CHEM 205 sections 01

LECTURES #23-24 Nov.21st-26th, 2013

ASSIGNED READINGS:

THIS WEEK: Ch.11 overview

study theory at level of notes
but do problems from textbook

TIMELINE OF TOPICS IN CHEM205 ↓

1ST: WHAT HAPPENS (practical)

2ND: UNDERSTANDING WHY (theoretical)

Classifying matter & making measurements

Atoms & elements

Molecules, ions & their compounds

Chemical equations & reaction outcomes

Typical reactions that occur in water

Understanding how molecules are "put together" & how this affects the substance's properties

How to realistically think about atoms

Why different elements have different properties

Understanding the behaviour of gases

Chapter 11: Gases & their Properties

You are only responsible for learning Ch.11 at the level presented in the lectures and notes.
(*i.e.*, you can read Ch.11 superficially… but must do Ch.11 problems!)

11.1 Properties of Gases: Pressure
11.2 Gas Laws: Experimental Basis
11.3 The Ideal Gas Law
11.4 Gas Laws & Chemical Reactions
11.5 Gas Mixtures & Partial Pressures
11.6 The Kinetic Molecular Theory
11.7 Diffusion & Effusion
11.8 Nonideal Behaviour: Real Gases

Chapter Goals:

- **Understand the basis of the gas laws** (Boyle, Charles, Avogadro, Dalton) **& how to use them**
- **Use the ideal gas law**
- **Apply the gas laws to stoichiometric calculations**
- **Understand kinetic-molecular theory of gases, especially the distribution of molecular speeds (energies)**
- **Recognize why real gases do not behave as "ideal gases" under some conditions**

11.1: Properties of gases

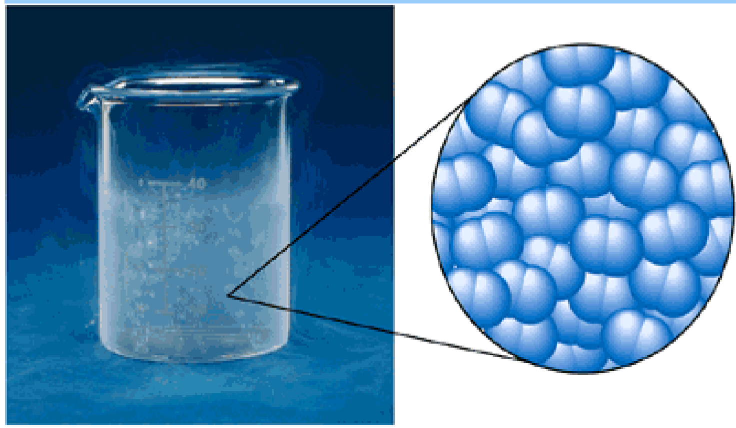
- **Referring to gas-phase substances:**
  - "gas" = substance normally in gaseous state *(phase)* at ambient temperature (T) & pressure (P)
    Table 12.1: He, H_2, O_2, N_2, F_2, Cl_2, CO_2, C_3H_8, etc…
  - "vapour" = gaseous form of substance that is normally a solid or liquid at ambient T & P
    water vapour (steam), gasoline vapour (gas fumes)

Particles (atoms/molecules) in a gas are VERY FAR APART

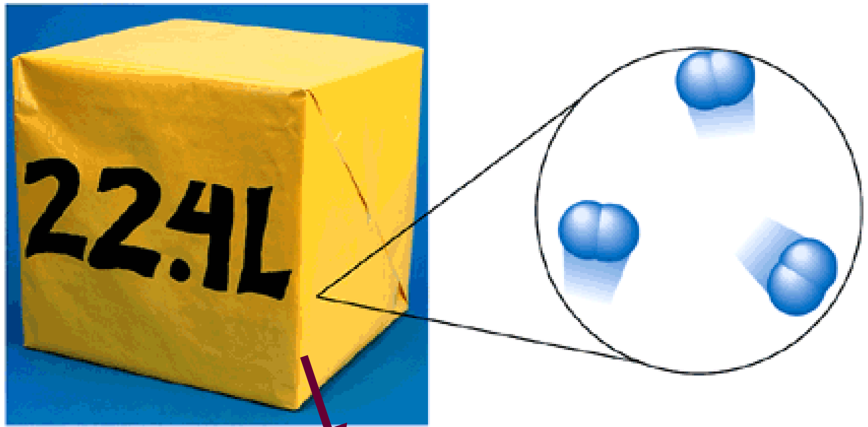
- lowest density state of matter
- particles can be squeezed closer together: compressible
- particles don't interact much with each other: simple behaviour!

Zumdahl's Figure 5.14:

(a) One mole of $N_2(l)$ = 28.13 g
V = ~ 35 mL = ~ (3.3 cm)^3
d = 0.81 g/mL



(b) One mole of $N_2(g)$ at STP
V = 22.4 L
d = 1.2 x 10^{-3} g/mL



22.4 L = 22.4 dm^3
= 22.4 x (10cm)^3
= 22.4 x 10^3 cm^3
≈ 28cm x 28cm x 28cm
≈ roughly 1 cubic foot

Volume of one mole of N_2, as LIQUID *vs.* GAS:

$$V(N_{2(g)}) / V(N_{2(l)})$$

= 22.4 L / 0.035 L
= 640x greater as gas!

(5)

Gases exert PRESSURE: what is it?

Force caused by gas particles colliding with container walls

PRESSURE is defined as:

FORCE per unit AREA

- Force = mass x acceleration
  = mass x (velocity change/unit time)
  1 N = 1 kg x (m/s^2) = **definition of "Newton"**

- **SI unit of Pressure: "Pascals" 1 Pa = 1 N/m^2**
  = 1 (kg·m·s^{-2})·m^{-2}
  = 1 kg·m^{-1}s^{-2}

- **Typical "atmospheric pressure" = 1 atm (COMMON unit)**
  = 101.325 kPa
  = 1.01325 bar
  = 760 mm Hg
  = 760 torr

NOTE:

- in everyday life: "psi" (pounds per square inch) where 1 atm = 14.7 psi

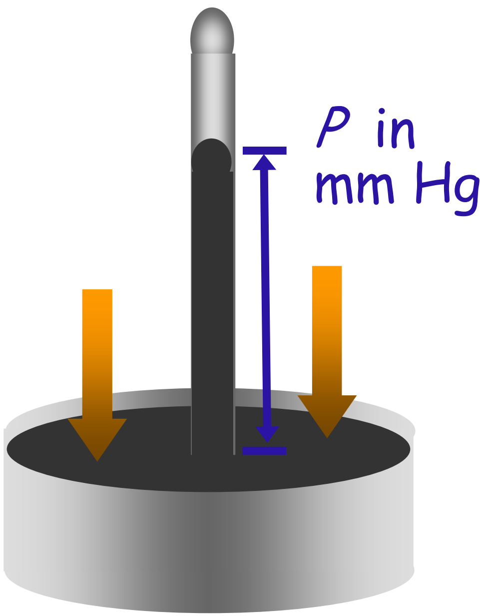
Measuring pressure: barometers & manometers

Kotz Fig.11.2

A barometer

- dish of Hg(*l*) open to atmosphere
- evacuated tube inverted into Hg
- Hg rises until $P_{Hg(gravity)} = P_{atm}$

