

Chemistry Unit 6
Primary reference: CHEMISTRY, Addis

B3 Sprt

Topic	Primary reference: CHEMISTRY, Addison-Wesley	
Atomic	Eccontial Vnovided	Study Suppor
Structure	Electronegativity is the measure of an atom attraction for electrons in a bond. Electronegativity	
2.6	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
SOL 2f		1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1
Nomencia	Exothermic reactions release heat whereas endothermic reactions absorb heat. Heat of reaction is the amount of energy absorbed or released during a charginal charge of the second during a charginal charge.	
ture,	the amount of energy absorbed or released during a chemical change.	Ch 11: Read
Formulas,	LAOUTETTIIC TEACTIONS Have a Medative AH Whereas and otherwise resetting to	pp. 303-304
and	. Source mile reaction equation are:	
Reactions		
	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_{20} + 890 \text{ kJ}$	
3.6	Or Or	Ch 16: Read
CO1 24	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$, $\Delta H_{rxn} = -890 \text{ kJ/mol}$	pp. 460-466
SOL 3d,	Polar covalent honds form between elements in	pp. 100 400
3e, 3f	Polar covalent bonds form between elements with very different electronegativities. The more electronegative atom will attract the electrons more electronegative.	
=	charge. The less electronegative atom then takes on a click and this will result in it having a slight negative	
	form between atoms of similar electronegativities.	
	A polar molecule has unequally distributed electrons around the	
	positive dipole and the negative end has a negative dipole. Polar molecules have dipole-dipole intermolecular attractions as well as londer dispersion in the positive end of the molecule has a intermolecular attractions as well as londer dispersion in the positive end of the molecule has a intermolecular attractions as well as londer dispersion in the positive end of the molecule has a londer dispersion in the positive end of the molecule has a londer dispersion in the positive end of the molecule has a londer dispersion in the positive end of the molecule has a londer dispersion in the positive end of the molecule has a londer dispersion in the positive end of the molecule has a londer dispersion in the positive end of the molecule has a londer dispersion in the positive end of the molecule has a londer dispersion in the positive end of the molecule has a londer dispersion in the positive end of the molecule has a londer dispersion in the positive end of the molecule has a londer dispersion in the positive end of the molecule has a londer dispersion in the positive end of the molecule has a londer dispersion in the molecule has a londer dispers	
	only have London dispersion only have London dispersion intermolecular attractions. Non-polar molecules	
	intermolecular attractions. EXOTHERMIC	
	with O-H, N-H or F-H bonds have activated Molecules	
	activation	Ch 19: Read
	Product energy Ea	pp. 533-538.
	Kinetics is the study of reaction rates, increase with increased temperature, (kJ) Reaction rates	
	reactant concentration, increased \ \ \ \ \ \ \ \ \ \ \ \ \ \ \ \ \ \ \	
	and the use of a catalyst. Activation Surface area	
	minimum energy needed to initiate a energy is the	
	High activation energies correspond to reaction.	,
Molar		
Relationsh	Stoichiometry can be combined with heat of reaction, ΔH_{pxn} , to calculate the amount of heat produced from a known amount of reactant.	
ips	and an editorial	
4.6		
Phases of	Forces of attraction (intermolecular forces) between molecules determine their state of matter at a given temperature. Forces of attraction include bydrogen bending disabilities their state of matter at a given	
Matter		Ch 10: Read
and	(van der Waals) forces.	pp. 269-280
Kinetic	Vapor pressure is the pressure of the vapor found directly above a liquid in a closed container. When the	and pp. 284-
Molecular	vapor pressure equals the atmospheric pressure, a liquid boils. Volatile liquids have high vapor pressures,	286.
Theory	intermolecular forces and high boiling points. Sublimate inquids nave low vapor pressures, strong	
5.6	passing through the liquid phase. A substance's triple point , is the phase change from solid to gas without where all three phases coexist in dynamic equilibrium.	
COL EL	where all three phases coexist in dynamic equilibrium.	
SOL 5b,	The following mathematical relationship between the pressure, volume and temperature of a	
5c, 5d	Sale as asserbe the behavior of gases:	
	$\underline{P_1}\underline{V_1} = \underline{P_2}\underline{V_2}$	
	An Ideal Coards	Ch 40 0
	An Ideal Gas does not exist, but this concept is used to model gas behavior. A Real Gas exists,	Ch 12: Read
	has intermolecular forces and particle volume, and can change states. The Ideal Gas Law states	pp. 350-353.
	PV = nRT.	
	R is the ideal gas law constant and has two values depending on the pressure units.	1
	They are $R = 8.314 \text{ L'kPa/mol'K}$ and $R = 0.0821 \text{ L'atm/mol'K}$	
1		
1	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.	
	and proceed of the gas mixture.	1
	$P_{total} = P_A + P_B + P_C$ and $n_A/n_{total} = P_A/P_{total} = V_A/V_{total}$	
1		
	Graham's Law says gas molecules with the lightest mass travel fastest. $[K.E = 0.5 \text{ mv}^2]$	
	VIII C.O — J.VIII	

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

 - Vapor pressure
 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

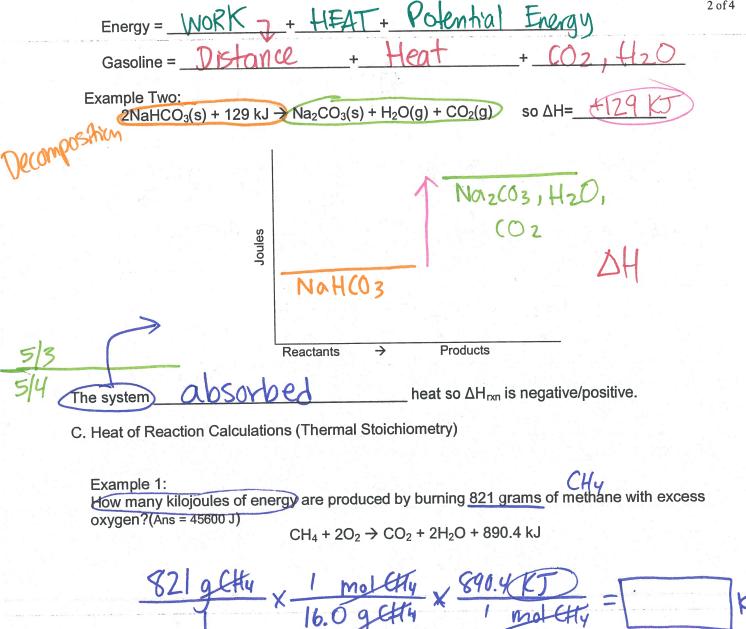
(SOL) Learning Objective

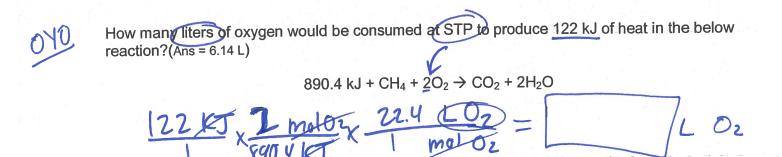
- (3e) Identify a reaction as endothermic or exothermic based on its thermochemical equation and/or sign of ΔH .
- (4b) Calculate the heat change of a reaction using stoichiometry and heats of reaction.
- (2f) Compare the electronegativity of two elements based on their position on the periodic table.
- (3d)Compare relative bond polarity based on the two elements position on the periodic table.
- (3d) Use VSEPR theory and electronegativity to identify polar and non-polar molecules.
- (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) Indentify 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

lassifying Reactions	as Endothermic or Exothermic
Exothermic Reactions	as Endothermic or Exothermic
Endothermic Reaction	
Heat of Reaction, ΔH + ΔH _{rxn}	reaction is the heat absorbed or released by a reaction. -ΔH _{rxn}
Thermochemical Equ	ations $\frac{110 \text{ reg. ean. but w}}{A + 30 \text{ KJ} \Rightarrow 2D + C}$
There are two genera	I ways they're written. Both are acceptable:
Heat as a Product (exo) of	or Reactant (endo): $A+B \rightarrow C+20 \text{ kJ}$
Heat shown as a change -ΔH means heat is los +ΔH means heat was	in Enthalpy (ΔH) $A + B \rightarrow C$ $\Delta H = -20 KJ$ absorbed (endothermic)
	Potential Energy Diagram (hast of combustion) always $CH_{4} + O_{2}$ $\Delta H = -890.4 kJ$
	$CO_2 + H_2O$
	Reactants → Products
e system <u>(eQQSe</u>	heat so ΔH is negative/positive.
The reaction is endothe	ermic/exothermic?
Law of Conservation of	f Energy:









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) CaO (s) + H_2O (l) \rightarrow Ca(OH)₂ (s) + 65.2 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)

2 NaHCO
$$_3$$
 + 129 kJ \rightarrow Na $_2$ CO $_3$ + H $_2$ O + CO $_2$

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)

2 NaHCO₃
$$\rightarrow$$
 Na₂CO₃ + H₂O + CO₂ Δ H_{rxn} = +129 kJ/mol



Comparing Endothermic and Exothermic Reactions

Exothermic Reactions	Endothermic Reactions
The reactionheat Example of exothermic reaction	The reaction heat Example of endothermic reaction
$A+B \rightarrow AB + heat$	$CD + heat \rightarrow C + D$
Potential Energy vs. Reaction Path for Exothermic Reactions	Potential Energy vs. Reaction Path for Endothermic Reactions
EA A+B AHrxn C AB	C+D EA CH CH CH CH CH CH CH CH CH C
An exothermic reaction has a net ΔH	An endothermic reaction has a net
In exothermic reactions, the product's energy is than the reactant's energy	In endothermic reactions, the product's energy is than the reactant's energy

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

EA = Activation Energy DH = amt of energy lost/gained

Chapter 16: Polar Bonds and Polar Mole	1 of 6 ecules and Intermolecular Forces
Electronegativity: Tendency of or attract e-	an atom to "hog"
A. Non-Polar and Polar Covalent Bonds 1) non-polar covalent bonds: atoms share bonding DRAW Example: Cl ₂ He-	Le-cloud
polar covalent bonds (polar bonds): bonding election	ctrons shared unequally.
S Example: HCI S+ H—CI:	less move
Electron sharing based on electronegativity difference a) more electronegative atom attracts the electronegative charge.	ons more closely and acquires a slight
 b) less electronegative atom then acquires a slig c) unequal sharing creates "polarized" bonds with d) Two ways to show polarity in structural formula 	th opposite charges
lower case greek deltas:	+) 8+ "partially (1)"
slashed arrows :	
The type of bond depends on electronegativity differ	ongos between the starre
Electronegativity Difference Guideline: Type of Bond	Example (electronegativity
>0.0 – 0.4 Non-polar Covalent	C-H in CH ₄ 2 C 2 1 - 10 11
0.4 – 2.0 Polar Covalent	HF 11.0 2.1 = [1.0]
>2.0 — onic	NaCl $3.0 - 0.9 = [2.17]$
Selected Electronegativity Values	Is the bond polar, non-polar or ionic?
H 2.1 Non-metals High	C=O in CO_2 ? $2.5-3.5 = 1.0$ PC
Li Be B C N O F 1.0 1.5 2.0 2.5 3.0 3.5 4.0	Si-H in SiH ₄ 1.8-2.1 = 0.3 NPC
Na Mg Al Si P S Cl 0.9 1.2 1.5 1.8 2.1 2.5 3.0	C-F in CF ₄ 2.5-4.0 = 1.5 PC
K Ca Ga Ge As Se Br 0.8 1.0 1.6 1.8 2.0 2.4 2.8	N-Cl in NCl ₃ $3.0-3.0 = 0$ NPC
1.0 2.0 2.4 2.0	1500 501 0

