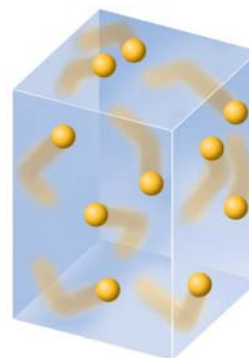


## Gases

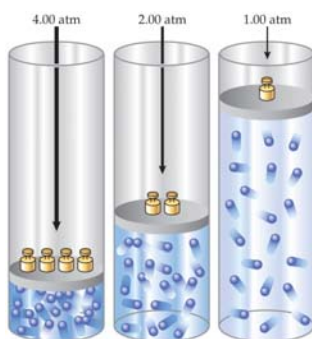
- Things you've already observed
  - What happens when you take a filled balloon outside in the winter?
  - What happens when you take a filled balloon and squeeze it?
- Kinetic-Molecular Theory of Gases helps to rationalize these observations
  - Particles in constant motion
  - Lots of empty space between particles
  - Particles aren't attracted to one another
  - When particles collide, they undergo *elastic* collisions
  - Temperature is a measure of kinetic energy of the gas ( $KE = mv^2/2$ )
- How does this fit with the observations above?



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## Quantifying Observations with Gas Laws

- Boyle's Law - volume inversely related to pressure ( $V_1P_1 = V_2P_2$ )
- Charles' Law - volume directly related to temperature ( $V_1/T_1 = V_2/T_2$ )
- Ideal gas law –  $PV = nRT$  (T in Kelvin)
  - $R = 0.0821 \text{ (L atm)/(mol K)}$
  - 1 mole of gas occupies 22.4 L at STP (STP= 0°C, 1 atm)
    - "Standard Molar Volume"



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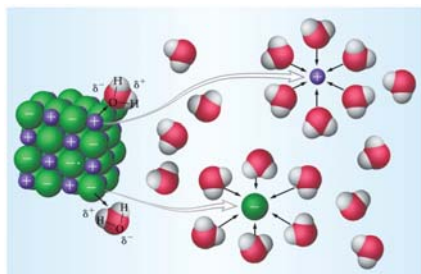
(a)

(b)

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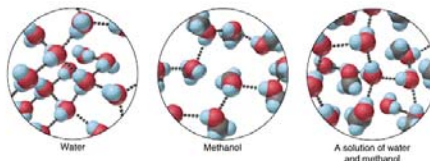
## Solutions and Concentrations

- Solutions consist of a **solute** dissolved in a **solvent**.
  - Why/how do solutions form? Intermolecular Forces!!
    - What forces must be disrupted and what new forces formed?
    - Must be energetically favorable



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Fig 6.9



<http://www.chem.ufl.edu/~itl/2045/change/C12F3.GIF>

- Solutions can be **electrolytic** or **nonelectrolytic**
  - Strong vs. weak

## Solution Concentrations

- Describes the amount of solute relative to the entire solution.
- Concentration units
  - % w/w, %w/v
  - ppm and ppb as mg/L and  $\mu\text{g/L}$
  - molarity
- Molarity is very useful in studying chemical reactions involving solutions:  $(\text{moles/L}) \times L = \text{moles!}$