- P = height of Hg column

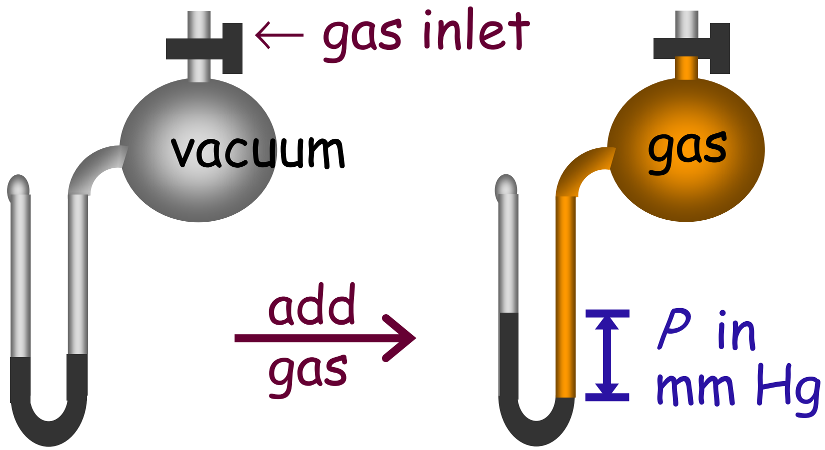


Kotz p.511: *A Closer Look*

A manometer

- U-tube, with Hg(*l*)
- sealed on one end
- gas flask on other end

- P = difference in height of two ends of Hg "column" in the U-tube



11.2-3 Gas Laws: Experimental Basis… ⇨ Ideal Gas Law

Physically: (NOT chemically) **All gases essentially behave same way!**

Gas behaviour is described by natural "laws"

Based on general observations made in the 17th & 18th centuries…by:

- **Avogadro:** more particles ⇨ larger volume $V = a\,n$
- **Boyle:** higher pressure ⇨ smaller volume $V = b/P$
- **Charles:** higher temperature ⇨ larger volume $V = c\,T$

SUMMARY: The volume (V) occupied by a sample of gas is:
directly proportional to n (#moles) & T (temperature)
inversely proportional to P (pressure)…

…combine laws into one equation: $PV = nRT$ The IDEAL GAS LAW

A proportionality constant: R = "the gas constant"
$= 8.314\ J \cdot mol^{-1}K^{-1} = 0.08206\ L \cdot atm \cdot mol^{-1}K^{-1}$

Charles' law & the absolute (Kelvin) temperature scale

If measure V vs T for various gases:

(Fig.11.5)

T (°C)	T (K)	Vol. H_2 (mL)	Vol. O_2 (mL)
300	573	47.0	21.1
200	473	38.8	17.5
100	373	30.6	13.8
0	273	22.4	10.1
−100	173	14.2	6.39
−200	73	6.00	—

Details about Avogadro's & Boyle's law exp'ts too in Kotz 11.2…

⇨ Extrapolate to "predict" absolute lowest possible temperature

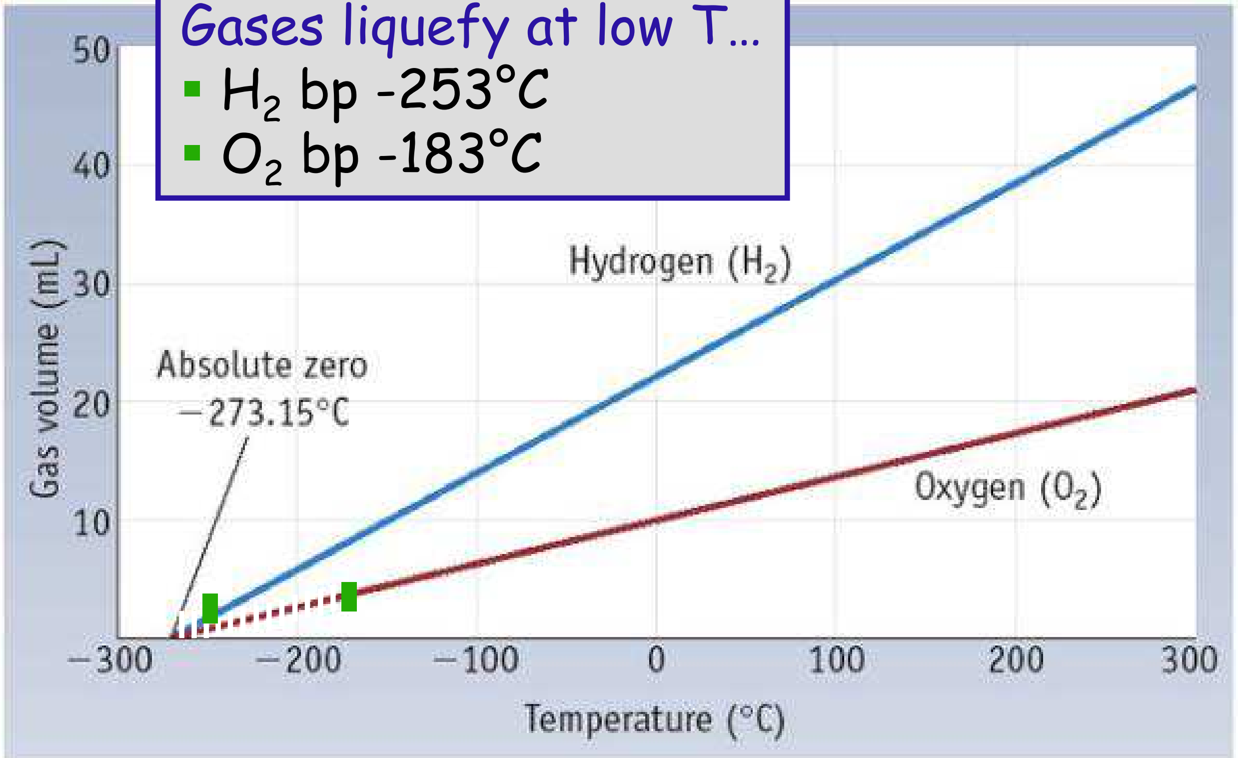
= 0 Kelvin

= -273.15 °C

The Kelvin scale was proposed by W.Thomson (aka Lord Kelvin).

Gases liquefy at low T…

- H_2 bp -253°C
- O_2 bp -183°C



Gases behave ideally if their particles don't interact much

In an IDEAL gas, the particles (atoms/molecules)

1) do not interact with each other ⇦ particles are very far apart
2) each have ~zero volume ⇦ most of container is truly empty
3) are in constant motion ⇦ enough KE to overcome all attractions, whether strong *(polar)* or weak *(nonpolar)*

THUS: **<u>Any</u> gas will behave ideally…**

- **at LOW pressures** ⇦ low concentration of particles
- **at "HIGH" temperatures** ⇦ particles moving quickly

Under these conditions, gases OBEY THE IDEAL GAS LAW

⇨ **PV=nRT accurately describes physical behaviour**

Standard Temperature & Pressure: "STP" = 1 atm, 0°C

= **conditions where most gases behave ideally** ***(i.e., "high T")***

Molar volume: 22.42 L at STP

1 mole of <u>any</u> gas occupies 22.42 L at STP

Why memorize? calculate easily using PV=nRT…

Can use the parent gas laws individually…
…to compare same sample of gas under different conditions

- **Boyle's law:** ↑P ⇨ ↓V $P_1V_1 = P_2V_2$ **when n & T are constant**
- **Charles' law:** ↑T ⇨ ↑V $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ **when n & P are constant**

CLICKER Q: Will your balloon expand or contract?

You inflate a 2.0 L balloon with He at room temperature, 25°C (298 K). You then take it outside on a hot summer day when T = 35°C (308 K). What will the balloon's new volume be? (assume P is constant)

A. 2.8 L (expanded)

B. 2.1 L (expanded)

C. 2.0 L (no change)

D. 1.4 L (contracted)

Can use PV=nRT for the two conditions to derive needed law (less memorizing…).
#n of gas does not change here:

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

Sample exam question: how many balloons?

