Chapter 5 + Section 1.6 (end) → Free Energies

Enthalpy → energy plus the work to make room for the system under constant pressure from the environment

$$H \equiv U + PV$$ ← total energy to create the system and put it into an environment

But, the system can get its thermal energy from the surroundings, so all we really need to do is provide to the system any additional work

Gibbs Free Energy

$$G \equiv H - TS$$
$$G \equiv U + PV - TS$$

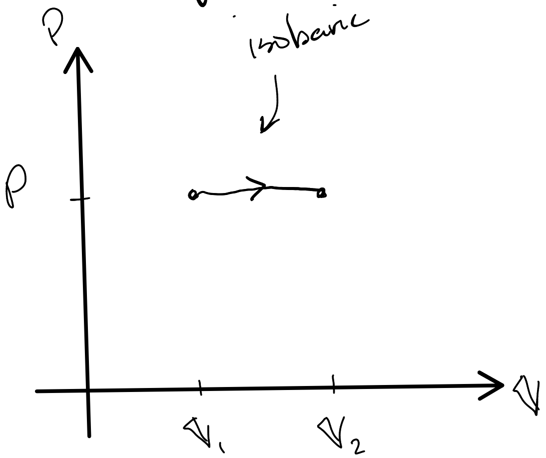
← energy to make the system and make room for the system minus the heat we get to extract from surroundings.

Helmholtz Free Energy

→ total energy to create the system minus the heat we can extract from the surroundings

$$F = U - TS$$

Enthalpy



#1.34

$$W = -P(V_2 - V_1)$$

$$\Delta U = \frac{f}{2} N k_B \Delta T = \frac{f}{2} P \Delta V = \frac{f}{2} P (V_2 - V_1)$$

$$N k_B \Delta T = P \Delta V$$

— or —

$$= \Delta P \cdot V$$

— but not —

$$= \underline{\underline{\Delta P \cdot \Delta V}}$$

$$Q = \Delta U - W = \Delta U + P \Delta V$$

$$Q = \frac{f}{2} P (V_2 - V_1) + P (V_2 - V_1)$$

$$Q = \left(\frac{f}{2} + 1\right) P \Delta V$$

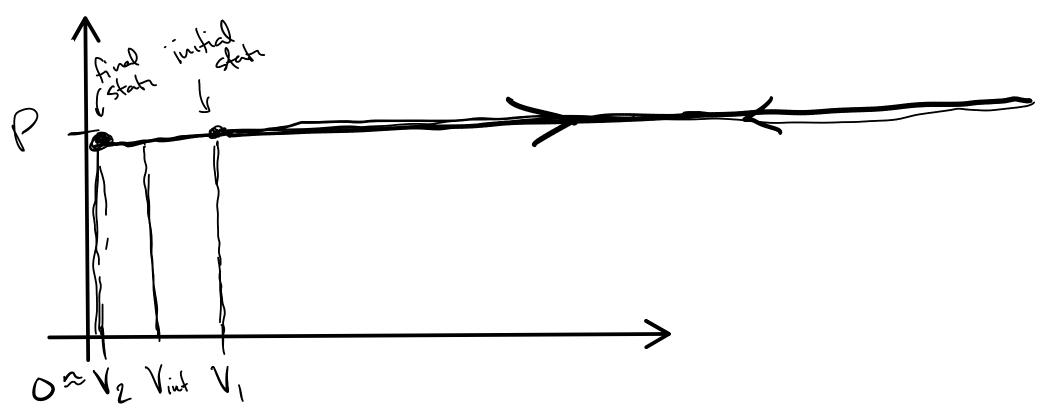
$$H \equiv U + PV$$

$$\Delta H = \Delta U + P \Delta V = Q$$

Another example, what if we cause something to happen inside the tube & we measure the heat given off at constant pressure.

$$H_2(g) + \frac{1}{2} O_2(g) \longrightarrow H_2O(l)$$

$$\Delta H = -286 \text{ kJ/mol}$$



1.4a)

1 mol $H_2(g)$
0.5 mol O_2

1.5 mol → n_m

$$PV = n \cdot R T$$

$P = 10^5$ Pa, $n = 1.5$ mol, $R = 8.31$, $T = 298$ K

1 mL = 1 cm^3

1000 L = 1 m^3

$$V = \frac{nRT}{P} = \frac{1.5 \cdot 8.31 \cdot 298}{10^5} = 0.037\ m^3 = 37\ L$$

n_m after the reaction?

