KEY

Atomic Structure and Periodicity

1. Calculate the energy of electromagnetic radiation with a wavelength of 310 nm.

$c = \lambda\nu$

$E = h\nu$

$$\nu = \frac{c}{\lambda} = \frac{2.998 \times 10^{8}\ m/s}{310\,nm \ \frac{1\ m}{1\times10^{9}\ nm}} = 9.67\times10^{14}\ s^{-1} \rightarrow Hz$$

$$E = h\nu = (6.626\times10^{-34}\ Js)(9.67\times10^{14}\ s^{-1}) = \boxed{6.41\times10^{-19}\ J/photon}$$

2. Carbon absorbs energy at a wavelength of 150. nm. The total amount of energy emitted by a carbon sample is 1.98×10^{5} J. Calculate the number of carbon atoms in the sample, assuming that each atom emits one photon.

$c = \lambda\nu$

$E = h\nu$

$$\nu = \frac{c}{\lambda} = \frac{2.998\times10^{8}\ m/s}{150\,nm\ \frac{1\ m}{1\times10^{9}\ nm}} = 2.00\times10^{15}\ s^{-1}$$

$$E = h\nu = (6.626\times10^{-34}\ Js)(2.00\times10^{15}\ s^{-1}) = 1.32\times10^{-18}\ J/photon$$

$$\#photons = \frac{1.98\times10^{5}\ J}{1.32\times10^{-18}\ J/photon} = \boxed{1.5\times10^{23}\ photons\ (C\ atoms)}$$

3. A carbon-carbon double bond in a certain organic molecule absorbs radiation that has a frequency of $6.0 \times 10^{13} s^{-1}$.

   a. What is the wavelength of this radiation?

$c = \lambda\nu$

$$\lambda = \frac{2.998\times10^{8}\ m/s}{6.0\times10^{13}\ s^{-1}} = 5.00\times10^{-6}\ m = \boxed{5.00\times10^{3}\ nm}$$

   b. What is the energy of this radiation per photon? Per mole of photons?

$$E = h\nu = (6.626\times10^{-34}\ Js)(6.0\times10^{13}\ s^{-1}) = \boxed{3.98\times10^{-20}\ J/photon}$$

$$(3.98\times10^{-20}\ J/photon)(6.022\times10^{23}\ photons/mole) = \boxed{2.39\times10^{4}\ J/mole}$$

4. Use the periodic table and other information concerning atomic structure and bonding to explain the following observations:

   a. The radii of Cl^{-} ions are larger than the radii of Cl atoms.

- ***There is greater electron - electron repulsion in Cl^{-} than Cl because Cl^{-} has more electrons in the valence shell***
- ***Also, the electrons outnumber the protons in Cl^{-} which reduces the effective nuclear charge compared to Cl atoms.***

   b. The electron affinity of fluorine is higher than the electron affinity of oxygen.

- ***Electron affinity is the change in energy when an electron is added to an atom. When an electron is added to fluorine, more energy is released than when an electron is added to oxygen because fluorine has a greater nuclear charge and a greater attraction for the electron.***

c. Explain the trend in electronegativity from P to S to Cl.

- ***As you move across the period from P to S to Cl, the effective nuclear charge increases so the attraction for electrons increases as well. (An increase in the effective nuclear charge is seen when there is an increase in the number of protons without an increase in shielding effect.)***

d. Explain the trend in electronegativity from Cl to Br to I.

- ***As one goes down the group from Cl to Br to I, more energy levels are added which increases the shielding effect. The results in a decreasing attraction for electrons down the group.***

e. Rank the following atoms from smallest to largest and explain your choice.
Al, Mg, Na, Si

- ***Si<Al<Mg<Na Atomic radius decreases from left to right across a period because of increasing effective nuclear charge across the period.***

f. Place the following in order of increasing first ionization energy and explain your choice. Mg, Ca, Si, Al

- ***Expected: Ca<Mg<Al<Si Actual: Ca<Al<Mg<Si***

  Ionization energy decreases going down a group because of increasing shielding effect; ionization energy increases across a period because of increasing effective nuclear charge. The IE of Al is slightly lower than Mg due to the fact that p electrons are shielded by s electrons.

g. Why is the second ionization energy of Ca greater than the first ionization energy of Ca?

