1

Chemistry

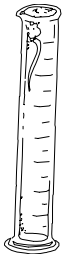
2.

Matter is anything that occupies space and has mass.

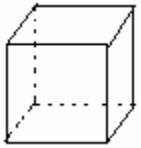
Examples:

- Air
- Oxygen
- Table
- Chair
- Water

Find mass using a balance



Find the volume of a liquid and an irregular solid using a graduated cylinder

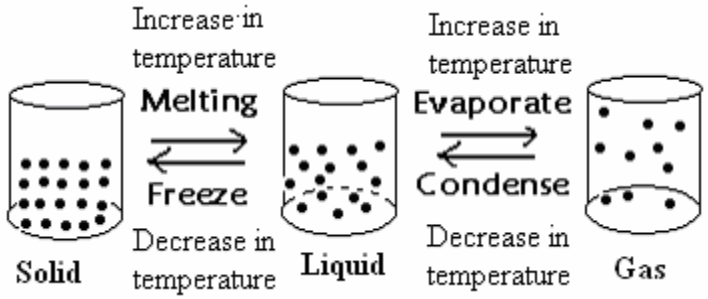


Find the volume of a regular solid by l x w x h

Matter

3.

- Solid- has definite shape and volume and is not compressible.
- Liquid- (fluid) Flows; it has a fixed volume, and takes the shape of its container.
- Gas – (fluid) Flows, takes the shape and volume of the container, and is compressible.



Three states of matter

4.

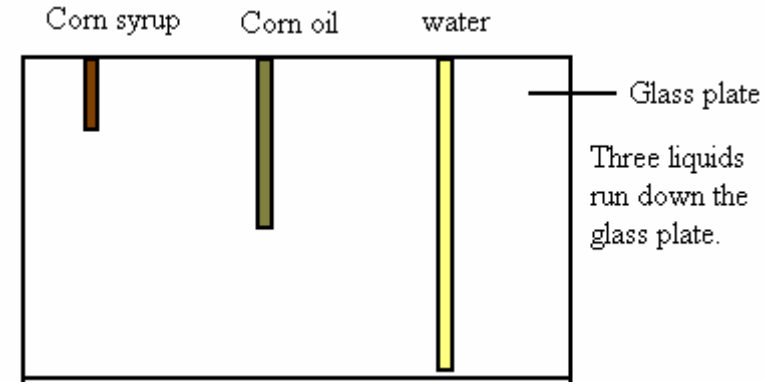
Fluid -A substance in which the atoms or molecules can freely move past one another.

Examples: Gases and Liquids

Fluid

5.

- A fluid's resistance to flowing.
- Heat caused liquids to flow faster.
- Loss of heat caused liquids to flow slower.



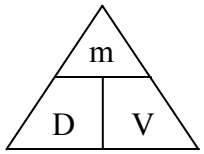
The liquid with the highest viscosity flows the slowest

Viscosity

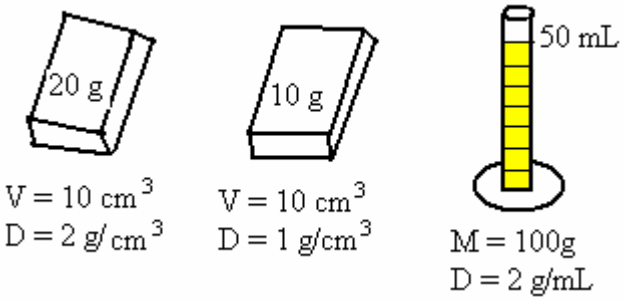
6.

Density is the mass per unit of volume.
It is affected by a change in temperature.

Formula: Density = $\frac{\text{mass}}{\text{Volume}}$ $D = \frac{m}{V}$



Units: g/mL or g/cm^3

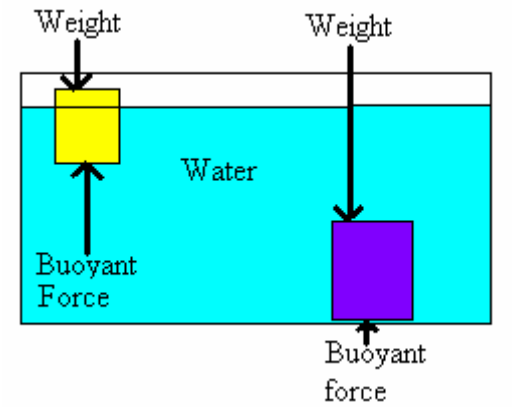


Short cut – **if the volumes are equal, the one with more mass has more density.**

Density

7.

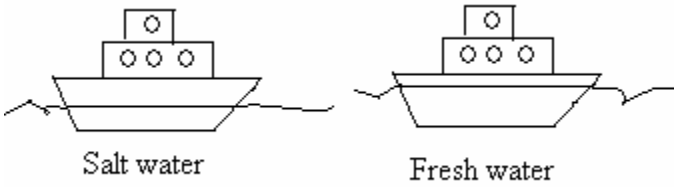
The ability of a fluid to exert and upward force on an object immersed in it.

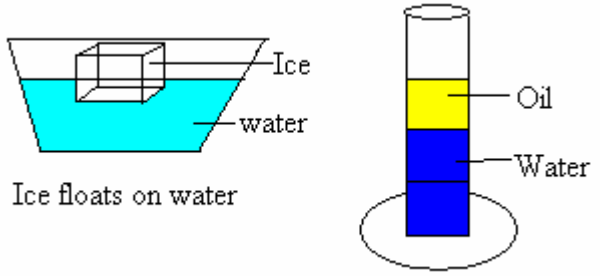


Buoyancy

8.

Boats float higher on salt water than on fresh water.





In order to float the overall density of the object or liquid must be less than the fluid they are in.

Why does something float?

9.

