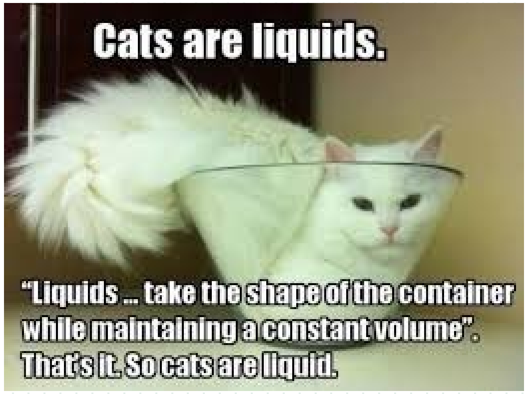
Chapter 3

Everyday Chemistry of Life

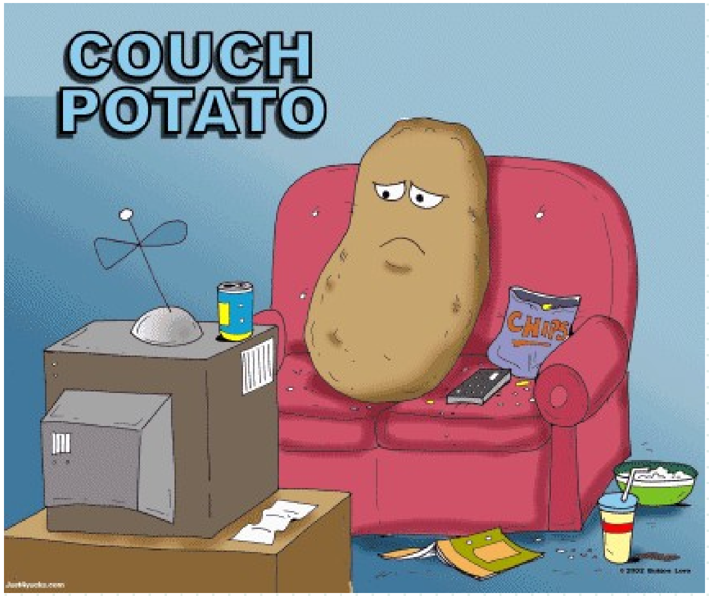
Matter

- **Matter**—anything that has mass and occupies space
  - Weight—pull of gravity on mass
- 3 states of matter
  - Solid—definite shape and volume
  - Liquid—changeable shape; definite volume
  - Gas—changeable shape and volume



Energy

- Capacity to do work or put matter into motion
- Types of energy
  - **Kinetic**—energy in action
  - **Potential**—stored (inactive) energy
- Energy can be transferred from potential to kinetic energy



Forms of Energy

- Chemical energy
  - Stored in bonds of chemical substances
- Electrical energy
  - Results from movement of charged particles
- Mechanical energy
  - Directly involved in moving matter
- Radiant or electromagnetic energy
  - Travels in waves (e.g., visible light, ultraviolet light, and x-rays)

Major Elements of the Human Body

- Four elements make up 96.1% of body mass

Element	Atomic symbol
Carbon	C
Hydrogen	H
Oxygen	O
Nitrogen	N

Lesser Elements of the Human Body

9 elements make up 3.9% of body mass

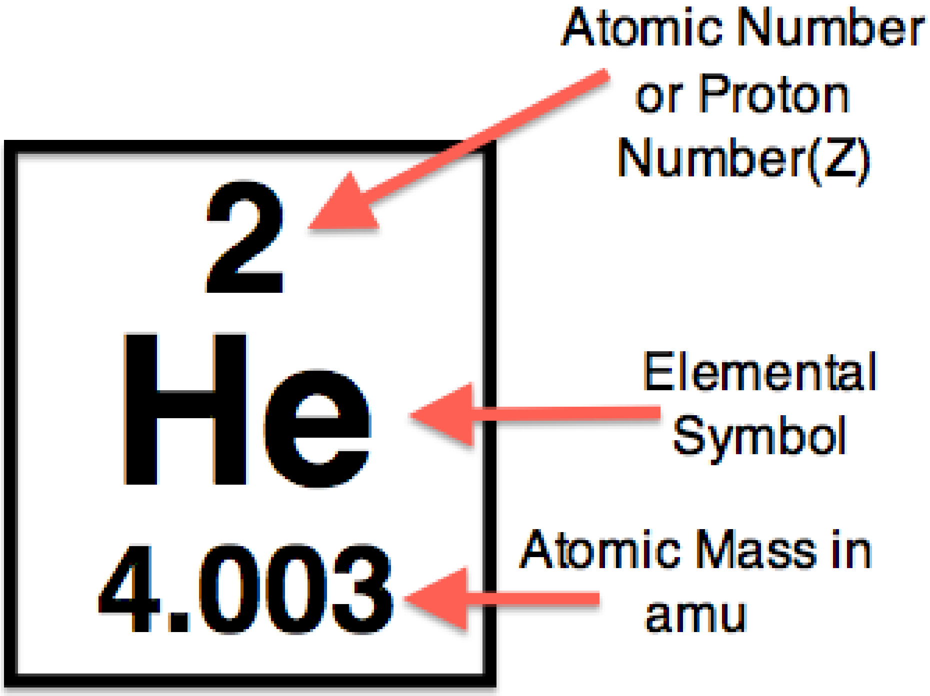
Element	Atomic symbol
Calcium	Ca
Phosphorus	P
Potassium	K
Sulfur	S
Sodium	Na
Chlorine	Cl
Magnesium	Mg
Iodine	I
Iron	Fe

Composition of Matter

- Atoms
  - Unique building blocks for each **element**
  - Give each element its physical & chemical properties
  - Smallest particles of an element with properties of that element

- Atomic symbol
  - One- or two-letter che

shorthand for each elem



Composition of Matter: Elements

- Elements **are made of atoms**
  - Each atom is the smallest possible unit of a particular element
  - Elements cannot be broken into simpler substances by ordinary chemical methods
  - Each has unique properties
    - Physical properties
      - Detectable with our senses, or are measurable
    - Chemical properties
      - How atoms interact (bond) with one another

Atomic Structure

Atoms are composed of **subatomic particles**

- **Protons, neutrons, electrons**

- Protons and neutrons found in nucleus
- Electrons orbit nucleus in an electron cloud

Neutrons

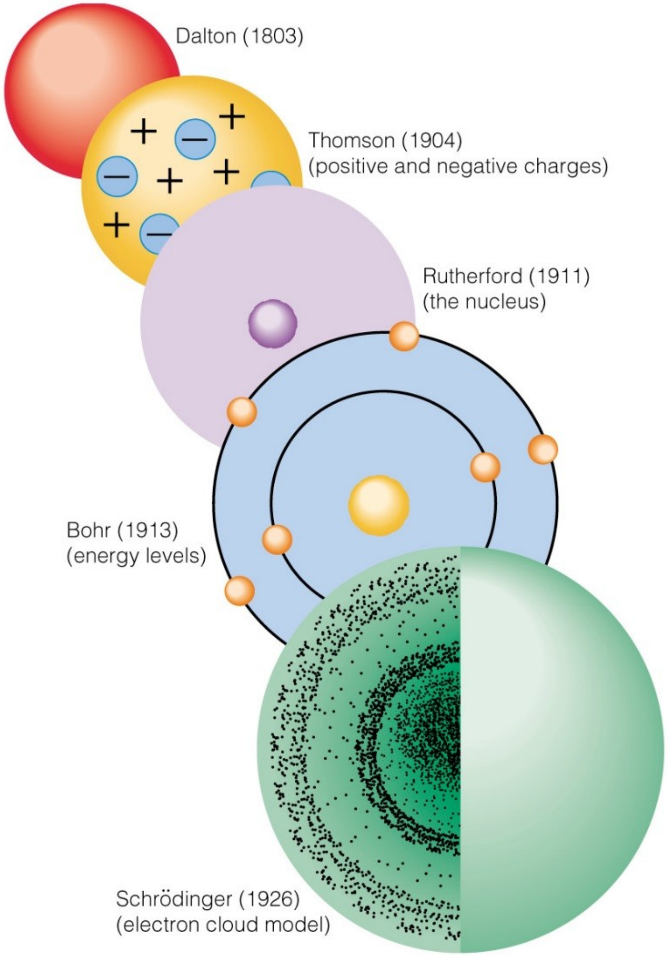
- Carry no charge

Protons

- Carry positive charge

Electrons in orbitals within electron cloud

- Carry negative charge
- Number of protons and electrons always equal



Periodic Table of the Elements

Oxford LABS

atomic number
atomic weight

14 28.09
Si
Silicon

name

symbol:

black	solid
blue	liquid
red	gas
white	synthetically prepared
	most stable isotope

⬜	alkali metals
🟩	alkaline earth metals
🟨	transitional metals
🟪	other metals
🟦	nonmetals
🟧	noble gases