- ***After the removal of the first electron, the protons outnumber the electrons (effective nuclear charge is increased), so the second electron is bound more tightly to the nucleus and requires more energy to remove.***

5. Correctly state the following in your own words.

a. Hund's Rule

- ***Every orbital in a sublevel is singly occupied with one electron before any one orbital is doubly occupied. All electrons in singly occupied orbitals have the same spin. Electrons will repel each other and will spread out across the orbitals if possible. (Note: Paramagnetism is when an atom, molecule or ion is weakly attracted in a magnetic field because of the presence of at least one unpaired electron.)***

b. Heisenberg Uncertainty Principle

- ***The position and momentum of an electron cannot be simultaneously known to high precision. The more precisely on property is measured, the less precisely the other can be measured.***

c. Pauli Exclusion Principle

- ***No two electrons in the same atom can be in the same quantum state. This means that no two electrons can have the same set of quantum states: 1)energy, 2) sublevel, 3) orbital, 4) spin. The number of electrons that can occupy each orbital is limited by the Pauli Exclusion Principle.***

d. Aufbau Principle

- ***Electrons fill orbitals starting at the lowest available energy states before filling higher states.***

6. A. What are isotopes?

Isotopes are atoms of the same element which have different numbers of neutrons (different atomic mass).

B. The average atomic mass of rhenium is 186.2. Given that 37.1% of natural rhenium is rhenium-185, what is the other stable isotope?

$186.2 = (.371 \times 185) + (.629 \times X)$

$186.2 = 68.64 + .629X$

$117.6 = .629X$

$X = 186.9$

$\boxed{Rh-187}$

7. Complete the table.

Element	Electron configuration	Most common oxidation state	Lewis Structure
Calcium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	Ca^{2+}	Ca :
Iron	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ or $[Ar] 4s^2 3d^6$	Fe^{2+}	· Fe :
Sulfur	$1s^2 2s^2 2p^6 3s^2 3p^4$	S^{2-}	· S :

Bonding and Intermolecular Forces

1. A. An ionic compound is one in which the attractions between atoms is electrostatic and obeys Coulomb's Law. What is Coulomb's Law (the equation)? What do the symbols in the equation represent?

Coulomb's Law is $F = k(q_1q_2/r^2)$ or simply the relationship $F = q_1q_2/r$, where q_1 and q_2 are the ion charges and r is the internuclear distance.

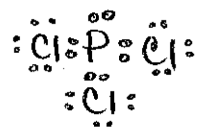
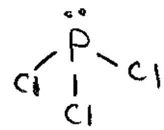
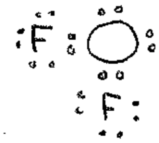
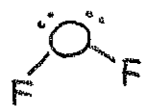
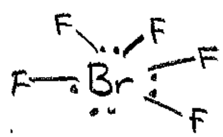
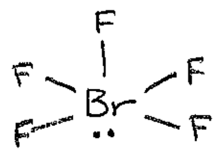
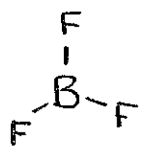
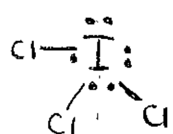
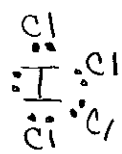
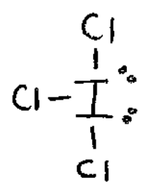
B. i) Use Coulomb's Law to explain the difference in energy required to break up a CaO(s) crystal into ions compared to a KF(s) crystal.

- ***The positive ions, Ca^{2+} and K^+, are similar in radius with Ca^{2+} being slightly smaller due to its greater nuclear charge. The negative ions, O^{2-} and F^-, have similar radii with F^- being slightly smaller due to greater nuclear charge. Therefore the value for r in CaO is similar to the value of r in KF.***
- ***Ca and O have 2+ and 2- charges, compared to K and F which are 1+ and 1-. As the charges are greater in CaO, the attractive force is greater and so more energy is required to break the CaO bond than the KF bond.***

ii) Use Coulomb's Law to explain the difference in energy required to break up a KF(s) crystal compared to a NaF(s) crystal.