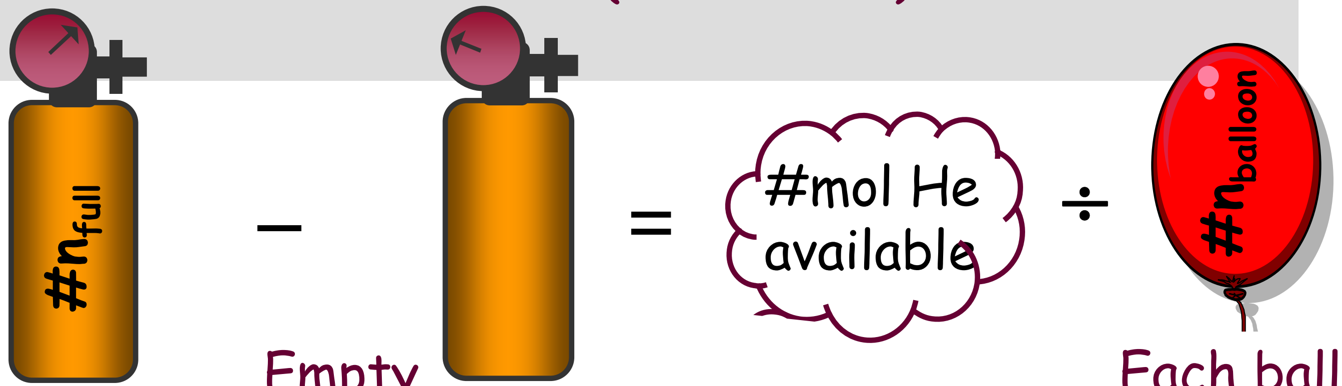
Imagine you are filling balloons with helium from a 30.0 L gas cylinder. The "full" cylinder contains 125 atm of He at a temperature of 26°C, and the cylinder is "empty" when the He pressure reaches 1.05 atm (*i.e.*, atmospheric P) at the same temperature. *Assuming ideal behaviour:* how many balloons can be <u>filled</u> to a 3.00 L volume at 1.05 atm & 23°C?

Approaches: conceptually simplest one uses only PV=nRT

1.) Calculate #moles He in "full" *vs.* "empty" tank ⇨ available He

2.) Calculate #moles He in full balloon (note new T)

3.) Divide…



Full	Empty	Each balloon
P = 125 atm	P = 1.05 atm	P = 1.05 atm
V = 30.0 L	V = 30.0 L	V = 3.00 L
T = 26°C	T = 26°C	T = 23°C

ANS: 1168.6 balloons (many ways to reach this answer)

To 3SF : 1170 (= more than can actually <u>fill</u>… ∴ better to report 1168…

11.4 Gas laws & chemical reactions: An explosive example…

What would happen if 25.0g of nitroglycerin ($C_3H_5N_3O_9$, MM 227.1 g/mol) were detonated inside an evacuated, rigid 10.0 L container at 500°C?

$$4\ C_3H_5N_3O_9(s) \rightarrow 12\ CO_2(g) + 6\ N_2(g) + 10\ H_2O(g) + O_2(g)$$

(a) What would the total pressure be inside the container?

ANS: 0.1101 mol $C_3H_5N_3O_9$ ⇨ 0.7982 mol gas ⇨ 5.06 atm

(b) If the hot gases were released out into the atmosphere, where the pressure is 1.00 atm, what would the total volume of gas be? Assume the temperature remains at 500°C.

ANS: 50.6 L

(c) Now consider that the original solid occupied approximately 15.0 cm^3 ($1cm^3 = 1mL$, or $1dm^3 = 1L$). How many times greater is the volume of the products compared to reactants (at 1 atm P)?

- ⇨much of destructive force of explosives comes from sudden formation of large volumes of gas
- ⇨forces everything out of the way…

ANS:

$$50.6\ L \times \frac{(10cm)^3}{1L}$$

$= 50600\ cm^3$

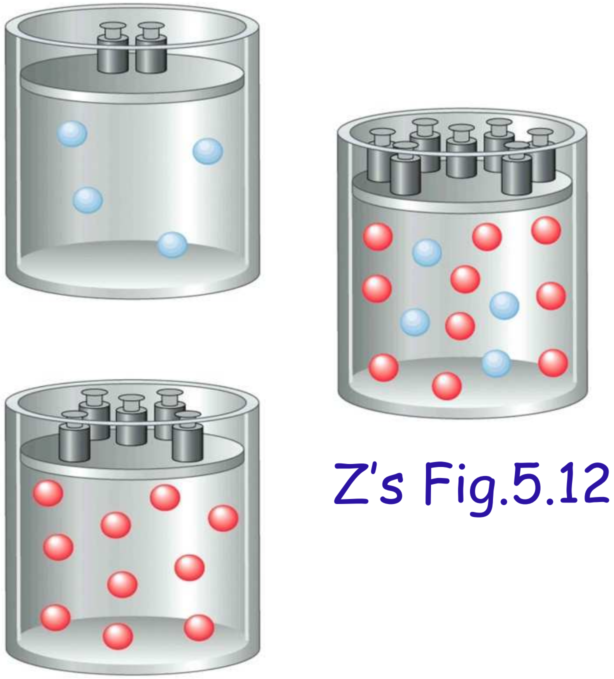
≈ 3400 x larger!

11.5 Gas mixtures – Dalton's Law of Partial Pressures

For mixture of gases in container:

$$P_{Total} = P_1 + P_2 + \ldots + P_n$$

Each gas contributes to the total pressure…



Z's Fig.5.12

BUT…

the partial pressure of each gas is **SAME AS IF ALONE** in container!

- P_1 depends only on # moles of gas *1* (use PV=nRT)
- If know P_{total} & mole ratio of gases, can calculate partial pressures

CLICKER Q: What is the typical partial pressure of oxygen in air?

Air is ~21 % O_2 and ~78% N_2 by mole, plus other trace components.
If atmospheric pressure is 1.00 atm, what is the partial pressure of O_2?

A. 0.99 atm
B. 0.78 atm
C. 0.21 atm
D. 0.01 atm

Our explosion example: now think about partial pressures

25.0g of nitroglycerin ($C_3H_5N_3O_9$, MM = 227.1 g/mol) detonated inside an evacuated rigid 10.0 L container at 500°C…

$$4\ C_3H_5N_3O_9(s) \rightarrow 12\ CO_2(g) + 6\ N_2(g) + 10\ H_2O(g) + O_2(g)$$

(d) Calculate the partial pressure of oxygen in the container.

ANS: 0.174 atm
(1/29[th] of P_{TOT})

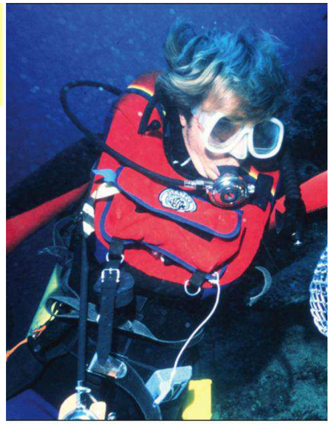
QUESTION:
Do we <u>need</u> to know the total P ?

ANS: no -- Could use stoichiometry to find $\#n_{O2}$ formed, then convert to P_{O2} using PV=nRT

(e) Imagine this explosion occurred on a large scale in a building, & the air in the room was displaced by the gaseous rxn products at 1 atm. Would a person who survived the explosion be able to breathe comfortably in that room?
(assume T = 25°C if use PV=nRT approach)

ANS: P_{O2} = (1/29) x 1.00 atm = 0.0345 atm
but we're used to P_{O2} = 0.21 atm (21% of atmospheric P)
so we would <u>not</u> get enough oxygen now…

Interesting real-world applications…



Deep sea diving (see Kotz p.534 A Closer Look)

- **At surface:** normal air, 21% O_2 ⇨ P_{O_2} = 0.21 atm
  78% N_2 ⇨ P_{N_2} = 0.78 atm
- **Using a SCUBA tank:** P's must be balanced inside & out…
  i.e., pressure of air = pressure exerted on body
  = 2 atm at 33 feet below water

  ⇨ because O_2 is still 21% of the air in the tank,
  P_{O_2} = 0.21 x 2 atm = 0.42 atm oxygen
  P_{N_2} = 0.78 x 2 atm = 1.56 atm nitrogen

- **Problems:** high P_{N_2} (depths ~below 100ft) leads to
  huge increase in N_2 dissolved in blood
  ⇨ problems with nerve conduction
  ⇨ poor judgement, giddyness…

& when return to surface: <u>bubbles</u> in blood ⇨ "the bends"

more in Chem206…

Air bag chemistry (ON YOUR OWN): exploiting a redox rxn

$$2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$$

…in the presence of SiO_2 & KNO_3, which convert the reactive Na into harmless salts.