$$V = \frac{n RT}{P}$$ ($\frac{RT}{P}$: constant) → $V \propto n$ → $\frac{V_{int}}{V_1} = \frac{n_2}{n_1} = \frac{1}{1.5}$

$$V_{int} = \frac{1 \cdot 8.31 \cdot 298}{10^5} = 0.0248\ m^3$$

$V_{int} = 0.0248\ m^3$

this condenses to give ~0 L of liquid, giving up heat along the way

So what total work has been done?

$$W = -\int_{0.037\,m^3}^{0\,m^3} p\,dV = -10^5\,Pa\,(0 - 0.037\,m^3)$$

$$= 3700\,J = 3.7\,kJ$$

→ $\Delta H = -286\,kJ = \Delta U - W$

$-286\,kJ = \Delta U - 3.7\,kJ$

$$\boxed{\Delta U = -282.3\,kJ}$$

1.50) $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$

O_2: 0 kJ/mol

a) $CH_4 \rightarrow C + 2H_2$

CH_4: −74.81 kJ/mol; C, H_2: 0 kJ

it would take 74.81 kJ/mol to convert $C + H_2$ to CH_4

b) $C + O_2 \rightarrow CO_2$

C, O_2: 0 kJ/mol; CO_2: −393.51 kJ/mol

$2H_2 + O_2 = 2H_2O(g)$

$\Delta H = -241.82$ kJ/mol

× 2

−483.64 kJ

c) $\Delta H_{reaction} = 74.81 - 393.51\,\text{kJ/mol} - 483.64\,\text{kJ}$

$\Delta H_{reaction} = -802.34\,\text{kJ}$ → heat given to the surroundings

d) 802.34 kJ

e) $\underline{\Delta H} = U + \underline{\underline{p\Delta V}}$ ← find this

$\Delta H = U$ when final mixture is a gas

$pV = nRT \rightarrow p\Delta V = \Delta n R \cdot T$

$= 0\,\text{mol} \cdot 8.31 \cdot 298\,\text{K}$

what if $H_2O(g) \rightarrow (l)$

$CH_4 + 2O_2 \Rightarrow CO_2 + \underline{\underline{2H_2O(l)}}$

3 mol of gas → 1 mol of gas

$p\Delta V = -2\,\text{mol} \cdot 8.31 \cdot 298\,\text{K}$

$= -4.9\,\text{kJ}$

$\Delta H_{reaction} = ?$

f) ?

also do 1.55

Gibbs Free Energy

$$G = U + PV - TS$$

$$G = H - TS$$

$$dG = dU - TdS + pdV$$

$$dG = đQ + đW - TdS + pdV$$

- for a reversible process (no new entropy is created)
  - $đQ_{rev} = TdS$
  - $đW = -pdV + đW_{other}$
    - ↳ electrical

$$dG = \cancel{TdS} \cancel{-pdV} + đW_{other} \cancel{-TdS} \cancel{+pdV}$$

$$dG = đW_{other}$$

not reversible

$$dG \leq đW_{other} \text{ at constant } T, p$$

Thermodynamic Potentials

$$dF = dU - TdS$$

$$dF = đQ + đW - TdS$$

for a reversible process (no new entropy is created)

$$đQ_{rev} = TdS$$

$$dF = Tds + đW - TdS$$

$$dF = đW$$

↳ amount of work done on the system by you or by surroundings

otherwise, (not a reversible process)

$$dF \leq đW$$

How is ΔG_f measured?

First measure $\Delta H \rightarrow$ heat absorbed or released for a reaction at constant pressure + no other work has been done

Then calculate ΔS from initial + final states using heat capacity data

$$\Delta S = \int_0^{T_f} \frac{C_p}{T} dT$$

$$\Delta G = \Delta H - T\Delta S$$

So lets do an example: (5.2)

$$N_2 + 3H_2 \rightarrow 2NH_3 \quad \text{at } 298K + 1 \text{ bar}$$

N_2: $\Delta H = 0 kJ$, $S = 191.61 J/K mol$

H_2: $\Delta H = 0 kJ$, $S = 130.68 J/K mol$

NH_3: $\Delta H = -46.11 kJ$, $S = 192.45 J/K mol$

$$\Delta H = 2 mol \cdot (-46.11 kJ/mol) = -92.22 kJ$$

$$\Delta S = S_f - S_i = 2 \cdot 192.45 J/K mol - 3(130.68) - 191.61$$

$$= -198.75 J/K$$

$$\Delta G = \Delta H - T\Delta S$$ watch out for kJ!

$$= -92.22\,kJ - 298(-0.19875\,kJ/K)$$

$$\underline{\underline{\Delta G = -32.99\,kJ}}$$

compare!