KCI

LOW LOW

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

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



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

<u>Dipole</u>: a molecule with two poles (one negative, one positive or δ -, δ +)

What makes a	A molecule is polar if the electrons are pulled to one side of the
molecule polar?	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
СО	st.C≡O: 8+	Polar
CO ₂	O=C=O: A Has 2 polar bonds But they're opposite legral	Mondar
SO ₃	Peam A Team A Team A Team A Team A	Hoyar





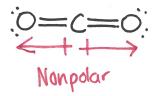
Molecule	Louis Structure and Connect	
SO ₂	Lewis Structure and Geometry	Polar or Non Polar
302		
		01.6
	:0-8-0. 0. 1\0.	() d() v
		YOU
	<u>a i kazalen 22 o</u> g 1 bilgija priber at 12.	
CF ₄	E	2.4
₩ w 1.	·F-C-F:	NON
		MINNI
	·F-C-F:	110,100
	F F	, Oplan
Transfer of	it:	(V)
CH ₂ F ₂	. 4	- 4.3
	H 8+ H	Y I I
	• • • • • • • • • • • • • • • • • • • •	=/01
	: F-C-F:	0/00
	H	No.
	H T	
NH ₃		
INFI3		2 11 10 10 1
	H-N-H	A Ch
and a		$\sim 10^{\circ}$
	H Miller H	
	H + W	
H ₂ O	H (+	
	1, 9,	
	H=0=H	Jav
A .		ω_{α}
100	8+4/0	Byan

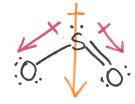
Which molecule is most polar: HCl or HI?

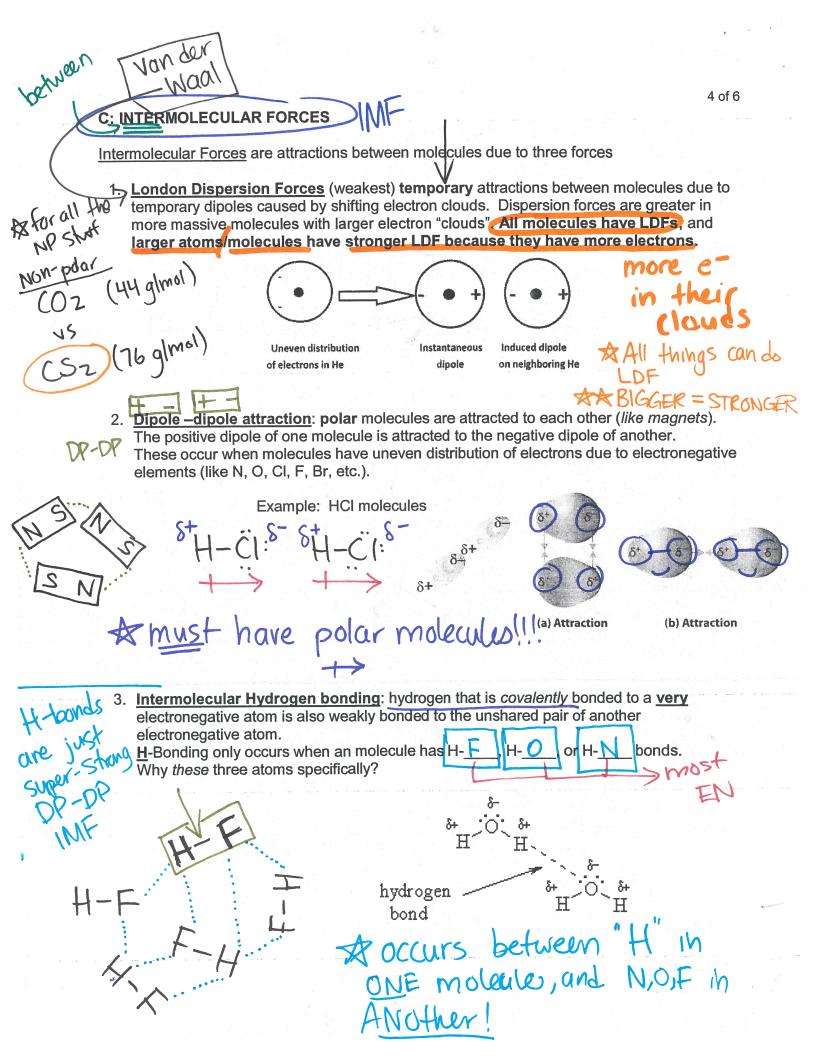


٧S

Which molecule is most polar; CO_2 or SO_2 ?) $CO_2 = 16$ ve $SO_2 = 18$ ve







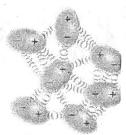
More Stronger (MF -> harder to separate. 5 of 6 4. Intermolecular Forces and Molecular Physical Properties More Stronger (MF -> harder to separate. 5 of 6

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 th	e Halogens and t	heir Physical States:		
1. 15 M	The Halogens are (Lewis dia		:X-X:		
* char		- 200	all pure X2 (halog	ens) are No	N-POLAR
, T	Are pure diatomic	halogen elemen	ts polar? Explain & draw dipol	e vectors & partial ch	arges.
	What is the only I	NO!!!	alan Main and A		
	vinacis the only i	wir triat air non-p	olar things can do?		
		Halogen (X)	Molar Mass (X ₂)	State at Room Te	mp.
			$F_2 = 38.0$ glmol	905	27
		Cl	Cl2 = 70.9 glmol	gas	en ?
	BIGG	Br	Brz=159.8 glmol	liquid	E380253
		T	Iz= 253.8g/mol	Solid	(11)
		LDF	increases as	s size	increase

Comparing Ionic and Covalent Compounds Molecular Compound Ionic Compound Characteristic TRANSFER of e-SHARE bond formation :C三O: Na VCI: - (Nat)[Ci] · Metal (cation) Types of Non-metals elements in · Non-metal (anion) compound Physical state at 25°C Melting point relatively low (but variable) Electrical Conductivity in (insulation) aqueous solution Sometimes, something be both. nitrate Sodium (chalent

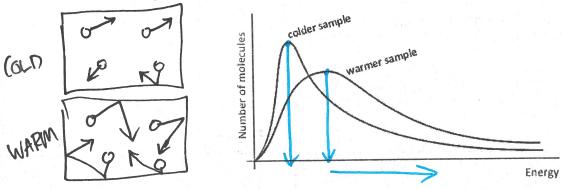
motion molecules Chapter 10 Kinetic Theory, IMFs, and Phase Changes I Kinetic Energy and Kelvin temperature scale A) Temperature measures average kinetic energy 100 colder sample Percentage of Molecules Intermediate Energy Number of molecules Molecules warmer sample Low Energy Molegules High Energy Molecules more -Kinetic Energy B) Gas particle's kinetic energy increases as C) Kelvin Temperature scale is **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	1	
Evaporation: Phase change at the non-boil	ing point.	
01 problemic		>

- 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 <u>dynamic equilibrium</u> is reached.

vopor of a vopor of a pressure pressure of a container! vopos in a container! vopos sealed container!
Volatile liquids Evaporate easily (low IMF between particles Nonvolatile liquids DOEAN't easily evaporate! (high IMF)
Nonvolatile liquids DOESNIT EASILY EVAPORATE. (MIGNIMI)
we volatile volatile of Less I vapor! when (Law mit) 00000 (high mf) 0000 (Might) - 233333333333333333333333333333333333
Boiling Point of a Liquid (open container) Reiling Point: Texas & Lilvich ALL Darticles Office SS
Boiling Point: Temp @ which All particles vess
have sufficient Knetic energy to
leave (D+0)G

Sea Level = 1 atm = 760 mm Hq (Normal)

A) Boiling point changes as external pressure changes/

Vapor Pressure of three liquids

60

Temperature (°C)

O = "normal" pressure Normal What is the vapor pressure of liquid A at

20°C?_

Which liquid represents water?

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

Which liquid is most volatile?

Which liquid has the strongest intermolecular forces?

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

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

III.Kinetic Energy and Solids

20

endothermix melting point:

100

120

80

sublimation:

800

600

400

200

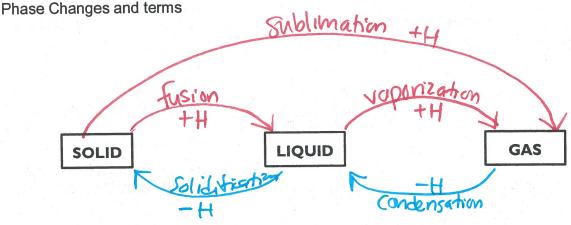
apor Pressure (mmHg)

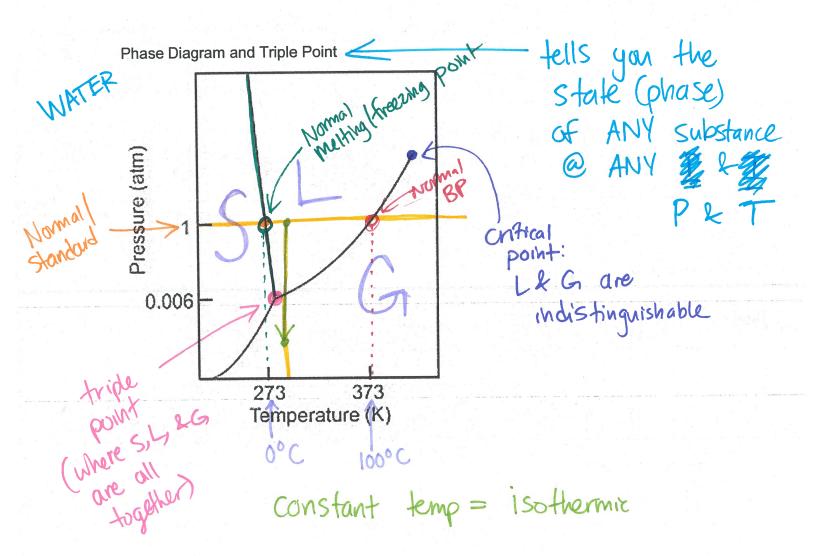
Examples:

 $l_2(s) \rightarrow l_2(g)$

Sublimation

IV. Phase Changes and Phase Diagrams

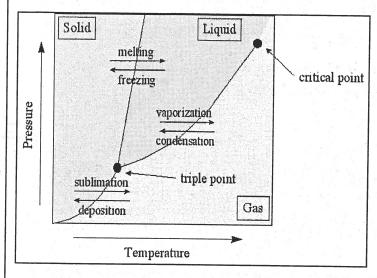




Phase Diagrams

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

The <u>CURVES</u> indicate the conditions of <u>temperature</u> and <u>pressure</u> 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