Your tasks:

1. Draw the Lewis structure of the azide anion, N_3^-. Include all equivalent (most-stable) resonance structures.
2. Identify the oxidant and the reductant in the above reaction.
3. Answer Kotz Ch.11 #33: Find the mass of sodium azide required to fill a 75.0 L airbag to 1.3 atm pressure of $N_2(g)$ at 25°C.

When a car decelerates in a collision, an electrical contact is made in the sensor unit. The propellant (green solid) detonates, releasing nitrogen gas, and the folded nylon bag explodes out of the plastic housing.

Driver-side air bags inflate with 35–70 L of N_2 gas, whereas passenger air bags hold about 60–160 L.

The bag deflates within 0.2 second, the gas escaping through holes in the bottom of the bag.

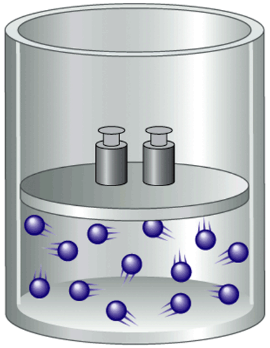
Fig.11.6

11.8 Real gases behave non-ideally at times
(things to consider if/when we must be VERY accurate)

Must correct for **non-ideal gas behavior** for samples at **high pressure** (smaller volume) and/or **low temperature**.
e.g., in compressed gas cylinders!

…WHEN INTERACTIONS ARE ACTUALLY SIGNIFICANT…

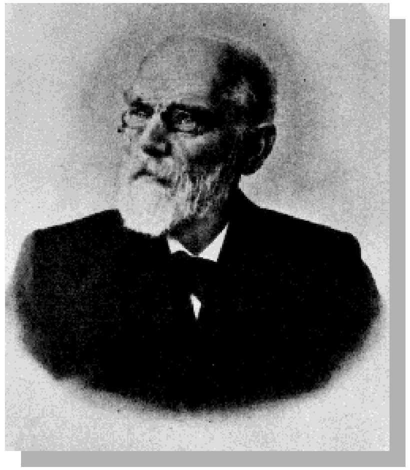
- **HIGH P *or* SMALL V:**
  - **Particles are closer together**
  - Polar molecules attract each other fairly strongly
  - Nonpolar molecules' interactions small but significant…



Zumdahl
Fig.5.29

- **LOW TEMPERATURE:**
  - **Particles are only moving slowly**
  - If not enough thermal energy to break away from intermolecular interactions, gas will NOT behave ideally!

To more accurately describe REAL GASES: use van der Waals' equation



J. van der Waals.
1837-1923, Physics
Prof., Amsterdam.
Nobel Prize 1910.

Pressure term × Volume term = nRT

$$\left(P + \frac{n^2 a}{V^2}\right) \times \left(V - nb\right) = nRT$$

Measured P

Measured V

Compensate for intermolecular forces…

Real gas particles "waste" some energy by interacting with each other

→ Observed pressure is lower than if behaving ideally

"Volume of particles" correction

Real gas particles do NOT each have zero volume

→ Total volume occupied by real gas sample therefore larger than predicted for "ideal" gas

Van der Waals' constants for various gases: Table 11.2

A real sample of Cl_2 gas: **8.00 moles in a 4.00 L tank at 27.0°C**

Ideally: P = nRT/V *note: 8.00 mol is ~ 284 g of Cl_2*

P = (8.00 mol)(0.08206 L·atm·mol^{-1}K^{-1})(300 K) / 4.00 L

= 49.2 atm ← *if the gas can behave ideally in this tank…*

For "real" Cl_2, experiments have shown: a = 6.49 atm·L²/mol²

b = 0.0562 L/mol

$$\left(P + \frac{n^2 a}{V^2}\right) \times (V - nb) = nRT$$

$\frac{n^2a}{V^2}$ = $\frac{(8.00\ \text{mol})^2 \times (6.49\ \text{atm·L}^2/\text{mol}^2)}{(4.00\ \text{L})^2}$

= 25.96 atm …interactions slow particles down…

…lost KE would be enough to cause this much more P

$V - nb$ = 4.00L – (8.00mol × 0.0562 L/mol)

= 4.00L – (0.4496 L of matter)

= 3.550 L actual empty space…

Thus: (P + 25.96 atm) × (3.550 L) = nRT

(3.550 L) P + 92.16 L·atm = (8.00 mol)(0.08206 L·atm·mol^{-1}K^{-1})(300 K)

(3.550 L) P + 92.16 L·atm = 196.94 L·atm

⇨ P = 29.5 atm

⇨ the P we'd observe in this case is MUCH lower than if behaving ideally… …implies that it must be relatively crowded in the tank!

11.6 Kinetic Molecular Theory of Gases

THEORY PROPOSED (and accepted)
TO EXPLAIN OBSERVED BEHAVIOUR OF GASES:

1) Gases consist of particles (atoms, molecules), separated by distances much greater than size of particles

2) Particles in constant, random, rapid motion…
…collide with each other and walls of container

3) Temperature determines the average kinetic energy of particles (thus: same for any gas at same T!)

$$\overline{K.E.} = \tfrac{1}{2} m \bar{u}^2$$

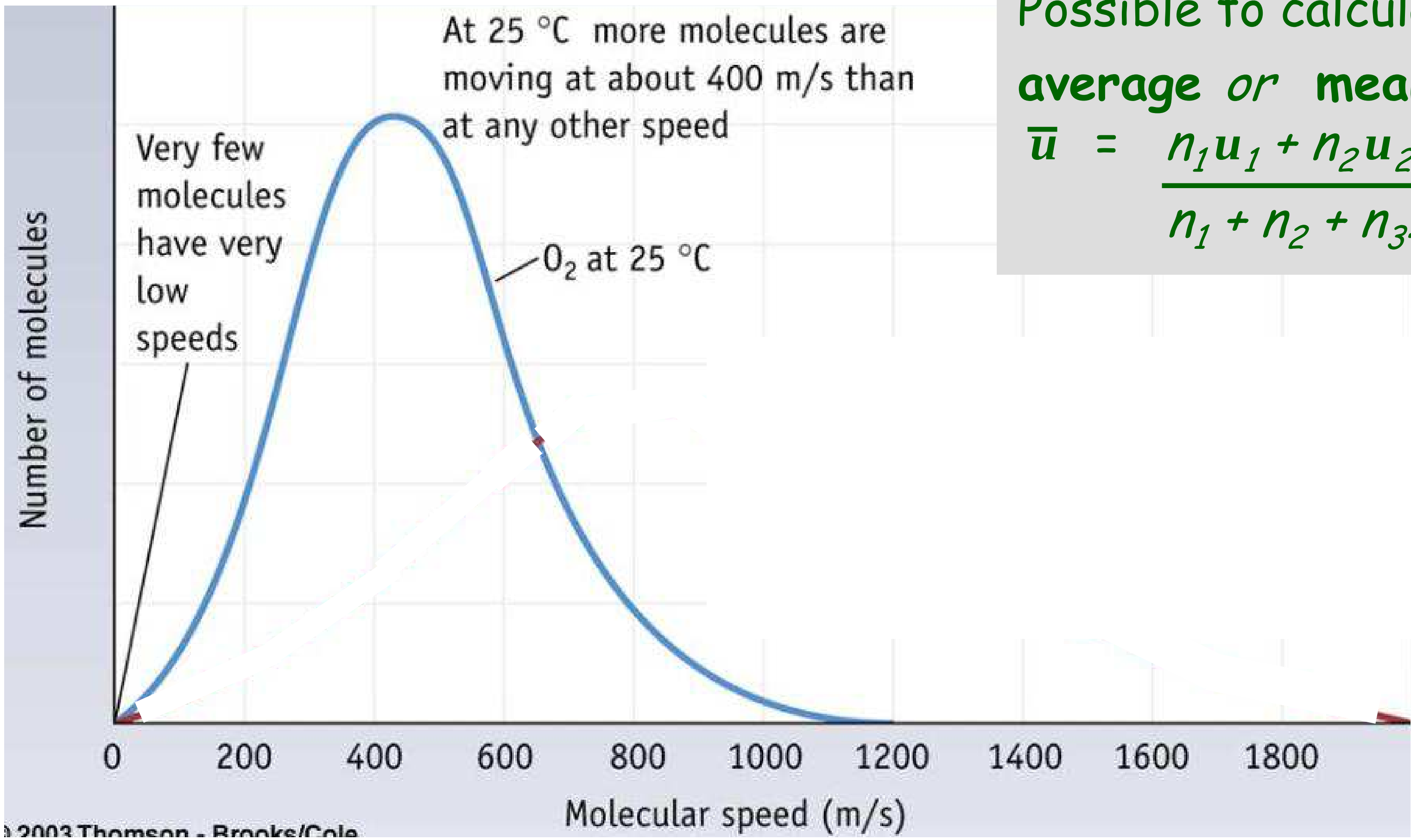
mass (m) — **Average velocity (speed)** ($\bar{u}$)