from the book

$$\Delta G = -16.45\,kJ/mol$$

for this reaction

$$\Delta G = 2\,mol \cdot (-16.45\,kJ/mol) = \underline{\underline{-32.9\,kJ}}$$

Electrolysis, Fuel Cells, Batteries (another example)

$$H_2O \rightarrow H_2 + \tfrac{1}{2}O_2$$

electrical work must be done

$\Delta H = \underline{286\,kJ}$ ⇝ heat out when we run in reverse

To run this reaction forward we need to put in 286 kJ of energy

→ Some of this will push aside the atmosphere

$$p\Delta V = \Delta n RT = 1.5\,mol \cdot (8.31\,J/K\,mol)(298\,K) = 3714\,J$$

gas form

$$= 3.7\,kJ$$

$$\sim 4\,kJ$$

→ so, 282 kJ will remain in the system in the form of potential energy in the chemical bonds

Do we need to do all of this work, or can we get some help from surroundings?

What is the change in entropy?

$$S_{H_2O} = 70\,J/K \qquad \underbrace{S_{H_2} = 131\,J/K \qquad S_{O_2} = 205\,J/K}_{\text{products}}$$ ← from the back of the book

$$S_{prod} = 131\,J/K + \tfrac{1}{2}(205\,J/K) = 233.5\,J/K$$

$$\Delta S = S_p - S_r = 233.5\,J/K - 70\,J/K = \underbrace{163.5\,J/K}_{\text{entropy increased}}$$

$$đQ \leq T\,dS$$

$$Q = T\Delta S = 298\,K \cdot 163.5\,J/K$$

$$Q = \underline{48.7\,kJ}$$

So of the 286 kJ that go in. 48.7 kJ can come from heat from surroundings

$$286\,kJ - 48.7\,kJ = \underline{237.3\,kJ}$$ → come from electrical work

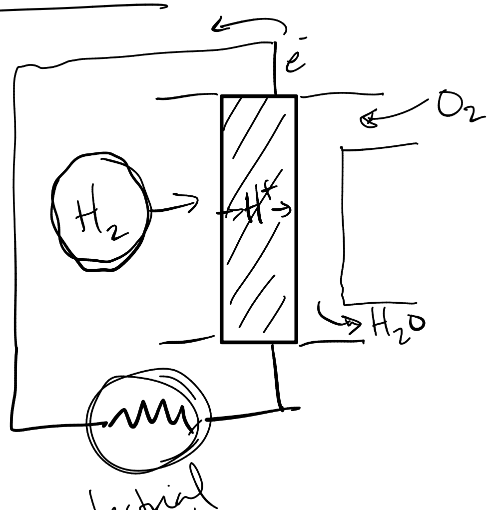
$$\Delta G = \Delta H - T\Delta S$$

$\Delta G_{H_2O} = -228.57\,kJ$ $\quad \Delta G_{H_2} = 0\,kJ$ $\quad \Delta G = 0\,kJ$

HW 5.1 + 5.5

$\Delta G = \Delta G_{prod} - \Delta G_{react} = 0 - 228.57$
$= \underline{228.57 \text{ kJ}}$

In reverse, we can extract 237 kJ of electical work



→ 49 kJ of waste heat goes to the environment

→ $\Delta U = 282$ kJ → comes from chemical reactions

→ $P\Delta V = 4$ kJ → comes from the collapse of gas to liquid

efficiency = $\frac{\text{get}}{\text{pay for}}$

fuel cell

$\frac{237 \text{ kJ}}{286 \text{ kJ}} = \underline{83\%}$

compare to efficiencies we would have gotten in chapter 4

max theoretical engine efficiency → 66%

max theoretical car engine efficiency → 50%
↳ practically → 20-30%

gray H_2 blue h_2 green h_2
↳ solar

Battery

$$Pb + PbO_2 + 4H^+ + 2SO_4^{-2} \longrightarrow 2PbSO_4 + 2H_2O$$

start w/ $\Delta H = -316\ kJ/mol$

energy out from reaction at constant pressure

what about $\Delta G = -394\ kJ/mol$

electrical energy out

ΔG is more than the ΔH. How? Heat is coming in to the system from surroundings!

How much?

$394\ \frac{kJ}{mol} - 316\ \frac{kJ}{mol} = 78\ \frac{kJ}{mol}$

$\Delta G = \Delta H - T\Delta S$

$\Delta G - \Delta H = -T\Delta S$

$-394 - (-316) = -298\ \Delta S$

$\Delta S > 0$

$78\ \frac{kJ}{mol} = T\underbrace{\Delta S}_{\text{increase in entropy}}$

means the heat is being absorbed from surroundings $\leftarrow$ $\Delta S = \frac{78\ kJ}{298\ K} = 261.7\ \frac{J}{K}$ $\leftarrow$ watch out for units

Energy/charge $\rightarrow$ potential (voltage)

gibbs

$$\text{voltage} = \frac{\Delta W_E}{q} = \frac{394\text{ kJ}}{2e \cdot N_A \cdot 1.6 \cdot 10^{-19}\, C/e} = \underline{2.04\text{ V}}$$

$\leftarrow$ per "cell" in a car battery

6 "cells" $\rightarrow$ $\underline{12\text{V}}$

need to know how many electrons are moved during the reaction

$2e^-$

Thermodynamic identity

$$dU = T\underline{dS} - P\underline{dV} + \mu \underline{dN}$$

but what about H, F, G? How do these change when conditions change?