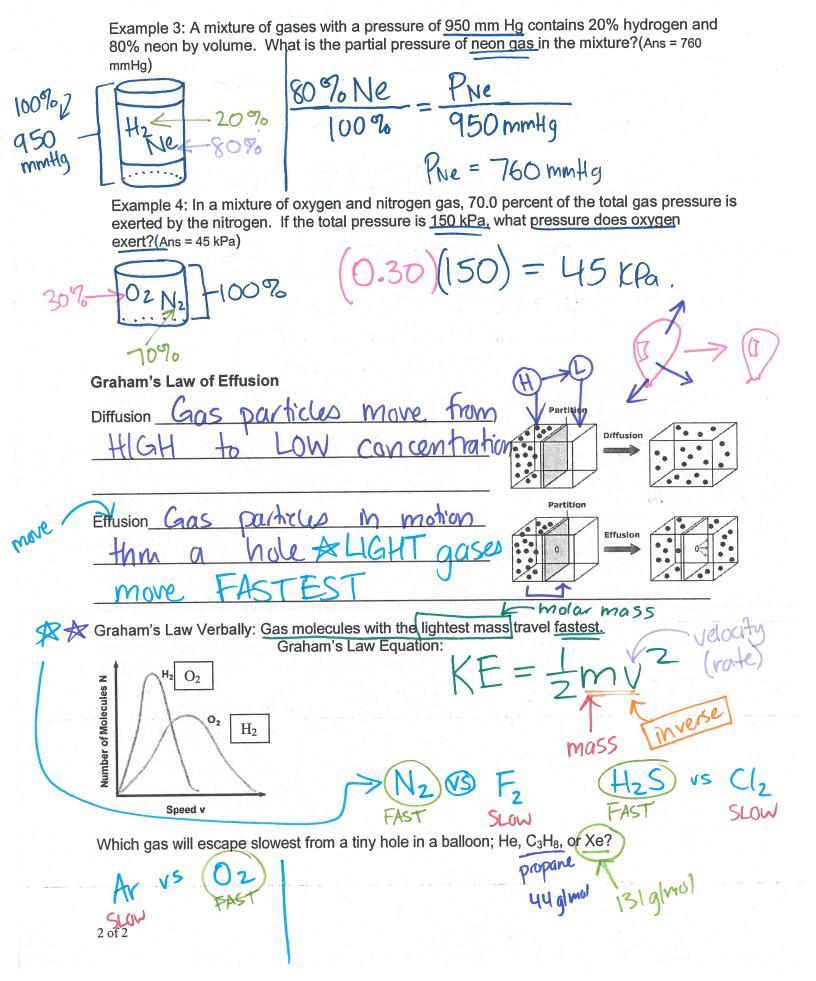
Refer to the phase diagram below when answering the questions. NOTE: "Normal" refers to STP - Standard Temperature and Pressure. Fus, then vap. A cond, then solid. 2.00 1.75 Pressure (atmospheres) 1.50 1.25 1.00 0.75 0.50 0.25 0.00 100 Temperature (degrees C)

- 1) What are the values for temperature and pressure at STP? T=
- 2) What is the <u>normal</u> freezing point of this substance? \\ \frac{100\circ}{\circ}
- 3) What is the *normal* boiling point of this substance?

	4) What is the normal melting point of this substance?
	5) What is the phase (s, I, g) of a substance at 2.0 atm and 100 °C?
	6) What is the phase (s, I, g) of a substance at 0.75 atm and 100 °C?
	7) What is the phase (s, I, g) of a substance at 0.5 atm and 100 °C?
	8) What is the phase (s, I, g) of a substance at 1.5 atm and 50 °C?
	9) What is the phase (s, I, 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=
	P_{i} $T_{i} \rightarrow \frac{573}{1}$
دما	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?
	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?
	그리고 얼마나 아무렇게 하는 아무지도 얼마나 맛이 되었다면 하는 사람들이 아무지 않는데 그 사람들이 되었다.
	14) If a quantity of this substance was at an initial pressure of 1.0 atm and a temperature of 200 ⁰ 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 ⁰ 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, <u>and</u> solid phase?
	21) At what pressure would it be possible to find this substante in the gas, inquis, and prices
	22) If I had a quantity of this substance at a pressure of 1.00 atm and a temperature of -100 ⁰ C, what phase change(s) would occur if I increased the temperature to 600 ⁰ 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)

Chapter 12 Dalton's Law and Graham's Law

Mixtu	ures of Gases: Dalton's Law of Partial Pressure
	Partial Pressure: The pressure due to a single gas In a mixture of gases.
	In a mixture of gases.
5	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 $P_{N2} = 79.1$ kPa, $P_{CO2} = 0.040$ kPa and $P_{others} = 0.94$ kPa? $P_{T} = P_{N2} + P_{CO2} + P_{others}$
	101.3 = 79.1 + Poz + 0.000 + 0.99
	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? $P_T = P_{02} + P_{H20(g)}$ $V_{120mmHg} = P_{02} + V_{15mmHg}$ ints of Dalton's Law $V_{120mmHg} = V_{15mmHg}$
025	420 mm Hg
Varia	nts of Dalton's Law Poz = 405 mm Hg
	$\frac{\text{mole}_{A}}{\text{mole}_{Total}} = \frac{P_{A}}{P_{Total}} = \frac{\%V_{A}}{100 \%V_{Total}}$
Ex.	cample 1: A tank contains 6.0 moles of a mixture of hydrogen, helium, and nitrogen at 102 a. If there are 2.0 moles of hydrogen in the tank, what is the partial pressure of hydrogen?
H ₂ , I N ₂	mol Total PT 6.6 mol 102 kga
OI C	ample 2: A gas cylinder contains 8.0 moles of argon, 2.0 moles of nitrogen, and 2.0 moles oxygen at 600. mmHg. What is the partial pressure of nitrogen in the cylinder?(Ans = 1.0 × mmHg)
	mol Nz PN23
	Ar 8 mol N2 mol (12 mol) N2 mol (12 mol) Ptotal
1 of 2	O2 - 7 mol @ 600



Partial Pressures Practice (Lee)

NDB→

1) Argon, oxygen, and nitrogen are mixed together and pressurized in a tank. P_{02} is 155 kPa, P_{N2} is 415 kPa, and P_{Ar} is 285 kPa.

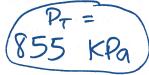


a. What is the total pressure of the gas mixture?

b. What is the % O2?

$$\frac{100\%}{100\%} = \frac{155 \text{ KPa}}{855 \text{ KPa}}$$
What is the % N22

c. What is the % N2?



185%

040

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: $1200 = P_{kr} + \frac{1}{2}P_{kr} + P_{kr}$

- 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.

$$330 = 2x + 3x + 0.5x + x$$

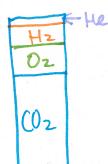
 $330 = 6.5 \times$ b. Determine the number of moles of each gas.

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

express the partial pressure of X, write
$$P_x$$
)
$$\frac{\text{mol } F_z}{\text{total mol}} = \frac{P_{Fz}}{P_T} \implies \frac{50.77}{330} = \frac{P_{Fz}}{46.2}$$

$$P_{Fz} = 7.1 \text{ atm}$$

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.



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 atm () 159 N2 159 C/2

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

15g Nz x 1 mol 28.0 g Nz

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

15gCl2 x 1 mol 70.9 g Cl2

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

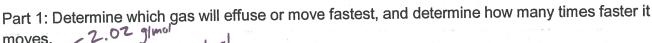
0.747 mol tola

d. What is the P_{N2} in kPa? $\frac{\text{mol Nz}}{\text{total mol}} = \frac{P_{N2}}{\text{total P}} \Rightarrow \frac{0.536}{0.747} = \frac{P_{N2}}{101.3}$

PNZ = 72.7 KPa

e. What is the P_{CI2} in kPa? $\frac{\text{Mol }(12 - \text{Potal Potal Po$

Gas A vs. Gas B
Graham's Law of Effusion (Applied Ownstit) Same temp Same energy
Graham's Law of Effusion (Applied Quantitative Practice)
Graham's Law can be derived from the equation for kinetic energy (K = 100 V)
Derivation: (1) KEA = KEB : they're the same temp!!!
2 KEA = 1 mava and KEB = 1 mBVB
3 $\frac{1}{2}$ $m_A V_A^2 = \frac{eqval!!!}{2} \frac{(same temp.)}{(ancels blc ita a constant.)} \frac{4}{2} m_A V_A^2 = m_B V_B^2$
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 Tips for Use:
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?)
$\frac{\text{rate A}}{\text{rate B}} = \sqrt{\frac{\text{MMB}}{\text{MMA}}} = \sqrt{\frac{70.9}{32}}$
Oz is 1.49 times Faster



- 1) H₂ vs. He 4.00 glmal
- 2) 02 vs. Ne 20.0 glmal
- 3) CH₄ vs. NCl₃
- 4) Ammonia vs. Hydrogen Sulfide (H₂S)

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 (SiH₄) 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 is reasonable to assume this gas is silicon tetrafluoride? Explain and justify your answer using mathematics and complete sentences.

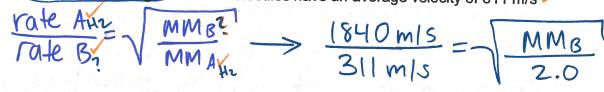
$$\frac{q}{1} = \sqrt{\frac{MM_B}{2.0}}$$

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

No, it's not

040

12)At a certain temperature, hydrogen molecules move at an average velocity of 1.84 x 10³ m/s. Estimate the molar mass of a gas whose molecules have an average velocity of 311 m/s



35 = MM8

MMB = 70.09 (mol

13) Nitrogen gas (N₂) effuses at a rate 2.17 faster than an unknown noble gas. Identify the noble gas

2.17) NMB 28.0

(Xe)

14)A sample of Br₂(g) take 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?

BOWAS long

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

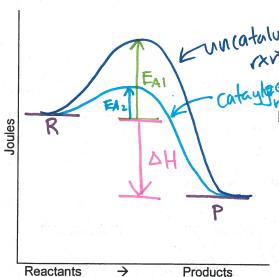
Chapter 19 Part 1: Reaction Rates

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.
Al a Produce the will work
only when they have enough [energy] & corre
Reaction Coordinate Diagram for Exothermic Process (forward direction)
Chapter 11: Thermochemistry Collision Theory
E _A
WOOD T
HA HA
R WSHES
endothem, x
Reactants → Products Reactants → Products
a + Aline He and 4
Activation Energy: <u>LA = Activation energy</u> . The amt of
(E) needed to START TXN.
Activation Complex: The top of the hill
(unstable, temporani)
P. 5. And Reflection Page 1991
B. Factors Influencing Reaction Rate
B. Factors Influencing Reaction Rate 1. Temperature 1 Increase Temp, Speeds up rxn. (foster wole wole wole with the contraction of the contract of the con
2. Concentration 1 increase concentration (either MOLARITY
2. Concentration / INCrease concentration (control control con
3. Particle Size
3. Particle Size Smaller particles react faster.
4. Catalysts
4. Catalysts Speeds up rxn by LOWERING act.en.
(E_A)

Page 1 of 4

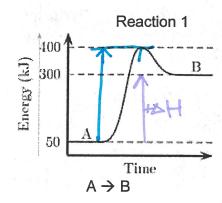
Catalyst Effect on Reaction Path

Why do catalysts increase reaction rate?