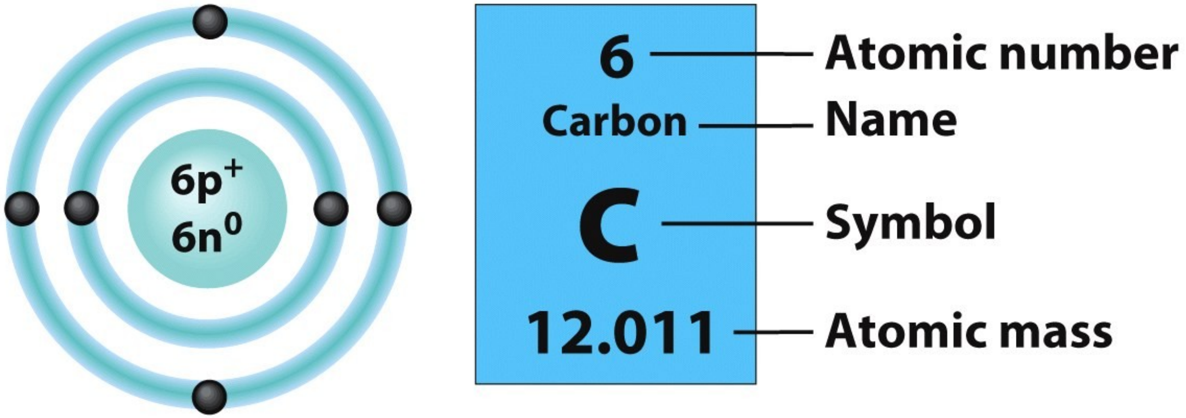
1 1.01 H Hydrogen																	2 4.003 He Helium
3 6.94 Li Lithium	4 9.01 Be Beryllium											5 10.81 B Boron	6 12.01 C Carbon	7 14.01 N Nitrogen	8 15.999 O Oxygen	9 18.998 F Fluorine	10 20.18 Ne Neon
11 22.99 Na Sodium	12 24.31 Mg Magnesium											13 26.98 Al Aluminum	14 28.09 Si Silicon	15 30.97 P Phosphorus	16 32.06 S Sulfur	17 35.45 Cl Chlorine	18 39.95 Ar Argon
19 39.10 K Potassium	20 40.08 Ca Calcium	21 44.96 Sc Scandium	22 47.90 Ti Titanium	23 50.94 V Vanadium	24 51.996 Cr Chromium	25 54.94 Mn Manganese	26 55.85 Fe Iron	27 58.93 Co Cobalt	28 58.70 Ni Nickel	29 63.55 Cu Copper	30 65.37 Zn Zinc	31 69.72 Ga Gallium	32 72.59 Ge Germanium	33 74.92 As Arsenic	34 78.96 Se Selenium	35 79.90 Br Bromine	36 83.80 Kr Krypton
37 85.47 Rb Rubidium	38 87.62 Sr Strontium	39 88.91 Y Yttrium	40 91.22 Zr Zirconium	41 92.91 Nb Niobium	42 95.94 Mo Molybdenum	43 (98) Tc Technetium	44 101.07 Ru Ruthenium	45 102.91 Rh Rhodium	46 106.40 Pd Palladium	47 107.87 Ag Silver	48 112.41 Cd Cadmium	49 114.82 In Indium	50 118.69 Sn Tin	51 121.75 Sb Antimony	52 127.60 Te Tellurium	53 126.90 I Iodine	54 131.30 Xe Xenon
55 132.91 Cs Cesium	56 137.33 Ba Barium	57 138.91 La ▸ Lanthanum	72 178.49 Hf Hafnium	73 180.95 Ta Tantalum	74 183.85 W Tungsten	75 186.21 Re Rhenium	76 190.20 Os Osmium	77 192.22 Ir Iridium	78 195.09 Pt Platinum	79 196.97 Au Gold	80 200.59 Hg Mercury	81 204.37 Tl Thallium	82 207.19 Pb Lead	83 208.98 Bi Bismuth	84 (209) Po Polonium	85 (210) At Astatine	86 (222) Rn Radon
87 (223) Fr Francium	88 226.03 Ra Radium	89 227.03 Ac ▸ Actinium	104 (261) Rf Rutherfordium	105 (262) Ha Hahnium	106 (266) Sg Seaborgium	107 (262) Bh Bohrium	108 (265) Hs Hassium	109 (266) Mt Meitnerium	110 (271)	111 (272)	112 (277)	(113)	114 (285)	(115)	116 (289)	(117)	118 (293)

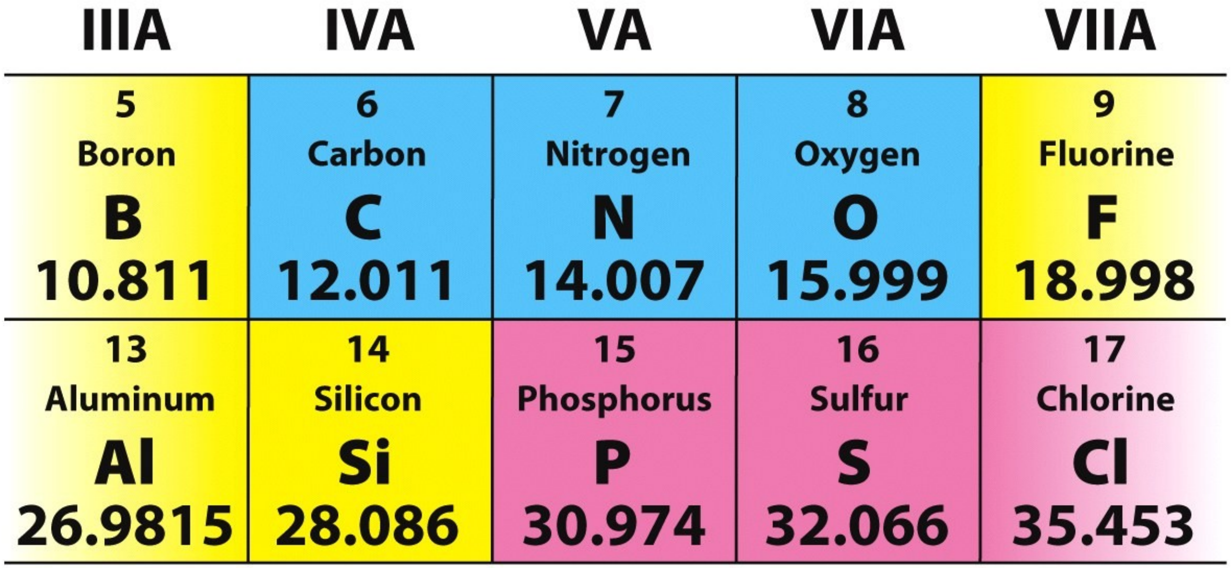
58 140.12 Ce Cerium	59 140.91 Pr Praseodymium	60 144.24 Nd Neodymium	61 (145) Pm Promethium	62 150.40 Sm Samarium	63 151.96 Eu Europium	64 157.25 Gd Gadolinium	65 158.93 Tb Terbium	66 162.50 Dy Dysprosium	67 164.93 Ho Holmium	68 167.26 Er Erbium	69 168.93 Tm Thulium	70 173.04 Yb Ytterbium	71 174.97 Lu Lutetium
90 232.04 Th Thorium	91 231.04 Pa Protactinium	92 238.03 U Uranium	93 237.05 Np Neptunium	94 (244) Pu Plutonium	95 (243) Am Americium	96 (247) Cm Curium	97 (247) Bk Berkelium	98 (251) Cf Californium	99 (252) Es Einsteinium	100 (257) Fm Fermium	101 (260) Md Mendelevium	102 (259) No Nobelium	103 (262) Lr Lawrencium

Identifying Elements:
Atomic Number and Mass Number

- **Atomic number** = Number of protons in nucleus
  - Written as subscript to left of atomic symbol
    - Ex. $_{3}Li$
- **Mass number**
  - Total number of protons and neutrons in nucleus
    - Total mass of atom
  - Written as superscript to left of atomic symbol
    - Ex. ^{7}Li

The Periodic Table

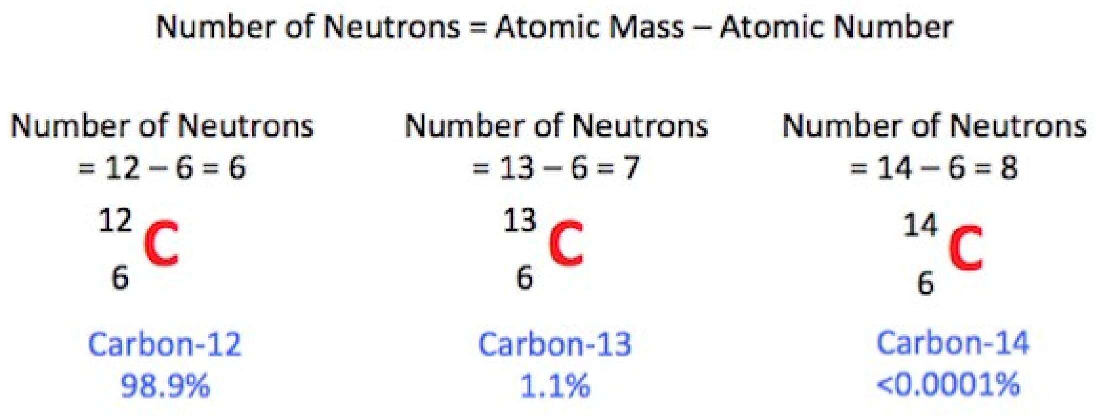




IIIA	IVA	VA	VIA	VIIA
5 Boron B 10.811	6 Carbon C 12.011	7 Nitrogen N 14.007	8 Oxygen O 15.999	9 Fluorine F 18.998
13 Aluminum Al 26.9815	14 Silicon Si 28.086	15 Phosphorus P 30.974	16 Sulfur S 32.066	17 Chlorine Cl 35.453

Identifying Elements: Isotopes and Atomic Weight

- **Isotopes**
  - Structural variations of atoms
  - Differ in the number of neutrons they contain
  - Atomic numbers same; mass numbers different
- **Atomic weight**
  - Average of mass numbers (relative weights) of all isotopes of an atom