$\rightarrow$ $H = U + pV$

$$dH = dU + \cancel{p\,dV} + V\,dp$$

$$\hookrightarrow T\,dS - \cancel{p\,dV} + \mu\,dN$$

$$dH = T\underline{dS} + V\underline{dp} + \mu\underline{dN}$$

So the variables have changed from dU

we tend to use enthalpy

when $dp = 0$

$dN = 0$

$dH = T\,dS$ $\leftarrow$ heat at constant pressure what about F?

$\rightarrow$ $F = U - TS$

$$dF = dU - TdS - SdT$$

$$dF = -Sd\underline{T} - pd\underline{V} + \mu d\underline{N}$$

what about G? $\rightarrow$ you derive this

We can use these identities to derive many useful statements.

<u>Ex</u>: $dF = -SdT - pdV + \mu dN$

<u>hold V + N constant</u> $\rightarrow$ dV + dN are 0

$$dF = -SdT$$

$$S = -\left(\frac{dF}{dT}\right)_{V,N}$$ $\leftarrow$ what does this mean?

entropy is the rate of change of F w.r.t. T

$$p = -\left(\frac{\partial F}{\partial V}\right)_{T,N} \quad , \quad \mu = \left(\frac{\partial F}{\partial N}\right)_{T,V}$$

What about Gibbs?

$$S = -\left(\frac{\partial G}{\partial T}\right)_{P,N}, \quad V = \left(\frac{\partial G}{\partial P}\right)_{T,N}, \quad \mu = \left(\frac{\partial G}{\partial N}\right)_{T,P}$$

How does the Gibb Free Energy change w/ pressure

volume of 1 mol

Use reference data → 1 mol of C (graphite) ⇒ $5.3 \cdot 10^{-6}\ m^3$

$\Delta G_f = 0\,kJ$ ← most stable form

Gibbs will increase by $5.3 \cdot 10^{-6}\,J$ for each additional Pa of pressure above atmospheric pressure

Ex: 1 mol of $H_2O(l)$ at 25°C (298K)

Gibbs at 30°C? (303K).

$G_{H_2O} = -237.13\,kJ$ at 298K and 1 atm

$$S = \left(\frac{\partial G}{\partial T}\right)_{P,N} \qquad S_{H_2O} = 69.91\,J/K$$

units warning!

$\Delta G = S \cdot \Delta T$

$= 69.91 \cdot 5K = 349.55\,J$

new Gibbs

$+ \; -237.13\,kJ$

$\boxed{-236.78\,kJ}$

5.14 → <u>Monday</u>

Now do <u>5.12</u> then come back to <u>5.14</u>.

<u>Thermodynamic identity</u>

$$f(x,y)$$

$$dU = \boxed{T}\, dS \boxed{-P}\, dV + \mu\, dN$$

$$dU = \boxed{\frac{\partial U}{\partial S}}\, dS + \boxed{\frac{\partial U}{\partial V}}\, dV + \frac{\partial U}{\partial N}\, dN$$

$$\left(\frac{\partial f}{\partial x}\right)_y \qquad \left(\frac{\partial f}{\partial y}\right)_x \qquad \frac{\partial^2 f}{\partial x \partial y}$$

$$\frac{\partial U}{\partial S} = T \qquad \frac{\partial U}{\partial V} = -P$$

$$\frac{\partial}{\partial V}\left(\frac{\partial U}{\partial S}\right) = \left(\frac{dT}{dV}\right)_S \qquad \frac{\partial}{\partial S}\left(\frac{\partial U}{\partial V}\right) = -\left(\frac{\partial P}{\partial S}\right)_V$$

$$\left(\frac{dT}{dV}\right)_S = -\left(\frac{\partial P}{\partial S}\right)_V$$

$$dS(U,V,N) = \underbrace{\left(\frac{\partial S}{\partial U}\right)_{V,N}}_{\frac{1}{T}} dU + \underbrace{\left(\frac{\partial S}{\partial V}\right)_{U,N}}_{\frac{P}{T}} dV + \underbrace{\left(\frac{\partial S}{\partial N}\right)_{U,V}}_{-\frac{\mu}{T}} dN$$