Examples: Catalytic Convertors

Interpret the following potential energy diagrams



Reaction 2 450 Energy (LJ) 350 250 Reaction coordinate

 $A + B \rightarrow c$

1. Which reaction is endothermic?

2. What is the activation energy of Reaction 1? + 5

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

4. What is the activation energy of reaction 2? + (00 KJ

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

6. Sketch the effect of a catalyst on both reactions a decrease of 50KJ

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

ow are they usually used?	stuff delwn or
BUILD them up	1.
is the process	s by which life is
metabol	lism
(meaning "to build up")	(meaning "to break down")
Anabolic ("constructive") processes are when larger, more complex biomolecules are created from ones.	Catabolic ("destructive") processes are when,,,
Example #1: When your body is building MASCL proteins using acids	Example #1: RESPIRATION & GLYCOLYSIS: Your body breaks down
Example #2: PHOTOSYNTHESIS Plants use <u>H20 & C02</u> to create <u>OWCOSE - C6H206</u> which is a larger molecule. Anabolism = Energy <u>REQUIRED</u>	Example #2: When you exercise, fats are Catabolism = EnergyRELEASED
	Activation energy without enzyme Reactants Reactants

Page 3 of 4

Lactose Intolerance:				
Lactose is a that is found no	aturally in <u>any</u>	_ products.		
People are who lactose intolerant have a difficult time I	breaking down the lactose mo	olecule.		
Lactose Galactose & Glucose	What suffix do carbohydrates ofte	n have?		
CHOH OH OH OH OH	-050			
OH CH OH OH OH OH	If you see dextr <u>ose</u> , malt <u>ose</u> , or st label, would you call them fat? Su			
CONO ON CONTROL ON CON	Sugar			
Lactose is a 2-ring sugar. It must be by the enzyme a (+a se				
This will turn it into 2 Simpler Sugar, which the body can then use.				
9 .	"сн,он но э—о	°сн,он		
People who are lactose intolerant don't have enough la	H	O TOH H		
enzymes in their to break down the lac	Galaciose	H OH Glucose		
	Lac	itose Lippingo		
THINK: What kinds of problems result from consuming insufficient or non-existent gut lactase?	lactose with	Laciase culon		
Other Francisco Bosses Letter 1	HO HO H	+ HO OH H		
Other Enzymes - Research their bodily functions o	n your own: H OH Galactose	и́ о́н Glucose		

DNA Helicase DNA Polymer**ase** Amylase _____ Protease ("PRO-tee-ase")_____ Lipase _____

-

Black 2 F16

Chemistry Unit 6
Primary reference: CHEMISTRY, Addison-Wesley

Topic	Timary reference: CHEMISTRY, Addison-Wesley			
Atomic	Essential Knowledge	Study Support		
Structure	Electronegativity is the measure of an atom's attraction for electrons in a bond. Electronegativity	Ch 14: Read		
2.6				
SOL 2f	atom is fluorine. The least electronegative element (excluding noble gases) is Francium, Fr.	p. 405		
Nomencla		Cl. 44 D		
ture,		- Late Incad		
Formulas	, Laborier file reactions have a negative ΔH _{co.} whereas endothermic reactions have a second secon	pp. 303-304		
and	Examples of writing an exothermic reaction equation are:			
Reactions				
	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_{20} + 890 \text{ k}$	4 0		
3.6	Or	47.4		
3.0	CH + 20 \ CO \ 200 \ M	Ch 16: Read		
601 24	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O_1 \Delta H_{rxn} = -890 \text{ kJ/mol}$	pp. 460-466		
SOL 3d,		PP. 100 100		
3e, 3f	Polar covalent bonds form between elements with very different electronegativities. The more			
	charge. The less electronegative atom then takes on a slight positive charge. A non-polar covalent bonds form between atoms of similar electronegativities.			
	A notar molecule has unmar electronegativities.			
	A polar molecule has unequally distributed electrons around the central atom. This is caused by unsymmetrical polar bands or a lone pair on the central atom. This is caused by	j- [-		
	intermolecular attractions as well as London dispersion intermolecular attractions. Non-polar molecules only have London dispersion			
	1. /			
	with O-H, N-H or F-H bonds have EXOTHERMIC activated activated activated activated activated activated			
	hydrogen bonding attractions.	Ch 40- D		
	activation E_	Ch 19: Read		
	Kinetics is the study of well as	pp. 533-538.		
	increase with increased temperature, (kJ) Reaction rates			
	reactant concentration increased \ AH Increased			
	and the use of a catalyst Activation Surface area	1		
	minimum energy needed to initiate energy is the			
	High activation energies correspond to	1		
	rates. Catalysts speed up reactions reaction pathway slow reaction			
	decreasing the activation energy. Potential process discussions are supported by rates by	74 64		
Molar	decreasing the activation energy. Potential energy diagrams are used to analyze reaction energy changes.			
Relationsh	Stoichiometry can be combined with heat of reaction, ΔH_{rxn} , to calculate the amount of heat produced from a known amount of reactant.			
ips	a known amount of reactant.			
4.6				
Phases of	Forese of all and the	1		
	Forces of attraction (intermolecular forces) between molecules determine their state of matter at a given	Ch 10. D.		
Matter	I will be a second and action include Hydrodell Donding, dipole-dipole attraction, and I and an include Hydrodell Donding, dipole-dipole attraction, and I and an include Hydrodell Donding, dipole-dipole attraction, and I and an include Hydrodell Donding, dipole-dipole attraction, and I and an include Hydrodell Donding, dipole-dipole attraction, and I a	Ch 10: Read		
and		pp. 269-280		
Kinetic	Vapor pressure is the pressure of the vapor found directly above a liquid in a closed container. When the	and pp. 284-		
Molecular		286.		
Theory	weak intermolecular forces, and low boiling points. Nonvolatile liquids have low vapor pressures, intermolecular forces and high boiling points. Sublimention intermolecular forces and high boiling points. Sublimention in the low vapor pressures, strong	200.		
5.6 ´				
		8		
SOL 5b,				
	The following mathematical relationship between the pressure, volume and temperature of a			
5c, 5d	gas is used to describe the behavior of gases:			
	$\underline{P_1V_1} = \underline{P_2V_2}$			
	T, T,			
	An Ideal Gas does not exist, but this concept is used to model gas behavior. A Real Gas exists,	Ch 12: Read		
	has intermolecular forces and particle volume and the forces and the forces and the forces and the force where the force with the force with the force and the force with the fo	pp. 350-353.		
	has intermolecular forces and particle volume, and can change states. The Ideal Gas Law states	pp. 550-555.		
		1		
	PV = nRT.	İ		
	R is the ideal gas law constant and has two values depending on the account.			
	They are R = 8.314 L'kPa/mol·K and R = 0.0821 L'atm/mol·K			
	Dalton's Law of Partial Pressures says the sum of the partial pressures says			
	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.			
	mindade equals the total pressure of the gas mixture.			
	$P_{\text{out}} = P_{\text{out}} + P_{\text{out}} + P_{\text{out}}$			
	$P_{\text{total}} = P_A + P_B + P_C$ and $n_A/n_{\text{total}} = P_A/P_{\text{total}} = V_A/V_{\text{total}}$			
	Graham's Law says gas molecules with the lightest mass travel fastest. $K.E = 0.5 \text{ mV}^2$			
	The state of the s			

Unit 6 Objectives

Chemistry, Addison-Wesley, 2002

- **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
 - 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

- (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.
- (2f) Compare the electronegativity of two elements based on their position on the periodic table.
- (3d)Compare relative bond polarity based on the two elements position on the periodic table.
- (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
- 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) Indentify 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

Classifying Reactions as Endothermic or Exothermic
Exothermic Reactions
Endothermic Reactions
Heat of Reaction, $\Delta H_{reaction}$ is the heat absorbed or released by a reaction. $+\Delta H_{rxn} -\Delta H_{rxn} -\Delta H_{rxn}$
Thermochemical Equations like regular egns, but Notuce heat flow (+ & H ov - A H) There are two reports were the low in There are two reports were the low in the l
There are two general ways they're written. Both are acceptable:
Heat as a Product (exo) or Reactant (endo): A+B -> C + 100 KJ
Heat shown as a change in Enthalpy (ΔH) $\Delta H = -100 \text{ KJ}$ - ΔH means heat is lost (exothermic) + ΔH means heat was absorbed (endothermic) endo $\Delta H = +100 \text{ KJ}$
Example one: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 890.4 \text{ kJ}$ so $\Delta H = -890.4 \text{ kJ}$
Potential Energy Diagram A CHu+0z $CO_2 + H_2 O$
Reactants -> Products
The system released produced heat so ΔH is negative/positive.
The reaction is endothermic/exothermic?
Law of Conservation of Energy:

2 of 4 **Example Two:** $2NaHCO_3(s) + 129 kJ \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$ Na2CO3+H2O+CO2 Naticos **Products** Reactants heat so ΔH_{rxn} is negative/positive. The system 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) $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + 890.4 \text{ kJ}$

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

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



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) CaO (s) + H_2O (l) \rightarrow Ca(OH)₂ (s) + 65.2 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)

2 NaHCO
$$_3$$
 + 129 kJ \rightarrow Na $_2$ CO $_3$ + H $_2$ O + CO $_2$

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)

2 NaHCO₃
$$\rightarrow$$
 Na₂CO₃ + H₂O + CO₂ Δ H_{rxn} = +129 kJ/mol

Comparing Endothermic and Exothermic Reactions

Exothermic Reactions	Endothermic Reactions
The reaction products heat	The reaction Obsorbs heat
Example of exothermic reaction	Example of endothermic reaction
Potential Energy vs. Reaction Path for Exothermic Reactions	Potential Energy vs. Reaction Path for Endothermic Reactions
F _A = Activation Energy An exothermic reaction has a net	h = AH = AH
An exothermic reaction has a net	An endothermic reaction has a net ΔH
	In endothermic reactions, the product's energy is
In exothermic reactions, the product's energy is than the reactant's energy	than the reactant's energy

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

Chapter 16: Polar Bonds and Polar Molecules and Intermolecular Forces

	Electronegativity:	Tendeny of an atom to or affact e- (FO)
	A. Non-Polar and Polar 1) non-polar covalent be Non- Example: Cl ₂	onds: atoms share bonding electrons equally.
	2) polar covalent bonds	(polar bonds): bonding electrons shared unequally.
N	S Example: HCl	8+ H-C1:) 8- 1055 more EN EN
	 a) more electronega negative charge. b) less electronegati c) unequal sharing c d) Two ways to show lower case greek 	tive atom attracts the electrons more closely and acquires a slight ve atom then acquires a slight positive charge. reates "polarized" bonds with opposite charges. v polarity in structural formulas. Parhally Oreganize deltas: Parhally Operative
	slashed arrows : _	- 68b.
	The type of bond depend Electronegativity Difference	Guideline: Type of Example (electronegativity difference) 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 MT LIM - NaCI
	Na Mg Al Si 0.9 1.2 1.5 1.8 K Ca Ga Ge	
	ION EN	wiah EN

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

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

<u>Dipole</u>: a molecule with two poles (one negative, one positive or δ -, δ +)

What makes a	A molecule is polar if the electrons are pulled to one side of the
molecule polar?	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?