Isotopes of Carbon

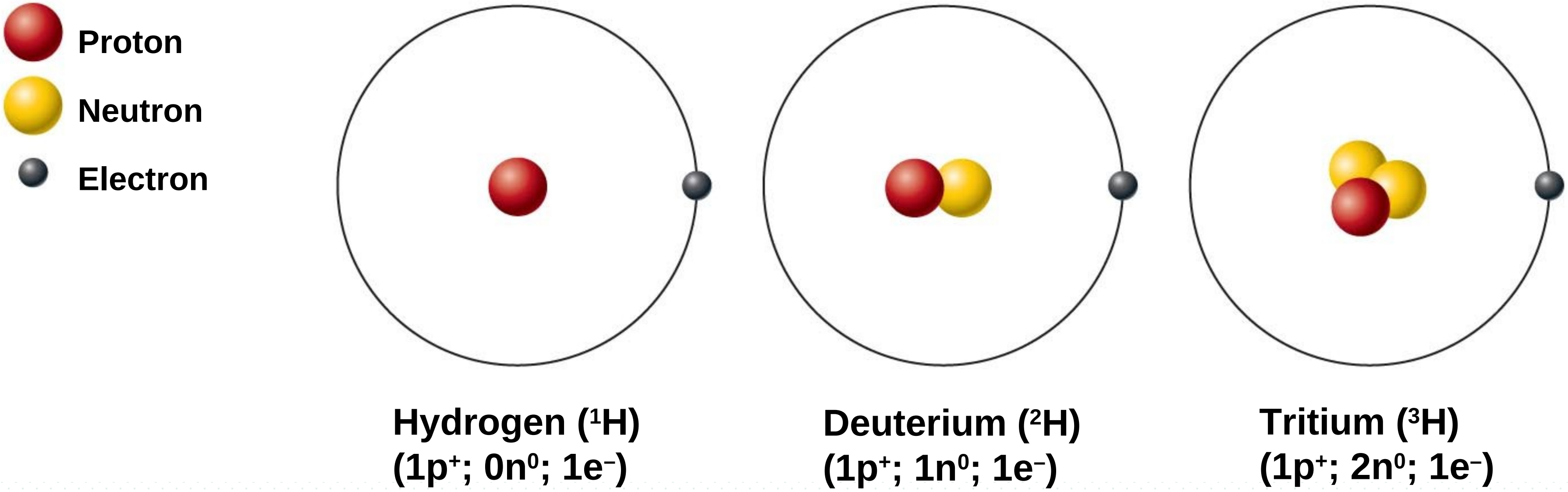
Radioisotopes

- Heavy isotopes decompose to more stable forms
  - Spontaneous decay called radioactivity
  - Similar to tiny explosion
  - Can transform to different element
- Can be detected with scanners

Radioisotopes

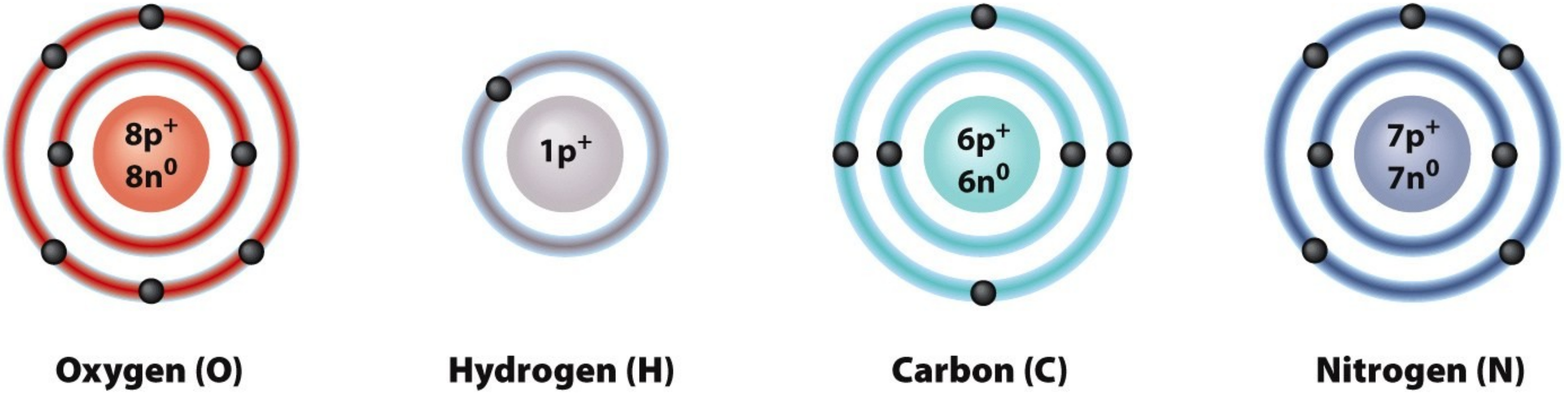
- Valuable tools for biological research and medicine
  - Share same chemistry as their stable isotopes
  - Most used for diagnosis
- All damage living tissue
  - Some used to destroy localized cancers
  - Radon from uranium decay causes lung cancer

Figure 2.3 Isotopes of hydrogen.



Orbitals

- **Orbitals are clouds of electrons that surround an atom’s nucleus**
  - Different orbitals have different levels of energy due to the electrons they contain
  - The outermost energy level is the valence shell
  - The electrons in the valence shell are the valence electrons



Chemical Bonds

- **Chemical bonds** are energy relationships between electrons of reacting atoms
- Electrons can occupy up to seven electron shells (energy levels) around nucleus
- Electrons in **valence shell** (outermost electron shell)
  - Have most potential energy
  - Are chemically reactive electrons
- Octet rule (rule of eights)
  - Except for the first shell (two electrons) atoms interact to have eight electrons in their valence shell

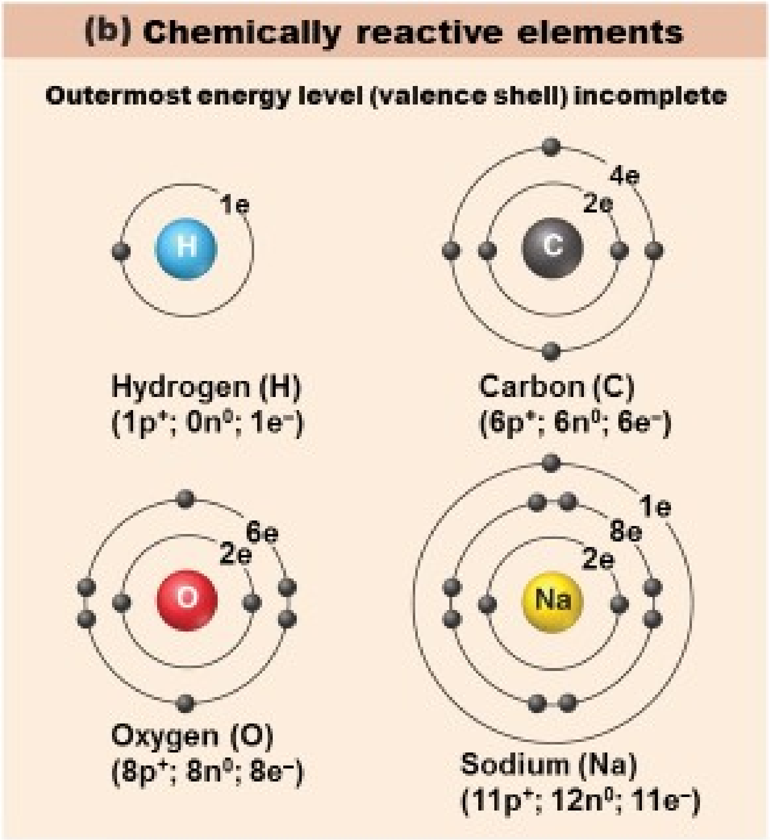
Combining Matter: Molecules and Compounds

- Most atoms chemically combined with other atoms to form molecules and compounds
  - **Molecule**
    - Two or more atoms bonded together (e.g., H_2 or $C_6H_{12}O_6$)
    - Smallest particle of a compound with specific characteristics of the compound
  - **Compound**
    - Two or more different kinds of atoms bonded together (e.g., $C_6H_{12}O_6$, but not H_2)

Chemically Inert Elements

- Stable and unreactive
- Valence shell fully occupied or contains eight electrons
- Noble gases

Figure 2.5b Chemically inert and reactive elements.



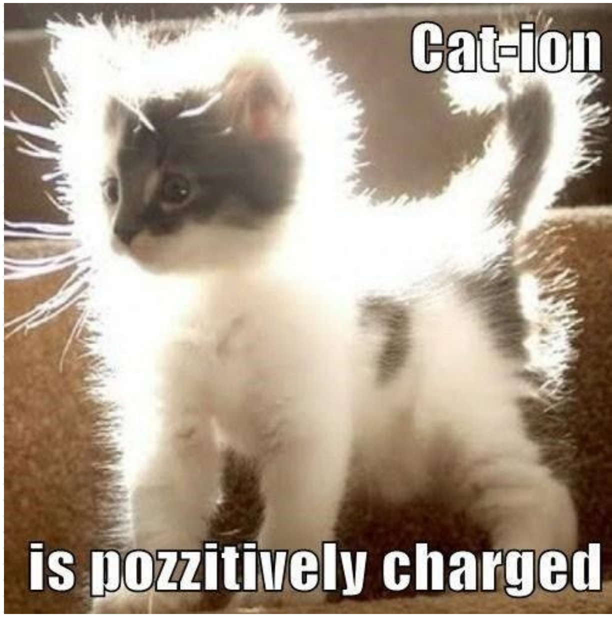
Types of Chemical Bonds

- Chemically reactive elements -- Valence shell not full!
- Tend to gain, lose, or share electrons (form bonds) with other atoms to achieve stability

- Three major types
  - **Ionic bonds**
  - **Covalent bonds**
  - **Hydrogen bonds**