Chemical properties

- How a substance behaves
- Examples:
  - Decomposes by electricity
  - Reacts with water/acids
  - Burns in oxygen
  - Non-reactive

Physical Properties

- Descriptive
- Can be observed without changing composition.
- Examples:
  - Mass
  - Volume
  - Odor
  - Color
  - Boiling point
  - Melting point
  - Freezing point

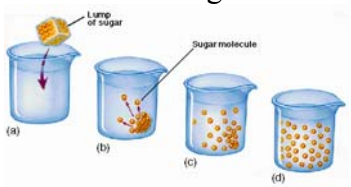
Kinds of Properties

10.

Alters a substance without changing its composition.

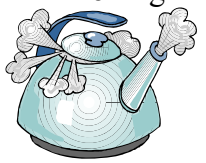
Examples: cutting, grinding, bending, condensation of steam

Dissolving

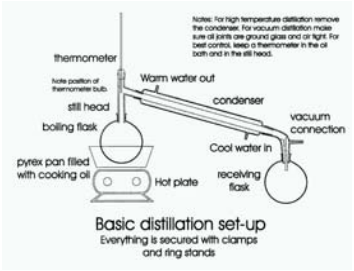


Evaporation and condensation

Boiling



Distillation



Breaking

Melting Ice

Cutting

Grinding

Physical Change

11.

During a chemical change a **new substance** is formed.

Examples:

Combustion

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Respiration

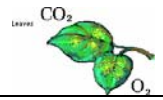
$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$

Rusting

$2Fe + 3\ O_2 \rightarrow Fe_2O_3$

Photosynthesis

$6CO_2 + 6H_2O \xrightarrow[\text{Light}]{} C_6H_{12}O_6 + 6O_2$



Digestion

Large food molecules → Smaller food molecules

Oxidation – reacting with oxygen

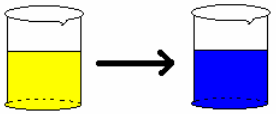
$2Mg + O_2 \rightarrow 2MgO$

Chemical Change

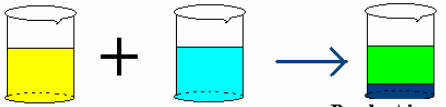
12.

- Increase or decrease in temperature.
- Color changes.
- Production of a new odor or the release of a gas
- Formation of a precipitate.

Color Change



Formation of a precipitate

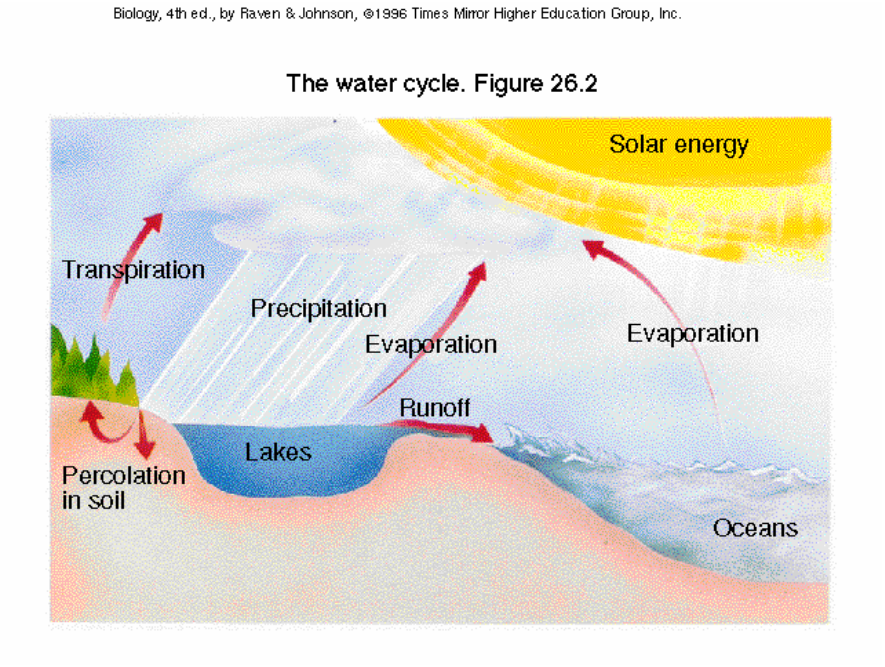


Release of heat or light

Production of a gas

Possible Indications of a Chemical Change

13.



Physical Change

14.

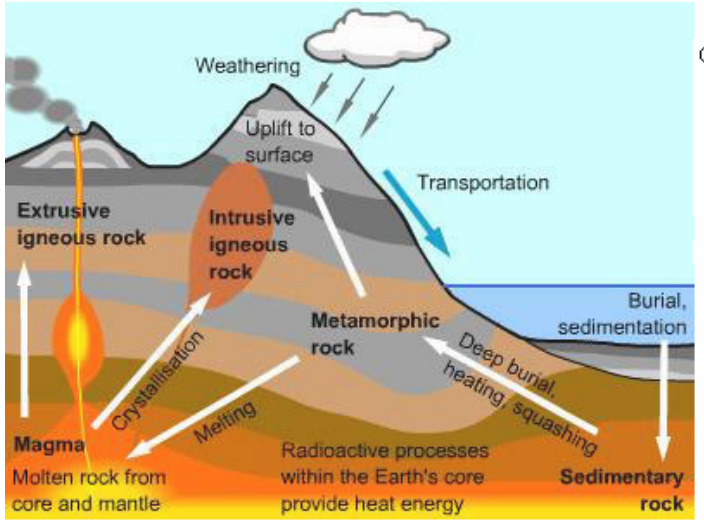
Weathering is breaking down of rock.

Chemical	Physical
CO_2 and H_20 react with calcium carbonate (limestone), which can then form stalactites and stalagmites in caves.	Water in the rock cracks can freeze and then break the rock apart.