ABOVE THREE POSTULATES: Exactly what we thought about to determine when gases would behave "ideally"…

Fig. 11.13: **Velocities of N_2 molecules at two temperatures**

In a collection of molecules: DISTRIBUTION OF ENERGIES

- Some have high K.E., most have average, some have low K.E.
- Implies: not all molecules in a sample move at same speed…



Possible to calculate (not us…):

average *or* **mean speed**

$$\bar{u} = \frac{n_1 u_1 + n_2 u_2 + n_3 u_3 \ldots}{n_1 + n_2 + n_3 \ldots}$$

Heavier molecules are slower at same temp…

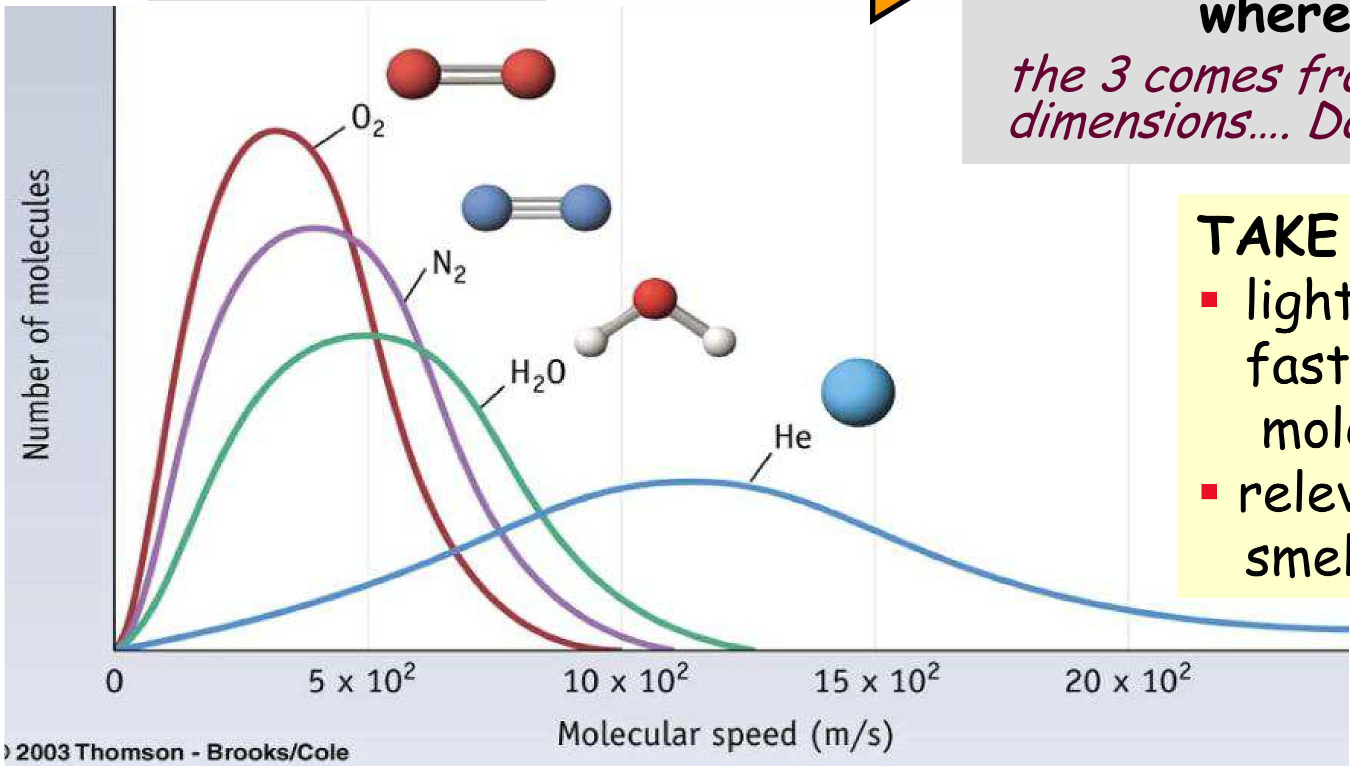
- In reality: velocity has direction *(a vector quantity)*
- Take SQUARE of average velocity then take square root
  - ⇨ removes sign: get directionless velocity called "root mean square velocity" $\bar{u}_{rms}$

- K.E. proportional to mean square velocity ($\bar{u}_{rms}^2$) & **MASS**
- If consider a mole of molecules:

$\overline{KE} = \frac{1}{2} m \bar{u}^2$ in 3D yields… $\bar{u}_{rms} = (3RT/M)^{1/2}$

where M = molar mass

the 3 comes from velocities in 3 dimensions…. Don't worry about how.

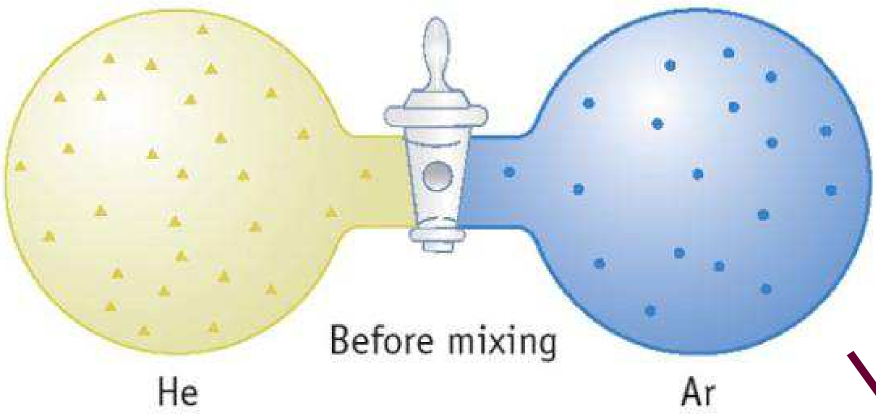


TAKE HOME MESSAGE:
- lighter molecules travel faster than heavier molecules…
- relevant to things like smelling odours!

