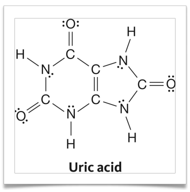
Chem 150, Spring 2015

Unit 2 - Molecular Interactions

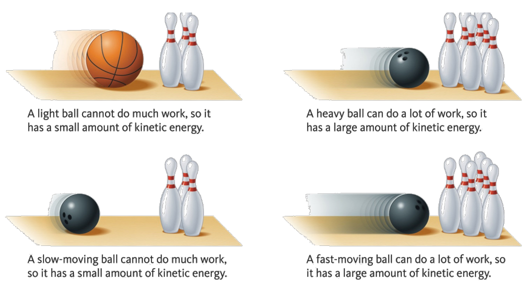
Chapter 4 - Introduction

- The painful medical condition know as **gout**, is caused by the accumulation of **uric acid** crystals in the joints.
  - This occurs when uric acid levels in the blood are high, because uric acid has a low solubility in water.



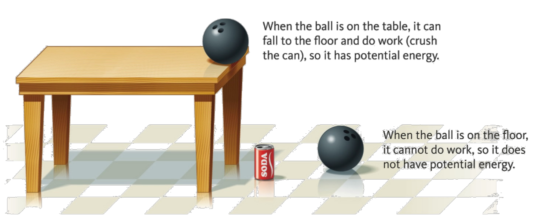
4.1 Heat and Energy

- Energy is the ability to do *work*.
- Energy of *motion* is called kinetic energy.



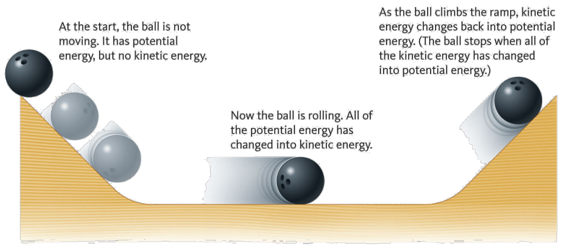
Potential Energy

- The energy an object has that is due to its *position* is called potential energy.



Law of Conservation of Energy

- The law of conservation of energy says that energy cannot be *created* or *destroyed*.
- We can change it from one form to another.



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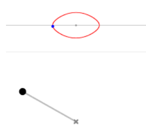

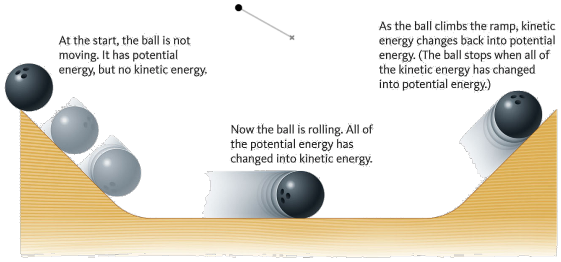

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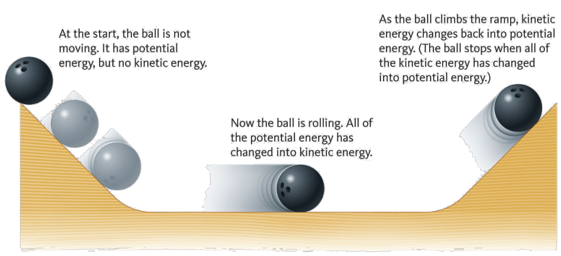
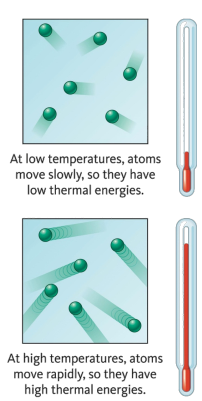


Thermal Energy

- Atoms are always in motion and the random kinetic energy of atoms is called thermal energy.
- Thermal energy depends on the *amount* of the substance, while temperature does not.



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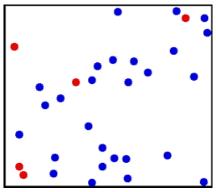
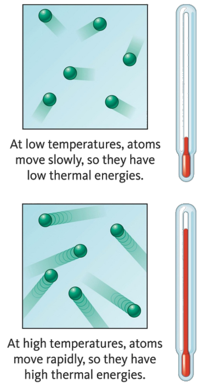




Heat and Calories

- Thermal energy added or removed from a substance is called heat.
  - When heat is added or removed from a substance it normally produces a temperature change.

    heat = *mass* x *temperature* x *specific heat*

  - The specific heat the conversion factor, which is a property of the substance.
  - A common unit for heat is the *calorie*.
    - A calorie is the amount of heat needed to raise the temperature of 1g of water by 1 °C

Heat and Calories

- The energy content of food is measured in calories
  - 1 food calorie = 1,000 calories of heat energy
  - Food calories are determined by burning a quatity of food and measuring how much heat energy is released.

Units of Energy

- We can convert from one unit of energy to another:
  - How many kcal are in 4,000. cal?

    Ans: 4.000 kcal
- A Joule is another unit of energy
  - How many Joules in 4,000. cal?