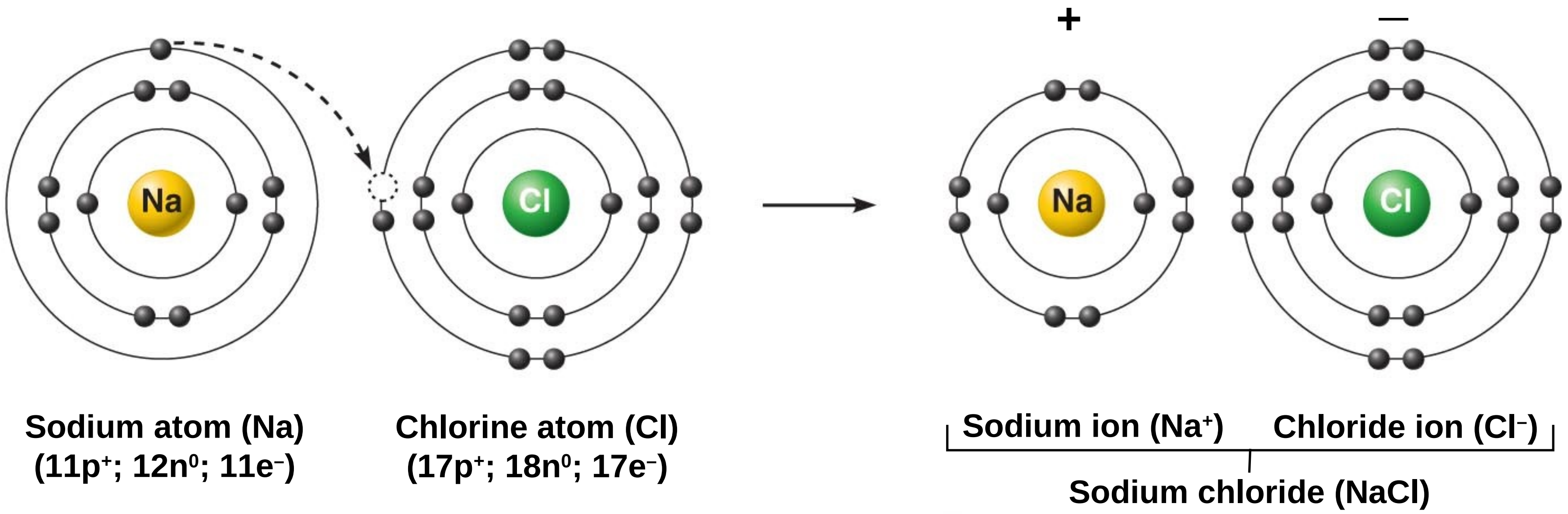
Ionic Bonds

- **Ions**
  - Atom gains or loses electrons and becomes charged
    - # Protons ≠ # Electrons
- *Transfer* of valence shell electrons from one atom to another forms ions
  - One becomes an **anion** (negative charge)
    - Atom that gained one or more electrons
  - One becomes a **cation** (positive charge)
    - Atom that lost one or more ***ELECTRONS***
- Attraction of opposite charges results in an ionic bond



This cat is now a cation because it's missing an electron

Figure 2.6a–b Formation of an ionic bond.



(a) Sodium gains stability by losing one electron, and chlorine becomes stable by gaining one electron.

(b) After electron transfer, the oppositely charged ions formed attract each other.

Ionic Compounds

- Most ionic compounds are salts
  - When dry salts form crystals instead of individual molecules
  - Example is NaCl (sodium chloride)

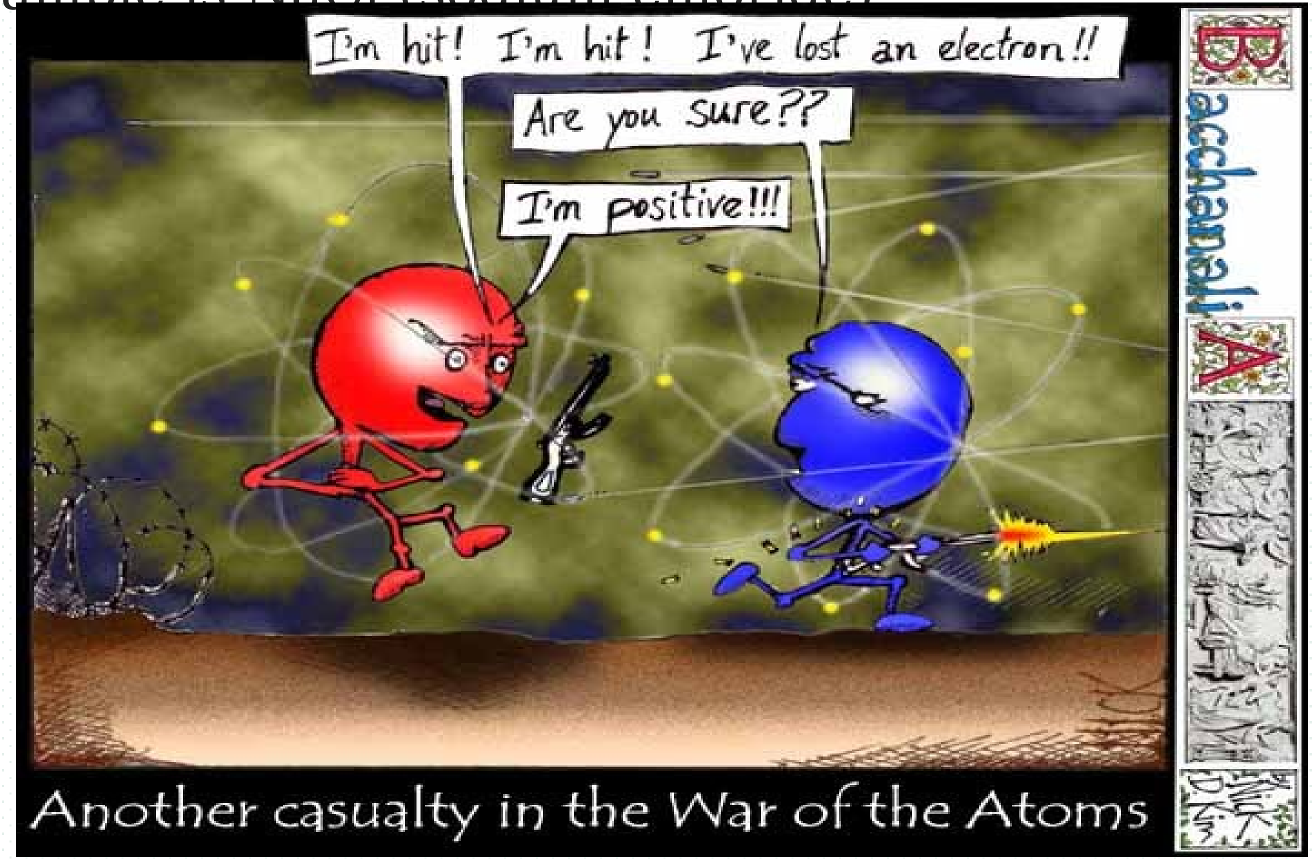
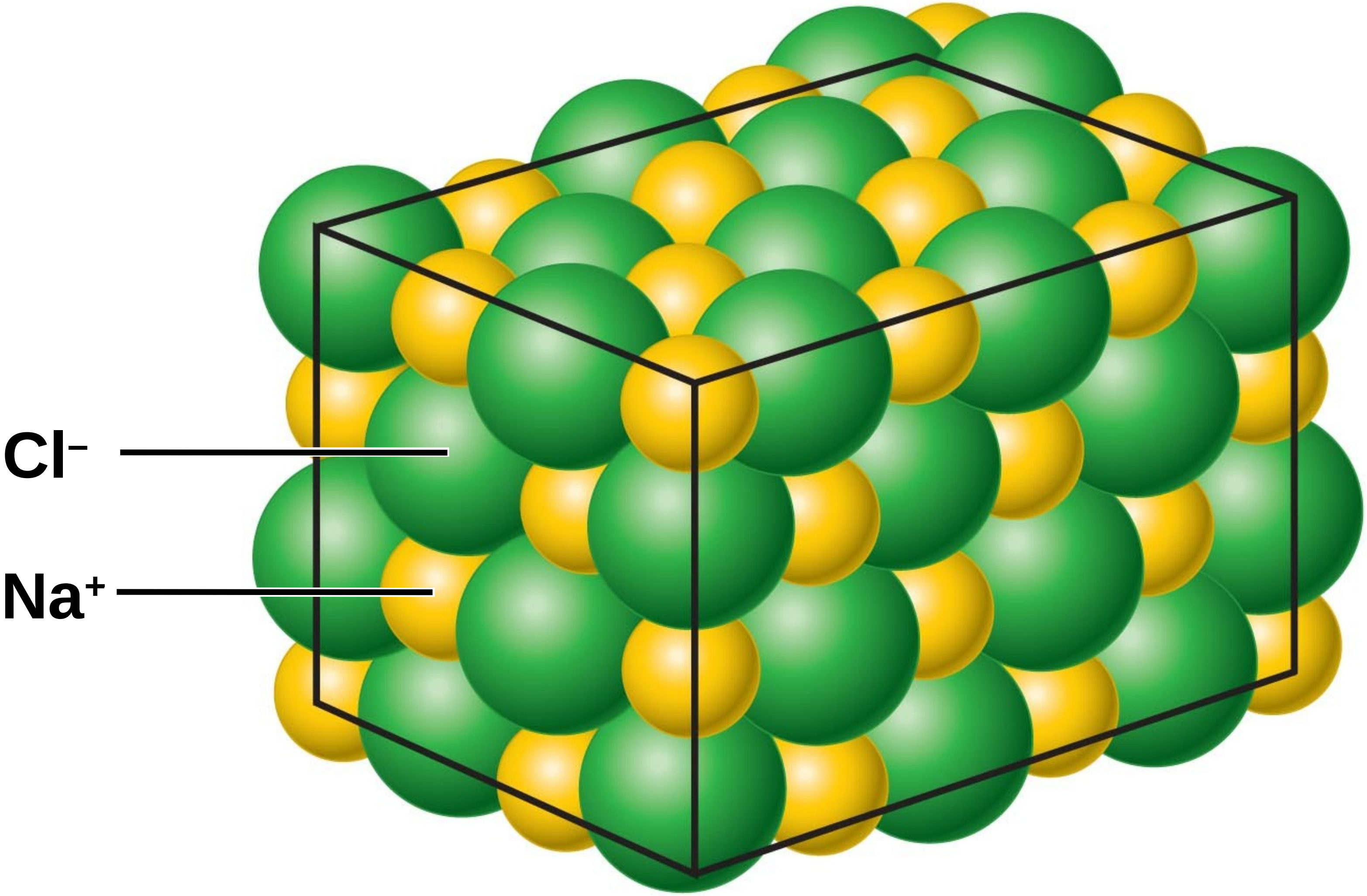