Erosion is the carrying away of rocks and dirt. (Physical change)

Weather and Erosion

15.

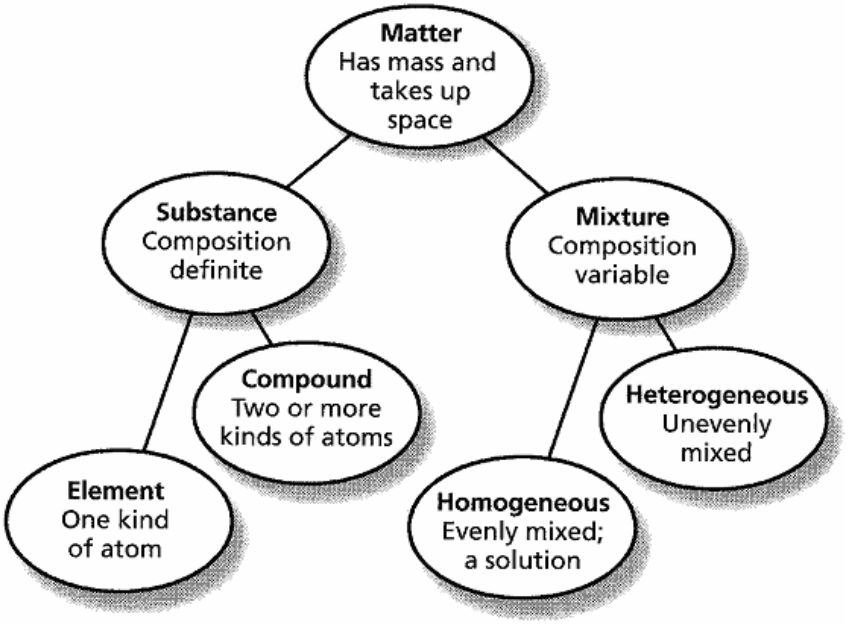


Chemical change
Sedimentary to metamorphic and metamorhpic to igneous

Physical changes
Sedimentation
Uplifting
Melting
Crystallization
Weathering

Rock Cycle and Physical and Chemical changes

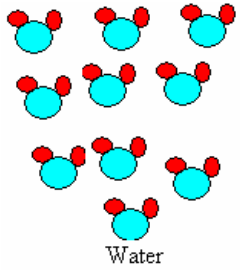
16.



Classification of Matter

17.

- Definite composition
- Elements and compounds
- Homogeneous (same throughout)

Element	Compound
	

Pure Substance

18.

Homogeneous	**Heterogeneous**
• Variable Composition • Uniform throughout • Called a solution Examples: • Air • Tap water • Salt water • Toothpaste • Metals and alloys (Brass = Zn and Cu)	• Variable composition • Non-uniform • Distinct phases Examples: • Vegetable soup • Concrete • Italian Salad Dressing • Silver-plated jewelry • Fruit salad

Mixtures can be separated into parts by physical means:
magnet, distillation, filtration, by hand, and solubility

Types of Mixtures and Means of Separation

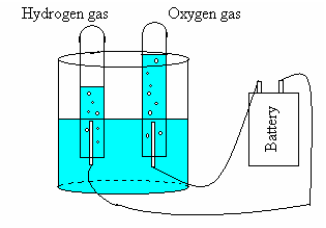
19.

- Compounds are **made of two or more kinds of atoms chemically combined**.
- They are **pure substances**.
- They are **homogeneous**.
- They are **represented by formulas**

Smallest unit of a covalent compound is a molecule.	**Smallest unit of an ionic compound is a formula unit**.
CO_2 = carbon dioxide H_2O = water	NaCl = sodium chloride CaF_2 = calcium fluoride $NaHCO_3$ = baking soda

Compounds can only be broken down by chemical means.

Electrolysis of water



Compounds

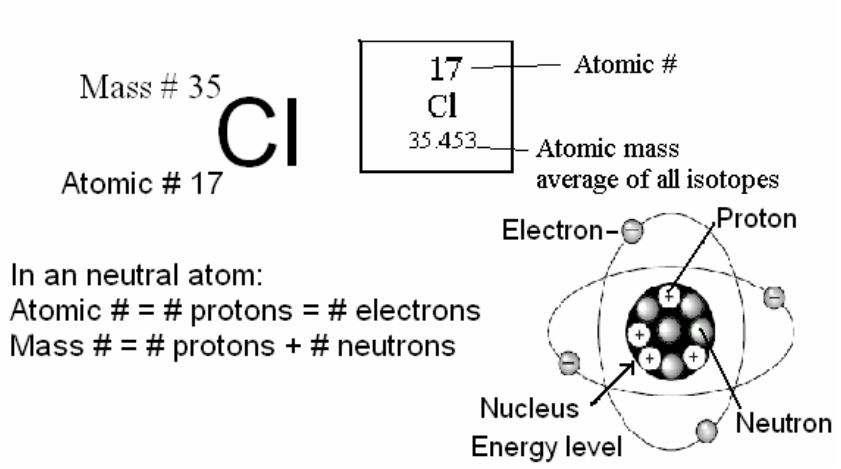
20.

- Made of one kind of atom
- Pure substance
- Homogeneous
- Represented by a symbol
- Smallest unit is called an atom
- Examples:
  - Sulfur -------- S
  - Iron ---------- Fe
  - Gold ---------- Au
  - Sodium ------- Na
  - Fluorine ------ F
  - Magnesium – Mg

Element

21

Particle	Location	Charge
Proton	Nucleus	+
Neutron	Nucleus	No charge
Electron	Energy levels	-



Parts and Relationships of an atom

22.

Atoms that have the same number of protons but a different number of neutrons.

