Chemistry 6A F2007

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Freezing Point Depression:

Similarly:

The *Freezing point* of a solution is always lower than the freezing point of the pure solvent of the solution.

$$t_f(\text{solvent}) < t_f(\text{solution})$$

The difference in freezing point between pure solvent and solution depends on the concentration of solute particles, and is calculated using the following equation:

$$\Delta t_f = nK_fM$$

$\Delta t_f = t_f(\text{solution}) - t_f(\text{solvent})$

K_f = fp constant *(depends on the substance)*

M = the molarity of the solution

n = total number of particles in solution

n = 1 for molecular compounds

n = 2 for NaCl (Na^+ and Cl^-)

Calculate the freezing point of a solution made up by adding and completely dissolving 4.52g sodium phosphate to 100.0 mL of water.

$$\Delta t_f = nK_fM$$

$$K_f = -1.86\frac{^\circ C}{M}$$

step 1: Calculate the molarity of the solution

step 2: Recognize that n = 4

$$Na_3PO_4(aq) \rightarrow 3Na^+(aq) + PO_4^{3-}(aq)$$

n = 3 + 1= 4

step 3: enter the values into the equaton

Calculate the freezing point of a solution made up by adding and completely dissolving 4.52g sodium phosphate to 100.0 mL of water.

$\Delta t_f = nK_fM$ n = 4 $K_f = -1.86\frac{^\circ C}{M}$

$$\Delta t_f = 4\times\left(-1.86\frac{^\circ C\cdot\cancel{L}}{\cancel{mol}}\right)\times 4.52\cancel{g\ Na_3PO_4}\times\frac{1\cancel{mol}Na_3PO_4}{163.94\cancel{g}}\times\frac{1}{100.0\cancel{mL}}\times\frac{10^3\cancel{mL}}{\cancel{1L}}$$

$$\Delta t_f = -2.05\ ^\circ C$$

$$t_f = 0.00\ ^\circ C + \Delta t_f = 0.00\ ^\circ C - 2.05\ ^\circ C = \boxed{-2.05\ ^\circ C}$$

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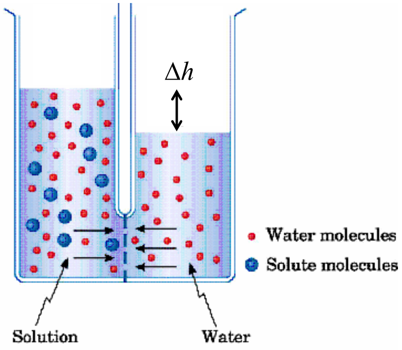
By spreading salt on the roadway, the freezing point of water is lowered, thus preventing ice formation or causing already solid ice to melt!

OSMOTIC PRESSURE OF SOLUTIONS

When solutions having different concentrations of solute are separated by a semi-permeable membrane, the solvent tends to flow across the membrane from a less concentrated solution towards a more concentrated solution in a process called ***osmosis***.



Solution opposite the water: The height of the more concentrated solution to rises!

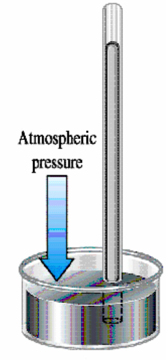
water on each side, same height:

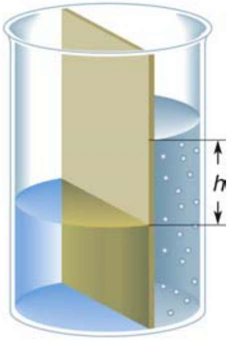
The solvent molecules (water) want to dilute the solution by moving across the membrane.

Recall that pressure is related to the height of a column of liquid:

Analogously, the height increase due to osmosis is like a pressure:

$$\pi = \text{nMRT}$$





π = osmotic pressure (torr)

M = mols/L

T = Kelvins

$R = 62.4 \dfrac{\text{L} \cdot \text{torr}}{\text{mol} \cdot \text{K}}$

A 1.00 liter aqueous solution containing 1.00 g of bovine insulin has an osmotic pressure of 3.1 torr at 25.0^{o}C. Calculate the molar mass of bovine insulin.

Need to find molar mass: $\dfrac{\text{g}}{\text{mol}}$

grams are given in the problem… you need moles!

Look at the definition of osmotic pressure:

$$\pi = \text{nMRT}$$

Where do you see moles? correct, Molarity!

$$\text{M} = \frac{\text{mols}}{\text{L}}$$

A 1.00 liter aqueous solution containing 1.00 g of bovine insulin has an osmotic pressure of 3.1 torr at 25.0°C.
Calculate the molar mass of bovine insulin.

So you need to solve for molarity.
Then, molarity and volume will give moles!

$$\pi = \text{nMRT}$$

(n = 1 as the protein does not ionize)

rearranging:

$$\text{M} = \frac{\pi}{\text{nRT}}$$

$$\text{M} \times \text{Vol} = \text{moles}$$

grams and moles yield molar mass

A 1.00 liter aqueous solution containing 1.00 g of bovine insulin has an osmotic pressure of 3.1 torr at 25.0°C.
Calculate the molar mass of bovine insulin.

$$\text{M} = \frac{\pi}{\text{nRT}}$$

$$\text{M} = \frac{3.1 \text{ torr}}{1 \times 62.4 \frac{\text{L} \cdot \text{torr}}{\text{mol} \cdot \text{K}} \times (25.0 + 273.15)\text{K}} = 1.7 \times 10^{-4} \text{ M}$$

$$\frac{1.7 \times 10^{-4} \text{ mols}}{1\text{L}} \times 1.00 \text{ L} = 1.7 \times 10^{-4} \text{ mols}$$

$$\frac{1.00\text{g}}{1.7 \times 10^{-4} \text{ mol}} = 6.0 \times 10^{3} \text{ g/mol}$$

Please read section 7.8 and 7.9 (colloids and dialysis)
This material will not be covered in class, however there may be HW questions.