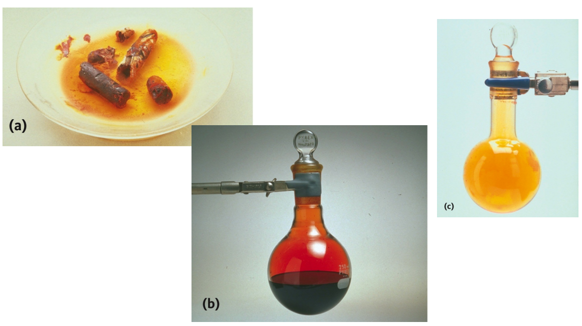
$$4{,}000\ \cancel{cal} \times \left(\frac{4.184\ J}{1\ \cancel{cal}}\right) = 16{,}740\ J$$

TABLE 4.3 Energy Units

Energy Unit	Relationship to the calorie
Kilocalorie (kcal)	1 kcal = 1000 cal
Joule (J)	1 cal = 4.184 J
Kilojoule (kJ)	1 kJ = 239 cal (1 kcal = 4.184 kJ)

4.2 The Three States of Matter

- The three states of matter are solid, liquid, and gas.



Comparison of Solids, Liquids, and Gases

- The state that prevails at a given temperature is determined by the strength of the intermolecular interactions.

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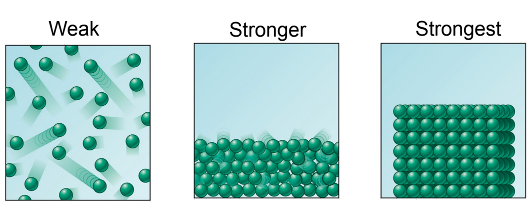


Gas: The particles are free to move throughout the container.

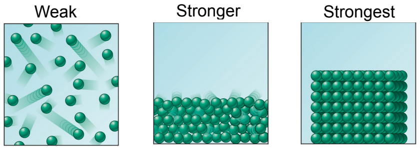
Liquid: The particles move about but remain in contact with one another.

Solid: The particles remain in fixed positions.

Molecular View of Solids, Liquids, and Gases

- At a given temperature, the molecules in all three states of matter have the same average kinetic energy.
- Because in solids and liquids the molecules are more strongly attracted to one another, the molecules have a stronger potential energy, which predominates over the kinetic energy.

Weak | Stronger | Strongest



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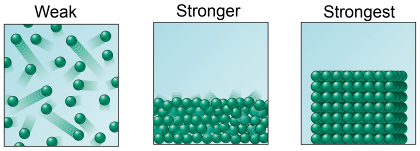
PV = nRT



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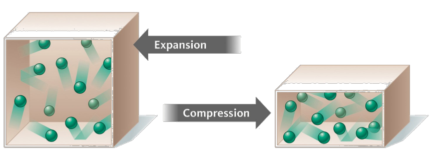
Gases can Expand or Compress

- Because the particles in gases are far apart, gases can be expanded and compressed.



In a larger volume, there is more empty space between atoms.

In a smaller volume, there is less empty space between atoms.

- This is not the case for liquids and solids where the particles are touching one another, consequently, liquids and solids are *incompressible.*

Molecular View of Solids, Liquids, and Gases

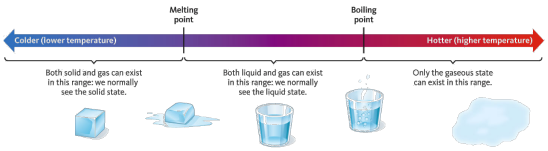
- The properties of the the three states of matter reflect the behaviors of the particles that make them up.

TABLE 4.4 The Three States of Matter

	Solid	Liquid	Gas
Shape	Fixed	Variable (liquids can be poured)	Variable (gases can be poured)
Volume	Fixed	Fixed	Variable (gases can expand and contract dramatically)
Typical density	Moderate to high (0.5 to 10 g/mL)	Moderate to high (0.5 to 10 g/mL)	Very low (0.0005 to 0.005 g/mL at room temperature)

Temperature and State

- As a substances is heated, the *kinetic energy* of its molecules increases.
- At some temperature the kinetic energy of the particles will disrupt some, but not all of the molecular interactions that hold the molecules fixed as a solid, the substance then melts to become a liquid.
  - The temperature at which a solid melts is called the melting point.



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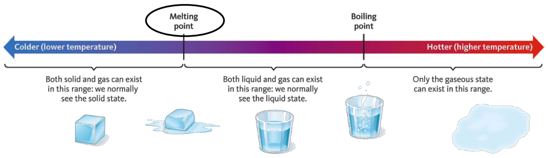
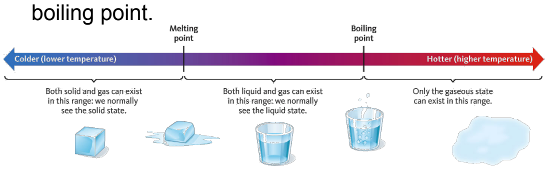


Temperature and State

- As a substances is heated further, and the *kinetic energy* of its molecules increases even more.
- A temperature will eventually be reached where the kinetic energy of the particles disrupt essentially all of the molecular interactions that were holding the molecules together and the liquid will boil and become a gas
  - The temperature at which a liquid boils, is called the boiling point.



