: Gas particles move from HIGH concentration to LOW conc.



Effusion: Gas particles escape from a hole with certain velocity (rate m/s)



Graham's Law Verbally: Gas molecules with the lightest mass travel fastest.



O_2 vs H_2
 32 g/mol vs 2 g/mol
 move fastest!

4 g/mol, 44 g/mol, 131.3 g/mol
 slowest

Which gas will escape slowest from a tiny hole in a balloon; He, C_3H_8 , or Xe?

3rd O_2 ← 32 g/mol
 2nd N_2 ← 28 g/mol
 4th Ar ← 39.9 g/mol
 1st He ← 4 g/mol