Isotopes

$^{35}_{17}Cl$ $^{36}_{17}Cl$

35 mass # — 36 mass #
-17 protons — -17 protons
18 neutrons — 19 neutrons

Not Isotopes

$^{36}_{17}Cl$ $^{36}_{18}Ar$

Chlorine has 17 protons — Argon has 18 protons

Since they have a different number of protons, they are two different elements not isotopes.

Isotopes

23.

Metals

- Luster (shiny)
- Malleable (will bend)
- Ductile (can be made into wire)
- Almost all have 1, 2, 3 electrons in their outer most energy level
- Located on the left and in the center of periodic table

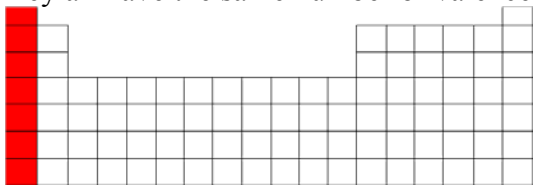
Non-metals

- Dull (no luster)
- Brittle
- Not ductile
- Most have 5, 6, 7 electrons in their outer most energy level
- Located on the right of the periodic table

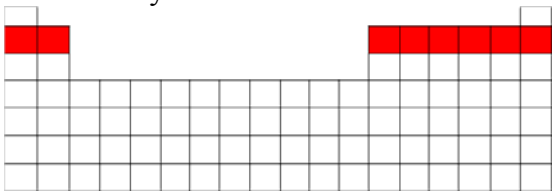
Metals vs. Non-metals

24.

Groups or Families are in vertical columns; elements in a family are chemically similar and react similarly in chemical reactions. They all have the same number of valence electrons.

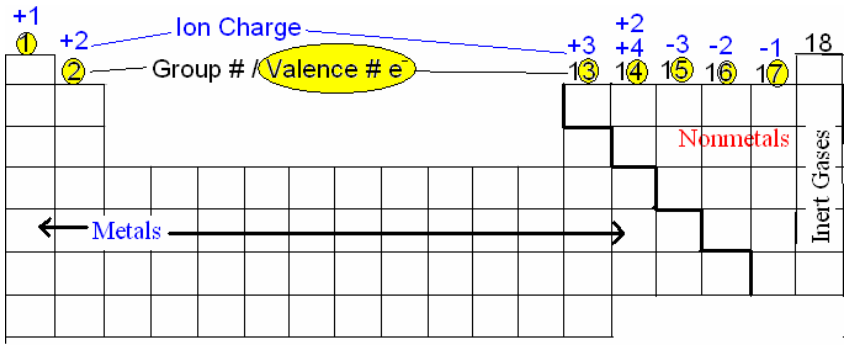


Periods are horizontal rows. Elements in a period are not chemically similar but they do have the same number of energy levels.



Periodic Table Organization - Groups and Periods

25.

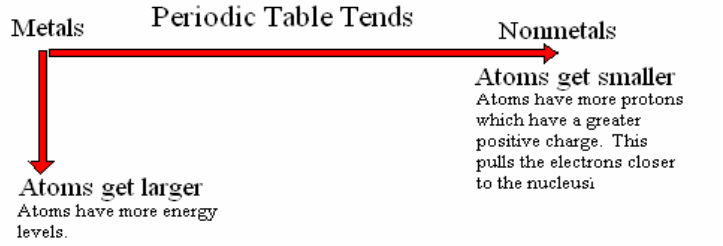


Metals are **losers!** They lose electrons.
Nonmetals are **gainers**. They gain electrons.
Inert gases **neither gain nor lose** electrons.

Periodic Table

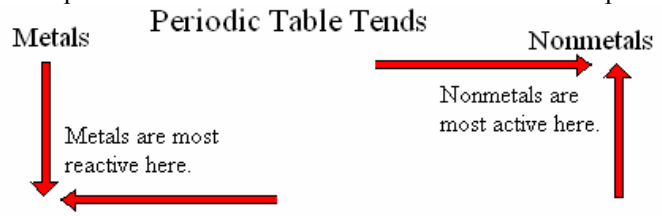
26.

Size of Atoms



Activity

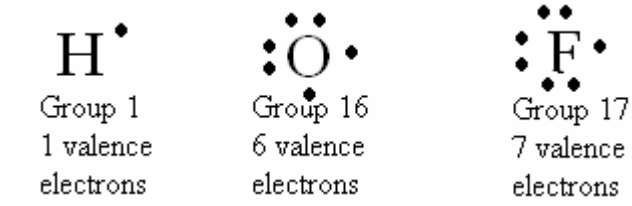
Group 1 — Group 17 — Group 18 Non-reactive



Periodic Table Patterns

27.

- Symbol represents nucleus and outer levels of electrons.
- Dots around symbol represent valence electrons.



Electron Dot Structure

28.

Cation:

- Positive ion
- Metals form cations by losing electrons.

Ca^{+2} lost 2 electrons

Anion

- Negative ion
- Nonmetals form anions by gaining electrons.

N^{-3} gained 3 electrons

Ions

29.

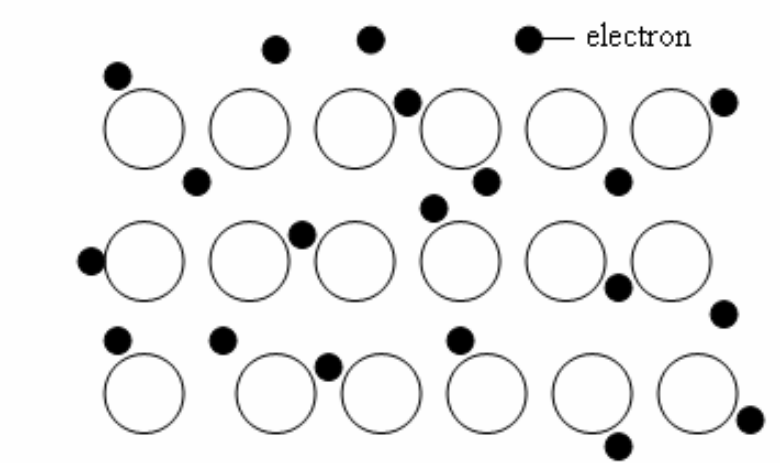
Bonds are an attraction between atoms.
There are 3 types of bonds.