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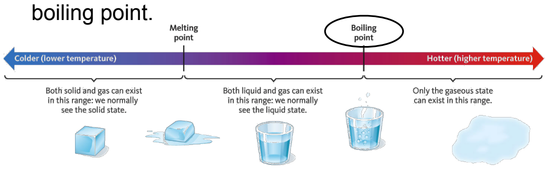

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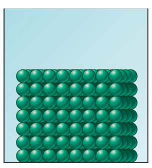

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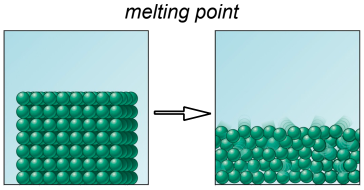
melting point



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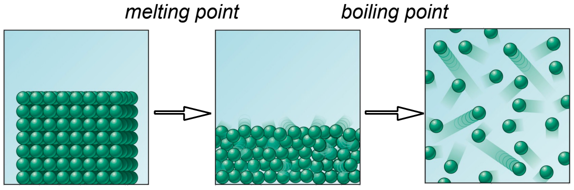
melting point

boiling point



Solid: The particles remain in fixed positions.

Liquid: The particles move about but remain in contact with one another.

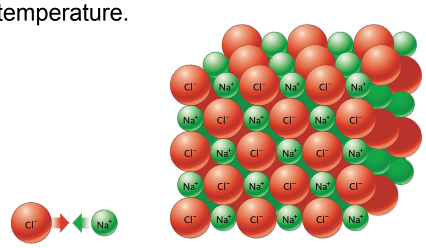
Gas: The particles are free to move throughout the container.

4.5 Attractive Forces and the Physical Properties of Matter

- There are a range of different attractive interactions that particles (molecules and ions) can have have with each other.
- We are going to focus on four of these
  - Dispersion interaction
  - Dipole/dipole interactions
    - Hydrogen bonding interactions
  - Ionic (charge/charge) interactions

Ionic (charge/charge) Interactions

- The ionic interactions are the strongest
- As a result of the strong attraction between positive and negative ions, all ionic compounds are solids at room temperature.



The powerful attraction between positive and negative ions . . . . . . overcomes their thermal energy and produces an organized array of ions.

Ionic (charge/charge) Interactions

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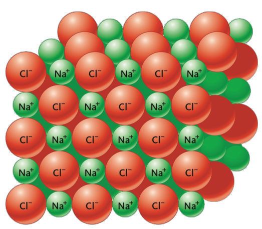
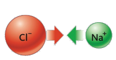
NaCl
Melting point = 801°C
Boiling point = 1,413°C





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Try It!

Question:
Is the bond between hydrogen and chlorine in H-Cl a polar covalent bond?

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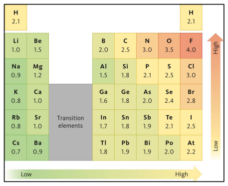
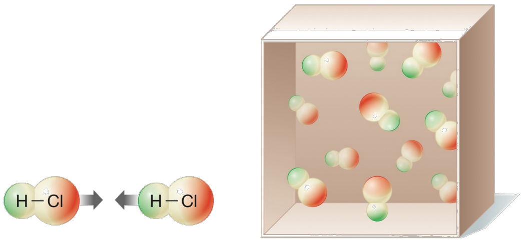


The Attraction Between Molecules

- The attraction between molecules is always weaker than the attraction between ions.



The weak attraction between HCl molecules . . . . . . cannot overcome their thermal energy. (HCl is a gas at room temperature.)

Attraction Between Molecules and Molecular Size

- All molecules are attracted to one another because the electrons of each molecule are attracted to the protons of nearby molecules.
- This interactions is called the dispersion interaction or dispersion force.
  - *All molecules* experience this interaction
  - Large molecules exert a stronger dispersion force than small molecules.
  - This is why larger molecules have higher melting and boiling points than smaller molecules

Molecular Size and Physical Properties

TABLE 4.9 The Effect of Atomic Size on Physical Properties

Substance	Strength of Dispersion Force	Melting Point	Boiling Point	State At 25°C
Chlorine (Cl_2) formula weight = 70.9 amu	Weakest	Lowest (−101°C)	Lowest (−34°C)	Gas
Bromine (Br_2) formula weight = 159.8 amu	Intermediate	Intermediate (−7°C)	Intermediate (59°C)	Liquid
Iodine (I_2) formula weight = 253.8 amu	Strongest	Highest (114°C)	Highest (185°C)	Solid

Increasing atomic size

TABLE 4.10 The Effect of Molecular Size on Physical Properties

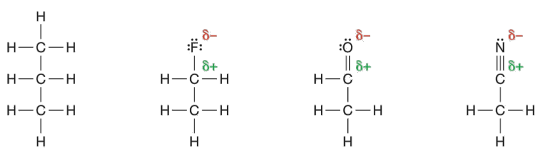
Substance	Strength of Dispersion Force	Melting Point	Boiling Point	State At 25°C
CH_4	Weakest	Lowest (−183°C)	Lowest (−161°C)	Gas
$C_{10}H_{22}$	Intermediate	Intermediate (−30°C)	Intermediate (174°C)	Liquid
$C_{20}H_{42}$	Strongest	Highest (37°C)	Highest (343°C)	Solid

Increasing numbers of atoms

Dipole-Dipole Attractions Raise Boiling Points

- Molecules that contain polar bonds tend to attract one another more strongly than molecules that are nonpolar, because the positively charged atoms in one molecule attract the negatively charged atoms of neighboring molecules.
- The attraction of molecules with polar bonds is called dipole-dipole attraction.

