Grade 11 Chemistry

Solutions and Solubility

Class 10

Think About It…

Which one would you drink?

- BP Oil Spill – April 10, 2010 in the Gulf of Mexico
- Lasted 87 days
- 200 million gallons of oil leaked

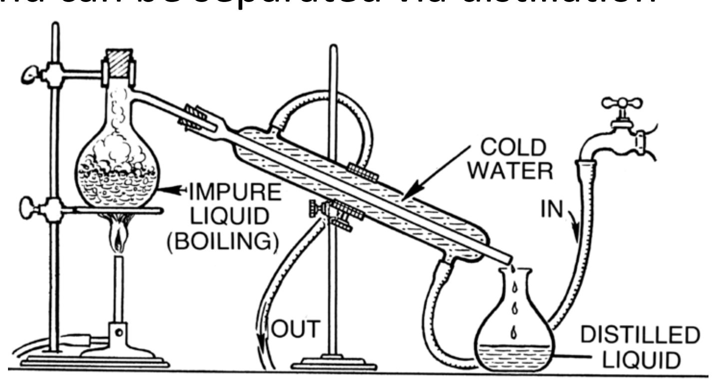
Overall Expectations

- Properties of solutions can be described qualitatively and quantitatively, and can be predicted
- Living things depend for their survival on the unique physical and chemical properties of water
- People have a responsibility to protect the integrity of Earth's water resources

Solutions

- Solution – a homogenous mixture
- Solvent – the substance present in the largest amount in a solution; substance that has other substances dissolved in it
- Solute – the substance present in the smaller amount in a solution; substance being dissolved

- Pure substances – fixed composition (i.e. water)
- Solutions have variable composition – different ratios of solvent to solute are possible
  - Physical Change: Water and ethanol have different boiling points and can be separated via distillation

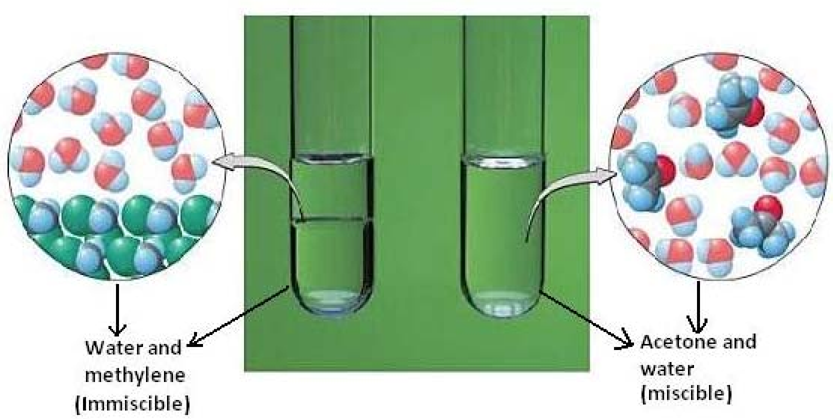


Types of Solutions

- Solutions can be gas, liquid or solid

Solute	Solvent	Examples
Gas	Gas	Air, natural gas
Gas	Liquid	Carbonated drinks, water in rivers and lakes
Gas	Solid	Hydrogen in platinum
Liquid	Gas	Water vapour in air
Liquid	Liquid	Alcohol in water
Liquid	Solid	Mercury in silver
Solid	Gas	Mothballs in air
Solid	Liquid	Sugar in water
Solid	Solid	Alloys

- Miscible – readily dissolve in each other
  - – Water and alcohol
- Immiscible – do not readily dissolve in each other
  - – Water and oil



Dissolution

- What is the difference between solubility and dissolving?
  - Solubility – the *amount* of solute that dissolves in the solvent
  - Dissolution – the *speed* of the solute dissolving in the solvent
- Depends on:
  - Temperature
  - Agitation
  - Particle Size

Temperature

- For solid solutes, the rate of dissolving is greater at high temperatures because the solvent molecules have higher kinetic energy

Agitation

- Agitating a mixture by stirring or shaking the container increases the rate of dissolving because fresh solvent makes contact with undissolved solid