1. **Metallic Bond**
2. **Covalent Bond**
3. **Ionic Bond**

Bonds

30.

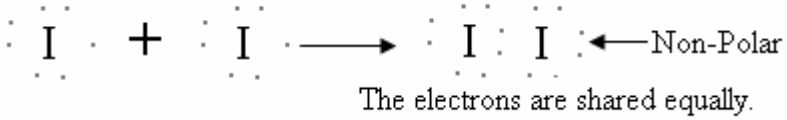
Occurs between atoms of a metal.
Electrons move freely between atoms (sea of electrons)
This is why metals are soft, malleable (bendable), ductile, (made into wires), and good conductors



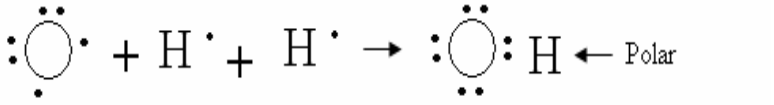
Metallic Bonds

31.

- Forms when atoms of different elements share electrons.



The electrons are shared equally.



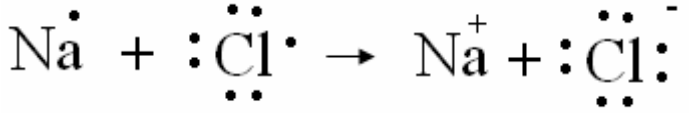
The electrons are not shared equally. The electrons are held closer to the oxygen than to the hydrogen.

- A bond between **non-metals** and **non metals**.
- It is a weak bond.

Covalent Bonds

32.

- Forms when an atom of one element transfers an electron to another element.



- Electrons from Na^+ Transferred to Cl^-.
- Forms an electrostatic bond – strong bond

- Is made of a **metal** and a **non-metal** or a **metal** and a **polyatomic ion**.

Ionic Bond

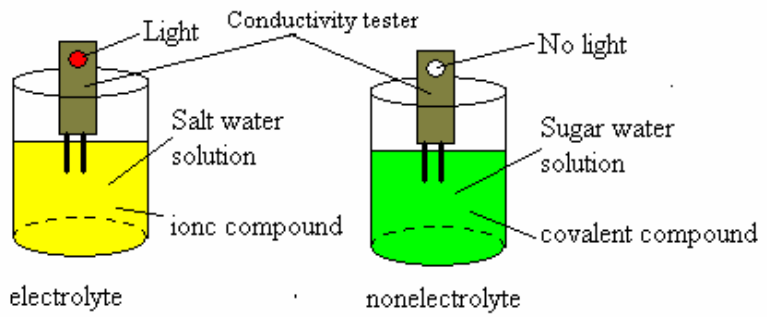
33.

Ionic	Covalent
• Exists as a formula unit • Contain positive and negative ions (e^- are transferred) • Usually solid • High melting and boiling point • Strong force of attraction between ions • Separates in water to form ions • Electrolytes (when dissolved in water or when molten) • Metal and non metal (NaCl)	• Exists as neutral molecules • Atoms share electrons • Can be a solid, liquid, or gas • Has low melting points • Weak force of attraction between molecules • Remains a molecule in water • Non-electrolyte • Made of non-metals (example: $C_6H_{12}O_6$)

Types of Compounds

34.

A compound that conducts electricity when dissolved in water or when molten.



Electrolyte

35.

Ionic	Covalent
○ Write formula of metal ion ○ Write formula of nonmetal ion (or polyatomic ion) ○ Balance the electrical charges to zero using subscripts ○ Example: sodium oxide Na^{+1} O^{-2} Na^{+1} Na^{+1} O^{-2} +2 and –2 = 0 Na_2O **or** Na^{+1} O^{-2} = Na_2O	Write the formula for the first non-metal Write the formula for the second non-metal Add the subscripts behind each formula that matches the prefix in front of the non-metal name. On the first non-metal mono or no prefix at all both mean one. Example: Carbon dioxide CO_2 mono- 1 hexa - 6 di – 2 hepta - 7 tri -3 octa - 8 tetra -4 nona - 9 penta – 5 deca - 10 Subscripts are used when writing formulas

Ionic vs. Covalent

36.

Ionic	Covalent
• Name the metal ion • Name the non-metal ion and change the ending to –ide • Example: $BaCl_2$ is named Barium Chloride	• Name the non-metal • Write the correct prefix in front of the first non-metal's name • If there is only on atom of the first non-metal the prefix mono is dropped • Write the second non-metal's name • Write the correct prefix in front of the second non-metal and make sure the second non-metal's name ends in ide • Example: • NO_2 is named nitrogen dioxide

Naming Binary Compounds

37.

All Ternary compounds

- All ternary compounds contain at least one group (polyatomic ion)
- **Name the metal ion** or the ammonia ion (NH_4^+)
- Second name the group or nonmetal ion and change the ending to –ide
- Almost all negative groups end in **ate** or **ite** (exceptions: hydroxide or cyanide)
- Example:
  - NH_4Cl = Ammonia Chloride
  - Na_2SO_4 = sodium sulfate

When to use Roman numerals

- Roman Numerals are used with metal ions that can form more than one ion.
- Group I and II metals never use Roman Numerals
- Most other metals do use Roman Numerals except for Aluminum, Zinc, and Silver

Examples: Chromium III Chloride; Copper I Sulfate

Naming Ternary Compounds

38.

Reactants are the starting materials. $C_6H_{12}O_6 + 6O_2$

Products are the ending materials. $6CO_2 + 6H_2O$

Respiration — Coefficients

$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$

sugar + oxygen reactants — Carbon dioxide + water product

Coefficients are used to balance the number atoms in a chemical equation.