Dipole-Dipole Attraction and Boiling Points



Boiling point: −42°C
Nonpolar molecule
Low boiling point due to weakness of dispersion force.

Boiling point: −38°C
Polar molecule
Low boiling point: weak dispersion force and little effect from dipole-dipole attraction.

Boiling point: 21°C
Polar molecule
Significant dipole–dipole attraction raises the boiling point.

Boiling point: 82°C
Polar molecule
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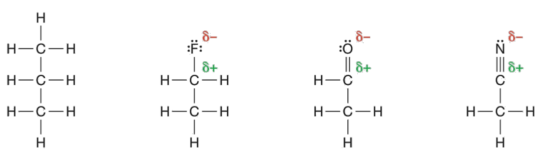
Increasing strength of dipole–dipole attraction
Increasing boiling points

Dipole-Dipole Attraction and Boiling Points

All of these molecules have about the same size and therefore similar dispersion interactions



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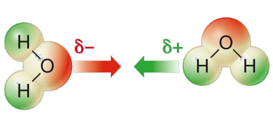
Increasing strength of dipole–dipole attraction
Increasing boiling points

Hydrogen Bonding

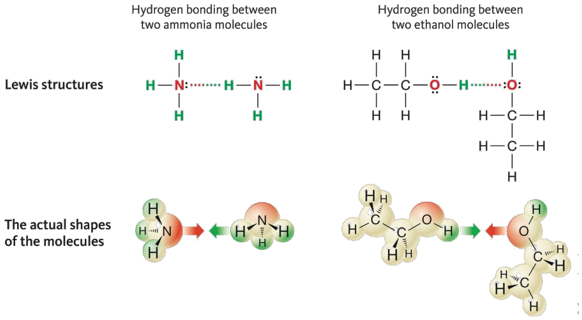
- Any molecule that contains O–H or N–H covalent bonds will form hydrogen bonds with a molecule that contains either an oxygen (:O) or nitrogen (:N) having non-bonding pair of electrons.



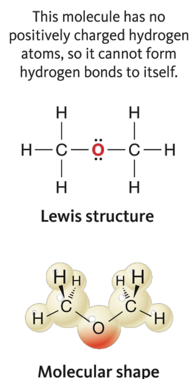
The attraction between the positively charged hydrogen and the negatively charged oxygen is called a **hydrogen bond.**

In the rest of this chapter, hydrogen atoms that can participate in hydrogen bonds are colored **green**. Negatively charged oxygen and nitrogen atoms that can participate in hydrogen bonds are colored **red**.

Example of Hydrogen Bonding

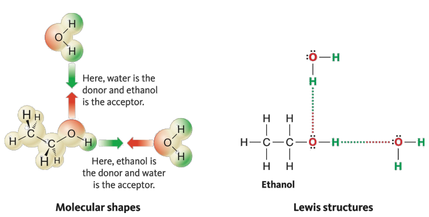


Some Polar Molecules Cannot Form Hydrogen Bonds



Hydrogen bond donors and acceptors

- The molecule that supplies the hydrogen atom is the hydrogen bond donor.
- The molecule that contributes the negatively charged atom with a non-bonded pair of electrons is the hydrogen bond acceptor.



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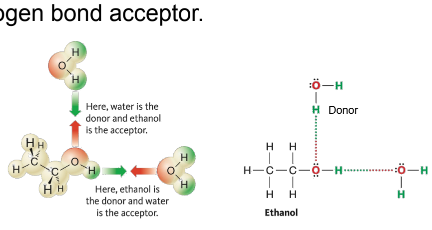

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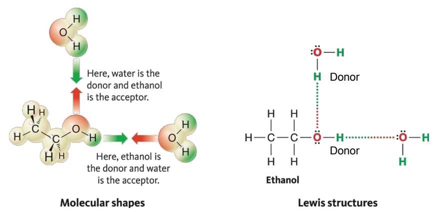

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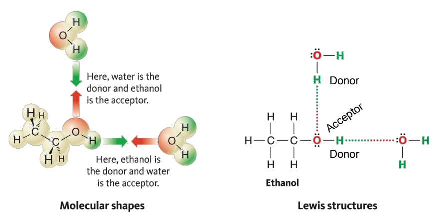

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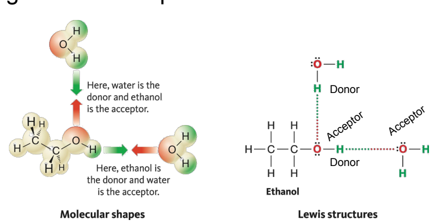


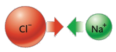
Summary of Attractive Forces

TABLE 4.11 Attractive Forces That Have an Impact on Boiling and Melting Points

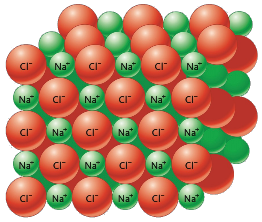
Type of Force	Types of Compounds That Exhibit This Force	Strength of This Force
Dispersion force	All molecular compounds	Weak, increases as the size of the molecule increases
Dipole–dipole attraction	Molecular compounds that contain polar bonds	Weak, primarily significant for molecules that contain N or O
Hydrogen bond	Molecular compounds that contain O—H or N—H groups	Weak, but always raises the melting and boiling point significantly
Ion–ion attraction	All ionic compounds	Very strong (ionic compounds have very high melting and boiling points)

Ionic (charge/charge) Interactions

- The ionic interactions are the strongest
- As a result of the strong attraction between positive and negative ions, all ionic compounds are solids at room temperature.