$$T\,dS = dU + P\,dV - \mu\,dN$$

Now do for Gibbs

$$dG = dU + p\,dV - T\,dS \quad \Big| \quad dN = 0$$

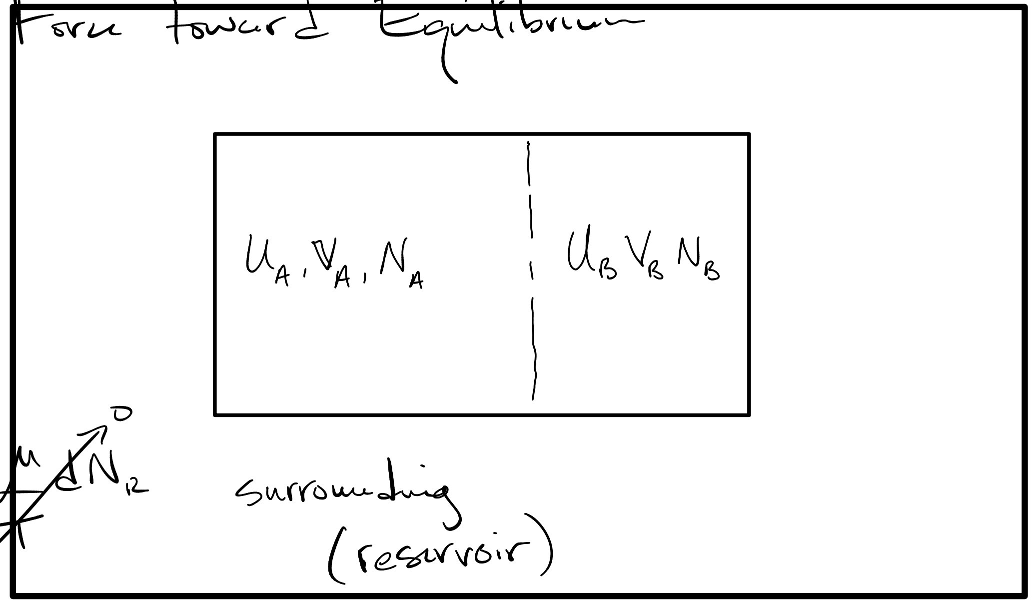
$$\frac{\partial}{\partial S}\left(\frac{\partial G}{\partial V}\right) = \frac{\partial}{\partial V}\left(\frac{\partial G}{\partial S}\right)$$

$$dG = \frac{\partial G}{\partial U}dU + \frac{\partial G}{\partial V}dV + \frac{\partial G}{\partial S}dS$$

$$\left(\frac{\partial p}{\partial S}\right)_V = -\left(\frac{\partial T}{\partial V}\right)_S$$

Now for <u>Helmholtz</u> + <u>Enthalpy</u> then back to <u>5.14</u>...

5.2 | Free Energy as a Force toward Equilibrium



$$dS_{total} = dS + \underbrace{dS_R}_{\text{rewrite in terms of system's variables}}$$

$$dS_R = \frac{1}{T}dU_R + \cancel{\frac{P}{T}dV_R}^{\,0} + \cancel{\frac{\mu}{T}dN_R}^{\,0}$$

$dV_R = 0$

$dN = 0$

$dU = -dU_R$ ← energy is conserved

$$dS_{total} = dS\frac{T}{T} - \frac{1}{T}dU$$

$$= \frac{1}{T}(T\,dS - dU)$$

$$= -\frac{1}{T}\underbrace{(dU - T\,dS)}_{\text{Helmholtz! } dF}$$

$$dS_{total} = -\frac{dF}{T}$$

increase this ↑   decrease that

Conditions: $dT = 0$, $dV = 0$, $dN = 0$

> So, the system will minimize its Helmholtz free energy

What about keeping Pressure, temperature and number constant and letting V change?

$$dS_{total} = dS + dS_{res}$$

$$\hookrightarrow dS_{res} = \frac{1}{T} dU + \frac{P}{T} dV - \cancel{\frac{\mu}{T} dN}^{\,0}$$

$$dS_{res} = -dS$$

$$dS_{total} = dS - \frac{1}{T} dU - \frac{P}{T} dV$$

$$= -\frac{1}{T}\underbrace{\left(dU - TdS + P dV \right)}_{\Delta \text{ Gibbs free energy} \rightarrow dG}$$

$$dS_{total} = -\frac{dG}{T}$$

system will spontaneously minimize Gibbs

Conditions:

$dT = 0$

$dP = 0$

$dN = 0$

At constant energy & volume, S tends to increase

At constant temperature & volume, F tends to decrease.