Parts of a Chemical Reactions

39.

Mass of reactants = Mass of products

Photosynthesis

$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

180 g + 192 g = 264 g + 108 g

372 g = 372 g

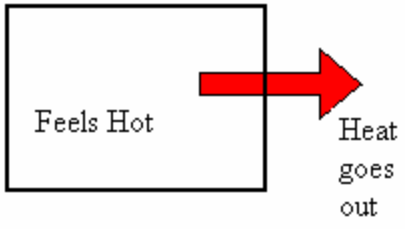
The number and kind of reactant atoms	=	The number and kind of product atoms
6 carbon		6 carbon
12 + 6 = 18 oxygen		6 + 12 = 18 oxygen
12 hydrogen		12 hydrogen

Law of Conservation of Mass

40.

$Fe_2O_{3(s)} + 2Al_{(s)} \rightarrow Al_2O_{3(s)} + 2\,Fe_{(l)} + \text{heat}$

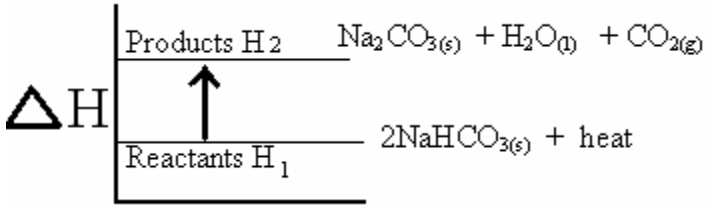
ΔH: Reactants H_1 $Fe_2O_{3(s)} + 2Al_{(s)}$ ↓ Products H_2 $Al_2O_{3(s)} + 2\,Fe_{(l)} + \text{heat}$

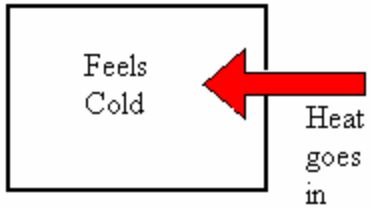


Exothermic
Exogonic

41.

$2NaHCO_{3(s)} + \text{heat} \rightarrow Na_2CO_{3(s)} + H_2O_{(l)} + 2CO_{2(g)}$





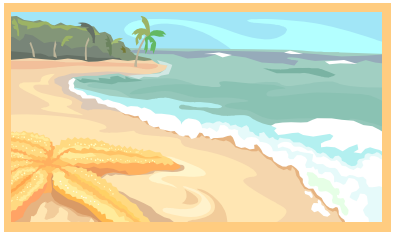
Endothermic
Endogonic

42.

Specific heat – the amount of heat necessary to change 1 gram of a substance by 1 ^{O}C.

Water has one of the highest specific heats. It takes a long time to heat up and it is slow to cool down.

Air will expand when it is warm and this causes the wind to blow from the warmer area to the cooler area.



During the day the land side is warmer than the water side. This causes the wind to blow from the sea to the land.

At night the water is warmer than the land so the wind blows from the land to the sea.

Specific Heat

43.

Solute is the part that gets dissolved.

Solvent is the part that does the dissolving.

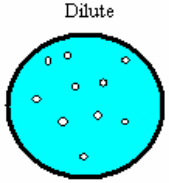
Alloys – mixtures of metals • Steel – iron & carbon • Brass – copper & zinc • 18 ct gold – gold & copper or silver • Sterling silver – silver & copper	**Air** is a mixture of gases. It is made up of: • Oxygen • Nitrogen • Hydrogen • Carbon dioxide • Water vapor	**Solutions** can be: • Liquids & liquids • Liquids & solid • Liquids & gas • Solid & solid • Gas and gas

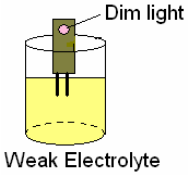
Parts and Types of Solutions

44.

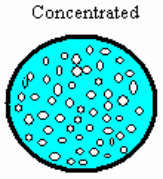
Ratio of amount of solute versus solvent; ratio can vary

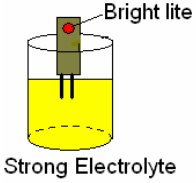
Dilute – small amount of solute versus a large amount of solvent.





Concentrated – larger amount of solute versus amount of solvent.



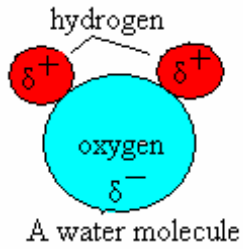


Concentration of Solution

45.

Water is called the **universal solvent**. It is Earths most important solvent for living organisms.
Water can dissolve both polar covalent and ionic compounds.

Water is a **polar molecule**. (It has a slightly positive charged end and a slightly negative charged end)



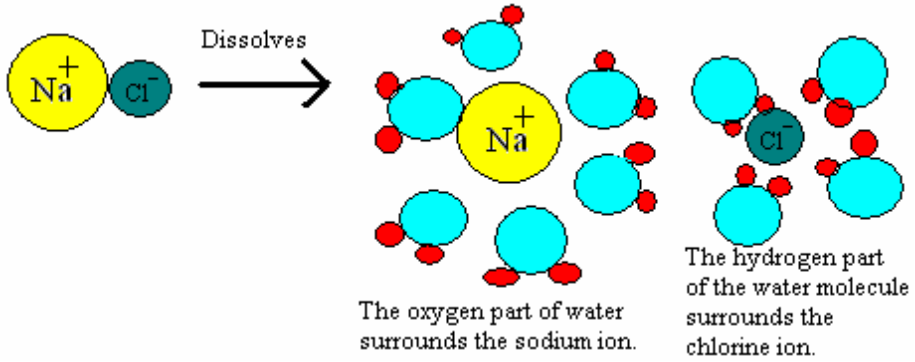
A water molecule

Note there is a slight electric charge on each end of the molecule.