The powerful attraction between positive and negative ions . . .



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NaCl
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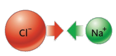

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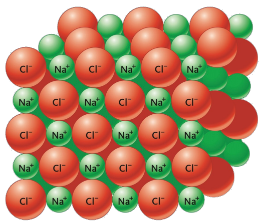

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Question:
Can you think of a simple way to disrupt the strong ion-ion interactions in NaCl?

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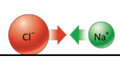
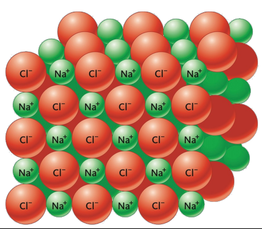
Melting point = 801°C
Boiling point = 1,413°C



The powerful attraction between positive and negative ions . . .

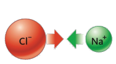


. . . overcomes their thermal energy and produces an organized array of ions.

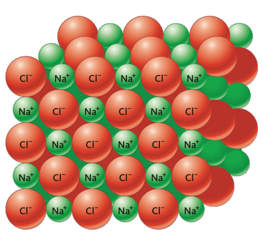
Ionic (charge/charge) Interactions

- The ionic interactions are the strongest
- As a result of the strong attraction between positive and negative ions, all ionic compounds are solids at room temperature.

NaCl
Melting point = 801°C
Boiling point = 1,413°C



The powerful attraction between positive and negative ions . . .



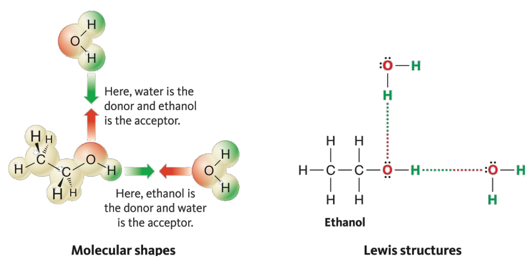
. . . overcomes their thermal energy and produces an organized array of ions.

4.6 Solutions and the Dissolving Process

- In a solution, the liquid in the greatest amount is the solvent, and the minor substance in the solution is the solute.
- Water is the most common solvent in biological systems; water solutions are called aqueous solutions.
- If a solid remains visible when added to water and settles when agitation stops, it is not a solution, but a suspension (example: sand in water).

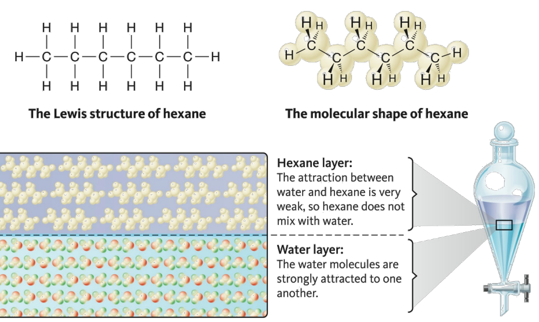
Compounds that Tend to Dissolve in Water

- Compounds that Form Hydrogen Bonds Tend to Dissolve in Water



Molecules that Cannot Hydrogen Bond with Water

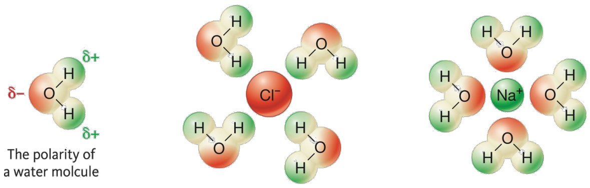
- Compounds that Cannot form Hydrogen Bonds Tend Not to Dissolve in Water



4.7 Electrolytes and Dissociation

- Solutions of electrolytes in water are conductive.
- To conduct electricity, a solute must form ions when it dissolves in water. This process is called dissociation.
- When an ionic compound dissolves in water, water molecules surround the ions and pull them away from one another (the solvation process).
- The positively charged hydrogen atoms in water are attracted to the negative ions, and the negatively charged oxygen atoms in water are attracted to the positive ions.

The Solvation of Ions

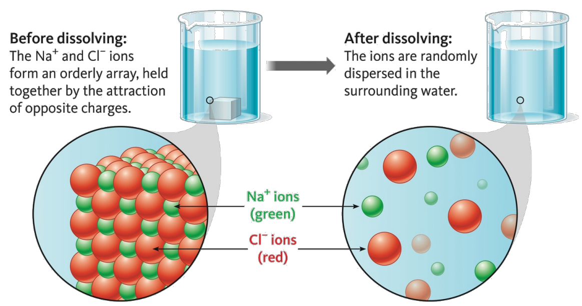


The polarity of a water molcule

When NaCl dissolves, the negative chloride ions are attracted to the positive hydrogen atoms of water . . .