- ***K and Na both form 1^+ ions so the magnitude of the Coulomb's law numerator is the same in both KF and NaF. K has one more energy level than Na, so the Coulomb's law denominator is smaller in NaF making the attractive force stronger in NaF than in KF. The NaF crystal will require more energy to break.***

2. Complete the following table:

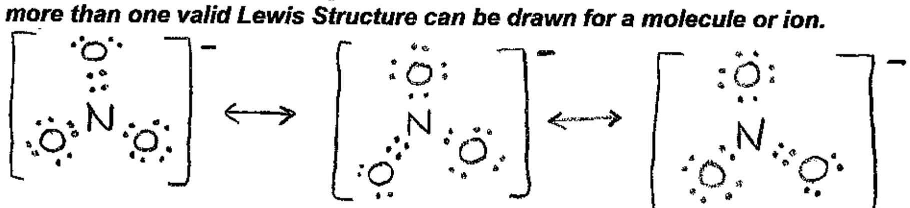
Compound (Needs, Available, Shared – if you need to)	Lewis structure	Structural formula (mimics shape)	Name of the shape of the molecule	Hybridization of the central atom
PCl_3			trigonal pyramid	sp^3
OF_2			bent	sp^3
BrF_5  Six areas of e^- density			Square pyramid	d^2sp^3
BF_3			trigonal planar	sp^2
ICl_3  five areas of e^- density			T-shaped	dsp^3

3. A. The nitrate ion exhibits resonance. What is resonance? Draw the resonance Lewis structures for the polyatomic ion.

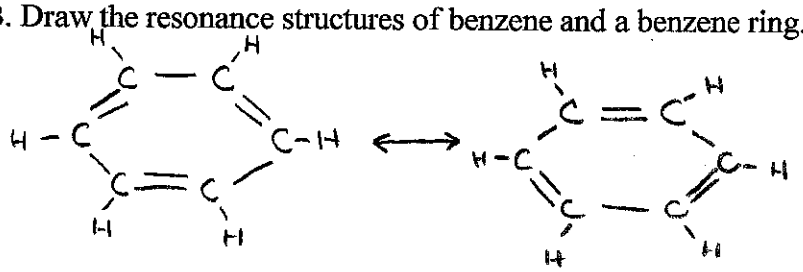
Resonance is when more than one valid Lewis Structure can be drawn for a molecule or ion.

NO_3^-

N 32
A 24
S 8
NS 16



B. Draw the resonance structures of benzene and a benzene ring.



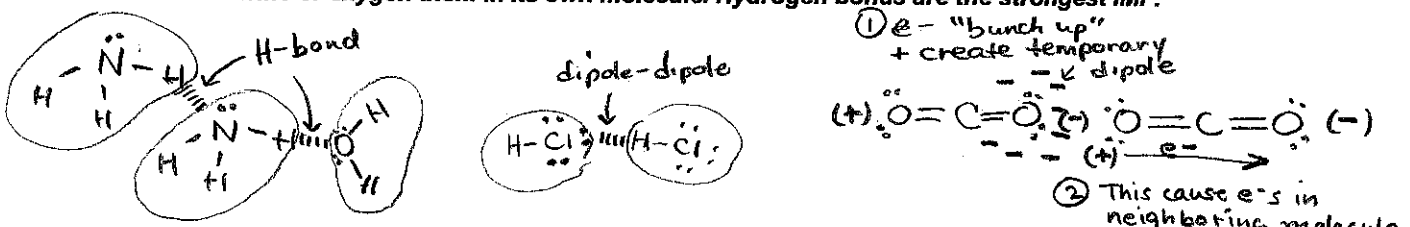
Benzene ring

4. Intermolecular forces are the forces of attraction between molecules. What are the three types of intermolecular forces? Which is weakest? Which is strongest? When and how are each formed? Give an example of a substance that forms each. Draw a sketch of how and where the attractions would form.