Importance of Water

46.

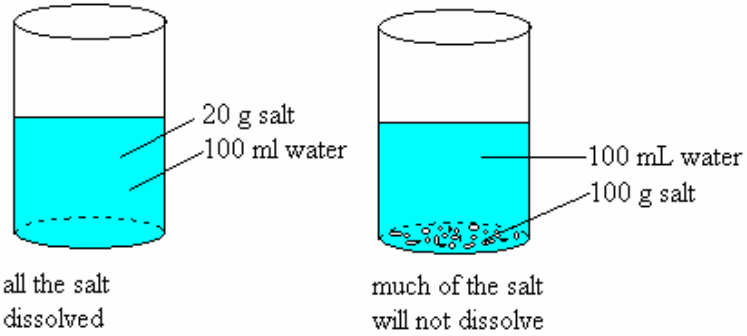
Water surrounds and causes the bond in an ionic compound to break.



Dissociation

47.

The **amount** of solute that will dissolve in a given amount of solvent at a specified temperature and pressure.



Units: <u>grams of solute</u>
100 Grams of solution

Effect of Nature of Solvent and Solute on Solubility

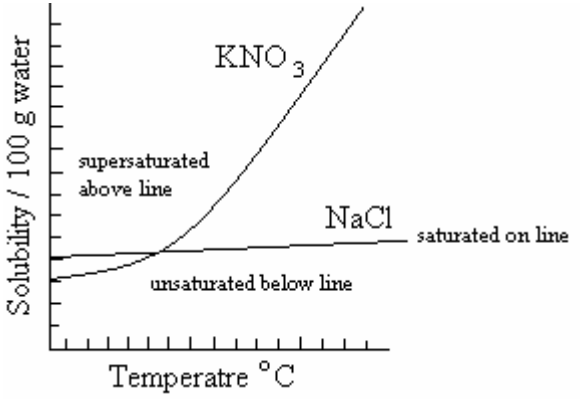
48.

1. **Unsaturated** (more will dissolve)
2. **Saturated** (no more will dissolve)
3. **Supersaturated** (unstable because it has more solute than normally dissolves in the solvent at a specified temperature)

Classification of Solutions

49.

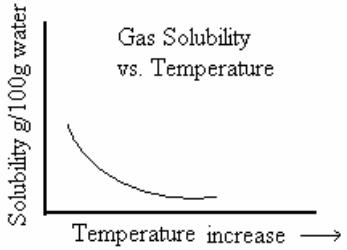
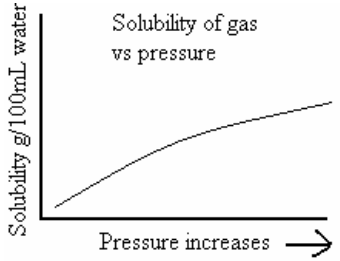
The solubility of a solid is affected by both temperature and the nature of solute and solvent.



Solubility of a Solid

50.

The solubility of a gas is affected by both temperature and pressure.

Increase in temperature causes less gas to dissolve.	Increase in pressure causes more gas to dissolve.
	

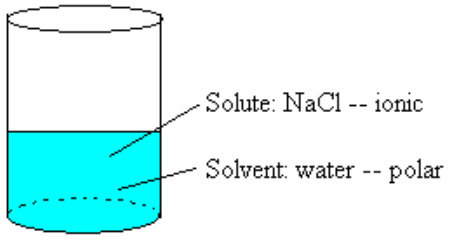
Solubility of a Gas

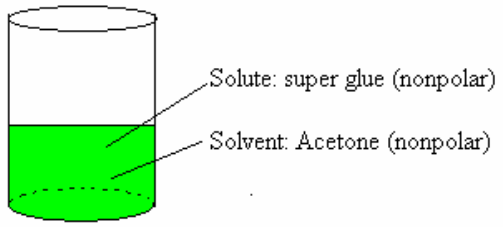
51.

Likes dissolve likes.

Polar solvents dissolve
- polar solutes
- ionic solutes

Nonpolar solvents dissolve only nonpolar solutes





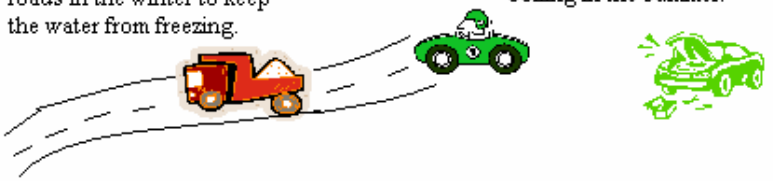
Effect of nature of solvent and Solute on Solubility

52.

Solutes dissolved in solvents can raise the boiling point of the solvent and reduce the freezing point of the solvent.

Example: Salt is put on roads in the winter to keep the water from freezing.

Example: Antifreeze is added to the water in a car to keep the water from freezing in the winter and to keep the water from boiling in the summer.



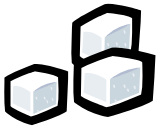
Effect of Solute on Solvents

53.

1. **Stirring**

2. **Temperature of solvent**

3. **Surface area of solute**
   (size of particles)

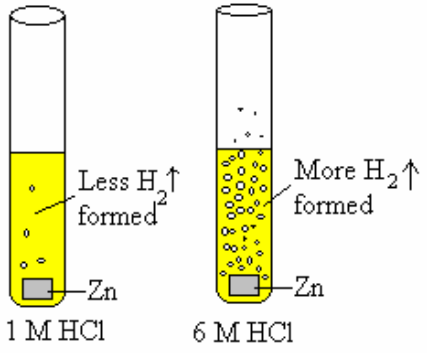


Factors Affecting Rate of Solution
(How fast something dissolves)

54.