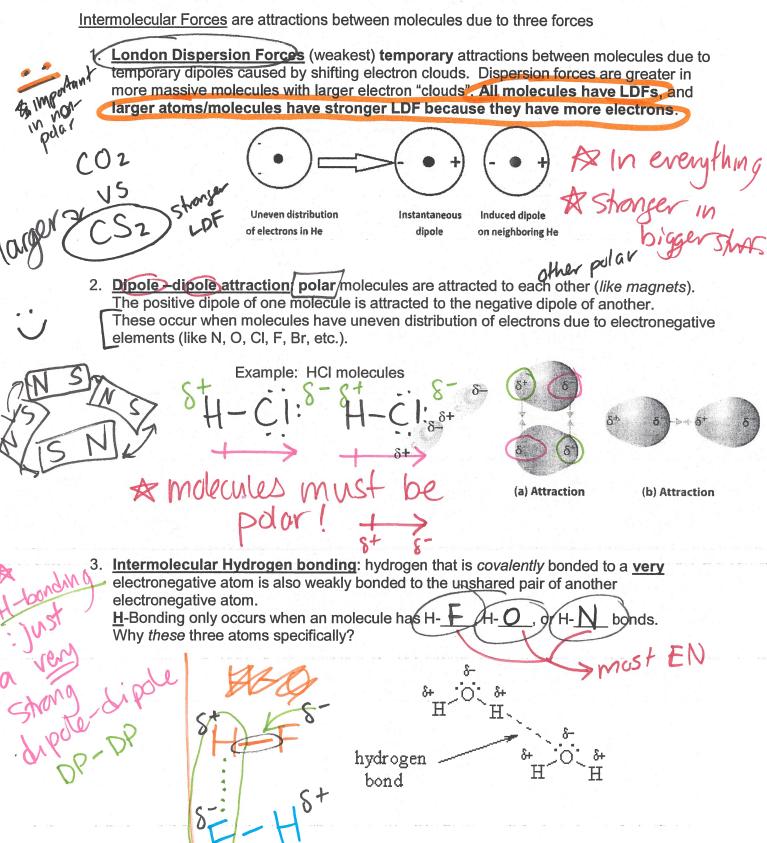
Molecule	Lewis Structure and Geometry	Polar or Non Polar
СО		
	:C≡O:	
CO ₂		
	:0=C=0:	
SO ₃		
503	:o: 	

Molecule	Lewis Structure and Geometry	Polar or Non Polar
SO ₂	그 그 아이를 이 사용하는 사용되는 사꾸 경기를 받는다.	
	:Ö————————————————————————————————————	
CF ₄		
	: F: : F: : F:	
CH ₂ F ₂	: F—C—F: H	
NH ₃	H-N-H H	
H ₂ O	н—ё—н	

Which molecule is most polar: HCl or HI?

Which molecule is most polar; CO_2 or SO_2 ?





more	Stronger	IMF	->	harder	to	Sen
	Stronger					34°

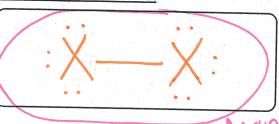
4. Intermolecular Forces and Molecular Physical Properties

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→)

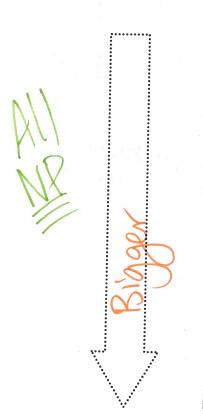




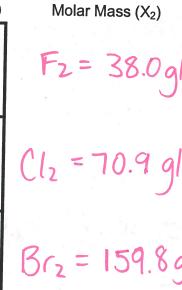
Are pure diatomic halogen elements polar? Explain & draw dipole vectors & partial charges.



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



Halogen (X)
F
Cl
Br
I





State at Room Temp.

moleular

co-valent Comparing Ionic and Covalent Compounds Molecular Compound Ionic Compound Characteristic TRANSFER of ebond formation Na·CI: → Nat:CI: Types of Non-metals elements in compound NON-ME+AL ion = Strong (MF!!! Physical state at 25°C Melting point relatively low Electrical Conductivity in aqueous solution

Sodium nitrate

[Nat] [OKO]

[Nat]

[Nat]

[okostatic

electrostatic

atraction

(ionic)

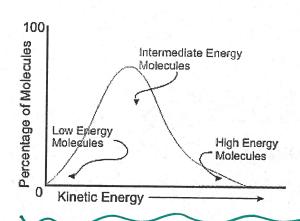
Motion 1

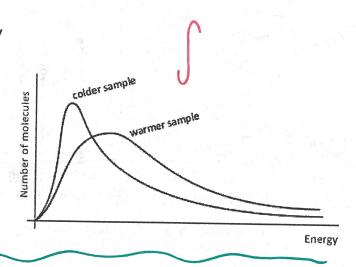


Chapter 10 Kinetic Theory, IMFs, and Phase Changes

I Kinetic Energy and Kelvin temperature scale

A) Temperature measures average kinetic energy





B) Gas particle's kinetic energy increases as

velocity

increases

absolute

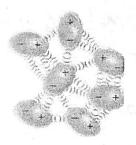
O->UP

OYO

K = °(

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

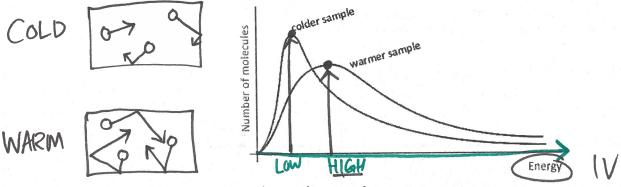
Evaporation:

Phase change

nange tra

mperature

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



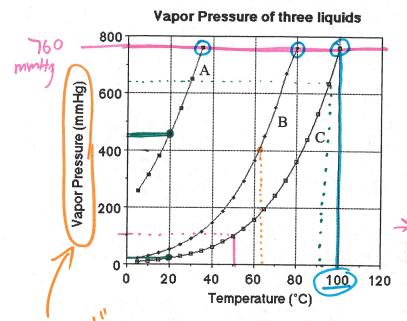
Evaporation in a closed container produces vapor pressure.

Increasing temperature increases vapor pressure over a liquid until a <u>dynamic equilibrium</u> is reached.

Smell: Vapors Vapors Vapors Gas in a container. Sealed Sealed
Volatile liquids <u>Evaporates easily. Low IMF.</u> Sometimes flammable Nonvolatile liquids <u>Dovit evap.</u> easily Higher IMF
Boiling Point of a Liquid (open container) NOVI AND AND AND AND AND AND AND AN
Boiling Point: Temp. @ Which all molecules In a Tignil have enough energy from L to G

0 = boiling point (@ sea level)

A) Boiling point changes as external pressure changes/



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

450 mm Ho

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?



Which liquid has the strongest intermolecular forces?

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

mmHg?

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

(Standard")

= latm = 760 mmHg

(y-axis)

492°C

III.Kinetic Energy and Solids

melting point: Temp @ Which S > L

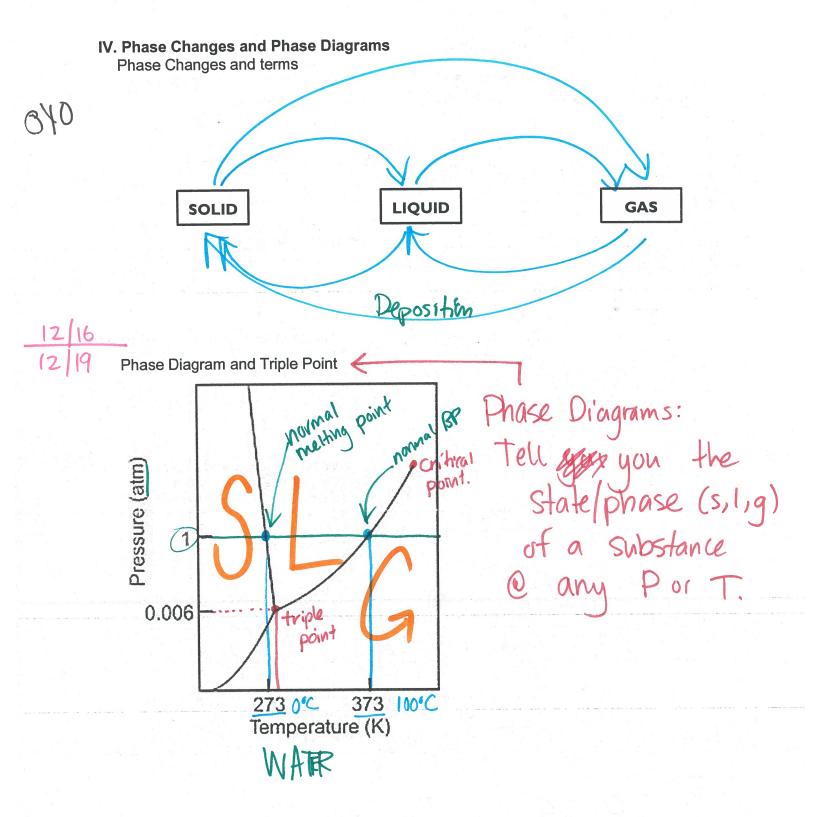
sublimation:

S > G (skips liquid)

Examples:

cos + co2 Dy Ice

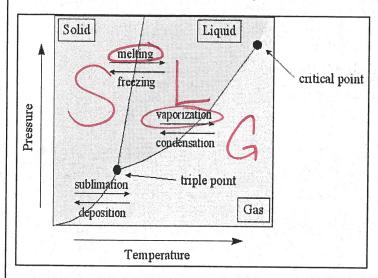
 $I_2(s) \to I_2(g)$



Phase Diagrams

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

The <u>CURVES</u> indicate the conditions of <u>temperature</u> and <u>pressure</u> 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)

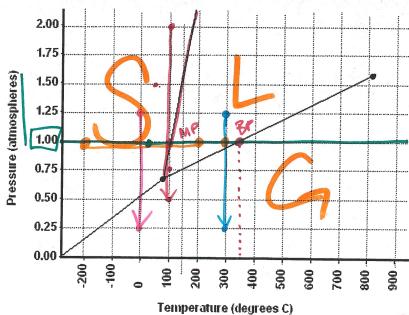
The TRIPLE POINT is the condition of temperature and pressure where ALL THREE phases exist in equilibrium (solid, liquid, gas)

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

Normali

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= 00, P= 1 atm
- 2) What is the *normal* freezing point of this substance?
- 3) What is the *normal* boiling point of this substance?

4) What is the <u>normal</u> melting point of this substance?
5) What is the phase (s, I, g) of a substance at 2.0 atm and <u>100 °C?</u>
6) What is the phase (s, I, g) of a substance at 0.75 atm and <u>100 °C</u> ?
7) What is the phase (s, I, g) of a substance at 0.5 atm and 100 °C?
(8) What is the phase (s, I, g) of a substance at 1.5 atm and 50 °C?
What is the phase (s, I, g) of a substance at 1.5 atm and 200 °C?
10) What is the phase (s, I, 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?
(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 <u>lowered to a temperature of -200° C</u> , 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 ⁰ 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, <u>and</u> solid phase?
ar y faragon ha wan agan ar kung ng makang alam kalang ang alam kalang an kalang an alam ng a
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 ⁰ 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

