Attractive forces between particles

1) **Covalent bonds**: involve the sharing of electrons between two or more atoms.
   - **Strength**: very strong, ranging from 200 kJ/mol to 1000 kJ/mol
   - **Seen in**: all molecular compounds, all network solids, all polyatomic ions

2) **Ion-ion attraction ("ionic bonds")**: the electrostatic attraction between oppositely charged ions.
   - **Strength**: very strong, ranging from 400 kJ/mol to 3000 kJ/mol for an individual pair of ions. (The overall lattice energies are larger than these numbers, because each ion is surrounded by several ions of the opposite charge.)
   - **Seen in**: all ionic compounds

3) **Metallic bonds**: involve the sharing of electrons throughout a set of metal atoms.
   - **Strength**: moderate to very strong, ranging from 70 kJ/mol to 800 kJ/mol
   - **Seen in**: all metals

The following three types of attraction are referred to collectively as “intermolecular forces”

4) **London dispersion forces (LDF’s)**: involve the random fluctuation of electron density in a molecule, creating dipoles which attract one another. The larger the atoms in a molecule, the stronger the LDF’s will be. In addition, the more atoms a molecule has, the stronger the LDF’s will be (if you keep the same elements).
   - **Strength**: weak to very weak, ranging from <1 kJ/mol to 20 kJ/mol.
   - **Seen in**: all molecular substances (and the inert gases)

5) **Dipole-dipole attraction**: involve the attraction of two molecules that have permanent dipoles.
   - **Strength**: weak to very weak, comparable to LDF’s
   - **Seen in**: all polar molecular substances

6) **Hydrogen-bonds**: involve the attraction between a molecule that has a N-H, O-H or F-H bond (which will be highly polar) and a second molecule that has a negatively polarized N or O.
   - **Strength**: weak, ranging from 5 kJ/mol to 40 kJ/mol.
   - **Seen in**: all substances that contain at least one N-H, O-H or F-H bond.

Note: hydrogen-bonded substances typically show noticeably higher melting and boiling points than substances that have similar composition but no ability to hydrogen-bond.
Types of solids

1) Molecular solid

Properties: low melting and boiling points (typically melt below 300°C, although higher temperatures are seen for very large compounds). Many are liquids or gases at room temperature. Solubilities vary widely: some dissolve in water, and many can dissolve in nonpolar liquids (no other type of substance dissolves in nonpolar solvents). Do not conduct electricity, and conduct heat poorly.

Bonding-intramolecular: covalent bonds

Bonding-intermolecular: London dispersion forces (all molecular solids). Polar substances also show dipole-dipole attraction. Substances with N-H, O-H or F-H bonds show hydrogen-bonding (a special case of dipole-dipole). All of these forces are relatively weak.

How to recognize: first, look at the physical properties, particularly the melting point and solubility. The chemical formula provides a good clue, too: any substance that is made entirely from nonmetals should be assumed to be molecular unless proven otherwise. This includes acids: all acids are molecular substances.

Examples: H₂O, CO₂, NH₃, CH₄, HCl, H₂SO₄

2) Ionic solid

Properties: high melting points and very high boiling points (typical ionic solids melt above 500°C and boil above 1000°C). All are solids at room temperature. Many dissolve well in water, and aqueous solutions conduct electricity well. Ionic compounds conduct heat fairly well. They do not conduct electricity when they are solid, but they conduct electricity very well when melted. Ionic compounds are normally crystalline in appearance, and the crystals are very brittle: they shatter when struck.

Bonding: the ions are attracted to one another by ion-ion attraction. (Many ionic compounds contain polyatomic ions such as CO₃²⁻ or NH₄⁺, which are held together by covalent bonds.)

How to recognize: any compound that contains a metal and one or more nonmetals should be assumed to be ionic until proven otherwise. A few seemingly ionic compounds are actually molecular: these are easily recognized by their low melting points. An example is AlBr₃, which looks ionic (Al is a metal and Br is a nonmetal) but melts at 97°C (far too low for an ionic compound).

Examples: NaCl, KNO₃, NaHCO₃, (NH₄)₂SO₄
3) Metallic solid

Properties: metals normally have high melting and boiling points, but there is a very wide range, from mercury (melting point -39°C) to tungsten (melting point 3410°C). Mercury is the only metal that is a liquid at room temperature: all other metals are solids. Metals are insoluble in all common solvents: they can only be dissolved in strong acids, which convert the metal into an ionic compound. Metals conduct heat and electricity very well. Most metals are silvery or gray (the exceptions are gold and copper), and all are reflective when polished. Metals are malleable (they deform when struck, rather than shattering) and ductile (they can be stretched into wires).

Bonding: metallic bonds. These “metallic bonds” are actually a special case of covalent bonding, in which the valence electrons of the metal atoms are delocalized over the entire sample of metal.

How to recognize: see if the element lies in the metal region of the periodic table! A substance that contains nothing but metallic elements is normally a metal: any substance that contains a nonmetal must be something else. A few elements have two forms, one of which is metallic (these elements are found on the border between the metals and nonmetals on the periodic table, and are called metalloids). To confirm that a particular form of an element is metallic, look for high electrical and thermal conductivity, malleability, and ductility.

Examples: Fe, Au, Na, Al, brass (a mixture of Cu and Zn)

4) Network covalent solid

Properties: network solids have very high melting and boiling points. They normally melt above 500°C, and many melt well over 1000°C. All are solids at room temperature. They are often very hard, but this varies. Network solids are brittle, so they shatter when struck. They normally conduct heat reasonably well, but are poor conductors of electricity (graphite being the exception, as it conducts electricity well). Network solids are insoluble in all solvents.

Bonding: covalent bonds. A diamond or a piece of quartz (SiO$_2$) is actually a single enormous molecule.

How to recognize: in general, if a substance is made entirely from nonmetals, it will be either molecular or network covalent: these are easily distinguished by their melting points.

Examples: C (diamond or graphite), B, SiO$_2$