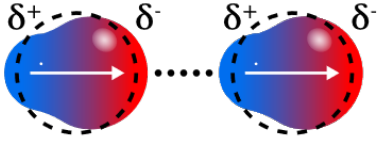
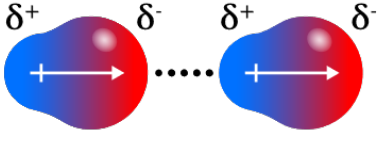
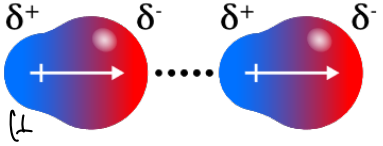
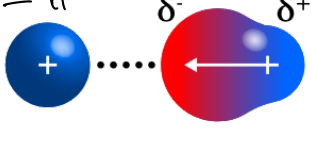
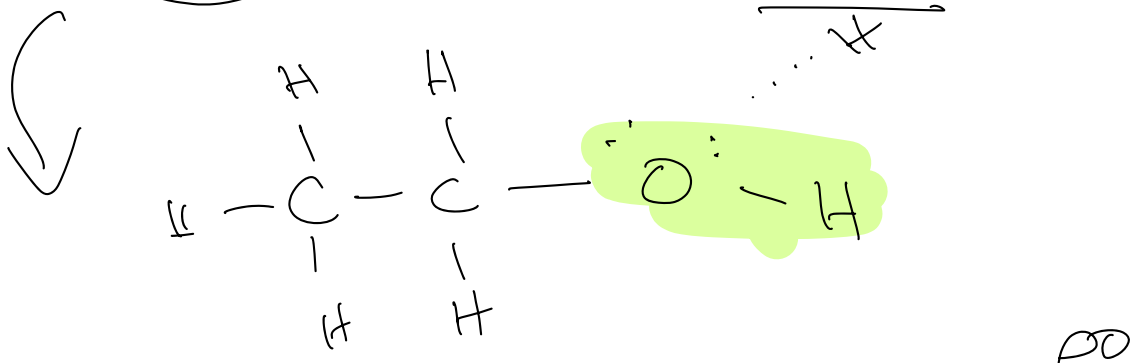
Ch 103 IMF review

Types of Intermolecular Forces	Present In	Molecular Perspective	Strength
Dispersion*	All molecules and atoms		0.05-40 kJ/mol
Dipole-Dipole	Polar molecules		5-25 kJ/mol
Hydrogen Bonding	Molecules containing H bonded to F, O, or N		10-40 kJ/mol
Ion Dipole	Mixtures of ionic compounds and polar compounds		40-600 kJ/mol

*Dispersion forces can become very strong (as strong and even stronger than the others) for molecules of high molar mass.

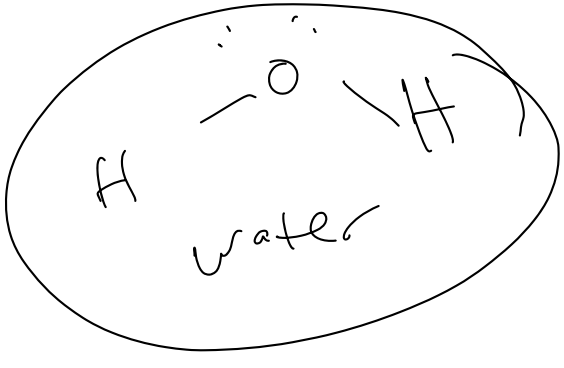
"mix"

Is ethanol miscible in water or hexane (C_6H_{14})?



polar

solvent



C_6H_{14} non-polar

b/c ethanol can't for H-bonds with water molecules.

Solutions

solute (example: NaCl)

solvent (water)

molarity (M) = $\frac{\text{mol solute}}{\text{L solution}}$

other ways to express concentration

- % m/v → mass → volume =
- % m/m =

mole fraction =

molality =

(M) Molarity vs. (m) molality

$$\frac{\text{mol solute}}{\text{L of solution}}$$

solute

solvent

3.0 M NaCl

$$\frac{3.0 \text{ mol of NaCl}}{1 \text{ L of solution}}$$

vs.

$$\frac{\text{mol solute}}{\text{Kg of solvent } (H_2O)}$$

3.0 m

3.0 mol NaCl

$$\frac{}{1 \text{ kg of } H_2O}$$

Molarity
molality

→ gen chem
analytical chemistry

ppm

environmental chem / public health

$$\frac{\text{mg of substance}}{1\text{ kg of } H_2O}$$

→ example Pb^{2+}

An IV solution has a glucose concentration of 0.556 M, what is the molal (m) concentration?