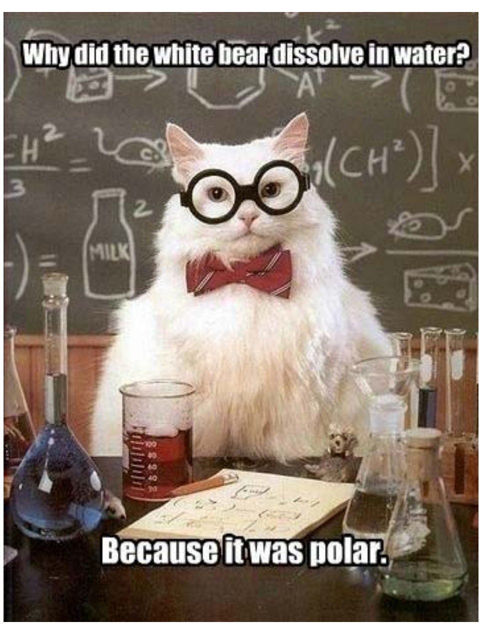
Size of Particles

- Decreasing the size of the particles increases the rate of dissolving

Why does something dissolve?

- The reasons why a solute may or may not dissolve in a solvent are related to the forces of attraction
- When forces between different particles in a mixture are stronger than the forces of attraction between like particles, a solution forms

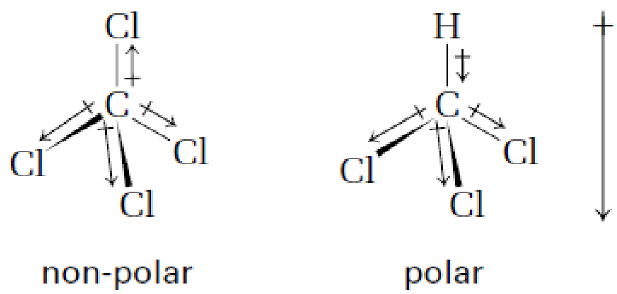
Polar and Non-Polar Substances



- Like dissolves like
  - Polar compounds dissolve in polar solvents
  - Non-polar compounds dissolve in non-polar solvents

Molecular Shape and Polarity

- Polarity can be determined by looking at electronegativity

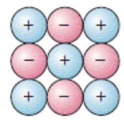
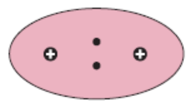
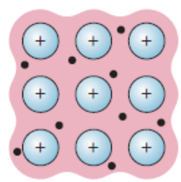


- Different molecular shapes can be non-polar or polar depending on the distribution of the atoms

Intramolecular Forces

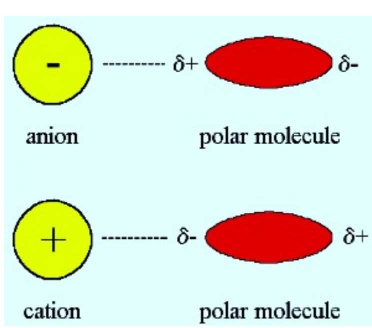
- **Ionic Solid** – held together by electrostatic attractions between cations and anions (ex: NaCl)
- **Covalent Network Solid** – atoms are connected in a lattice of covalent bonds (ex: Diamonds, Quartz)
- **Metallic Solid** – a lattice of cations (positive metals ions) surrounded by a "cloud of delocalized" electrons

Table 4.4 Comparing Intramolecular and Internolecular Forces

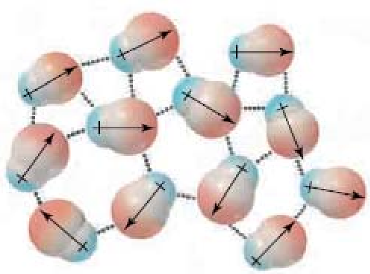
	Force	Model	Nature of Attraction	Energy (kJ/mol)	Example
Intramolecular	ionic		cation-anion	400–4000	NaCl
	covalent		nuclei-shared electron pair	150–1100	H—H
	metallic		cations-delocalized electrons	75–1000	Fe

Intermolecular Forces

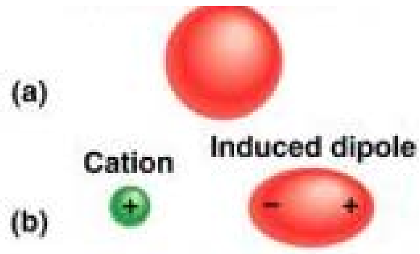
- Intermolecular forces (aka. Van der Waals forces) are relatively weaker than intramolecular forces