	Onapter 12 Datton's Law and Granam's Law
Mixtu	res of Gases: Dalton's Law of Partial Pressure
	Partial Pressure: Pressure due to a single gas
*	in a mixture of gases
Z Pax	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: $P_T = P_A + P_B + P_C \cdots$
P	Example 1: What is the partial pressure of oxygen in air at STP (101.3 kPa) if $P_{N2} = 79.1$ kPa, $P_{CO2} = 0.040$ kPa and $P_{others} = 0.94$ kPa?
	101.3 KPa = 79.1 KPa + 0.040 KPa + 0.94 KPa +
	Poz Poz = 21.2 KPa)
	Example 2: A sample of oxygen gas is collected over water at 20.°C. The vapor proseure
	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?
	$\frac{P_T}{P_T} = \frac{P_{02}}{P_{120(9)}} + \frac{P_{120(9)}}{P_{120(9)}}$
02 >	$420 = P_{02} + 15$: $P_{02} = 405 \text{ mmHz}$
Varian	ats of Dalton's Law 1 / Fraction
50 =	$\frac{\text{mole}_{A}}{\text{mole}_{Total}} = \frac{P_{A}}{P_{Total}} = \frac{\%V_{A}}{100 \%V_{Total}}$
Exa kPa	ample 1: A tank contains 6.0 moles of a mixture of hydrogen, helium, and nitrogen at 102 a. If there are 2.0 moles of hydrogen in the tank, what is the partial pressure of hydrogen?
0	Hz, Nz H mol Hz PHZ Z mol 7 PHz
P	= 102 KPO total Protal Protal
Exa of o	imple 2: A gas cylinder contains 8.0 moles of argon, 2.0 moles of nitrogen, and 2.0 moles exygen at 600. mmHg. What is the partial pressure of nitrogen in the cylinder?(Ans = 1.0 ×
10 ² r	mmHg)
	Smol Ar $-600 \text{ mmHg} = P_T$ $\frac{\text{mol N}_2}{\text{total mol}} = \frac{P_{N2}}{P_T}$ $P_{N2} = 100$
	2md N2 - 12mol, total
1 of 2	$\frac{2 \text{ mol Nz}}{12 \text{ mol Nz}} = \frac{f_{Nz}}{600 \text{ mmHg}}$

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)

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 Gras particles more from Partition Diffusion
HIGH to LOW
Concentration
Effusion Gas particles mong
Thu a hole & light
gases more/havel tastest
Graham's Law Verbally: Gas molecules with the lightest mass travel fastest. Graham's Law Equation: Mass
1 × 12
H ₂ Velocity ("speed") H ₂ VS F ₂ H ₂ VS Cl ₂ VS Cl ₂ VS Cl ₂ VS VS VS Cl ₂ VS V
Speed v Speed v VS Oz
Which gas will escape slowest from a tiny hole in a balloon; He, C ₃ H ₈ , or Xe?

- 1) Argon, oxygen, and nitrogen are mixed together and pressurized in a tank. P_{O2} is 155 kPa, P_{N2} is 415 kPa, and P_{Ar} is 285 kPa.
 - a. What is the total pressure of the gas mixture?
 - b. What is the % O2?
 - c. What is the % N2?
- 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:

PcH4 340 mmHg
Pkr 680 mmHg
Pol 180 mmHg

1200 = 680 + 340 + _

b. Determine the % volume of each gas:

% CI $\frac{150}{1200} \times 100 = 15\%$ % Kr $\frac{680}{1200} \times 100 = 56\%$ % CH₄ $\frac{340}{1200} \times 100 = 29\%$

- 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 mol = X + 2X + 3X + 0.5X

b. Determine the number of moles of each gas.

X = # of Ez modes = 50.77 mol

Clz = (52.31 md)

He = 101.53 mal Hz = 25.39 mal

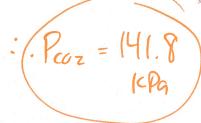
c. Determine the partial pressures of each gas. Use proper notation. (Example: To express the partial pressure of X, write P_X)

330 mul, total = PFZ 46.2 atm

· PF2 = 7.11 atm



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.





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? 0.5357 mol N_2

159 N2 x 1 mol N2

b. What is the number of moles of chlorine in the container? 0.216 mol

15g Clz x 1 mol az

- c. How many total moles of gas are in the container?
- d. What is the P_{N2} in kPa? = 72.6 Kfa $\frac{0.5357 \text{ mol N2}}{0.7473 \text{ mol, folial}} = \frac{P_{N2}}{101.3 \text{ Kfg}}$
- e. What is the P_{Cl2} in kPa?

Graham's Law of Effusion (Applied Quantitative Practice) Graham's Law can be derived from the equation for kinetic energy Derivation: $0 \text{ KE}_A = \text{KE}_B$ because they be the same temperature. $\text{TE}_B + \text{TM}_B \text{V}_B^2$ Equal energy bic same temp. $\text{TM}_A \text{V}_A^2 = \text{M}_B \text{V}_B^2$ Constants cancelsGraham'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 program is qualitable.)

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.

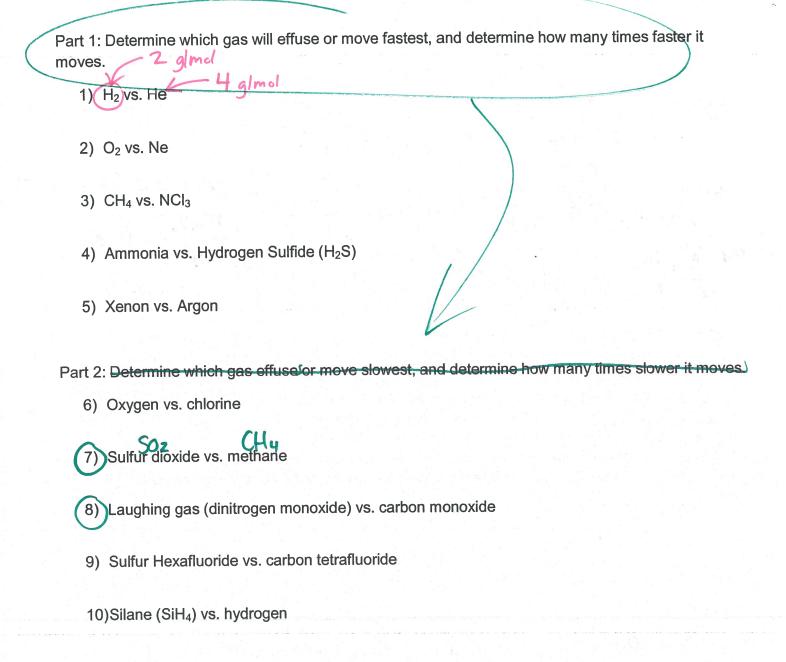
Analogy: A 60-kg girl eats 2 eggs and 3 slices of bacon. A 150 kg sumo wrestler eats the same thing. The girl will run faster because she's smaller, even though they had the same breakfast (energy).

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

Worked Example 1: The average velocity of oxygen (O₂) molecules will be faster than the average velocity of chlorine (Cl₂) 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?)

Cl2 has a velocity of 310. m/s

1.49 Dz moves 1 1.49 times faster.



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 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 x 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 Bro(g) take 10.0 min to offuso from one side of a record that is 20.5.
14)A sample of Br ₂ (g) take 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?
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

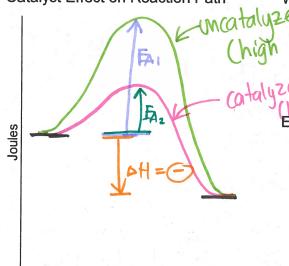
http://www.wwnorton.com/college/chemistry/gilbert/index/site_map.htm
Reaction rates are measured as mol/time units.

A. Collisi	on Theory:_	Atoms	moderate	lim c	partis	ipak	
m	rxns	only	When	they	have	enough	energy
Reaction	Coordinate	Diagram for E	Exothermic Pr	rocess (forward	ard direction)	A con-	201
Chapter	11: Thermoo	hemistry		Collision TI	neory	W COM	ect orien
					E		
		_		R			
Jonles	·		R	Moob Moob		H = () E	2/0
		DH=e	ndo (+)				
-	2				<u>V</u>	P	
						ASHES	
Read	ctants ->	Produc	ts	Reactan	ts →	Products	
Activation	on Energy:	Minimum	n energ	gy N	reded	to sta	A
F	eachier	1					_
Activation	on Complex:	The	high-	energi	1 "hil	1 top"	<u>.</u> 19
4	instab	l	tember	am)			
B. Facto	ors Influencin	g Reaction R	ate				
1. T	emperature	Incre	asing	Temp	$(T\Lambda)$	speeds	upran
2. 0	Concentration					Molarity 4	
	Particle Size		,				
4 (Smaller Catalysts	partic	iles rec	zet mi	me rap	oidly!!	
7.	, j					egy(11L) -
	(0000) ac			- yy (mu (p)
Page 1 o	f 4		7		1		

Catalyst Effect on Reaction Path

Reactants

Why do catalysts increase reaction rate?



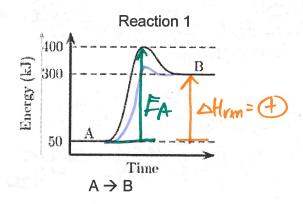
 \rightarrow

Examples: Catalytic Convertors

ater

Interpret the following potential energy diagrams

Products



Reaction 2

350

A+B

Reaction coordinate

 $A + B \rightarrow c$

- 1. Which reaction is endothermic?
- 2. What is the activation energy of Reaction 1? $\frac{1}{2}$ $\frac{1}{2$
- 3. What is the ΔH_{rxn} of reaction 1? (4) 250 (4)
- 5. What is the ΔH_{rxn} of reaction 2?
- 6. Sketch the effect of a catalyst on both reactions
- 7. Does a catalyst effect the ΔH_{rxn} ?

Biologically (in terms of biochemistry),	124mls are catalysts.
How are they usually used? Nelp build new :	stuff a UP
is the process	s by which life is
Metabo	lism
Į.	
Ahabolism	Catabolism
(meaning "to build up")	(meaning "to break down")
Anabolic ("constructive") processes are when larger, more complex biomolecules are created from ones.	catabolic ("destructive") processes are when(arge_,campleaded molecules are broken down into smaller ones biomolecules.
Example #1: When your body is building MUSCL proteins using amino a Cit Pephro	Example #1: RESPIRATION & GLYCOLYSIS: Your body breaks down Carbohydrates / Sugars to produce energy, CO2 & H2O.
Example #2: PHOTOSYNTHESIS Plants use H2O & CO2 to create C6H12O6, which is a larger molecule. Anabolism = Energy Yequired	Example #2: When you exercise, fats are
Enzymes (which are biological	Activation energy without enzyme Activation energy without enzyme Activation energy with enzyme

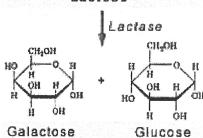
Progress of reaction -

Page 3 of 4

Enzymes usually end in the suffix $\underline{-ase}$.

Carbohydrates often end in ___OSe__.

