Big Idea #5

Thermodynamics & Changes in Matter

heat and temperature

temperature the amount of kinetic energy in a substance
heat energy flow between 2 substances @ different temps
first law of thermodynamics energy can be neither created nor destroyed

reaction ⟷ surroundings
energy flow
→ exothermic
← endothermic

state functions depend only on change between initial and final states (independent of pathway)
enthalpy (ΔH), entropy (ΔS), and free-energy (ΔG)

Standard state conditions
gases at 1 atm pressure
liquids and solids are pure
solutions are 1M concentration
energy of formation of an element in its normal state is 0
room temperature (25°C, 298 K)

enthalpy of formation (ΔH°)

ΔH°_f for a pure element is 0 (also true for diatomics)
$-\Delta H^\circ_f$ energy is released when compound is formed (product is more stable) exothermic
$+\Delta H^\circ_f$ energy is absorbed to create product (product is less stable) endothermic
$\Delta H^\circ = \sum \Delta H^\circ_{products} - \sum \Delta H^\circ_{reactants}$ - multiply by stoichiometry coefficients
enthalpy of combustion describes amount of energy released for one mole of combusted hydrocarbon (ΔH° should be negative)

bond energy $\Delta H^\circ = \sum$ bond energies of bonds broken $- \sum$ bond energies of bonds formed

Hess' Law
↳ flip equation, flip sign of enthalpy value
↳ multiply or divide equation by integer, multiply or divide enthalpy value by same integer
↳ enthalpy value of summed equation = sum of individual enthalpy values

enthalpy of solutions

(1) solute bonds are broken (equal to lattice energy)
ΔH is positive
(2) solvent molecules (water) are separated (breaking H-bonds & intermolecular forces)
ΔH is positive
(3) new attractions are created (ions attracted to dipoles of water molecules)
ΔH is negative

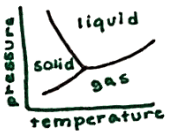
hydration energy
always negative
increases as ion charge increases

enthalpy of solution
lattice energy > hyd. energy
positive ΔHsoln
hyd. eng > lattice energy
negative ΔHsoln

phase changes

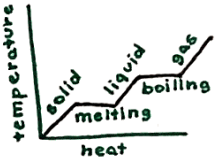
a liquid boils when its heat of vaporization (vapor pressure) is equal to the atmospheric pressure
heat of fusion energy necessary to melt a solid
heat given off by a substance when it freezes
heat of vaporization energy necessary to turn a liquid into a gas
heat given off by a substance when it condenses
the temperature of a substance remains constant during a phase change (both temp & phase never change @ same time)

phase diagram of water



solid-liquid equilibrium line is typically sloped upward (pressure increase causes change from liquid to solid)
for water, pressure increase causes change from solid to liquid.

heating curves



during phase change
heat = (heat of fusion/vap) kJ/mol · (mol of substance)
during temperature change
$q = mc\Delta t$

entropy a measure of the amount of disorder in a system
$\Delta S^\circ = \sum \Delta S^\circ_{products} - \sum \Delta S^\circ_{reactants}$
ΔS gases > ΔS liquid > ΔS solid
particles in solution have greater entropy than particles in a solid

voltage and favorability

$\Delta G = -nFE^\circ$
ΔG = Gibb's Free Energy change (J/mol)
n = moles of electrons exchanged in reaction
F = Faraday's constant = 96,500 C/mol
E° = standard reduction potential (V or J/C)

calorimetry

specific heat the amount of energy required to raise the temp of 1g of a substance by 1°C
↳ larger specific heat → difficult to change temp

$q = mc\Delta T$

q = heat added (J or cal)
m = mass of substance (g or kg) or total mass of solution
c = specific heat
ΔT = temperature change (K or °C)

Gibb's Free Energy

$\Delta G^\circ = \sum \Delta G^\circ_{products} - \sum \Delta G^\circ_{reactants}$
$+\Delta G^\circ$ reaction is thermodynamically unfavored
$-\Delta G^\circ$ reaction is thermodynamically favored
$\Delta G^\circ = 0$ reaction is at equilibrium

favorability of reactions

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

ΔH	ΔS	ΔG	
−	+	−	always favored
+	−	+	never favored
+	+	−	favored at high temps
−	−	−	favored at low temps