. . . and the positive sodium ions are attracted to the negative oxygen atoms of water.

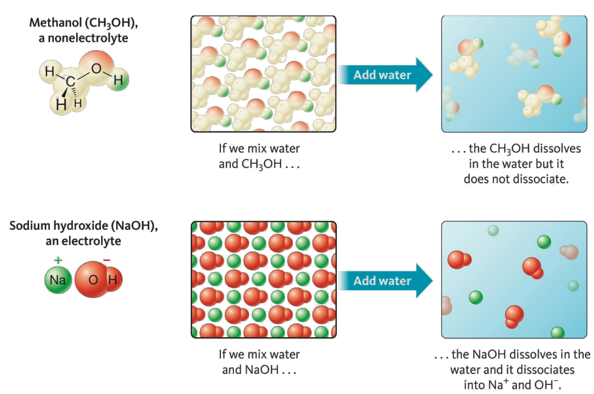
Dissociation of Ions in Water



Nonelectrolytes

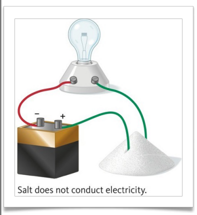
- When a molecular compound dissolves, the molecules move away from one another but generally do not break apart into ions.
- Nonelectrolytes are substances that do not conduct electricity when they dissolve in water.
- Most molecular solutes are nonelectrolytes .

Electrolytes and Nonelectrolytes in Water

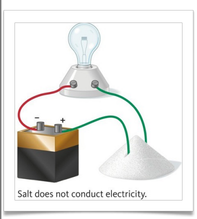


Electrolytes and Conductivity

Electrolytes and Conductivity

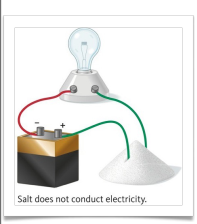


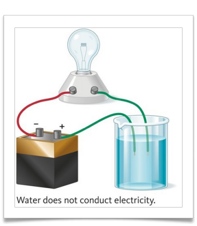
Electrolytes and Conductivity



Why?

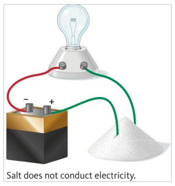
Electrolytes and Conductivity



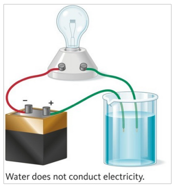


Why?

Electrolytes and Conductivity



Salt does not conduct electricity.

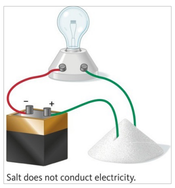


Water does not conduct electricity.

Why?

Why?

Electrolytes and Conductivity



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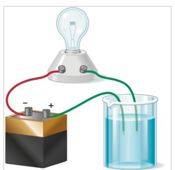

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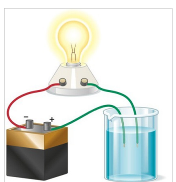


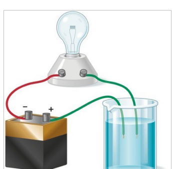
Salt water conducts electricity, so salt is an electrolyte.

Why?

Why?

Electrolytes and Conductivity

Salt does not conduct electricity.



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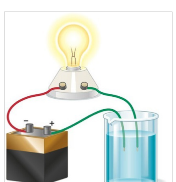


Salt water conducts electricity, so salt is an electrolyte.

Why?

Why?

Why?

Common Examples

TABLE 4.13 Common Electrolytes and Nonelectrolytes

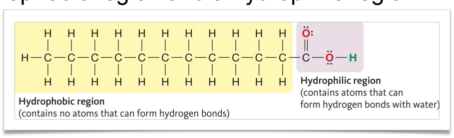
Electrolytes (contain a metallic element or the NH_4 group)	Nonelectrolytes (do not contain a metal or NH_4)
$NaCl$ (table salt: a major source of sodium and chloride ions)	$C_{12}H_{22}O_{11}$ (table sugar)
KI (potassium iodide: added to table salt as a source of iodide ions)	C_2H_5OH (ethanol, also called ethyl alcohol or grain alcohol)
$(NH_4)_2CO_3$ (ammonium carbonate: an ingredient in smelling salts and some leavening agents)	C_3H_6O (acetone: an ingredient in many paint thinners and in nail polish remover)
KH_2PO_4 (monobasic potassium phosphate: used in sports beverages as a source of potassium)	C_2H_6OS (dimethyl sulfoxide: used to reduce inflammation and transport medications through the skin; also called DMSO)*
$Ca(C_3H_5O_3)_2$ (calcium lactate: used in sports beverages as a source of calcium)	$C_{10}H_{19}O_6PS_2$ (malathion: an insecticide)

5.3 Solubility and Molecular Structure

- The ability to hydrogen bond makes many compounds soluble in water, but that ability is not the only factor that matters. The entire structure of the molecule plays a role in solubility.
- Fats are a good example.
  - They have a hydrophobic (water-fearing) region (cannot hydrogen bond)
  - AND they have a hydrophilic (water-loving) region (can hydrogen bond)

Lauric Acid- A fat

- Lauric Acid is a good example of a molecule with a hydrophobic region and a hydrophilic region.