Lactose Intolerance:
Lactose is a Sygar that is found naturally in dairy products.
People are who lactose intolerant have a difficult time breaking down the lactose molecule.
Lactose Galactose & Glucose What suffix do carbohydrates often have?
If you see dextrose, maltose, or sucrose on a food label, would you call them fat? Sugar? Or protein?
on epithelial cell surface Absorption SVG M
Lactose is a 2-ring sugar. It must be by the enzyme (a ctase.
This will turn it into 2 Simpler Sygars, which the body can then use.
"снон "снон но жо и жо о
People who are <i>lactose intolerant</i> don't have enough lactase
enzymes in their to break down the lactose sugar. to break down the lactose sugar. Galactose Glucose Lactose
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 Glucose



DNA Polymerase		
Amylase		
Protease ("PRO-tee-ase")		
Lipase		

	*
والمرابع والمرابع والمستعلق والمنافع والمستعدد والمستعد والمستعدد والمستعد والمستعدد والمستعد والمستعدد والمستعد والمستعدد والمستعدد والمستعدد والمستعدد والمستعدد والمستعدد وال	
	1 1 2 20 10 20

Chemistry Unit 6

Atomic Stricture Strictur		Primary reference: CHEMISTRY, Addison-Wesley			
Structure 2.6 SOL 24 Nomencia ture, 1	Topic	Essential Knowledge	Study Support		
Nomencia ture. Nomencia ture. Formulas, and Reactions 3.6 SOL 3d, 3e, 3f Polar covalent bonds from between elements with very different electronegativities. The more electronegative atom the telectronegative atom that the telectronegative atom th	_	increases across a period toward the halogens and docreases down a group. The			
Homencia ture, by the amount of energy absorbed or released during a chemical charge. Formulas, and smouth of energy absorbed or released during a chemical charge. Scamples of writing an exchemic reaction save a positive □H _{nos} gradient or seasons be the absorbed or released during a chemical charge. Scamples of writing an exchemic reaction equation are: Ch 4 + 20₂ → C0₂ + 2h₂0, □H _{rine} − 890 kJ/mol Polar covalent bonds from between elements with very different 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 dipice and the negative end has a negative dipice. Polar molecular attractions as well as London dispersion intermolecular attractions. With O+I, N+I or F+I bonds have hydrogen bonding attractions. Kinetics is the study of reaction rates. Increased with increased temperature, reactint concentratory, increased and the use of a catalysts. Activation in rates increased with increased temperature, reacting concentratory, increased and the use of a catalysts speed up reactions. Kinetics is the study of reaction rates. Increased with increased temperature, reacting concentratory, increased and the use of a catalysts speed up reactions. Kinetics is the study of reaction rates. Increased with increased temperature, reacting on the central and the produced and the use of a catalysts speed up reaction. High activation energics correspond to rates. Catalysts speed up reactions. Forces of attraction (intermolecular forces) between molecules determine their state of matter at a given the produced and the use of a catalysts. Activation energy charges are all under the produced and the use of a catalysts. Activation energy of the produced and the use of a catalysts are used to make the amount of heat produced and the use of the produced and the use o		atom is fluorine. The least electronegative element (excluding noble gases) is Francium. Fr			
the amount of energy absorbed or released during a chemical change. Earth of the standard of					
Examples of writing an exothermic reaction as a chemical change. Sol. 3d, 3e, 3f Sol. 3d, 3e, 3f 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 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 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 form between atoms onegative atom then takes on a slight positive drage. A non-polar covalent bonds A polar molecule has unequally despetivities. A polar molecule has unequally despetivities. The more electronegative atom will attract the electrons reactive positive dipole and the negative end has a negative dipole. Polar molecules that molecules a traction intermolecular attractions as well as tondon dispersion intermolecular attractions as well as tondon dispersion intermolecular attractions. With O-H, N-H or FH bonds have hydrogen bonding attractions. With O-H, N-H or FH bonds have hydrogen bonding attractions and the use of a catalyst. A chill attom minimum energy needed to initiate a High activation energies correspond to rasks. Catalysts speed up reactions energy. Potential energy diagrams are used to analyze reaction energy pits the reaction rates increased surface area high activation energies correspond to rasks. Catalysts speed up reactions energy diagrams are used to analyze reaction energy pits the pressure and the expension of the energy diagrams are used to analyze reaction energy pits the reaction. Solve reactions energy pits the pressure sit the pressure of the vapor found directly above a liquid in a closed container. When the advance was a sight possible to the molecular forces, and low bioling points. Solving the energy dia	1	Exothermic reactions release heat whereas endothermic reactions absorb heat. Heat of reaction is	Ch 11 · Poad		
And Reactions 3.6 SOL 3d, 3e, 3f Polar covalent bonds form between elements with very different elegatives. The more electronegaptive atom will attract the electronegaptive atom the reactive properties of similar electronegaptive atom will attract the electronegaptive atom will attract the electronegaptive atom that the positive dipole and the repative recompagnation of the molecule intermolecular attractions as well as London dispersion intermolecular attractions as well as London dispersion intermolecular attractions are well as tondon dispersion intermolecular attractions are well as tondon dispersion intermolecular attractions. With OH, NH or PH or Ho hords have hydrogen bonding attractions. EXOTHERMIC Exchanged Physics P		the amount of energy absorbed or released during a chemical change	1		
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Granam's Law says gas molecules with the lightest mass travel fastest.	1				
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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) Indentify 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

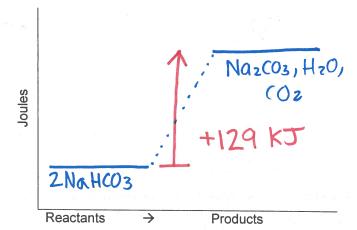
Classifying Reactions as Endothermic or Exothermic
Exothermic Reactions products (releases heat (-DH)
Endothermic Reactions absorbs heat (+ AH)
Heat of Reaction, $\Delta H_{reaction}$ is the heat absorbed or released by a reaction.
+ DHrxn endo - DHrxn exo
Thermochemical Equations are balanced stoichiometric équations showing the heat change.
Example one: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 890.4 \text{ kJ exo}$ so $\Delta H = -890.4 \text{ K) mc}$
Potential Energy Diagram
Ct
B. CH4, O2: - T
(Hu, Oz : ,
CO2, H2O
Reactants → Products
The quetare ICOLOGGIA O
Heat 30 ATT is flegative/positive.
The reaction is endothermic/exothermic?
Law of Conservation of Energy: (E) IS never created destroyed
only transformed
Energy = Nork + heat + potential energy
Gasoline = miles + heat + coz + H20
C8H18 (Work) Co, Sox, Nox
Catalytic
con verters

(64.5 Kj/mol



 $2NaHCO_3(s) + 129 kJ → Na₂CO₃(s) + H₂O(g) + CO₂(g)$

so ∆H=



The system

absorbina is

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)

 $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 890.4 \text{ kJ}$

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

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

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) CaO (s) + $H_2O(I) \rightarrow Ca(OH)_2$ (s) + 65.2 kJ

25.7g (a0) x Imol (aU) 65.2 KJ
56.0 g 1 mol (ac)

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)

2 NaHCO $_3$ + 129 kJ \rightarrow Na $_2$ CO $_3$ + H $_2$ O + CO $_2$

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)

2 NaHCO₃ \rightarrow Na₂CO₃ + H₂O + CO₂ Δ H_{rxn} = +129 kJ/mol

Comparing Endothermic and Exothermic Reactions

Exothermic Reactions	Endothermic Reactions
The reaction heat	The reaction heat
Example of exothermic reaction	Example of endothermic reaction
Potential Energy vs. Reaction Path for Exothermic Reactions	Potential Energy vs. Reaction Path for Endothermic Reactions
	1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1
A	
An exothermic reaction has a net	An endothermic reaction has a net
ΔΗ	ΔΗ
In exothermic reactions, the product's energy is	In endothermic reactions, the product's energy is
than the reactant's energy	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 otom 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: HCI
)	 a) Electron sharing based on electronegativity differences. a) more electronegative atom attracts the electrons more closely and acquires a slight negative charge. b) less electronegative atom then acquires a slight positive charge.
d	c) unequal sharing creates "polarized" bonds with opposite charges. d) Two ways to show polarity in structural formulas.
	lower case greek deltas: 8 & 8+ (partially Charges) slashed arrows: (points to more electroneg atom)
	The type of bond depends on electronegativity differences between the atoms
	Difference Guideline: Type of Example (electronegativity difference)
	0.0 – 0.4 Non-polar Covalent C-H in CH ₄ (2.5–2.1) = 0.1
>	0.4 - 2.0 Polar Covalent HF $(4.0 - 2.1) = 1.9$
	>2.0 Ionic NaCl (3.0 - 0.9) = 2.1
	Solooted Floatron and it is Not 1
	Selected Electronegativity Values Is the bond polar, non-polar or ionic? C=O in CO ₂ ? 2 5 - 2
	5.5 2.57 1
	Li Be B C N O F Si-H in SiH ₄ $\begin{bmatrix} 1.6 & -2.1 & -2.3 \\ 2.5 & 3.0 & 3.5 \end{bmatrix}$ Si Ore Si-H in SiH ₄ $\begin{bmatrix} 8 & -2.1 & -2.3 \\ 2.5 & 3.0 & 3.5 \end{bmatrix}$
	Na Mg Al Si P S Cl C-F in CF ₄
	0.9 1.2 1.5 1.8 2.1 2.5 3.0 K Ca Ga Ge As Se Br N-Cl in NCl ₃ (D. N-Cl in NCl in NCl ₃ (D. N-Cl in NCl i
	$0.8 \mid 1.0 \mid 1.6 \mid 1.8 \mid 2.0 \mid 2.4 \mid 2.8$
	KCI 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?

C-N

MASK C-F

B: POLAR MOLECULES

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

<u>Dipole</u>: a molecule with two poles (one negative, one positive or δ -, δ +)

What makes a

A molecule is polar if the electrons are pulled to one side of the molecule polar?

Molecule is polar if the electrons are pulled to one side of the molecule polar?

Nacl with H20

Determining if a molecule is polar.

Draw the Lewis structure

2. Determine the molecular geometry

3. Look for lone pairs on central atom (automatically polar)

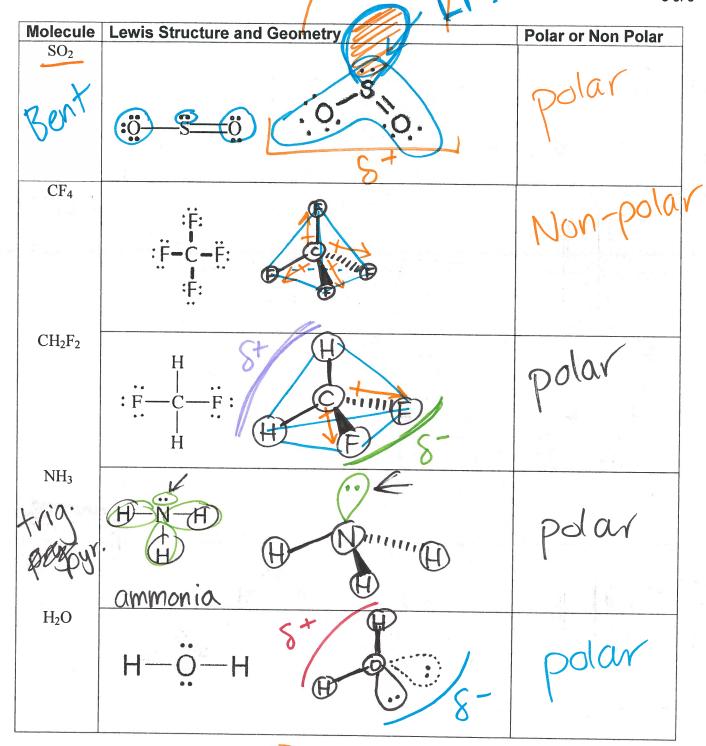
4. Are there polar bonds?

5. If yes, are the polar bonds unsymmetrical in 3-D around the molecule's center?

Molecule			Polar or Non Polar	
Mokoxige Carpan CO	81/80° 8- :C≡O:	END = 1.0 (PC)	polar	
CO2	\$- \$+ \$- :0=C=O: *+ +->		Non-Polar	
Non on a	10: 10: 10: 10: 10: 10: 10: 10: 10: 10:	allectors (dipoles) cancel.	Non-Pol	ar

A Will be polar if theres LP on central ortom

Padays things things things



Which molecule is most polar HCl or HI?

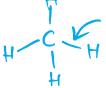




HW but 5/3?
P.122

→ Which molecule is most polar; CO₂ or SO₂?

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.