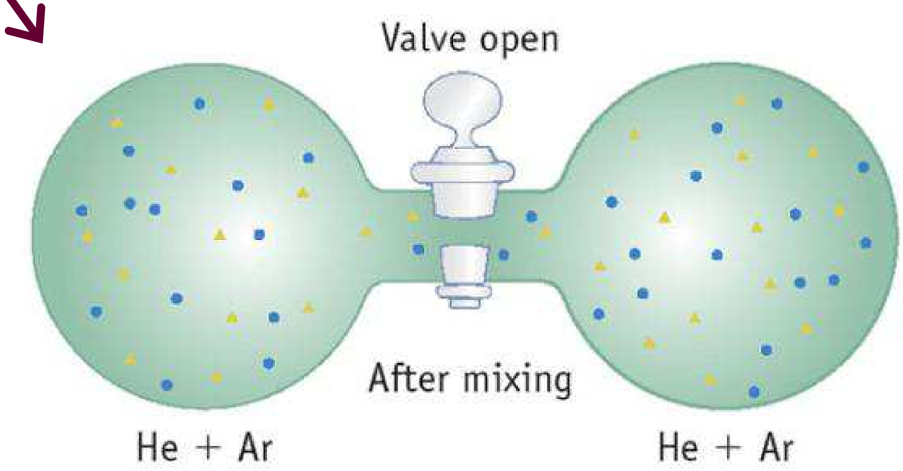
Fig. 11.14

11.7 Diffusion & Effusion: Movement of gas molecules

1. **Diffusion** = mixing of gases due to random molecular motions
   - eventually will mix completely & randomly
   - heavy particles move more slowly ⇨ diffuse more slowly



Open valve & eventually…



See Kotz Fig.11.16

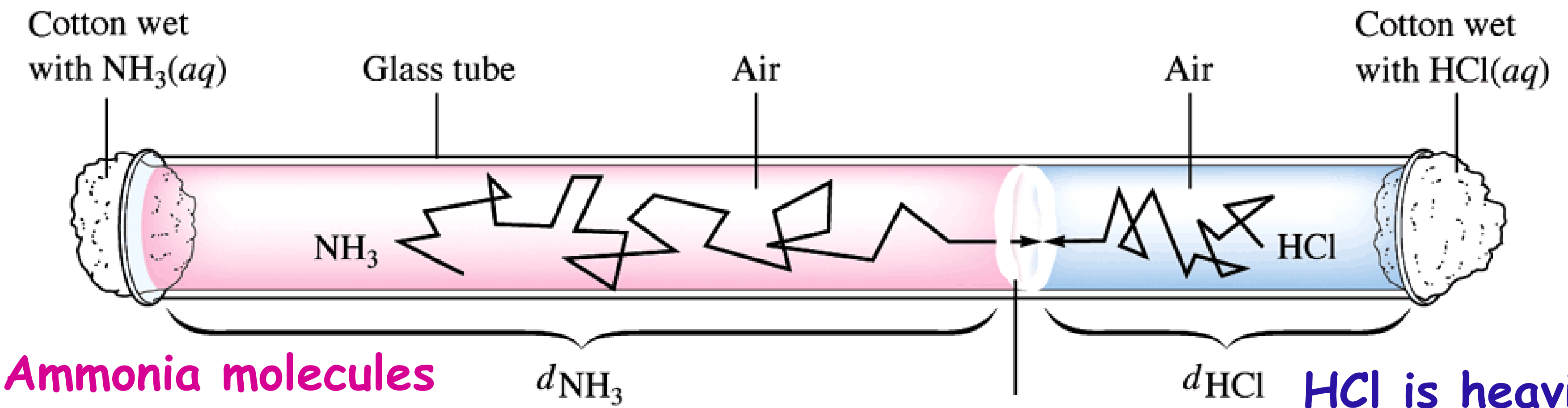
Graham's law governs effusion & diffusion

$$\frac{\text{Rate for A}}{\text{Rate for B}} = \sqrt{\frac{\text{M of B}}{\text{M of A}}}$$

Rate of diffusion/effusion is inversely proportional to gas's molar mass.

Thomas Graham
1805-1869

An experiment to demonstrate different rates of motion:

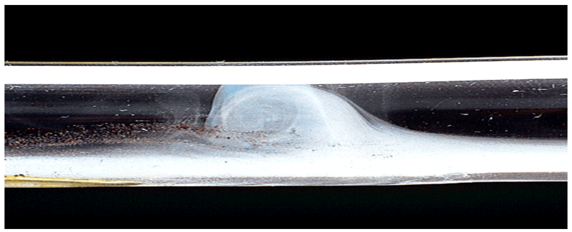


Ammonia molecules are lighter, so it diffuses more quickly

$NH_4Cl(s)$

HCl is heavier, so diffuses more slowly

HCl(*g*) + NH_3(*g*) form white ring of NH_4Cl(*s*)... →

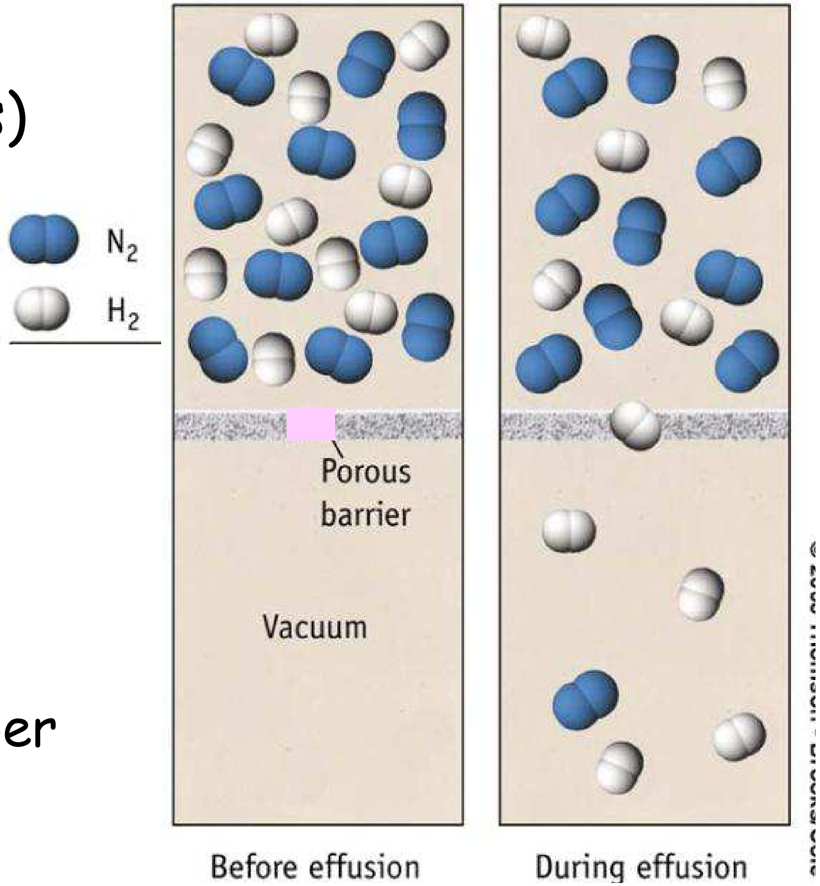


See Kotz Fig.11.17

2. Effusion

= movement of a gas through tiny hole(s)
= faster for lighter molecules…
since they hit barrier more often!

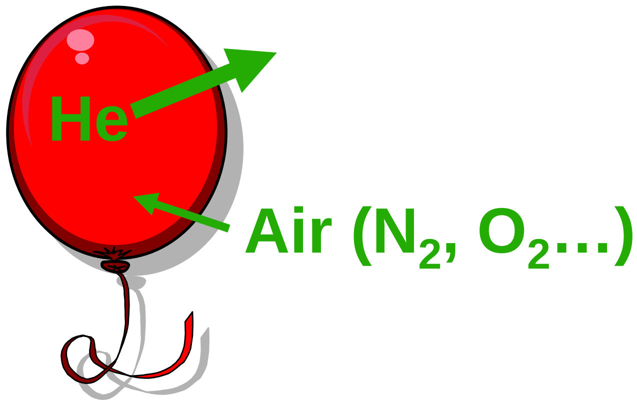
WHY DOES IT HAPPEN?
molecules collide with container walls…
but if "hit" the hole: go through!



BALLOONS:

- molecules effuse through holes in rubber
- rate (= moles/time) is:
  proportional to temperature
  inversely proportional to molar mass.