At constant temperature & pressure, G tends to decrease

Is this intuitive from their definitions

$$\underset{\downarrow}{G} = \underset{\downarrow}{U} + PV - T\underset{\uparrow}{S}$$

Derive chemical potential for an ideal gas, as a function of pressure

$$\mu = \frac{G}{N} \quad \leftarrow \text{find this} \qquad\qquad \mu = \left(\frac{\partial G}{\partial N}\right)_{T,P}$$

$$\frac{\partial \mu}{\partial p} = \frac{1}{N}\frac{\partial G}{\partial p} \qquad\qquad \left(\frac{\partial G}{\partial p}\right)_{TN} = \mathbb{V}$$

$$\frac{\partial \mu}{\partial p} = \frac{1}{N}\cdot \mathbb{V} \qquad \text{ideal gas} \rightarrow \quad \frac{\mathbb{V}}{N} = \frac{k_B T}{p}$$

$$\frac{\partial \mu}{\partial p} = \frac{k_B T}{p}$$

reference point

another reference

$$\int d\mu = \int \frac{k_B T}{p}\, dp \Rightarrow \mu(T,P) - \mu^\circ(T) = \underbrace{k_B T \ln p - k_B T \ln p^\circ}$$

$$\Delta \mu = k_B T \ln \frac{p}{p^\circ}$$

$$\mu(T,P) = k_B T \ln\left(\frac{p}{p^\circ}\right) + \mu^\circ(T)$$

We already have an expression for $\mu(T,P)$ from chapter 3.5 notes (problem 5.22)

$$\mu = -k_B T \ln\left[\frac{V}{N}\left(\frac{2\pi m k_B T}{h^2}\right)^{3/2}\right]$$

$$\mu = -k_B T \ln\left(\frac{V}{N}\right) - \frac{3}{2}k_B T \ln\left(\frac{2\pi m k_B T}{h^2}\right)$$

$$\frac{k_B T}{P}$$

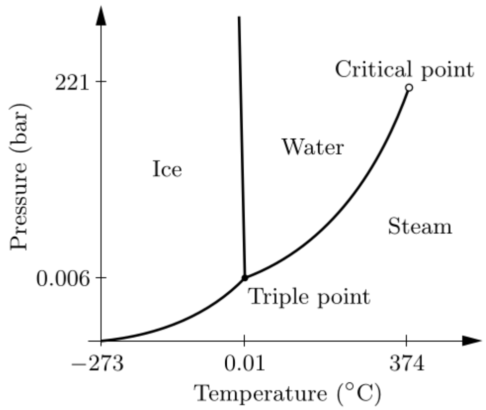
$$\mu = -k_B T \ln\left(\frac{k_B T}{P}\right) - \frac{3}{2}k_B T \ln\left(\frac{2\pi m k_B T}{h^2}\right)$$

$$\mu = k_B T \ln\left(\frac{P}{k_B T}\right) - \underbrace{\frac{3}{2}k_B T \ln\left(\frac{2\pi m k_B T}{h^2}\right)}_{\text{only a function of } T \,=\, \mu^\circ(T)}$$

(monoatomic ideal gas)

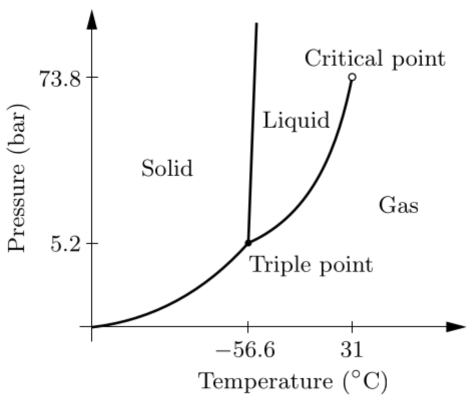
does this mean the P° is $k_B T$?

5.3



T (°C)	P_v (bar)	L (kJ/mol)
−40	0.00013	51.16
−20	0.00103	51.13
0	0.00611	51.07
0.01	0.00612	45.05
25	0.0317	43.99
50	0.1234	42.92
100	1.013	40.66
150	4.757	38.09
200	15.54	34.96
250	39.74	30.90
300	85.84	25.30
350	165.2	16.09
374	220.6	0.00

Figure 5.11. Phase diagram for H_2O (not to scale). The table gives the vapor pressure and molar latent heat for the solid-gas transformation (first three entries) and the liquid-gas transformation (remaining entries). Data from Keenan et al. (1978) and Lide (1994).