- ***London Dispersion Forces are temporary attractive forces between atoms or molecules as a result of an unsymmetrical distribution of electrons. Because of the constant motion of electrons, an atom or molecule can develop a temporary (instantaneous) dipole. Larger and heavier particles exhibit stronger dispersion forces as they have greater polarizability due to a greater number of electrons. LDF's occur in all atoms and molecules and are the weakest IMF.***

- ***Dipole-Dipole Forces are attractive forces between the positive end of one polar molecule and the negative end of another polar molecule. These forces are typically stronger than LDFS.***

- ***Hydrogen Bonds are attractive forces between the hydrogen of one molecule and a nitrogen, fluorine or oxygen in another molecule. The hydrogen must be covalently bonded to a nitrogen, fluorine or oxygen atom in its own molecule. Hydrogen bonds are the strongest IMF.***



5. Use the principles of bonding and molecular structure to explain the following statements:

A. The boiling point of argon is -186°C, whereas the boiling point of neon is -246°C.
Ar and Ne are noble gases and have only LDFs as IMFs. Since Ar has more electrons than Ne, Ar has greater polarizability (potential for electrons to become unevenly distributed) and therefore forms more LDFs which gives it a higher boiling point than Ne.

B. Solid sodium melts at 98°C, but solid potassium melts at 64°C.
Na and K are metals and have metallic bonds as the primary attractive force. Na atoms have a smaller radius than K atoms and so the delocalized electrons are closer to the nuclei in Na. This increases the strength of the metallic bond of Na and gives Na a higher boiling point than K.

C. HCl has a lower boiling point than either HF or HBr.
The order of boiling points is HF>HI>HBr>HCl. HF forms H-bonds so has the greatest bp. The other three exhibit dipole-dipole attractions. While HCl does have the largest magnitude dipole, current logic is that the London dispersion forces of HI and HBr contribute enough attractions to raise the bp of these compounds. HI has more electrons than HBr. HBr has more electrons than HCl. This gives HCl the weakest LDFs. (Note: Dipole-dipole forces and LDFs are collectively referred to as van der Waals forces.)

6. Matching.

C	1. Composed of macromolecules held together by strong bonds.	5	A. Molecular solid containing polar molecules
D	2. Composed of atoms held together by delocalized electrons.	4	B. Molecular solid containing nonpolar molecules
E	3. Composed of positive and negative ions held together by electrostatic forces.	1	C. Covalent network solid
B	4. Composed of molecules held together by London Dispersion forces.	2	D. Metallic solid
A	5. Composed of molecules held together by dipole-dipole forces.	3	E. Ionic solid

Solutions

1. What is the molarity of a solution when 1.00 g of HCl is completely dissolved in water to make 500. mL of solution? What is the pH of the solution?

$$\frac{1.00\,g\,HCl}{.500\,L\,soln} \times \frac{1\,mol\,HCl}{36.46\,g\,HCl} = \boxed{.0549\,M\,HCl}$$

$$\underset{.0549\,M}{HCl} \xrightarrow{100\%} \underset{.0549\,M}{H^+} + Cl^-$$

$$pH = -\log[H^+] = -\log .0549 = \boxed{1.260}$$

2. 1.00 g of ethanol is mixed with 100.0 g of pure water. What is the mole fraction of the ethanol?

$$\frac{1.00\,g\,C_2H_5OH}{1} \times \frac{1\,mole}{46.08\,g} = .0217\,moles$$

$$\frac{1.00\,g\,H_2O}{1} \times \frac{1\,mole}{18.02\,g} = 5.549\,moles$$

$$\chi_{eth} = \frac{moles\ eth}{total\ moles} = \frac{.0217}{5.5707} = \boxed{.0390}$$

3. How would you prepare 200.0 mL of 0.100M NaCl solution from solid NaCl?

$$.100\,M\,NaCl = \frac{X\,moles}{.2000\,L}$$

$$X = .0200\,moles\,NaCl$$

$$\frac{.0200\,moles}{1} \times \frac{58.44\,g}{1\,mol} = \boxed{1.17\,g\,NaCl}$$

* Dissolve 1.17g NaCl in enough distilled water to make 200.0 mL of solution.

4. How would you prepare 1.0 L of .20M HNO_3 from an 18M stock solution?

$$M_1V_1 = M_2V_2$$

$$(1.0)(.20) = (18)(V_2)$$

$$V_2 = .0111\,L = \boxed{11.1\,mL\ HNO_3\ \text{stock solution}}$$

$$1000 - 11.1 = 988.9\,mL\,H_2O$$

* Measure out 988.9 mL of distilled water & slowly add 11.1 mL of 18M HNO_3 to it.