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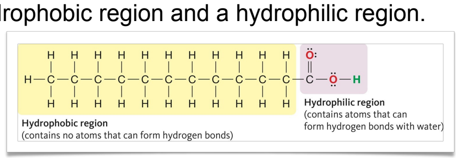
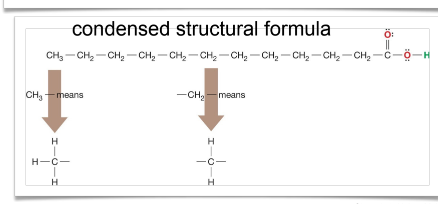




Hydrogen Bonding

- The more atoms that can participate in hydrogen bonds, the higher the solubility in water.
- The more carbon and hydrogen atoms, the lower the solubility.

TABLE 5.3 The Effect of Adding Hydrogen-Bonding Atoms on Water Solubility

Compound	Solubility in Water
$CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3$ **Heptane: no hydrogen bonding is possible**	Lowest (0.3 g/L)
$CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-C(=O)-OH$ **Heptanoic acid: three atoms can participate in hydrogen bonds**	Intermediate (2.4 g/L)
$H-O-C(=O)-CH_2-CH_2-CH_2-CH_2-CH_2-C(=O)-O-H$ **Pimelic acid: six atoms can participate in hydrogen bonds**	Highest (25 g/L)

Carbon chain length and solubility

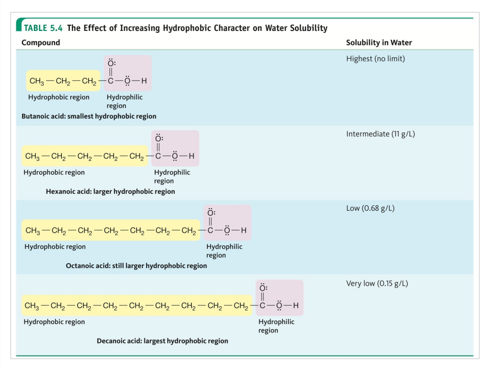


TABLE 5.4 The Effect of Increasing Hydrophobic Character on Water Solubility

Compound	Solubility in Water
$CH_3-CH_2-CH_2-C(=O)-O-H$ (Hydrophobic region: $CH_3-CH_2-CH_2$; Hydrophilic region: $C(=O)-O-H$) **Butanoic acid: smallest hydrophobic region**	Highest (no limit)
$CH_3-CH_2-CH_2-CH_2-CH_2-C(=O)-O-H$ (Hydrophobic region; Hydrophilic region) **Hexanoic acid: larger hydrophobic region**	Intermediate (11 g/L)
$CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-C(=O)-O-H$ (Hydrophobic region; Hydrophilic region) **Octanoic acid: still larger hydrophobic region**	Low (0.68 g/L)
$CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-C(=O)-O-H$ (Hydrophobic region; Hydrophilic region) **Decanoic acid: largest hydrophobic region**	Very low (0.15 g/L)

Vitamins and Solubility

- Solubility effects how our bodies use and store vitamins.
  - Water soluble vitamins (for example, Vitamin C and all of the B vitamins) dissolve in water and are not stored, therefore they must be a regular part of our diet.
  - Fat soluble vitamins (A,D,E, and K) are stored along with the fats in our bodies. Excessive amounts can be dangerous.

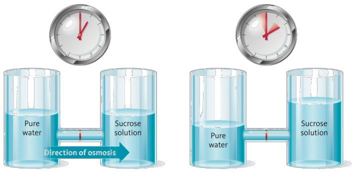
5.5 Osmosis, Dialysis and Tonicity

- Diffusion is the tendency of particles to distribute evenly in a mixture.

When you add a drop of coloring to water . . . . . . the solute and solvent molecules move about . . . . . . until they are evenly distributed.

Osmosis

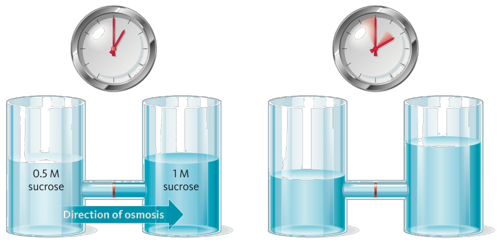
- Osmosis is the net movement of water through a semipermeable membrane.
  - A semipermeable membranes allows water, but not the solutes to pass through the membrane
- When the concentration of solutions on the two sides of the membrane are different, osmosis will occur.



At the start of the experiment, the two liquid levels are equal. At the end of the experiment, the level of the sucrose solution is higher.

Osmosis

- Water will move from the lower concentration side to the higher concentration side, trying to make the concentrations of each side equal.



Water flows from the solution with the lower molarity to the solution with the higher molarity.

At the end of the experiment, the two liquid levels are different.

Osmosis

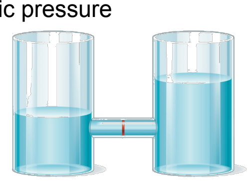
- If there is more than one solute in the solution, add up the molarities of each solute to determine which direction water will flow.
- Example: Solution A contains 0.1 M glucose and 0.05 M sucrose, solution B contains 0.12 M glucose. If these solutions are separated by a semipermeable membrane, which direction will water flow?