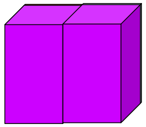
Chapter 8: LEARNING OBJECTIVES

After completing this chapter, you should be able to:

1. Use energy and entropy considerations to predict the spontaneity of a process or reaction.
2. Calculate reaction rates from experimental results.
3. Use the basic assumptions of reaction mechanisms to explain reaction characteristics.
4. Represent and interpret the energy relationships for reactions by using energy diagrams.
5. Explain how factors such as concentration, temperature, and catalysts influence reaction rates.
6. Perform calculations based on equilibrium expressions for reactions.
7. Explain the concept of equilibrium and use Le Châtelier's principle to predict the influences of concentrations of reactants, concentrations of products, catalysts, and reaction temperature on the position of equilibrium.

Why are certain processes or reactions spontaneous?
Why are others not?

Heat spontaneously flows from a hot object to a cold object



after some time, both blocks are at the same temp

Potassium will spontaneously react with water! (click here)

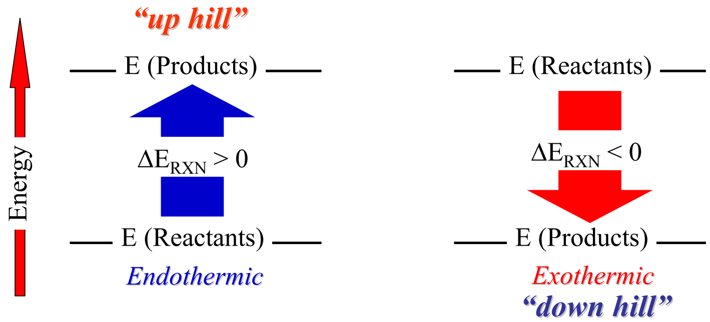
Aluminum however does not react with water.

Any process that occurs spontaneously must be related to energy:

Recall out discussion of energy…

Reactions that give off heat are… Exothermic

Reactions that absorb heat are… Endothermic



"up hill"

E (Products)

$\Delta E_{RXN} > 0$

E (Reactants)

Endothermic

E (Reactants)

$\Delta E_{RXN} < 0$

E (Products)

Exothermic

"down hill"

From this, you might think that all exothermic reactions (down hill) are spontaneous…

This is not true!

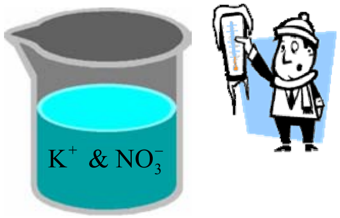
The reaction of hydrogen and oxygen is quite exothermic:

$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g) \quad \Delta H = -241.7\,kJ/mol$$

However, when hydrogen is released into the atmosphere, it does not spontaneously react.

We know that $\mathbf{KNO_3}$(s) dissolves in water.

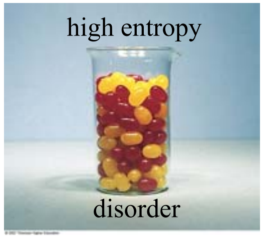
When this happens, the solution gets cold. *(endothermic)*



This means that there must be some other force that plays a part in determining whether or not a process is spontaneous.

Energy does play a part, but in addition, an increase "**randomness**" or "**Entropy**" can help or hinder a process to be spontaneous.

Entropy measures the *disorder* of a system.
The more disordered the system, the higher the entropy.



$$\underbrace{H_2(g) + \frac{1}{2}O_2(g)}_{\substack{1.5\ \text{mol} \\ \text{(more disorder)}}} \rightarrow \underset{\substack{1\ \text{mol} \\ \text{(less disorder)}}}{H_2O(g)} \qquad \Delta H = -241.7\ \text{kJ/mol}$$

Water does not spontaneously form from the reaction of H_2 and O_2 because there are fewer moles of product than reactants.

Potassium nitrate spontaneously dissolves even though the process is endothermic (up hill).

$$\underset{\substack{1\ \text{mol} \\ \text{(less disorder)}}}{KNO_3(s)} \xrightarrow{H_2O(l)} \underbrace{K^+(aq) + NO_3^-(aq)}_{\substack{2\ \text{mol} \\ \text{(more disorder)}}}$$

FACTORS THAT INFLUENCE SPONTANEITY

- A process will always be spontaneous if it is exothermic and the entropy of the system increases.
- A process that is endothermic will be spontaneous only if an increase in entropy of the system overwhelms the "up hill" energy climb.
- A process that causes an entropy decrease in the system will be spontaneous only if the exothermisity overwhelms the loss of entropy.

With some entropy increases, the increase prevents the system form returning to its original state…