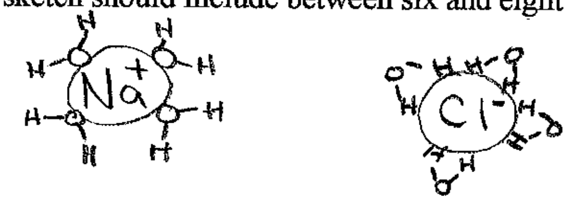
5. What are the three steps in making a solution and what are the energy requirements (endothermic or exothermic) for each step?

Step?	Endo or Exo?
Separation (expansion) of the solute	**Endothermic**
Separation (expansion) of the solvent	**Endothermic**
Attraction of solvent particles to solute particles	**Exothermic**

6. a) What is the enthalpy or heat of hydration? What happens during the hydration of ions?

Heat of hydration is the sum of the energy needed to expand the solvent and the energy released when the solvent-solute attractions are formed. When an ion is surrounded by water molecules it is referred to as a hydrated ion.

b) Draw a sketch that shows one sodium ion and one chloride ion after they are hydrated. The sketch should include between six and eight water molecules.



7. The heat of solution of NaBr is –1.0 kJ/mol. The lattice energy of NaBr is 735 kJ/mol. Determine the heat of hydration of NaBr .

Step 1 } Lattice E (ΔH_{LE})
Step 2
Step 3 } Hydration E (ΔH_{hyd})
Total ΔH_{soln}

$\Delta H_{soln} - LE = \Delta H_{hyd}$

$-1.0 - 735 = \boxed{-736 \text{ kJ/mol}}$

Gases

1. A 4.0 L elastic weather balloon travels from sea level, at 1.0 atm pressure, to a higher altitude, where the pressure is 0.20 atm. What is the new volume of the balloon?

$P_1V_1 = P_2V_2$

$(1.0)(4.0) = (.20)(V_2)$

$V_2 = \boxed{20.\text{ L}}$

2. A gas occupies 2.0 L at 300. K. What is the volume of the gas at 200. K, assuming the pressure is constant?

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ $\qquad \frac{2.0\text{ L}}{300\text{ K}} = \frac{V_2}{200\text{ K}}$ $\qquad V_2 = \boxed{1.3\text{ L}}$

3. A gas in a rigid container exerts 6.0 atm at 300.K. What is the pressure that the gas exerts at 500.K?

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$ $\qquad \frac{6.0}{300} = \frac{P_2}{500}$ $\qquad P_2 = \boxed{10.\text{ atm}}$

4. A rigid container holds a combination of nitrogen and oxygen gas at a total pressure of 2.4 atm. If the mole fraction of nitrogen gas is 0.16, what is the partial pressure exerted by the nitrogen gas?

$P_{N_2} = P_T X$ $\qquad P_{N_2} = (2.4)(.16) = \boxed{.38\text{ atm}}$

5. Suppose you were given 8.00 moles of a gas occupying a volume of 4.00 L at constant pressure and temperature. What volume of gas would 16.0 moles occupy at the same temperature and pressure?

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \qquad \frac{4.00}{8.00} = \frac{V_2}{16.0} \qquad V_2 = \boxed{8.00\ L}$$

6. What is the pressure exerted by 3.0 moles of gas a 17.0°C in a 2.0 liter container?

$$PV = nRT$$

$$P = \frac{nRT}{V} = \frac{(3.0)(.08206)(290)}{2.0} = \boxed{36\ atm}$$

7. A gas with a density of 1.67g/L exerts a pressure of 2.0 atm at a temperature of 299 K. What is the molar mass of the gas?

$$MM = \frac{dRT}{P} = \frac{(1.67\ g/L)(.08206)(299)}{2.0} = \boxed{20.\ g/mole}$$

8. A mixture of helium and carbon dioxide form a mixture in a rigid container. A small leak is created in the container; how much faster will the helium exit the container than the carbon dioxide?

$$\frac{r_{He}}{r_{CO_2}} = \sqrt{\frac{M_{CO_2}}{M_{He}}} = \sqrt{\frac{44.01}{4.003}} = \boxed{3.316 \times \text{faster}}$$

9. Ideal gases are gases that behave according to the following assumptions:
   a. The volume of the gas molecule is negligible compared to the space between the molecules.
   b. There is negligible intermolecular attraction between the gas molecules.