Soln A: total molarity = 0.15 M Soln B: total molarity = 0.12 M.

Water will flow from solution B to solution A.

Osmotic Pressure

- Osmotic pressure is pressure caused by the flow of water during osmosis.
- The greater the difference in concentration between the two solutions, the greater the osmotic pressure.
- If one of the solutions is ionic, dissociation effects osmotic pressure



At the end of the experiment, the two liquid levels are different.

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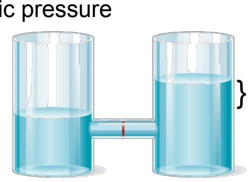

} The extra solution on this side of the membrane creates a downward pressure that opposes the flow of the water

At the end of the experiment, the two liquid levels are different.

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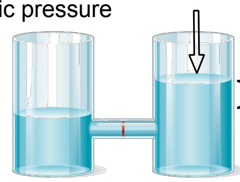

At the end of the experiment, the two liquid levels are different.

Tonicity

- Cell membranes are semipermeable
- Tonicity is the relationship between the overall concentration of the solution and the normal solute concentration within a cell, such as a blood cell.
  - Isotonic solutions contain a solute concentration **equal to** that within the cells.
  - Hypertonic solutions contain a solute concentration that is **higher than** what is inside cells.
  - Hypotonic solutions contain a solute concentration that is **lower than** what is inside cells.

Tonicity

TABLE 5.5 The Effect of Tonicity on Red Blood Cells

Tonicity of the solution	Hypotonic	Isotonic	Hypertonic
Total solute concentration	Less than 0.28 M	0.28 M	Greater than 0.28 M
Direction of osmosis	Water flows into the cell.	No osmosis occurs.	Water flows out of the cell.
Effect on a red blood cell	The cell swells, and it will burst *(hemolyze)* if the solute concentration is much lower than 0.28 M.	The cell is unaffected.	The cell shrinks, and it will shrivel up *(crenate)* if the solute concentration is much higher than 0.28 M.

Tonicity

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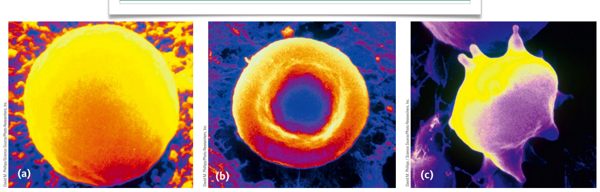
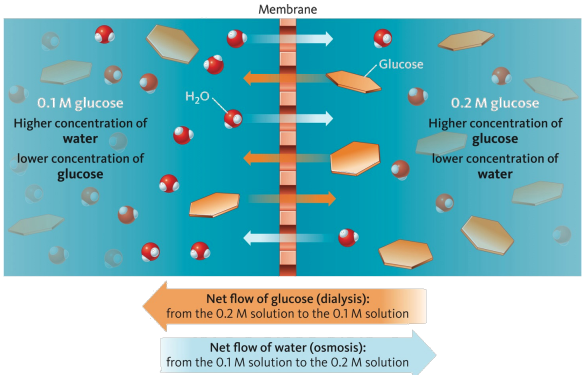


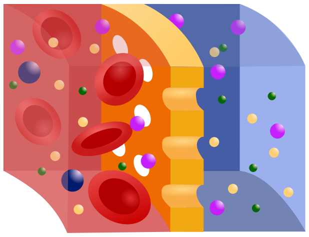
Dialysis

- Dialysis is the movement of *solute* particles through a membrane.
- Semipermeable membranes are materials that allow only some particles to pass.
  - Most allow small particles to pass through but not large particles.
- Osmosis and dialysis will occur in opposite directions, such that both the water and the solute are moving from lower concentration to higher concentration

Osmosis and Dialysis Comparison



Hemodialysis



5.6 Equivalents

- An equivalent is the amount of any ion that has the same total charge of 1 mol of hydrogen ions (H^+).
- Practically, the number of equivalents is equal to the number of moles times the charge.

1 mol of K^+	1 Eq of K^+
1 mol of NO_3^-	*1 Eq of* NO_3^-
1 mol of Mg^{2+}	*2 Eq of* Mg^{2+}
1 mol of S^{2-}	*2 Eq of* S^{2-}
1 mol of Fe^{3+}	*3 Eq of* Fe^{3+}
1 mol of PO_4^{3-}	*3 Eq of* PO_4^{3-}

Example

- A solution contains 0.31 mol of phosphate ion. How many equivalents of phosphate ions does the solution contain?
- 1 mol of PO_4^{3-} = 3 Eq of PO_4^{3-}
  - Translates to two conversion factors:

$$\frac{1\text{ mol}}{3\text{ Eq}} \qquad \frac{3\text{ Eq}}{1\text{ mol}}$$

$$0.31\ \cancel{\text{mol}}\ PO_4^{3-} \times \frac{3\text{ Eq}}{1\ \cancel{\text{mol}}} = 0.93\text{ Eq}$$

Try It!

Question:

How many grams sodium citrate are needed to make 1.0 L of a 25 mEq/L solution of citrate.

Next Up

- Unit 3 - Chemical Reactions
  - Reading Assignment: Chapter 6-4,5,6 & 7
  - Mastery Assignment due Feb. 10
  - Problem Assignment due Feb. 10