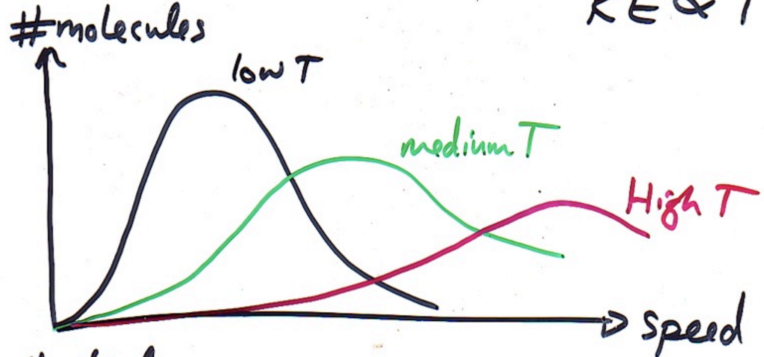
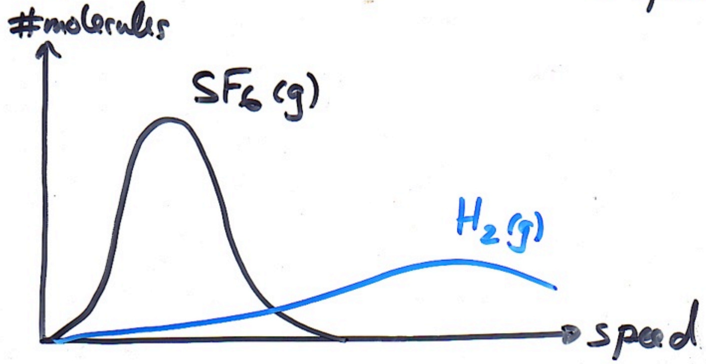
ARIS

Maxwell-Distribution Curve

$KE \propto T\,(K)$





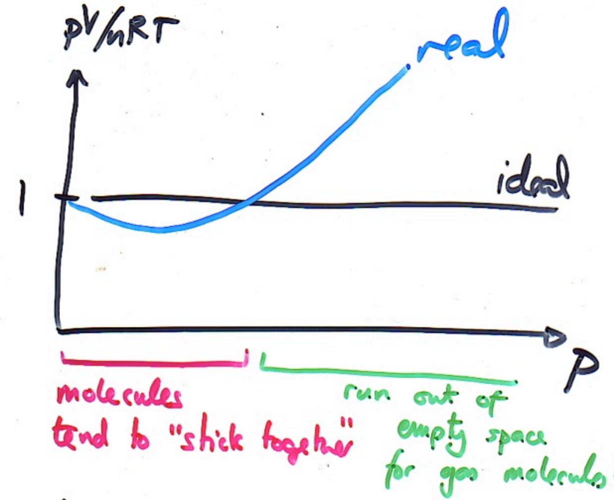
Ideal Gas Equation

Only works for IDEAL gases!