T (°C)	P_v (bar)
−120	0.0124
−100	0.135
−80	0.889
−78.6	1.000
−60	4.11
−56.6	5.18
−40	10.07
−20	19.72
0	34.85
20	57.2
31	73.8

Figure 5.12. Phase diagram for carbon dioxide (not to scale). The table gives the vapor pressure along the solid-gas and liquid-gas equilibrium curves. Data from Lide (1994) and Reynolds (1979).

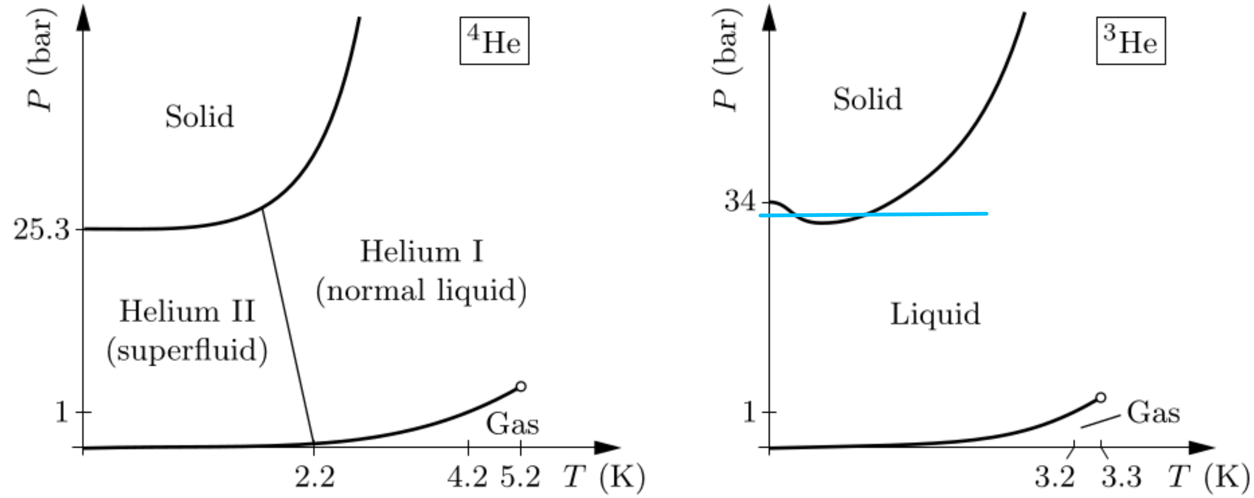


Figure 5.13. Phase diagrams of ^{4}He (left) and ^{3}He (right). Neither diagram is to scale, but qualitative relations between the diagrams are shown correctly. Not shown are the three different solid phases (crystal structures) of each isotope, or the superfluid phases of ^{3}He below 3 mK.

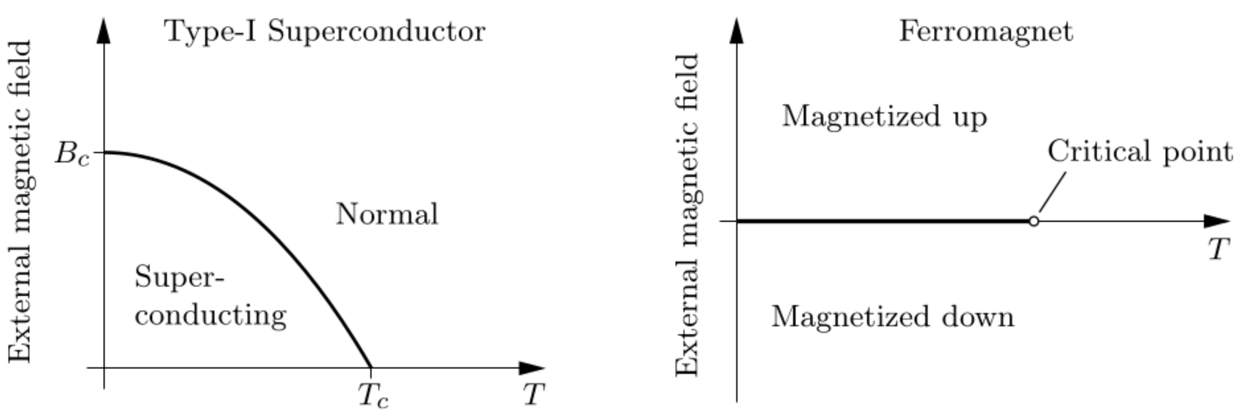
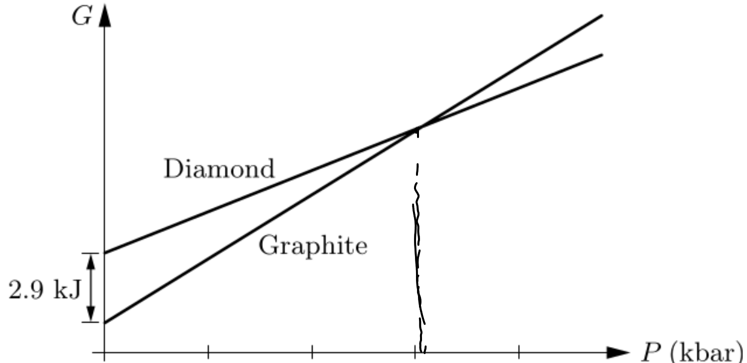