The density of the solution is $1.04 \frac{g}{mL}$

$M = \frac{1 \text{ mol solute}}{1 \text{ L solution}}$

$m = \frac{1 \text{ mol solute}}{1 \text{ kg solvent}}$

$M \rightarrow m$

$$\frac{0.556 \text{ mol glucose}}{1 \text{ L solution}}$$

① $0.556 \text{ mol glucose} \times \frac{180.1 \text{ g}}{1 \text{ mol}} = 100.1 \text{ g glucose}$

② $1{,}000 \text{ mL solution} \times \frac{1.04 \text{ g}}{1 \text{ mL}} = 1{,}040 \text{ g of solution}$

③ mass of solvent

$1{,}040 \text{ g solution} - 100.1 \text{ g glucose} =$

939.9 g of solvent

37.1 (m/m)
1.19 g/mL

mass percent
(m/m)

$$\frac{\text{mass of solute}}{100\ \text{g of solution}}$$

$$\frac{37.0\ \text{g of HCl}}{100\ \text{g of solution}}$$

? what is the molarity

$$M = \frac{\text{mol of solute}}{\text{L of solution}}$$

① moles of HCl

$$37.0\ \text{g HCl} \times \frac{1\ \text{mol}}{36.0\ \text{g}} =$$

1.03 mol of HCl

② $100\ \text{g solution} \times \frac{1\ \text{mL}}{1.19\ \text{g}} = 84.0$ mL of solution

$$84.0\ \text{mL} \times \frac{1\text{L}}{1000\ \text{mL}} =$$

$$0.08402\text{L}$$

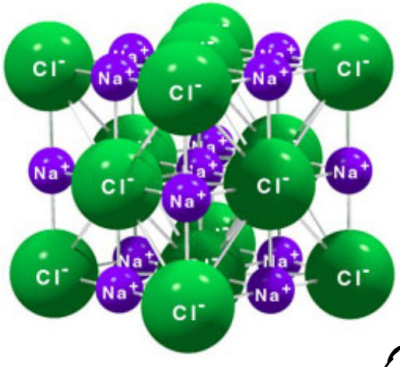
(3)

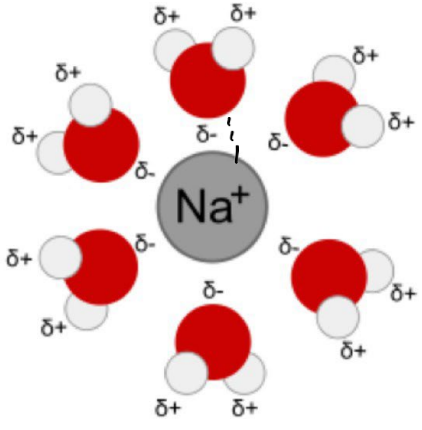
$$1.03\ \text{mol HCl} / 0.08402\text{L}$$

12.3 M

(A) $m = 0.556\ mol / 0.9399\ kg\ H_2O =$ 0.592 m

Thermodynamics of Dissolving





hydration of ions

exothermic

step 2

step 1

ion-ion

endothermic

ΔH_{diss} = step 1 + step 2

Example: hot pack vs. cold pack what is sign of ΔH?