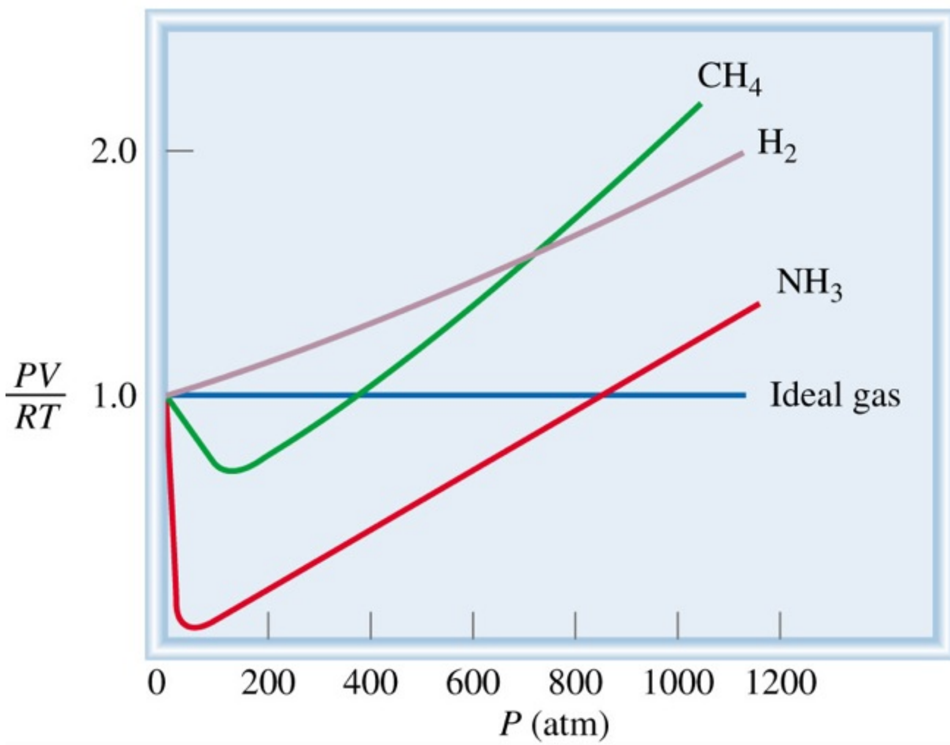
$$pV = nRT$$

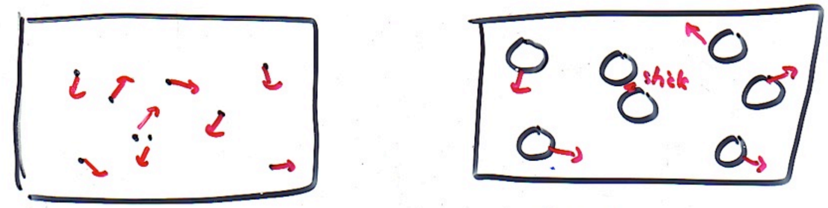
$$\frac{pV}{nRT} = 1$$

Deviations from ideal behavior.



IDEAL GAS	REAL GAS
- no attractions between the gas molecules.	- are attractions between gas molecules.
molecules themselves have no volume	- real molecules have size.





IDEAL REAL

$pV = nRT$

<u>van der Waals</u> equation.

$$\left(p + \frac{an^2}{V^2}\right)\left(V - nb\right) = nRT$$

effective pressure — effective volume

a ~ "stickiness" or intermolecular forces

b ~ "size" of molecules

ex: $H_2O(g)$

$a = 5.46 \frac{atm \cdot L^2}{mol^2}$ $b = 0.0305$ L/mol

Q. What is the pressure of 1.20 mol $H_2O(g)$ @ 382 K, when V = 1.50 L

~ use the ideal + vdW equations

Table 5.3

van der Waals Constants of Some Common Gases

Gas	a $\left(\frac{atm \cdot L^2}{mol^2}\right)$	b $\left(\frac{L}{mol}\right)$
He	0.034	0.0237
Ne	0.211	0.0171
Ar	1.34	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0266
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
Cl_2	6.49	0.0562
CO_2	3.59	0.0427
CH_4	2.25	0.0428
CCl_4	20.4	0.138
NH_3	4.17	0.0371
H_2O	5.46	0.0305

IDEAL

$$pV = nRT \Rightarrow p = \frac{nRT}{V} = \frac{1.20\,mol \times 0.08206\,\frac{atm \cdot L}{mol \cdot K} \cdot 382\,K}{1.50\,L}$$

$$= 25.1\,atm.$$

vdW $a = 5.46\ \frac{atm \cdot L^2}{mol^2}$ $b = 0.0305\ L/mol$

$$\left(p + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

$$\Rightarrow p + \frac{an^2}{V^2} = \frac{nRT}{V - nb}$$

$$\Rightarrow p = \frac{nRT}{V - nb} - \frac{an^2}{V^2}$$

$$= 22.2\,atm$$

REAL pressure: 20.15 atm.

ideal value: off by ≈25 %.

vdW : off by ≈10%.

Ch 6 Thermochemistry

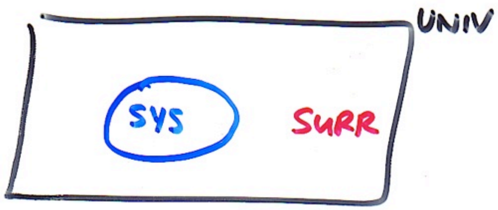
Energy ~ more forms
~ one form is heat

we're going to study heat gain/loss in chemical rxns.