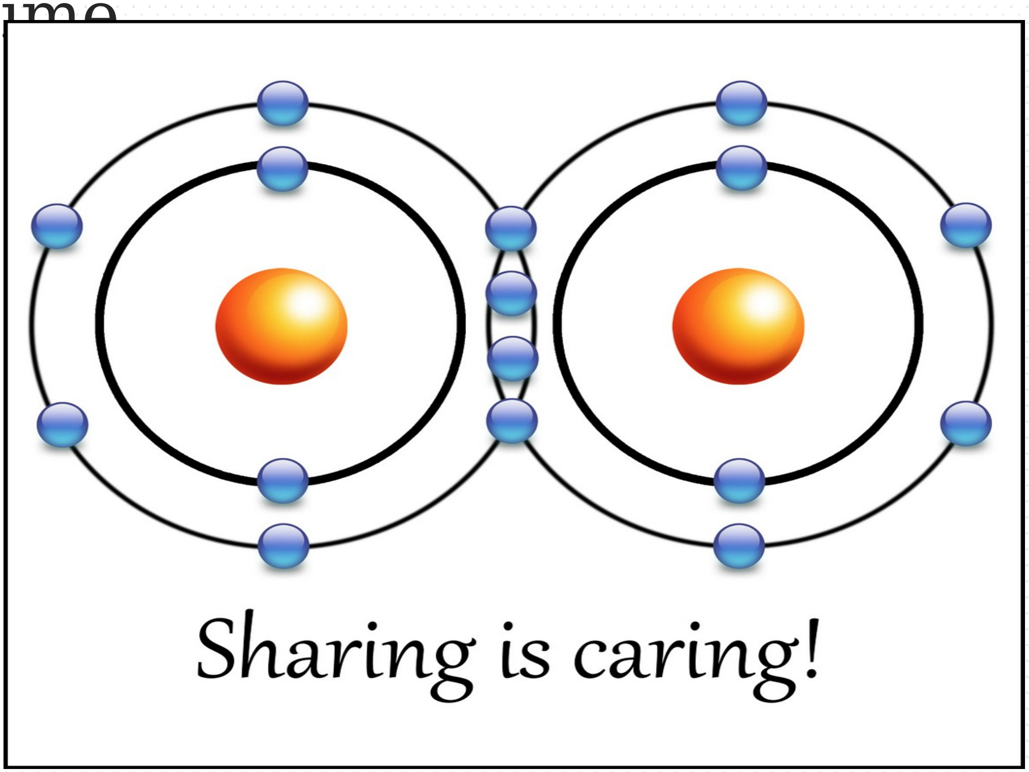
Figure 2.6c Formation of an ionic bond.



(c) Large numbers of Na^+ and Cl^- ions associate to form salt (NaCl) crystals.

Covalent Bonds

- Formed by sharing of two or more valence shell electrons
- Allows each atom to fill its valence shell at least part of the time.



Sharing is caring!

Atoms Share Electrons in Covalent Bonds

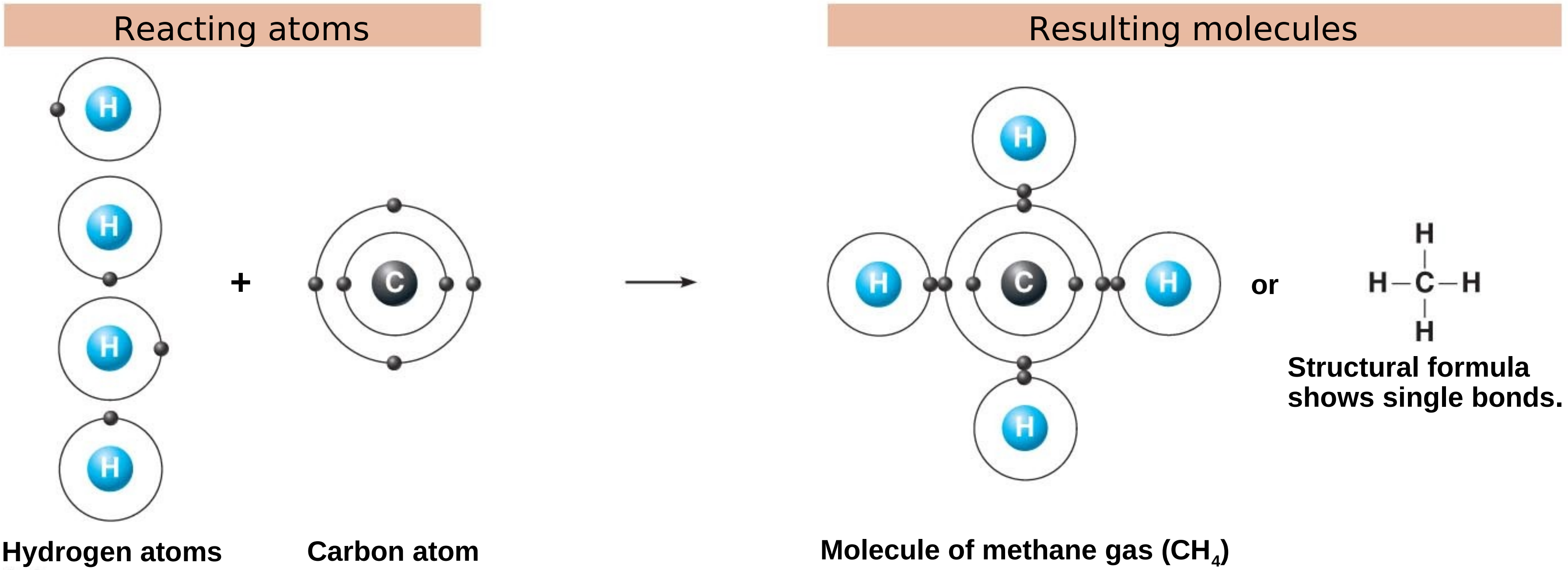
- **In a covalent bond, atoms share valence electrons so that each atom gets to fill its valence shell**
  - bonds **strongly** hold atoms together
  - Covalent bonds commonly involve **carbon, oxygen, nitrogen, or hydrogen atoms**

- **In a single covalent bond – one pair of valence electrons is shared**

- **In a double covalent bond – two pairs of valence electrons are shared**

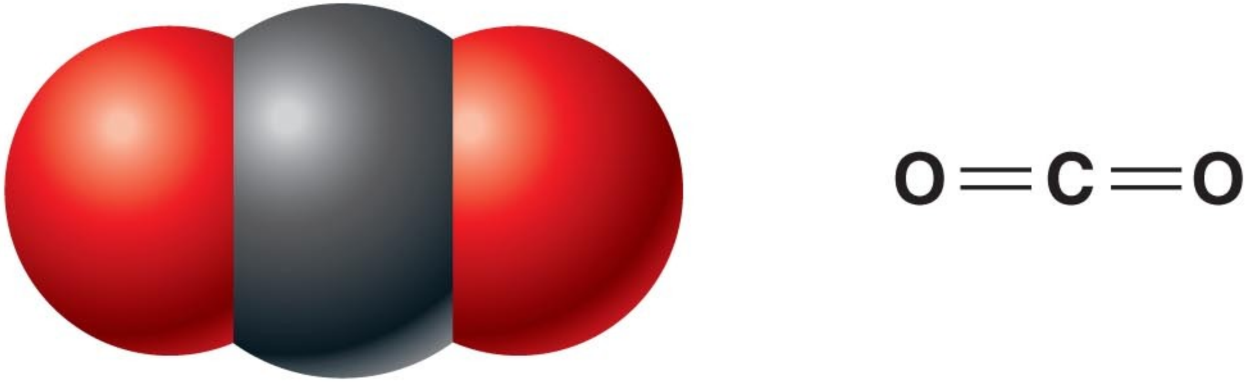
- **In a triple covalent bond – three pairs of valence electrons are shared**

Figure 2.7a Formation of covalent bonds.



(a) Formation of four single covalent bonds: Carbon shares four electron pairs with four hydrogen atoms.

Nonpolar Covalent Bonds



(a) Carbon dioxide (CO_2) molecules are linear and symmetrical. They are nonpolar.