Henry's Law

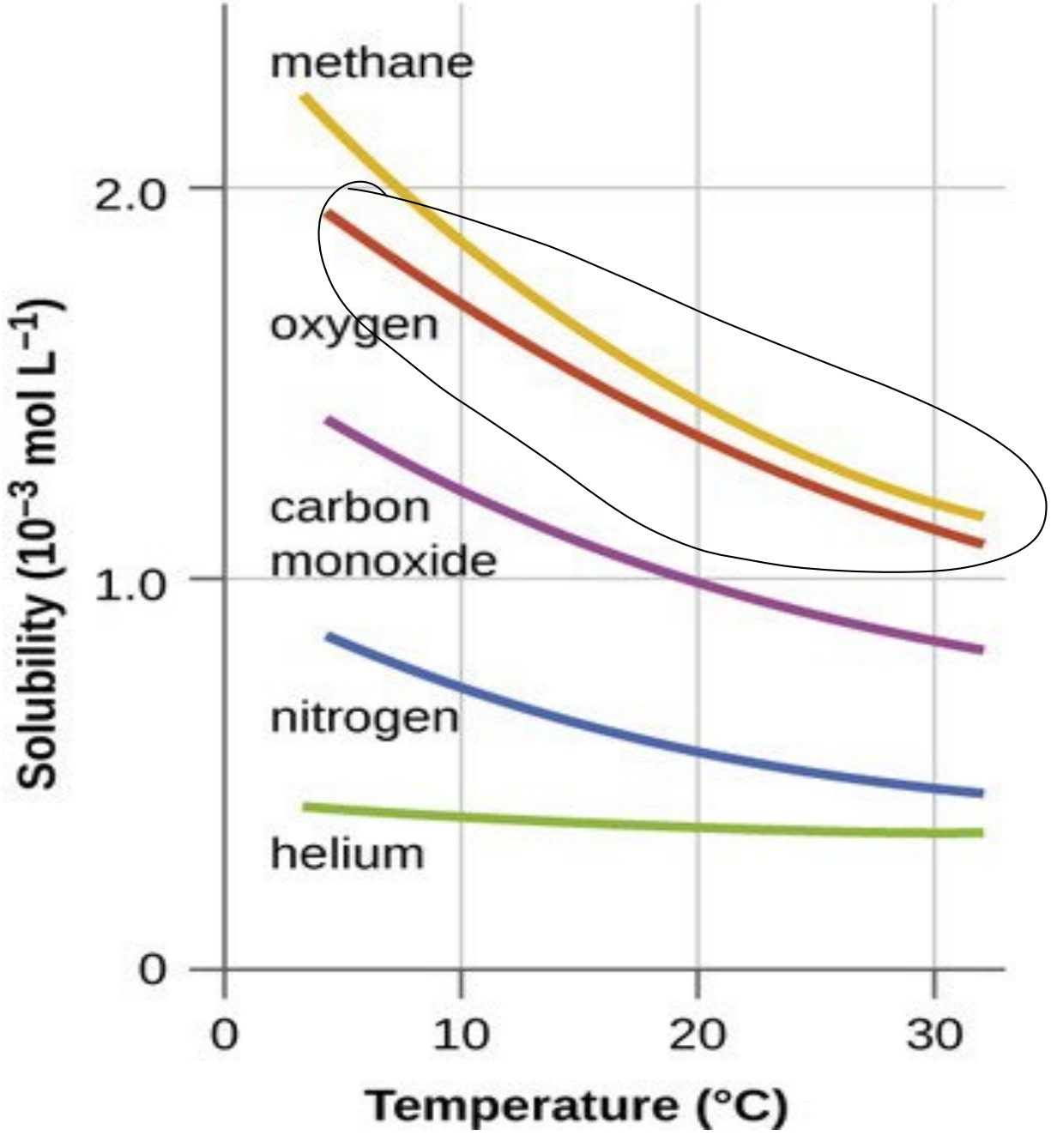
$= k_H P_g$

Henry's law constant

units $\frac{M}{atm}$

solubility of a gas in M or g/L

pressure of gas


methane
oxygen
carbon monoxide
nitrogen
helium
2.0
1.0
0
0
10
20
30
Solubility (10^{-3} mol L^{-1})
Temperature (°C)

Henry's Law example problem

Soda is bottled under a pressure of 2432 mm Hg. How many grams of carbon dioxide can dissolve in two liters of soda with this pressure? (Henry's law constant , = 3.1 x 10-2 M/atm) (760 mm Hg = 1 atm) ?

$$S_{gas} = k_H P_{gas}$$

$$2{,}432\,mm\,Hg \times \frac{1\,atm}{760\,mm\,Hg} = 3.2\,atm$$

$$S_{gas} = \left(3.1 \times 10^{-2}\,\frac{M}{atm}\right)(3.2\,atm)$$

$S_{gas} = 0.092\ M$

$$\frac{0.092\ \text{mol}\ CO_2}{1\ L\ \text{solution}} \times \frac{2\ L}{1} \times \frac{44.0\ g}{1\ \text{mol}\ CO_2}$$

$$= \boxed{8.73\ g}$$