Some Definitions

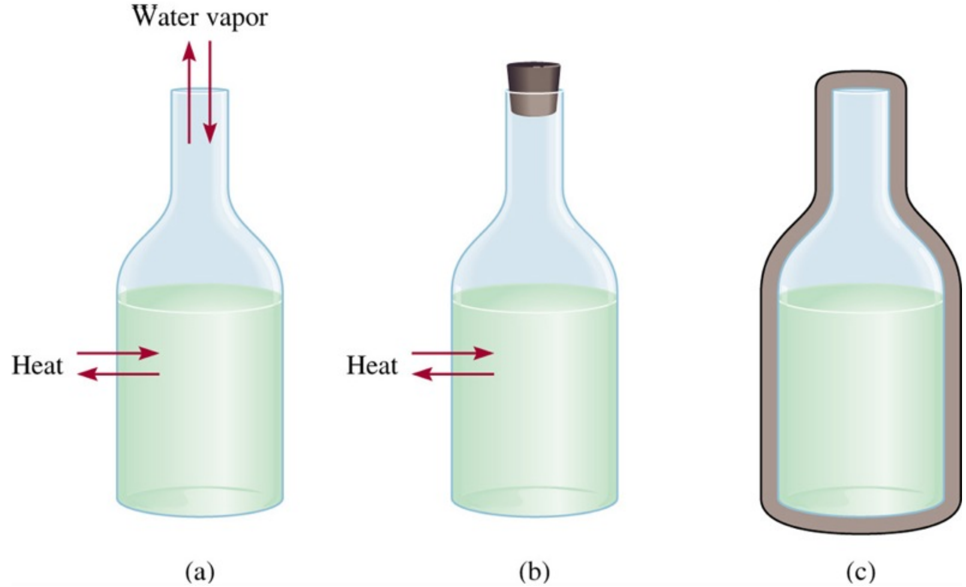
UNIVERSE = SYSTEM + SURROUNDINGS

(SYSTEM → chemical rxn.)



3 types of systems:

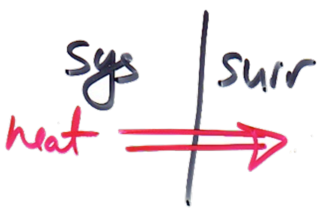
(1) OPEN heat + matter can flow sys ↔ surr
(2) CLOSED heat can flow, but not matter "
(3) ISOLATED Nothing can flow "



More definitions...

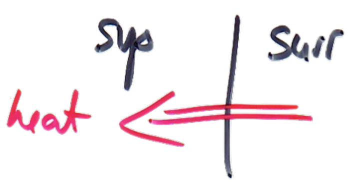
if heat is lost (released, evolved, ...) from the system to the surroundings:

EXOTHERMIC



if heat is absorbed by the system from the surroundings:

ENDOTHERMIC



1st Law of Thermodynamics

$$E_{system} + E_{surroundings} = E_{UNIVERSE}$$

$$\Delta E_{sys} + \Delta E_{surr} = 0$$

energy can neither be created nor destroyed.

Sign convention: if energy leaves,
$\Delta E = -ve$

if energy enters,
$\Delta E = +ve$

if only form of energy is heat, then 1st Law becomes:

$$q_{sys} + q_{surr} = 0$$

ex: if the system loses 10J of heat:
$q_{sys} = -10J$, then surroundings gain

Table 6.1 Sign Conventions for Work and Heat

Process	Sign
Work done by the system on the surroundings	−
Work done on the system by the surroundings	+
Heat absorbed by the system from the surroundings (endothermic process)	+
Heat absorbed by the surroundings from the system (exothermic process)	−

10J of heat : $q_{surr} = +10J$.

$$q_{surr} + q_{sys} = 0$$

$$+10J - 10J = 0$$

when Heat flows into objects, it increases its temperature, and vice-versa

How to measure q

- Calorimetry

Heat capacity = heat req'd to increase an object's temp by 1°C (or 1 K)

C

$$\Rightarrow q = C \cdot \Delta t$$

ex:

A Gold crown with a heat capacity of 522 J/°C increases by 15°C in temperature, then how much heat does it absorb?

$$q = C \cdot \Delta t$$

$$= 522\ J/\cancel{°C} \times +15\ \cancel{°C}$$

$$= 7800\ J$$

1 Joule = 1 J

↑

SI unit of energy

(James Joule, ≈1750)

mid 1800s actually!