Have you been reading my "notes"

Two topics

Relating ΔH and ΔS for a phase transition

What is special about water

Free Energy Change

What is the sign of the change in free energy for me dropping an eraser?

A. $\boxed{\text{negative (free energy decreases)}}$

B. positive (free energy increases)

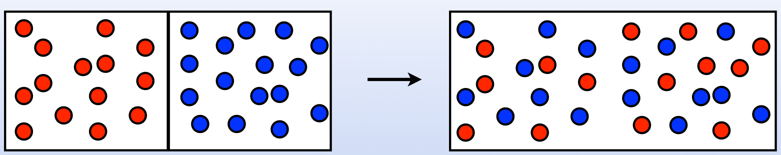
C. zero (free energy is constant)

D. it depends on the temperature

Things that happen decrease lower free energy (equilibrium is zero change)

Mixtures

What is different than pure substances?



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it happens. therefore free energy decreases

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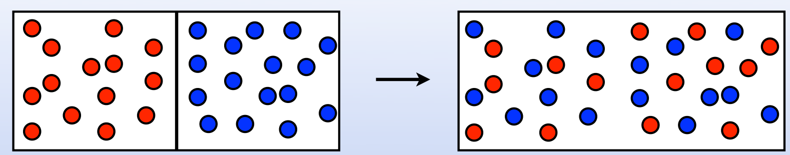

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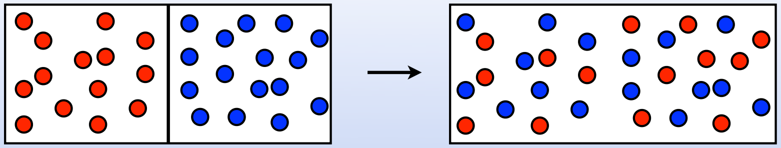


Why does the free energy decrease? ($\Delta G = \Delta H - T\Delta S$)

A. ΔH is positive, ΔS is zero

B. $\boxed{\Delta H \text{ is zero, } \Delta S \text{ is positive}}$

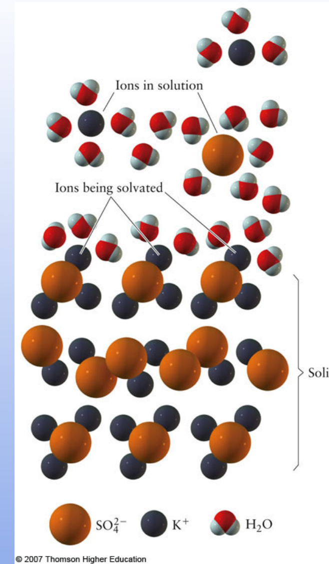
C. ΔH is negative, ΔS is positive

D. ΔH is negative, ΔS is zero

Gases ~ no IMF. Therefore $\Delta H = 0$

The volume of each gas increases therefore $\Delta S > 0$

Take home lesson generally entropy increases with mixing



When we think of mixtures we typically think about solutions

Solvent: the majority of the molecules

IMF only slightly changed
(most solvent molecules interacting with solvent molecules)

Solute: the minority substance
the “stuff that is dissolved”

could be a solid, liquid, or a gas

IMF total different
in solution, solute molecules only interacting with solvent molecules

Entropy of Solution $\Delta S_{solution}$
usually easy to predict

Solutions typically have a higher entropy than the unmixed compounds

Therefore

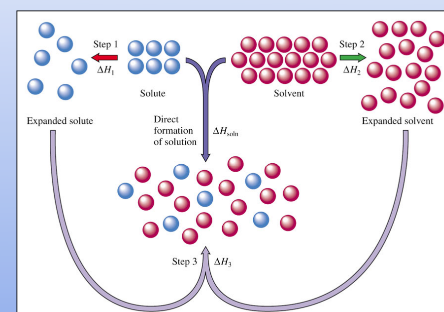
$$\Delta S_{solution} > 0$$

For most cases

Since entropy almost always favors mixing, the differences between different substances are the result of enthalpy (intermolecular forces)

What is enthalpy change for making a solution?

Lose solute-solute interactions (IMF)
Lose solvent-solvent interactions (IMF) (this is small)
Gain solute-solvent interactions



How we generally think of this

What is $\Delta H_{solution}$?
this is the enthalpy change for making the solution?

Two steps

First break up the solute $\Delta H_{Lattice\ Energy}$
(loss of solute-solute interactions) "costs energy" positive

Next put solute into solvet $\Delta H_{solvation}$
(add of solute-solvent interactions) "releases energy" negative

$$\Delta H_{solution} = \Delta H_{Lattice\ Energy} + \Delta H_{solvation}$$

Enthalpy of Solution $\Delta H_{solution}$ hard to predict

$\Delta H_{solution} = 0$
Ideal solution
Solute-solvent interactions are identical to solute-solute (and solvent-solvent)

$\Delta H_{solution} > 0$
Typical
Solute-solvent interactions are weaker than solute-solute (and solvent-solvent)

$\Delta H_{solution} < 0$
Unusual but possible
Solute-solvent interactions are stronger than solute-solute (and solvent-solvent)

Which do you think has a stronger interactions with a sodium ion?