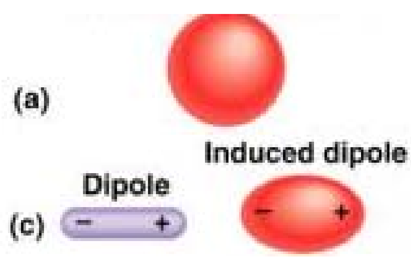
1) Ion-Dipole Forces – the electrostatic interactions between an *ion* and the partial charges on a *polar molecule*



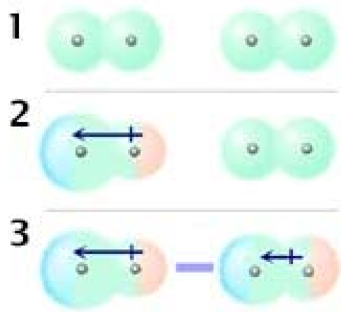
2) Dipole-Dipole Forces – electrostatic interactions between neutral but *polar molecules*; forces can be attractive or repulsive



3) Ion-Induced Dipole – electrostatic interactions between an *ion* and a *non-polar molecule* to produce a momentary dipole

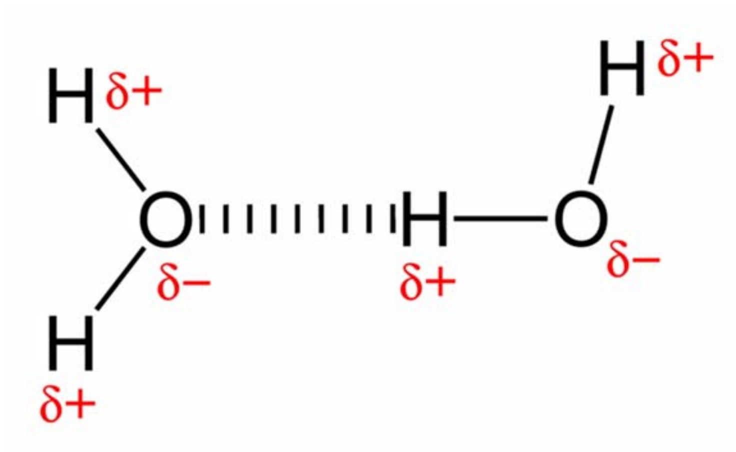


4) Dipole-Induced Dipole Forces – electrostatic interactions between a *neutral, polar molecule* and a *nonpolar molecule* to produce a momentary dipole



5) London Dispersion Forces – attractive forces from the temporary dipoles induced in atoms or molecules; due to the random shift in electrons

5) Hydrogen Bonding – attractive interaction between a hydrogen atom bonded to a very electronegative atom (HO, HN, HF) and an unshared electron pair on another electronegative atom (O, N, F); between polar molecules

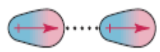
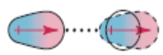
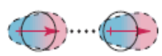


Checkpoint

Identify the binary mixture that cannot experience hydrogen bonding with each other:

a. NH_3 / H_2O
b. H_2O / HF
c. HF / CO_2
d. H_2S / HCl

Intermolecular

Force	Model	Nature of Attraction	Energy (kJ/mol)	Example
ion-dipole		ion charge-dipole charge	40-600	Na^{+}----$O(H)_2$
hydrogen bond	$\overset{\delta^-}{—A}—\overset{\delta^+}{H}\cdots\cdots:\overset{\delta^-}{B}—$	polar bond to hydrogen-dipole charge (lone pair, high *EN* of N, O, F)	10–40	$:\ddot{O}(H)—H$----$:\ddot{O}(H)—H$
dipole-dipole		dipole charges	5–25	I—Cl----I—Cl
ion-induced dipole		ion charge-polarizable electrons	3–15	Fe^{2+}----O_2
dipole-induced dipole		dipole charge-polarizable electrons	2–10	H—Cl----Cl—Cl
dispersion (London)		polarizable electrons	0.05–40	F—F----F—F

Comparing Strengths

Ion-Dipole

Hydrogen Bond

Dipole-Dipole

Ion-Induced Dipole

Dipole-Induced Dipole

London Dispersion

Increasing Strength
- Melting Point increases
- Boiling Point increases