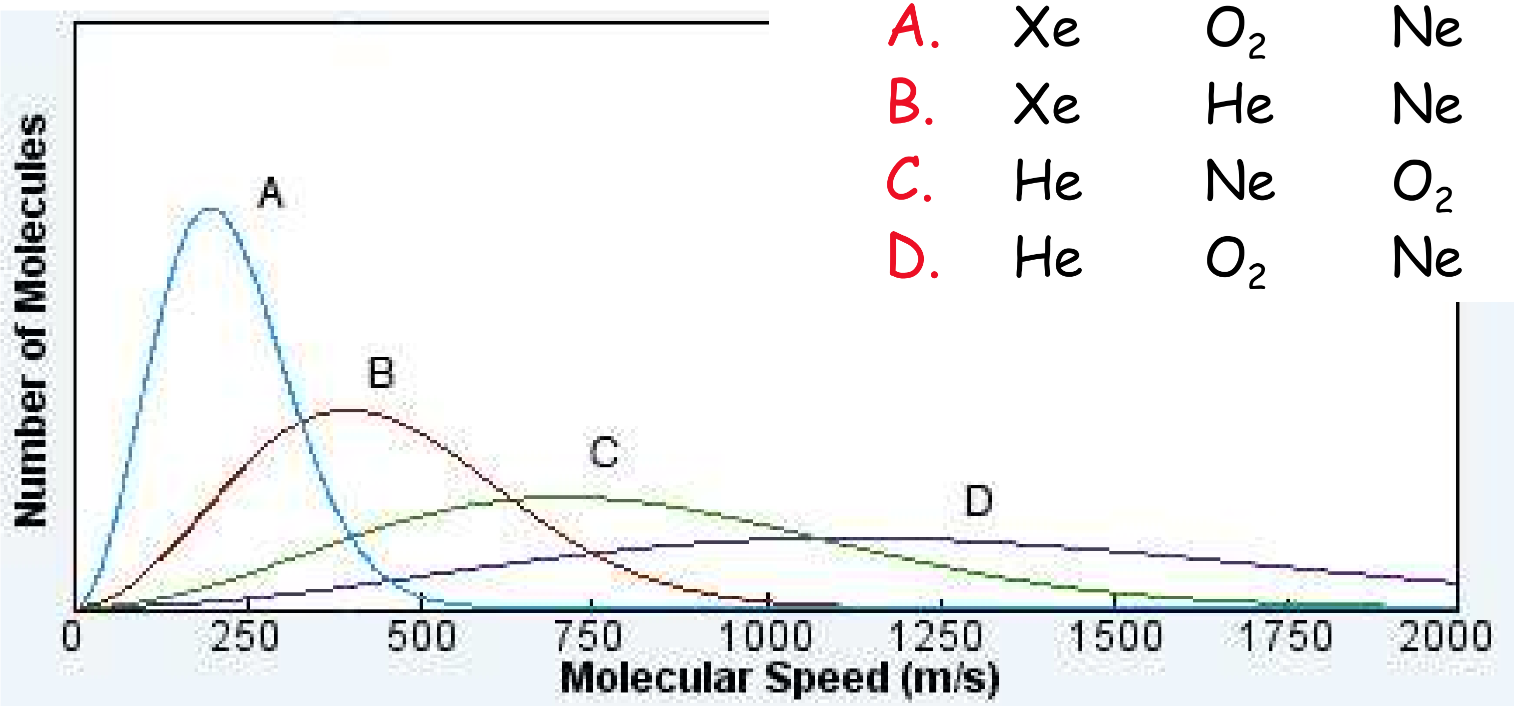
Thus: a He balloon deflates after a while…
He effuses out more rapidly than
N_2 & O_2 from air effuse in.



Compare to Fig.11.17

CLICKER Q: identifying gases based on relative molecular speed

The "Boltzmann plot" (molecules *vs* molecular speed at particular temp.) below was prepared using the Simulation in the Gases chapter in General*ChemistryNow*. Which species gives rise to which curve?



Species:	A	B	C	D
A.	Xe	O_2	Ne	He
B.	Xe	He	Ne	O_2
C.	He	Ne	O_2	Xe
D.	He	O_2	Ne	Xe

See Kotz Review & Check for Section 11.6

To end CHEM 205: Sustainability & green chemistry (1.2)

Reducing greenhouse gases by capturing $CO_2(g)$

Case Study
Ch.11 p.536

- Greenhouse gases: prevent heat leaving atmosphere ⇨ **global warming**
- Industrial activity produces more CO_2 than plants (CO_2 sinks) can consume

What if we could trap it via reaction, then bury it deep underground?

$$CO_2(g) + (NH_4)_2CO_3(s) + H_2O(\ell) \rightleftharpoons 2\ NH_4HCO_3(s)$$

- Pass fossil-fuel combustion gases through scrubber to remove SO_2

- Pass gas stream through slurry
- Captures CO_2 but not N_2

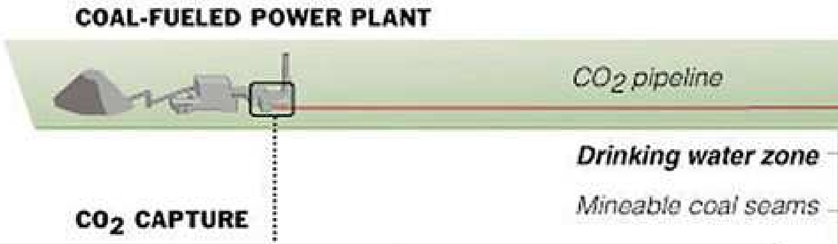
- Slurry isolated & heated…
- Reaction reverses to regenerate $(NH_4)_2CO_3$ & release pure CO_2

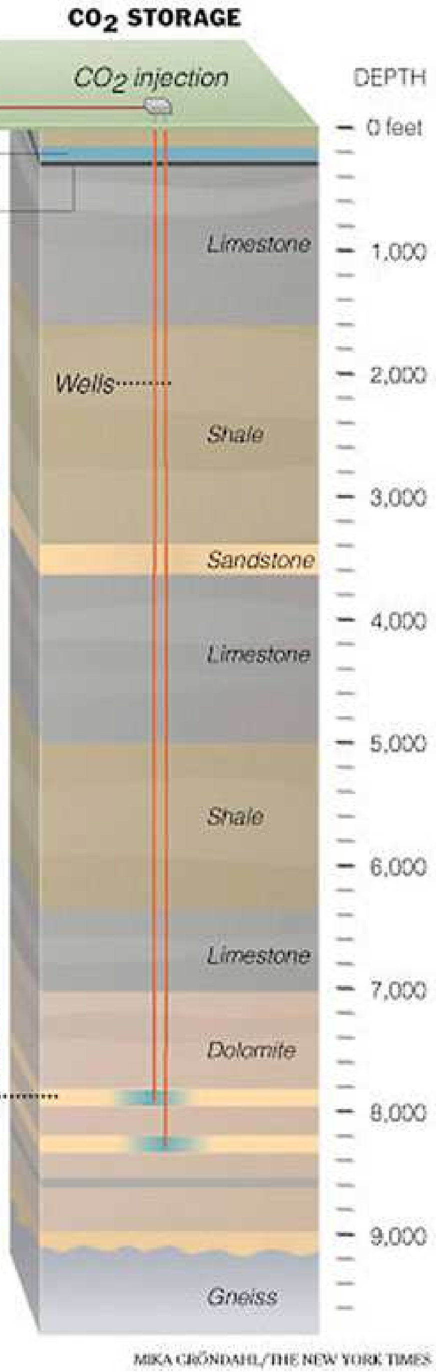
Then: **pure CO_2 "geologically sequestered"**…
- injected into porous rock layers, 2400m deep
- buried under impermeable rock layers
- trapped for 1000's of years…