- Electrons shared equally
- Produces electrically balanced, nonpolar molecules such as CO_2

Polar Covalent Bonds

- Unequal sharing of electrons produces polar (AKA **dipole**) molecules such as H_2O
  - Atoms in bond have different electron-attracting abilities
- Small atoms with six or seven valence shell electrons are **electronegative,** e.g., oxygen
  - Strong electron-attracting ability
- Most atoms with one or two valence shell electrons are **electropositive,** e.g., sodium

Hydrogen Bonds are Weak Chemical Bonds
But are Vital To Biology

- **The slight positive charge of a hydrogen atom within a polar covalent bond**
  - Can form weak attractive bonds (hydrogen bonds) with adjacent slightly negative atoms
- **Hydrogen bonds can**
  - Hold water molecules together
  - Join the two strands of DNA
  - Help shape proteins

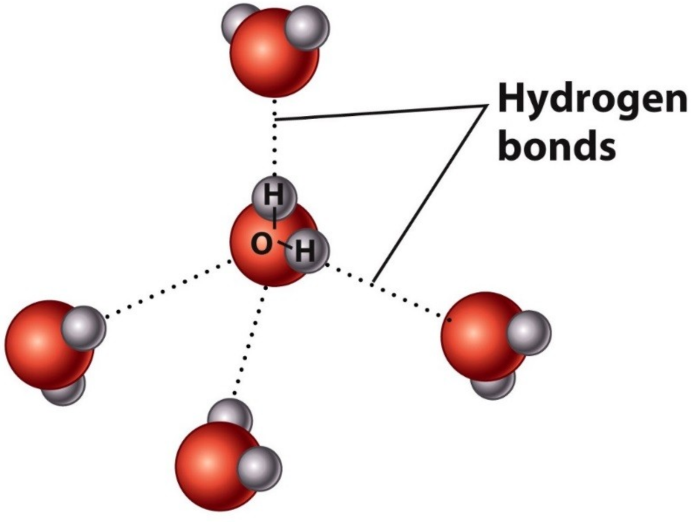
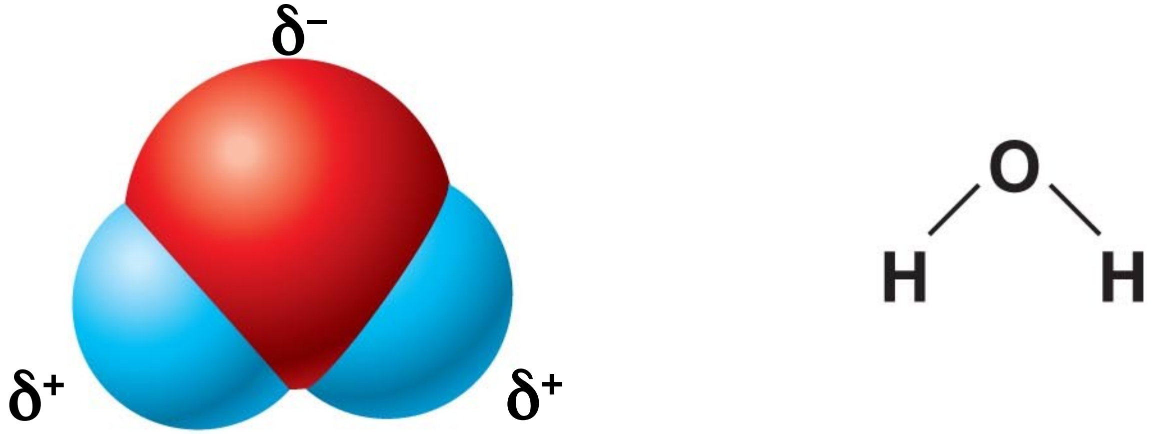
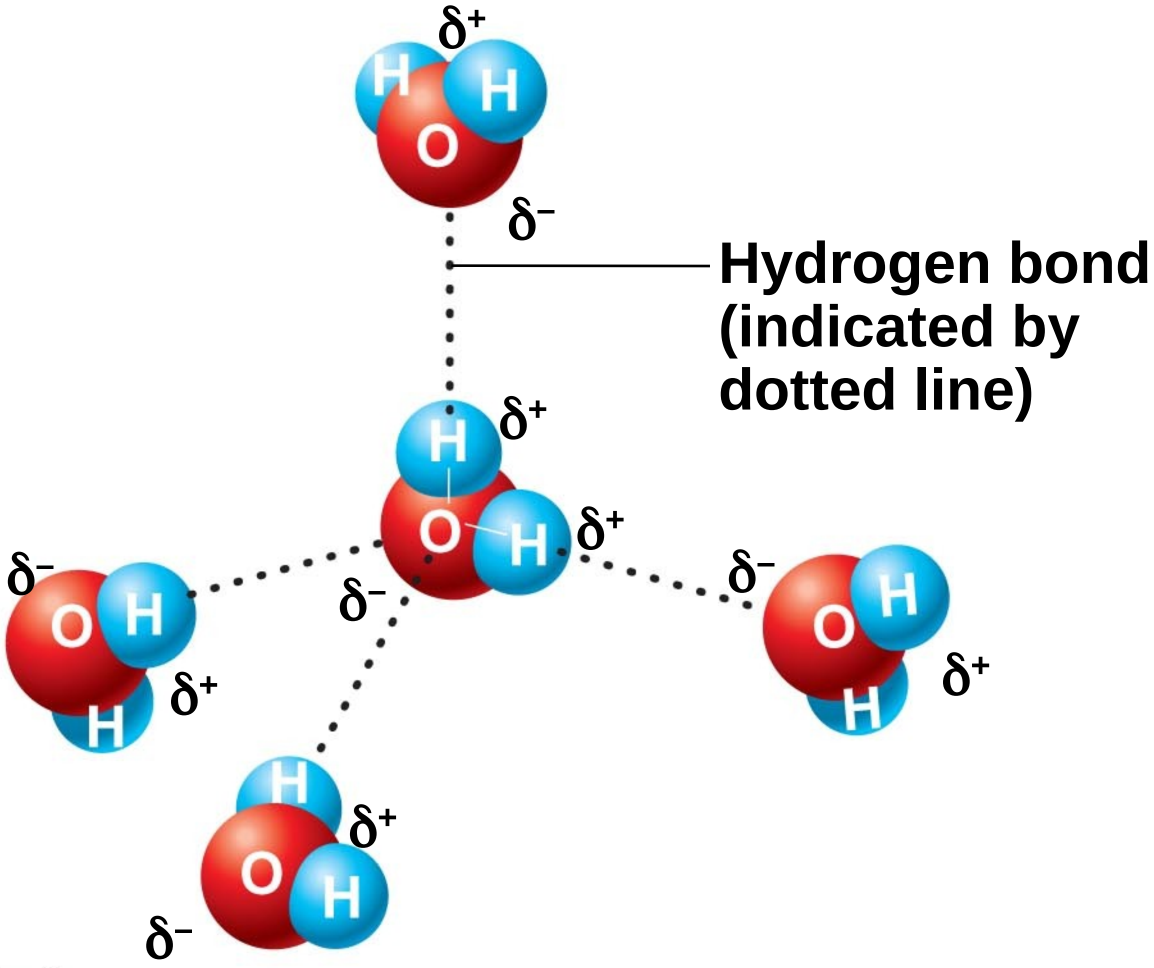


Figure 2.8b Carbon dioxide and water molecules have different shapes, as illustrated by molecular models.



(b) V-shaped water (H_2O) molecules have two poles of charge—a slightly more negative oxygen end (δ^-) and a slightly more positive hydrogen end (δ^+).

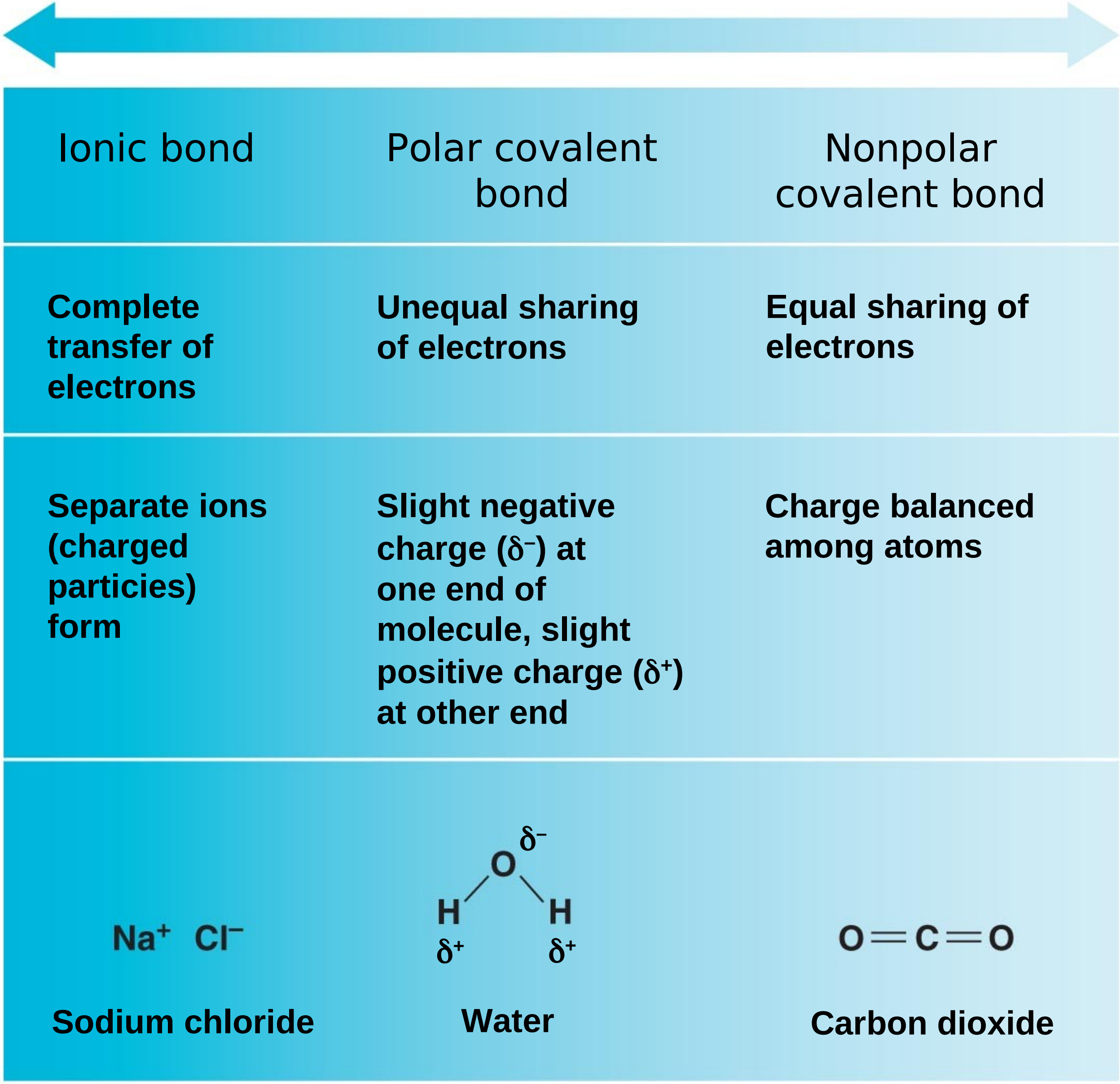
Figure 2.10a Hydrogen bonding between polar water molecules.