· Present in everything,

· Bigger things = stronger

Uneven distribution In:

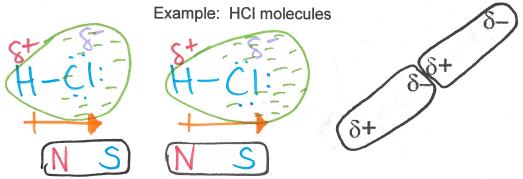
of electrons in He

Instantaneous

Induced dipole on neighboring He

more electrons =stronger LDF Bigger=1 LDF)

2. <u>Dipole – dipole attraction</u>: <u>polar molecules</u> are attracted to each other. The positive dipole of one molecule is attracted to the negative dipole of another.



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

The H in one molecule stracted to the hydrogen bound

molecules)

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.



Covalent

Comparing Ionic and Covalent Compounds

Comparing Ioni	c and Covalent Compounds	
Characteristic	Ionic Compound	Molecular Compound
bond formation	electrostatic	Shanny of
	attraction by	Shanny of valence e-
Types of elements in compound	M + NM	NM + NM
Physical state at 25°C		most: gases
100m Jomes		most: gases - liquids - Solids
		- Solida
Melting point		
	high at	100
Electrical Conductivity in aqueous solution	Tionic	molectum

Yours electrolytes NHy+ → ← CI-

NHyCI

non-electrolytes

C12H22O11 (Sucrose)

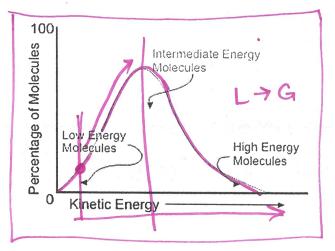


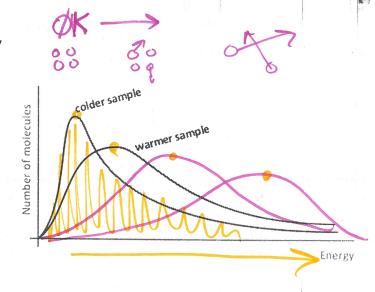
Chapter 10 Kinetic Theory, IMFs, and Phase Changes

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

C) Kelvin Temperature scale is absolute & divectly prop

273 K= 0 °C

directly propurtional

II.Kinetic Energy and Liquids

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



Stronger IMF take move 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:_

he

& L->G

transistion

that

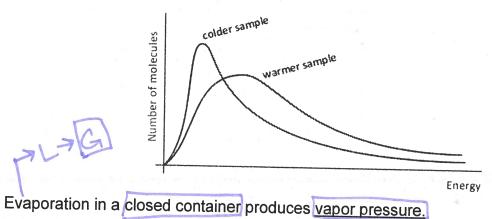
happens

BELOW

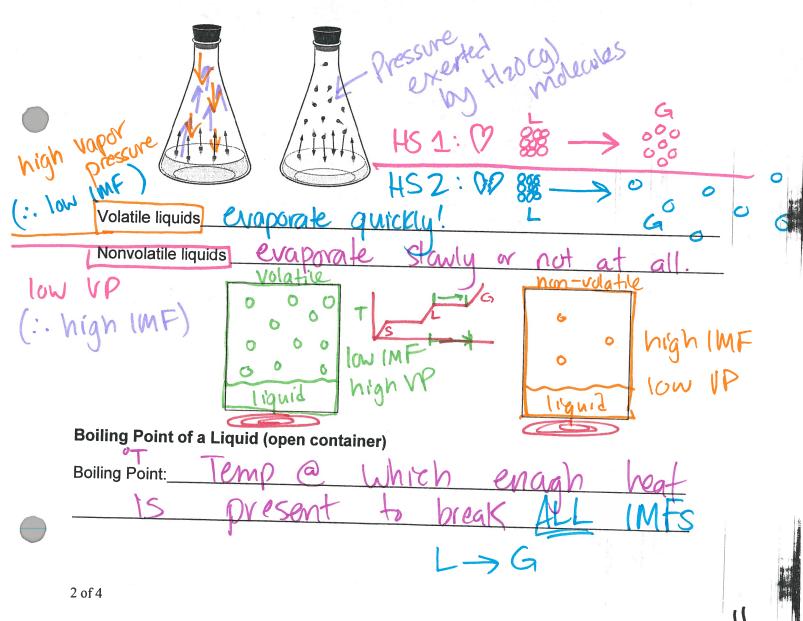
the

BP.

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

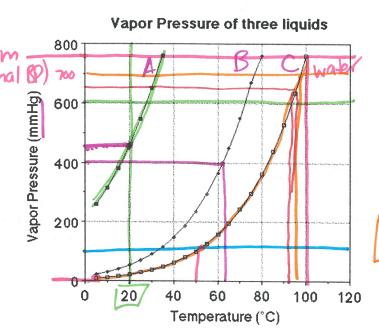


Increasing temperature increases vapor pressure over a liquid until a <u>dynamic equilibrium</u> is reached.



high IMF -> low VP -> high BP

A) Boiling point changes as external pressure changes/



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?

Which liquid has the strongest intermolecular forces?

Strong high IMF = low VF

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

BP goes V when presure when presure lids

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

497°C

III.Kinetic Energy and Solids

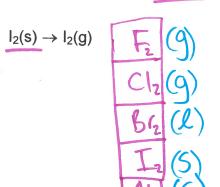
melting point: temp of S-> L transition e Normal

S->L: Break most IMF that hold in 88

sublimation: 5 > G (slaps liquid)

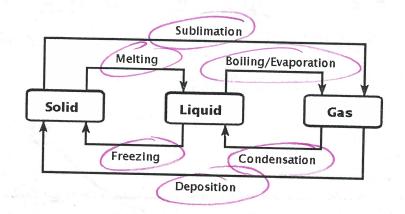
Examples:

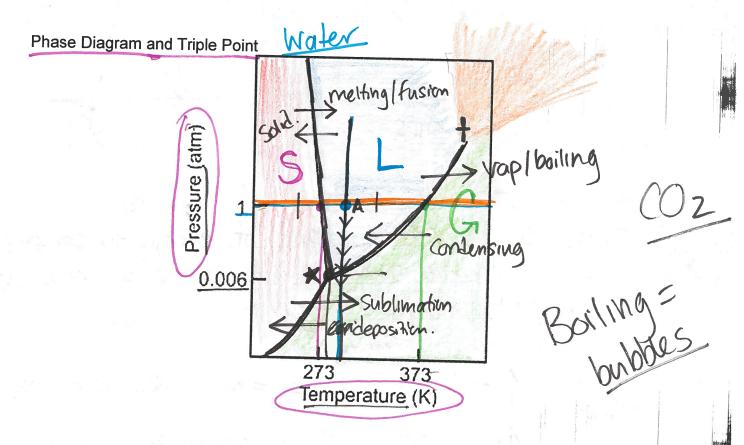
 $CO_2(s) \rightarrow CO_2(g)$ DRY CE



IV. Phase Changes and Phase Diagrams

Phase Changes and terms





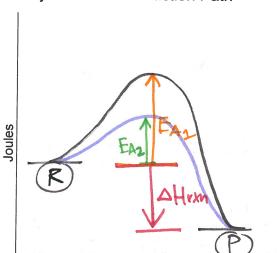
Chapter 19 Part 1: Reaction Rates
Reaction rates are measured as mol/time units. [And a late of the
A. Collision Theory: Atoms [Molecules react w each other only
if there's enough (E) & correct orientation
Reaction Coordinate Diagram for Exothermic Process (forward direction)
Chapter 11: Thermochemistry Collision Theory Activated activation
activation
energy energy
(R)
Anon Mondo
3 - AHM
*
ASMES
Reactants → Products Reactants → Products
Activation Energy: (Ea) The amt of (E) needed to
Start a reaction (amt of energy to comple)
Activated Activation Complex: The arrangement of ortens & moleculus
at highest energy stage
B. Factors Influencing Reaction Rate Increas
B+G, 1. Temperature 1 = faster molecules = more collisions = move poduct
2. Concentration 1 = increase # of moderales=more collisions
3. Particle Size increase Surface area = more cultisions) (Smaller Peoplesians)
4. Catalysts A speeds up reaction (but not reactants)

Page 1 of 2

14 40

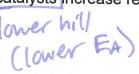
Catalyst Effect on Reaction Path

Reactants



 \rightarrow

Why do catalysts increase reaction rate?



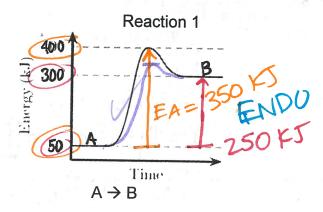
Examples: Catalytic Convertors

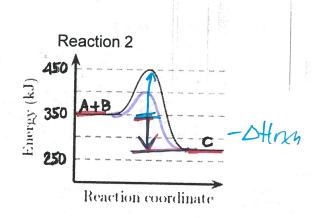


 $A + B \rightarrow c$

Interpret the following potential energy diagrams

Products





Which reaction is endothermic?

2. What is the activation energy of Reaction 1?

3. What is the ΔH_{rxn} of reaction 1?__

4. What is the activation energy of reaction 2?___

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

6. Sketch the effect of a catalyst on both reactions

7 Does a catalyst effect the ΔH_{rxn} ?

Chapter 12 Dalton's Law and Graham's Law

Mixtures of Gases	Dalton's	Law of	Partial	Pressure
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Partial Pressure: Pressure due to ONE gas in

a confined mixture of gases

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:
$$P_{A} = P_{A} + P_{B} + P_{C} \cdots$$

Example 1: What is the partial pressure of oxygen in air at STP (101.3 kPa) if $P_{N2} = 79.1$ kPa, $P_{CO2} = 0.040 kPa$ and $P_{others} = 0.94 kPa$?

0.5 atm 02

PT = PA + PR + PC + PD ... Example 2: A sample of oxygen gas is collected over water at $20.^{\circ}$ 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?

at 20.°C is 15 mm Hg. II unc.

of the oxygen?

H₂O(g) = 15mmHg $P_T = P_{H20} + P_{02}$ $P_{02} = 15mmHg + P_{02}$ $P_{02} = 405mmHg$

Variants of Dalton's Law

 $\frac{\text{mole}_{\overline{A}}}{\text{mole}_{\overline{\text{total}}}} = \frac{P_{\overline{A}}}{P_{\overline{\text{total}}}} = \frac{\% V_{\overline{A}}}{100 \% V_{\overline{\text{total}}}}$

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?

 $\frac{\text{Mol Hz}}{\text{He,Nz}} = \frac{P_{Hz}}{\text{Mol Total}} = \frac{P_{Hz}}{P_{T}} \rightarrow \frac{2 \text{ mol Hz}}{\text{Comol total}} = \frac{P_{Hz}}{P_{T}}$ $\frac{P_{Hz}}{P_{T}} = 33.9$

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 = $1.0 \times$ $10^2 \, \text{mmHa}$

2 mol Nz = KNz 12 mol total = 600 mmHq

H22 Ne 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) 142 PNe = 760 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) 4N2 = 105 KPa 70% N2 = 105 KPa 100% total = 150 KPa total Poz = 45 KPa aw of Effusion Diffuse is > L gas patrices more from Graham's Law of Effusion Diffusion Diffusion conc. Effusion Graham's Law Verbally: Gas molecules with the lightest mass travel fastest. O_2 Number of Molecules N H_2 Which gas will escape slowest from a tiny hole in a balloon; He, C₃H₈, or Xe?

312 O 2 32 9 mol

2nd N 2 39 9 mol

2 of 2 4th Ar A 39.9 g/mol