A. a chloride ion

B. water

C. they are the same

Ion-Ion interactions will be stronger than
ion-dipole interactions
(but ion dipole interactions are still strong)

What do you predict for the sign of the enthalpy of solution of NaCl in water?

A. positive

B. negative

C. zero

Because the solute/solvent interactions (ion-dipole)
are weaker than the solute/solute (ion-ion)
it will "cost" energy to get the salt into the water

For dissolving salt in water at room temperature

$\Delta H_{solution} > 0$

$\Delta S_{solution} > 0$

which is larger?

A. $\Delta H > T\Delta S$

B. $\Delta H = T\Delta S$

C. $\Delta H < T\Delta S$

This actually happens. $\Delta G < 0$.

This means that $\Delta H < T\Delta S$

For dissolving water in olive oil at room temperature

$\Delta H_{solution} > 0$

$\Delta S_{solution} > 0$

which is larger?

A. $\Delta H > T\Delta S$

B. $\Delta H = T\Delta S$

C. $\Delta H < T\Delta S$

This does not happen. $\Delta G > 0$.

This means that $\Delta H > T\Delta S$

When things will not dissovle

$\Delta H_{solution}$ is too large (bigger than $T\Delta S$)

IMF between the solvent/solute are much less favorable than solute/solute (solvent/solvent)

When will this happen?

Very different IMF

Oil (dispersion/nonpolar) and water(H-bonding, polar)

Very strong ion/ion (MgO)

Other problems for small high charge density ions

TABLE 17.2 Values of ΔS°_{soln} for Several Salts Dissolving in Water

Process	ΔS° ($J\ K^{-1}\ mol^{-1}$)
$KCl(s) \rightarrow K^+(aq) + Cl^-(aq)$	75
$LiF(s) \rightarrow Li^+(aq) + F^-(aq)$	−36
$CaS(s) \rightarrow Ca^{2+}(aq) + S^{2-}(aq)$	−138

Gibb's Free Energy of Solvation ΔG_{soln}

If $\Delta G_{soln} < 0$ solution strongly favored

If $\Delta G_{soln} > 0$ undissolved state is strongly favored

$$\Delta G_{soln} = \Delta H_{soln} - T\,\Delta S_{soln}$$

Typically $\Delta S_{soln} > 0$, $\Delta H_{soln} > 0$

need $|T\Delta S| > |\Delta H|$

What makes an ideal solution?

Same IMF for solute-solvent and solute-solute and solvent-solvent

"like dissolves like"

Polar compounds dissolve polar compounds (ionic)

Nonpolar compound dissolve nonpolar compounds

This minimizes $\Delta H_{solution}$

Miscibility Demo

Definitions:

Miscible: capable of being mixed
Immiscible: incapable of being mixed

Which is most likely to dissolve best in water?

A. methanol CH_3OH ←

B. butanol C_4H_9OH

C. octanol $C_8H_{17}OH$

D. didodecanol $C_{12}H_{25}OH$

Which is most likely to dissolve best in hexane (C_6H_{14})?

A. methanol CH_3OH

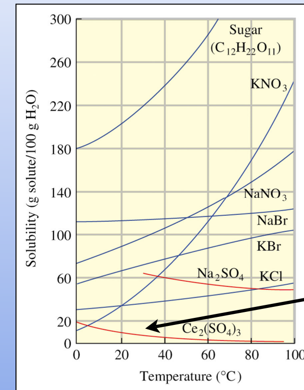
B. butanol C_4H_9OH

C. octanol $C_8H_{17}OH$

D. didodecanol $C_{12}H_{25}OH$ ←

Temperature Dependence

Generally at T goes up solubility increases



$\Delta H_{soln} < 0$
(reaction with water)

Gas Dissolved in a Liquid

Henry's Law

TABLE 17.3 The Values of Henry's Law Constants for Several Gases Dissolved in Water at 298 K

Gas	k_H (atm)
CH_4	4.13×10^2
CO_2	1.64×10^3
O_2	4.34×10^4
CO	5.71×10^4
H_2	7.03×10^4
N_2	8.57×10^4

$P_{solute} = K_{solvent} X_{solute}$

↑

mole fraction

n_{gas}/n_{total}

In General

Henry's Law constants increase with increasing Temperature

Less gas is dissolved at higher temperatures

$\Delta H < 0$

going from no attractions to being in a liquid

= bad news for fish in hot water (less dissolved O_2)