Hydrogen bond (indicated by dotted line)

(a) The slightly positive ends (δ^+) of the water molecules become aligned with the slightly negative ends (δ^-) of other water molecules.

Figure 2.9 Ionic, polar covalent, and nonpolar covalent bonds compared along a continuum.



Ionic bond	Polar covalent bond	Nonpolar covalent bond
Complete transfer of electrons	**Unequal sharing of electrons**	**Equal sharing of electrons**
Separate ions (charged particies) form	**Slight negative charge (δ^-) at one end of molecule, slight positive charge (δ^+) at other end**	**Charge balanced among atoms**
Na^+ Cl^- **Sodium chloride**	**Water**	O=C=O **Carbon dioxide**

Chemical Reactions

- Occur when chemical bonds are formed, rearranged, or broken
  - **Reactants**
    - Number and kind of reacting substances
  - Chemical composition of the **product(s)**
  - Relative proportion of each reactant and product in balanced equations

Examples of Chemical Equations

Reactants	Product
$H + H \rightarrow$	H_2 (Hydrogen gas)
$4H + C \rightarrow$	CH_4 (Methane)

Note: CH_4 is a molecular formula

Patterns of Chemical Reactions

- **Synthesis** (combination) reactions
- **Decomposition** reactions
- **Exchange** reactions

Synthesis Reactions

- A + B → AB
  - Atoms or molecules combine to form larger, more complex molecule
  - Always involve bond formation
  - **Anabolic**

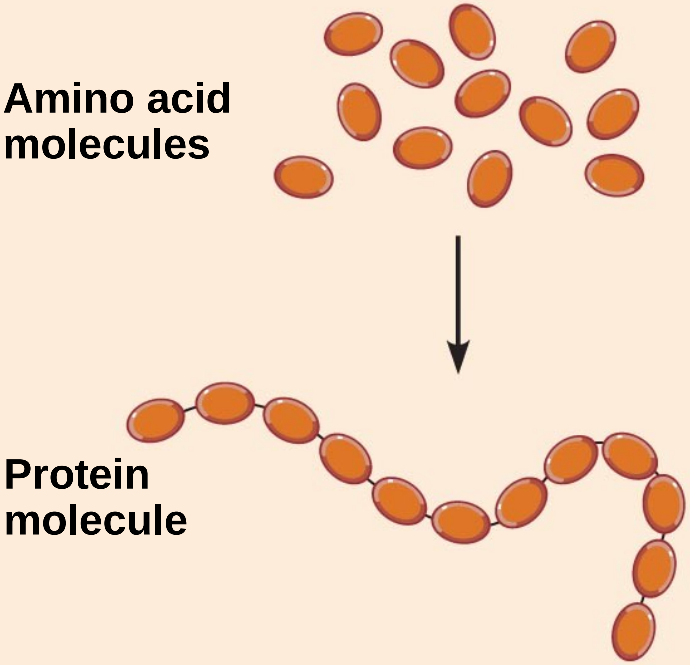
Figure 2.11a Patterns of chemical reactions.

(a) Synthesis reactions

Smaller particles are bonded together to form larger, more complex molecules.

Example

Amino acids are joined together to form a protein molecule.



Decomposition Reactions

- $AB \rightarrow A + B$
  - Molecule is broken down into smaller molecules or its constituent atoms
    - Reverse of synthesis reactions
  - Involve breaking of bonds
  - **Catabolic**

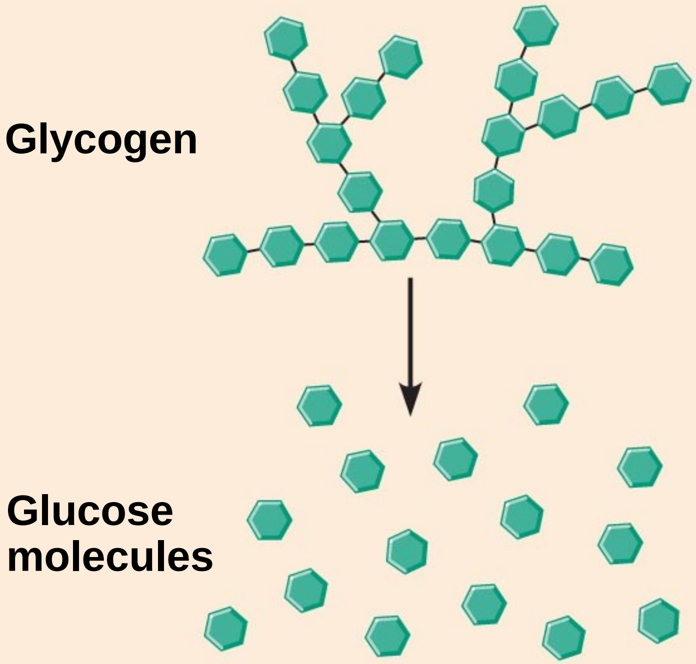
Figure 2.11b Patterns of chemical reactions.

(b) Decomposition reactions

Bonds are broken in larger molecules, resulting in smaller, less complex molecules.

Example

Glycogen is broken down to release glucose units.



Exchange Reactions

- $AB + C \rightarrow AC + B$
  - Also called displacement reactions
  - Involve both synthesis and decomposition
  - Bonds are both made and broken

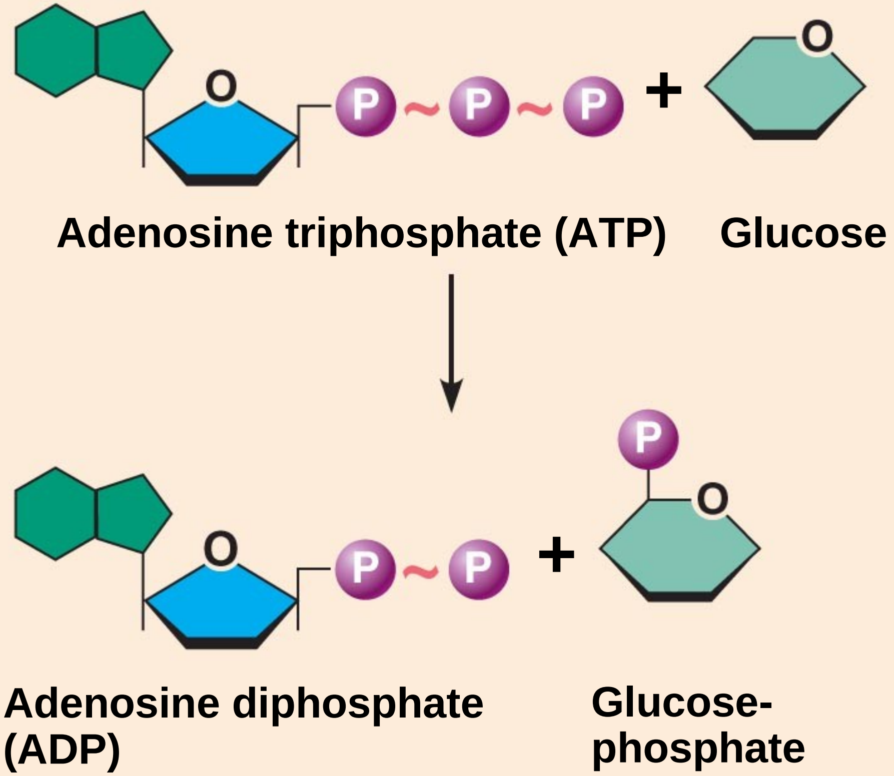
Figure 2.11c Patterns of chemical reactions.

(c) Exchange reactions

Bonds are both made and broken (also called displacement reactions).

Example

ATP transfers its terminal phosphate group to glucose to form glucose-phosphate.