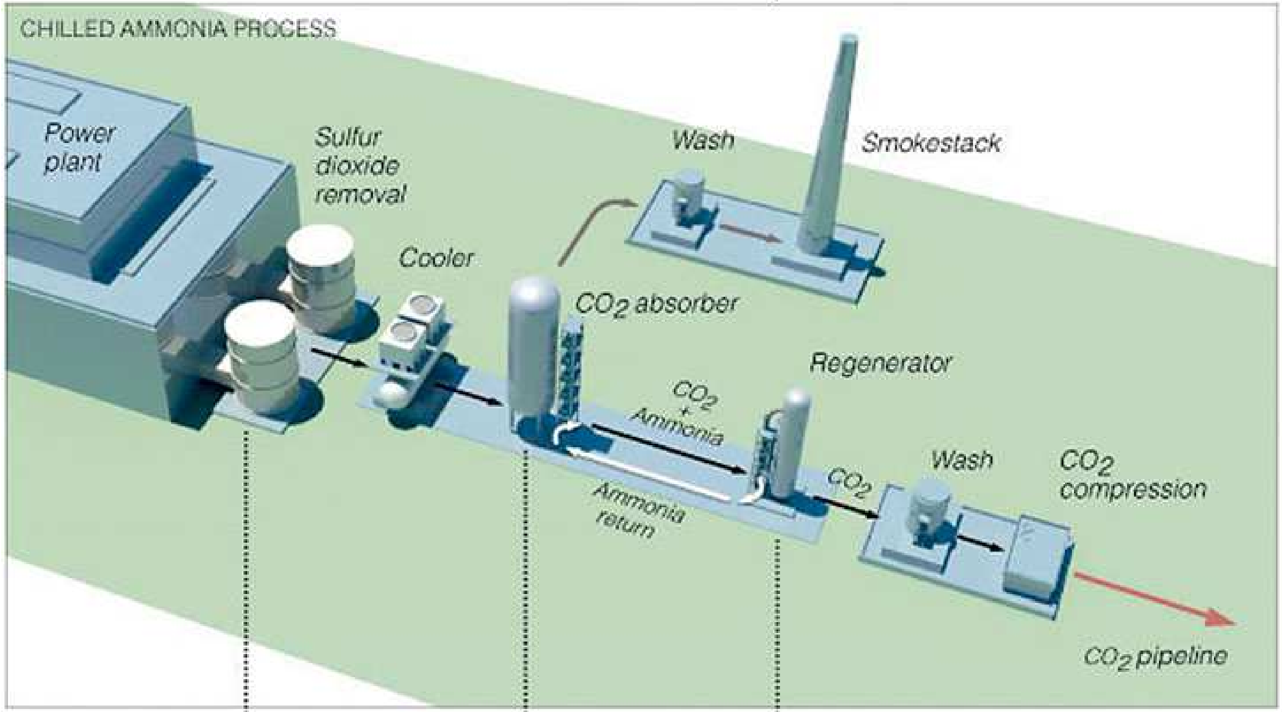
?....

Captured, Then Buried

A coal-fired electricity plant in West Virginia plans to capture some of the carbon dioxide it emits and inject it underground. Some environmentalists are wary.







1 A scrubber removes sulfur from the plant's exhaust. The remaining exhaust is cooled to reduce its volume in preparation for removal of carbon dioxide.

2 The exhaust enters a chemical reactor where it is mixed with an ammonia-based chemical, also chilled, that binds with the carbon dioxide. Then the exhaust is washed and sent to a smokestack.

3 The ammonia solution is reheated and gives off its carbon dioxide. The ammonia is rechilled and sent for reuse; the carbon dioxide is squeezed down to a fluid.

4 The carbon dioxide is pumped into porous rock layers deep underground, beneath impermeable layers that should prevent its escape. Underground, it spreads horizontally.

Schematic diagram

Case Study p.536

Source: American Electric Power; Alstom; Battelle

MIKA GRÖNDAHL/THE NEW YORK TIMES

To end CHEM 205: Sustainability & green chemistry (1.2)

Green chemistry principles to live by: (much of this applies to everyday life…)

- **PLAN AHEAD:** Prevent waste rather than cleaning it up afterwards.
- **BE EFFICIENT:** Design synthetic methods to maximize the incorporation of starting materials into the final product.
- **USE LESS ENERGY:** Perform reactions at ambient temperatures and pressures, to minimize cost & environmental impact of energy usage.
- **BE CAUTIOUS:** Choose substances that minimize the potential for chemical accidents, including releases, explosions and fires.
- **LIMIT TOXICITY:**
  - Design synthetic methods to use & generate substances that possess little/no toxicity to human health or the environment.
  - Design chemical products to function effectively while still reducing toxicity.
- **GO RENEWABLE:** Choose renewable raw materials whenever technically and economically practical.
- **GO BIODEGRADABLE:** Design chemical products so that, at the end of their function, they do not persist in the environment or break down into dangerous products.

FINAL EXAM INFORMATION

- **Final exam (tentatively): Tues. Dec.10th (7-10pm) @SGW…**
  (verify DATE & location!)

- **Covers entire course:** Ch.1-4, 20.1, 6-8, 13.4, 9.1-2, 11

- The examination room invigilators are VERY STRICT:
  - Student ID card mandatory
  - No programmable calculators
  - **No electronic dictionaries, cell phones, pagers, blackberries, *etc***
  - Book-format translation dictionaries (word-to-word only) allowed, but they will be inspected.
  - Arrive to the exam room early !

FINAL EXAM FROM PREVIOUS YEARS ON WEBSITE

Extra examples…

Stoichiometry involving gases: *cf* Kotz 6th Ed. Ch.12 #62

Dimethyldichlorosilane $(CH_3)_2SiCl_2$ is a starting material used to make silicones, which are polymeric substances used as lubricants, anti-stick agents and water-proofing caulk. It is formed by the following reaction:

$$Si(s) + 2\ CH_3Cl(g) \rightarrow (CH_3)_2SiCl_2(g)$$

If you place 2.25g of solid silicon in a 6.56 L flask containing CH_3Cl at a pressure of 585 mm Hg at 25°C…

a) what mass of dimethyldichlorosilane can be formed?
b) what pressure would be exerted by this $(CH_3)_2SiCl_2(g)$ at 25°C?
c) what would be the total pressure in the flask after the reaction?

Please provide FULL details in written calculations on exam, AND comments at each step to explain what you are doing.

Additional questions:

1. What type of reaction is this? How can you tell?
2. How would you describe the reactants?
   a) Element (metal, nonmetal, metalloid?) *vs* compound?
   b) For compounds: ionic *vs* covalent? polar *vs* nonpolar?
   c) Soluble in water?

Stoichiometry involving gases: *cf* Kotz 6th Ed. Ch.12 #62

ANS: provide FULL details on exam – outline given here only

$\#n_{Si}$ = 0.08011 mol *vs.* $\#n_{CH3Cl}$ = 0.2065 (using PV=nRT)
= 1 Si *to* 2.6 CH_3Cl

THUS: have excess CH_3Cl, since rxn requires 1:2 ratio ⇨ Si is limiting.

a) Yield of $(CH_3)_2SiCl_2$ (based on 1:1 stoich.) = $\#n_{Si}$ = 0.08011 mol
$MM_{(CH3)2SiCl2}$ = 129.06 g/mol ⇨ $m_{(CH3)2SiCl2}$ = 10.3 g

b) Using PV=nRT: this gas would exert a P of 0.299 atm = 227 mm Hg
= the partial pressure of $(CH_3)_2SiCl_2$ in the flask
= $P(CH_3)_2SiCl_2$ …if other gases present, total P will be greater than this…

c) What would be the <u>total</u> pressure in the container after the rxn?

TOTAL pressure: $P_{TOT} = P(CH_3)_2SiCl_2 + PCH_3Cl(excess)$

Excess $\#n_{CH3Cl}$ = $\#n_{initial}$ - $\#n_{reacted}$ ← = 2x $\#n_{Si}$
= 0.2065 mol - 0.1602 mol
= 0.04628 mol

Thus: $PCH_3Cl(excess)$ = 0.173 atm = 131 mm Hg ⇨ P_{TOT} = 227 + 131
= 358 mm Hg