Acids

- Taste sour
- Turn litmus red
- Neutralizes base to form salt and water
- Reacts with active metals to release H_2 gas
- $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

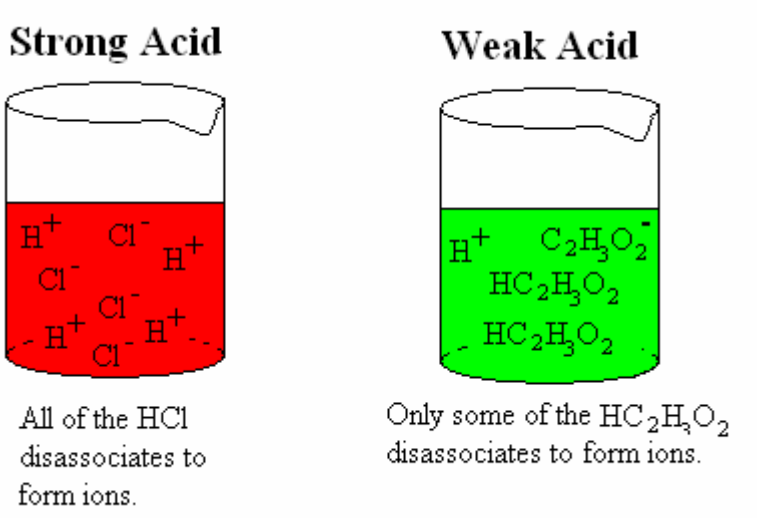


More concentrated

Properties of Acids

55.

A compound that produces hydrogen ions (H^+) in a water solution.



Acid

56.

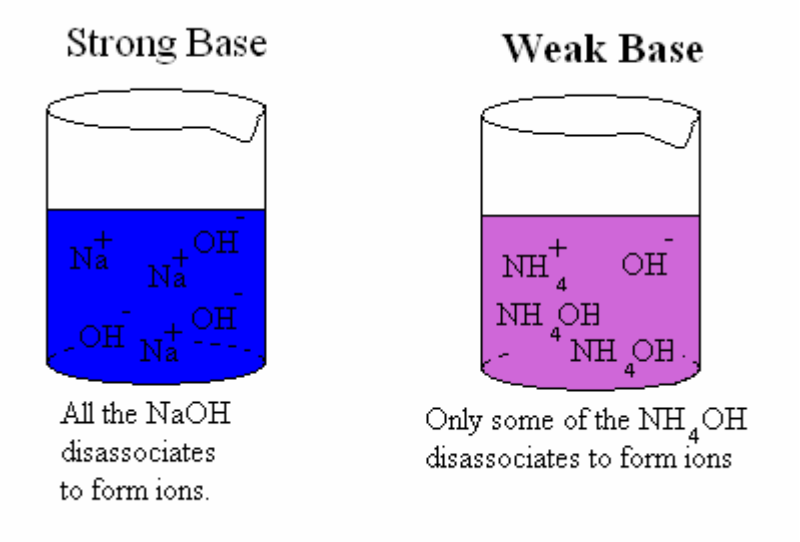
Bases:

- **Taste Bitter**
- **Turn litmus Blue**
- **Neutralizes acids** to form salt and water
  $NaOH + HCl \rightarrow H_2O + NaCl$
  Base Acid Water Salt

Properties of Bases

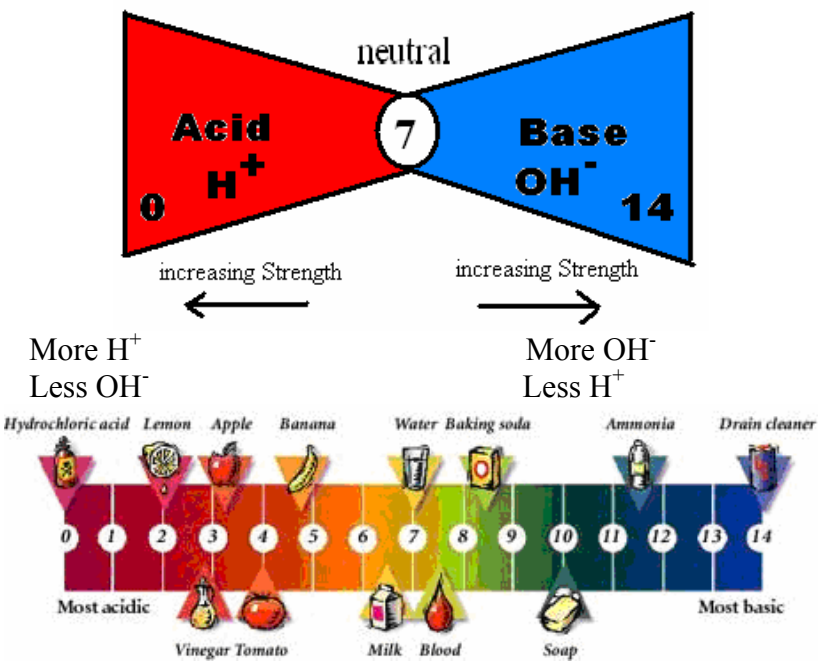
57.

A substance when dissolved in water produces **OH^-** (hydroxide ions).



Base

58.

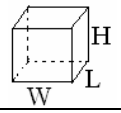


pH Scale

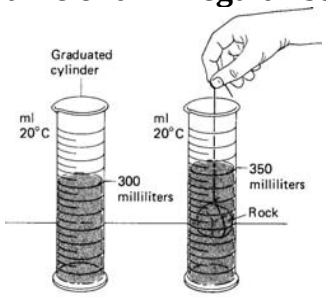
59.

Volume of regular solid

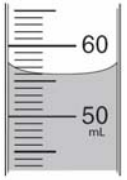
V= L x W x H



Volume of an irregular solid



Volume of a liquid



Volume