Figure 5.14. Left: Phase diagram for a typical type-I superconductor. For lead, $T_c = 7.2$ K and $B_c = 0.08$ T. Right: Phase diagram for a ferromagnet, assuming that the applied field and magnetization are always along a given axis.



15,000 bar = $15000 \cdot 10^5$ Pa

Figure 5.15. Molar Gibbs free energies of diamond and graphite as functions of pressure, at room temperature. These straight-line graphs are extrapolated from low pressures, neglecting the changes in volume as pressure increases.

$$\left(\frac{\partial G}{\partial P}\right)_{T,N} = V$$

$V_{graphite} = 5.31 \cdot 10^{-6}\, m^3$

$V_{diamond} = 3.42 \cdot 10^{-6}\, m^3$

STP

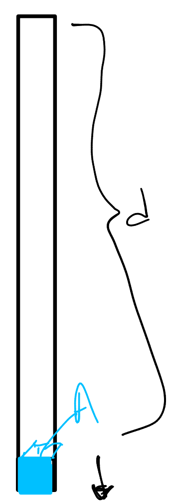
Temperature dependence follows entropy

$$-\left(\frac{\partial G}{\partial T}\right)_{P,N} = S$$

$S_{graphite} = 5.74\ J/K$

$S_{diamond} = 2.38\ J/K$

As the temperature increases, Gibbs goes down, but it goes down more quickly for graphite



$$\text{Pressure} = \frac{Mg}{A} = \frac{\rho_{rock} \cdot V \cdot g}{A} = \frac{\rho_{rock} \cdot \cancel{A} \cdot d \cdot g}{\cancel{A}}$$

$P = \rho_{rock} \cdot g \cdot d$

$d = \dfrac{P}{\rho \cdot g}$

$\rho_{H_2O} = 1\ g/mL$

$= 1000\ kg/m^3$

$\rho_{rock} = 3000\ kg/m^3$

So what is the shape of the phase boundary line?

On the phase boundary line

$$G_l = G_g$$

so, if we increase the p & T a little bit but we stay on the phase boundary line, then,

$$dG_l = dG_g$$

$$dG = -SdT + VdP + \mu \cancel{dN}^{0}$$

$$-S_l dT + V_l dP = -S_g dT + V_g dP$$

$$\frac{dP}{dT} = \frac{S_g - S_l}{V_g - V_l} = \frac{\Delta S}{\Delta V}$$

$$\Delta S = \frac{Q}{T} = \frac{L}{T}$$

Latent Heat (Enthalpy)

$$\boxed{\frac{dP}{dT} = \frac{L}{T\Delta V}}$$

Clausius - Clapeyron relation (applies to any phase boundary)

$$d = \frac{1.5 \cdot 10^9 \text{ Pa}}{3 \cdot 10^3 \cdot 10 \frac{N}{kg}}$$

$$d = 0.5 \cdot 10^5 \text{ m}$$

$$d = 5 \cdot 10^4 \text{ m}$$

$$d = 50 \cdot 10^3 \text{ m}$$

$$\underline{d = 50 \text{ km}}$$

So lets apply this to diamond / graphite

$\Delta V = 1.9 \cdot 10^{-6}\ m^3$

$\Delta S = 3.4\ J/K$

$$\frac{dP}{dT} = \frac{3.4\ J/K}{1.9 \cdot 10^{-6}\ m^3} = 1.8 \cdot 10^{6}\ Pa/K = 18\ bar/K$$

We found 15 kbar at 298 K to be the phase change b/t graphite + diamond, but what about 100°C higher temperatures?

$100°C = \Delta T = 100\ K$

$$\frac{\Delta P}{\Delta T} = 1.8 \cdot 10^{6}\ Pa/K$$

$$\Delta P = 1.8 \cdot 10^{6}\ \frac{Pa}{K} \cdot \Delta T = 1.8 \cdot 10^{6}\ \frac{Pa}{K} \cdot 100\ K = 1.8 \cdot 10^{8}\ Pa = 1.8 \cdot 10^{3} = 1800\ bar = \underline{1.8\ kbar}$$

So more pressure is necessary to make diamond become the more stable form.

more depth is required

100 – 200 km

Problems: 5.32, 35, 36, 37