Oxidation-Reduction (Redox) Reactions

- Are decomposition reactions
  - Reactions in which food fuels are broken down for energy
- Are also exchange reactions because electrons are exchanged between reactants
  - Electron donors lose electrons and are oxidized
  - Electron acceptors receive electrons and become reduced
- $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + ATP$
- Glucose is oxidized; oxygen molecule is reduced

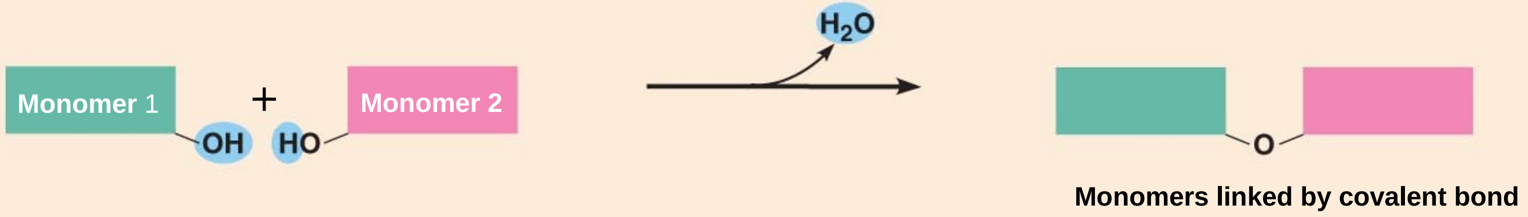
Rate of Chemical Reactions

- Affected by
  - ↑ Temperature → ↑ Rate
  - ↑ Concentration of reactant → ↑ Rate
  - ↓ Particle size → ↑ Rate
  - Catalysts: ↑ Rate without being chemically changed or part of product
    - Enzymes are biological catalysts

Figure 2.14 Dehydration synthesis and hydrolysis.

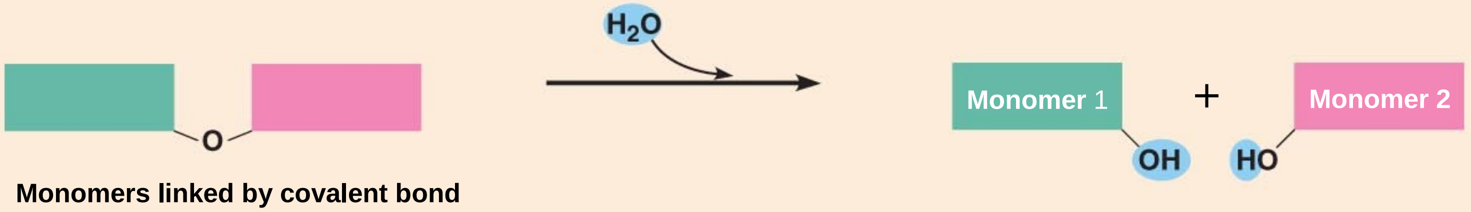
(a) Dehydration synthesis

Monomers are joined by removal of OH from one monomer
and removal of H from the other at the site of bond formation.



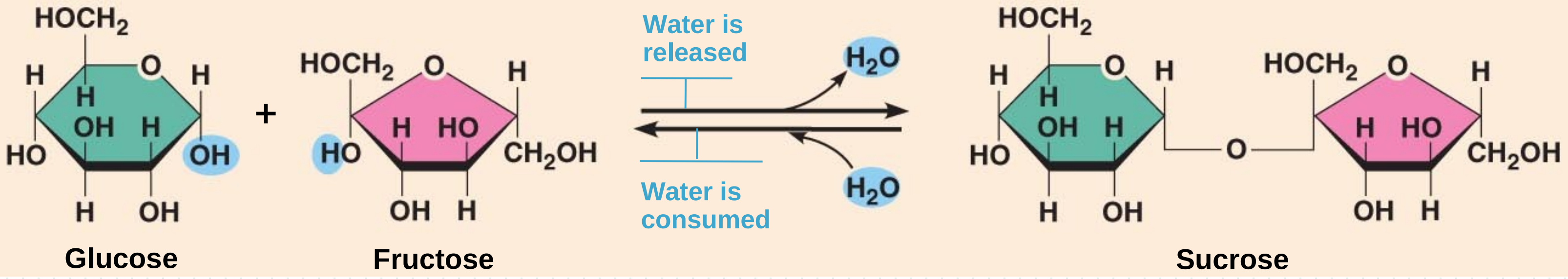
(b) Hydrolysis

Monomers are released by the addition of a water molecule, adding OH to one monomer and H to the other.



(c) Example reactions

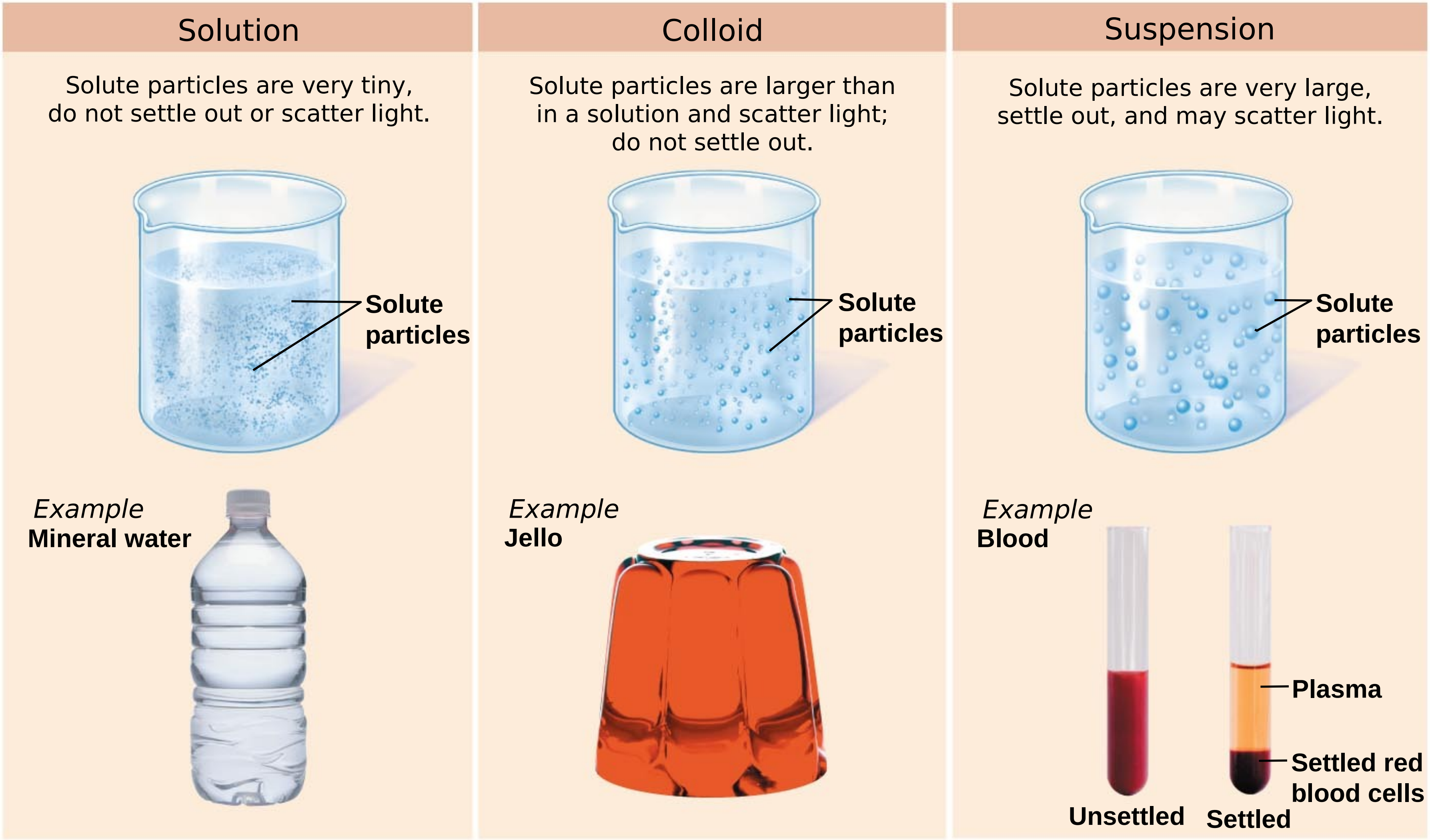
Dehydration synthesis of sucrose and its breakdown by hydrolysis



Mixtures

- Most matter exists as **mixtures**
  - Two or more components physically intermixed
- Three types of mixtures
  - **Solutions**
  - **Colloids**
  - **Suspensions**

Figure 2.4 The three basic types of mixtures.



Solution

Solute particles are very tiny, do not settle out or scatter light.

Colloid

Solute particles are larger than in a solution and scatter light; do not settle out.

Suspension

Solute particles are very large, settle out, and may scatter light.

Types of Mixtures: Solutions

- Homogeneous mixtures
- Most are true solutions in body
  - Gases, liquids, or solids dissolved in water
  - Usually transparent, e.g., atmospheric air or saline solution
- **Solvent**
  - Substance present in greatest amount
  - Usually a liquid; usually water
- **Solute**(s)
  - Present in smaller amounts
- Ex. If glucose is dissolved in blood, glucose is solute; blood is solvent

Concentration of True Solutions

- Can be expressed as
  - Percent of solute in total solution (assumed to be water)
    - Parts solute per 100 parts solvent
  - Milligrams per deciliter (mg/dl)
  - Molarity, or moles per liter (M)
    - 1 mole of an element or compound = Its atomic or molecular weight (sum of atomic weights) in grams
    - 1 mole of any substance contains 6.02×10^{23} molecules of that substance (Avogadro's number)

Colloids and Suspensions

- **Colloids** (AKA emulsions)
  - Heterogeneous mixtures, e.g., cytosol
  - Large solute particles do not settle out
  - Some undergo sol-gel transformations
    - e.g., cytosol during cell division
- **Suspensions**
  - Heterogeneous mixtures, e.g., blood
  - Large, visible solutes settle out

Mixtures versus Compounds

- Mixtures
  - No chemical bonding between components
  - Can be separated by physical means, such as straining or filtering
  - Heterogeneous or homogeneous
- Compounds
  - Chemical bonding between components
  - Can be separated only by breaking bonds
  - All are homogeneous