Most of the time gases behave as ideal gases, but under extreme conditions gases can deviate from ideal behavior. Under what conditions do gases deviate from ideal behavior?

Low Temperatures and High Pressures

$n = \frac{PV}{RT} = \frac{(2.0)(12.8)}{(.08206)(300)} = 1.0$ moles

MM = 159.70g

Stoichiometry and Types of Reactions

1. A. In the following reaction, a 12.8 L sample of CO at 2.0 atm and 27°C is combined with 6.33 g of Fe_2O_3. How many liters of carbon dioxide are formed at the same temperature, and which is the limiting reactant?

Fe_2O_3
$\frac{6.33g}{159.70 g/mol} = .0396$ mol

R	$Fe_2O_3(s)$ +	$3CO(g)$ →	$2Fe(s)$ +	$3CO_2(g)$
I	.0396	1.0	0	0
C	−.0396	−.1188	+.0792	+.1188
E	0	.8812	.0792	.1188

$\boxed{Fe_2O_3 \text{ is L.R.}}$

$V = \frac{nRT}{P} = \frac{(.1188)(.08206)(300)}{2.0}$

$\boxed{V = 1.5 L}$

B. How many grams of the excess reactant remain?

$\frac{.8812 \text{ moles CO}}{1} \times \frac{28.01 g \text{ CO}}{1 \text{ mol CO}} = \boxed{24.68 g \text{ CO}}$

2. A. What is the empirical formula of the hydrocarbon that contains 85.7% carbon?

$\frac{85.7gC}{1} \times \frac{1 \text{ mol}}{12.01g} = \frac{7.14}{7.14} = 1$

$\frac{14.3gH}{1} \times \frac{1 \text{mol}}{1.01g} = \frac{14.16}{7.14} \approx 2$

$\boxed{CH_2}$

B. Vapor pressure calculations determine that the molar mass of the hydrocarbon is 28.0 g/mole. What is the molecular formula of the compound?

$CH_2 \rightarrow 14$ g/mol

MM = 28.0 g/mol

$28 \div 14 = 2$

$\boxed{C_2H_4}$

3. How do you calculate the percent yield of a reaction?

$\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$

4. Write balanced equations for the following reactions. Include phases.

a. Solid sodium bicarbonate is gently heated

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

b. Solid sodium chlorate is gently heated

$2NaClO_{3(s)} \rightarrow 2NaCl_{(s)} + 3O_{2(g)}$

c. Solid sodium hydroxide is gently heated

$$2NaOH_{(s)} \longrightarrow Na_2O_{(s)} + H_2O_{(g)}$$

d. Solid sodium is placed in water

$$2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2\,(g)}$$

e. Chlorine gas is bubbled through a solution of potassium bromide

$$Cl_{2\,(g)} + 2KBr_{(aq)} \rightarrow 2KCl_{(aq)} + Br_{2\,(g)}$$

f. Sodium chloride solution is mixed with silver nitrate solution

$$NaCl_{(aq)} + AgNO_{3\,(aq)} \rightarrow NaNO_{3\,(aq)} + AgCl_{(s)}$$

g. Ethane (C_2H_6) is burned

$$C_2H_{6\,(g)} + O_{2\,(g)} \rightarrow 2CO_{2\,(g)} + 3H_2O_{(g)}$$

h. Hydrochloric acid solution is added to a solution of calcium hydroxide

$$2HCl_{(aq)} + Ca(OH)_{2\,(aq)} \longrightarrow CaCl_{2\,(aq)} + 2H_2O_{(l)}$$

i. Zinc metal is placed in a solution of copper (II) sulfate

$$Zn_{(s)} + CuSO_{4\,(aq)} \rightarrow ZnSO_{4\,(aq)} + Cu_{(s)}$$

j. Calcium oxide reacts with water

$$CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2\,(aq)}$$

k. Sulfur dioxide reacts with water

$$SO_{2\,(g)} + H_2O_{(l)} \longrightarrow H_2SO_{3\,(aq)}$$

l. Decompostion of carbonic acid

$$H_2CO_{3\,(aq)} \longrightarrow H_2O_{(l)} + CO_{2\,(g)}$$

m. Ammonia reacts with water $NH_{3\,(g)} + H_2O_{(l)} \rightleftarrows NH_4^+{}_{(aq)} + OH^-_{(aq)}$